The History of the Modern Periodic Table

Johann Dobereiner

In 1829, he classified some elements into groups of three, which he called triads. The elements in a triad had similar chemical properties and orderly physical properties.

(ex. Cl, Br, I and
Ca, Sr, Ba)

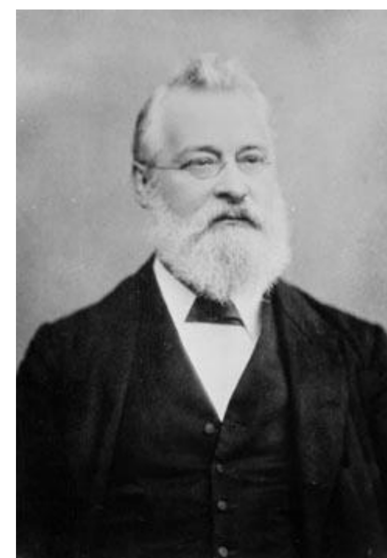
Model of triads

1780 - 1849

John Newlands

In 1863, he suggested that elements be arranged in "octaves" because he noticed (after arranging the elements in order of increasing atomic mass) that certain properties repeated every 8th element.

Law of Octaves



1838 - 1898

John Newlands

H 1	Li 7	Be 9	B 11	C 12	N 14	O 16
F 19	Na 23	Mg 24	Al 27	Si 28	P 31	S 32
Cl 35	K 39	Ca 40	Cr 52	Ti 48	Mn 55	Fe 56

Newlands' claim to see a repeating pattern was met with savage ridicule on its announcement. His classification of the elements, he was told, was as arbitrary as putting them in alphabetical order and his paper was rejected for publication by the Chemical Society.

1838 - 1898

Law of Octaves

John Newlands

His law of octaves failed beyond the element calcium. WHY?

Would his law of octaves work today with the first 20 elements?

1838 - 1898

Law of Octaves

During the nineteenth century, chemists began to categorize the elements according to similarities in their physical and chemical properties. The end result of these studies was our modern periodic table.

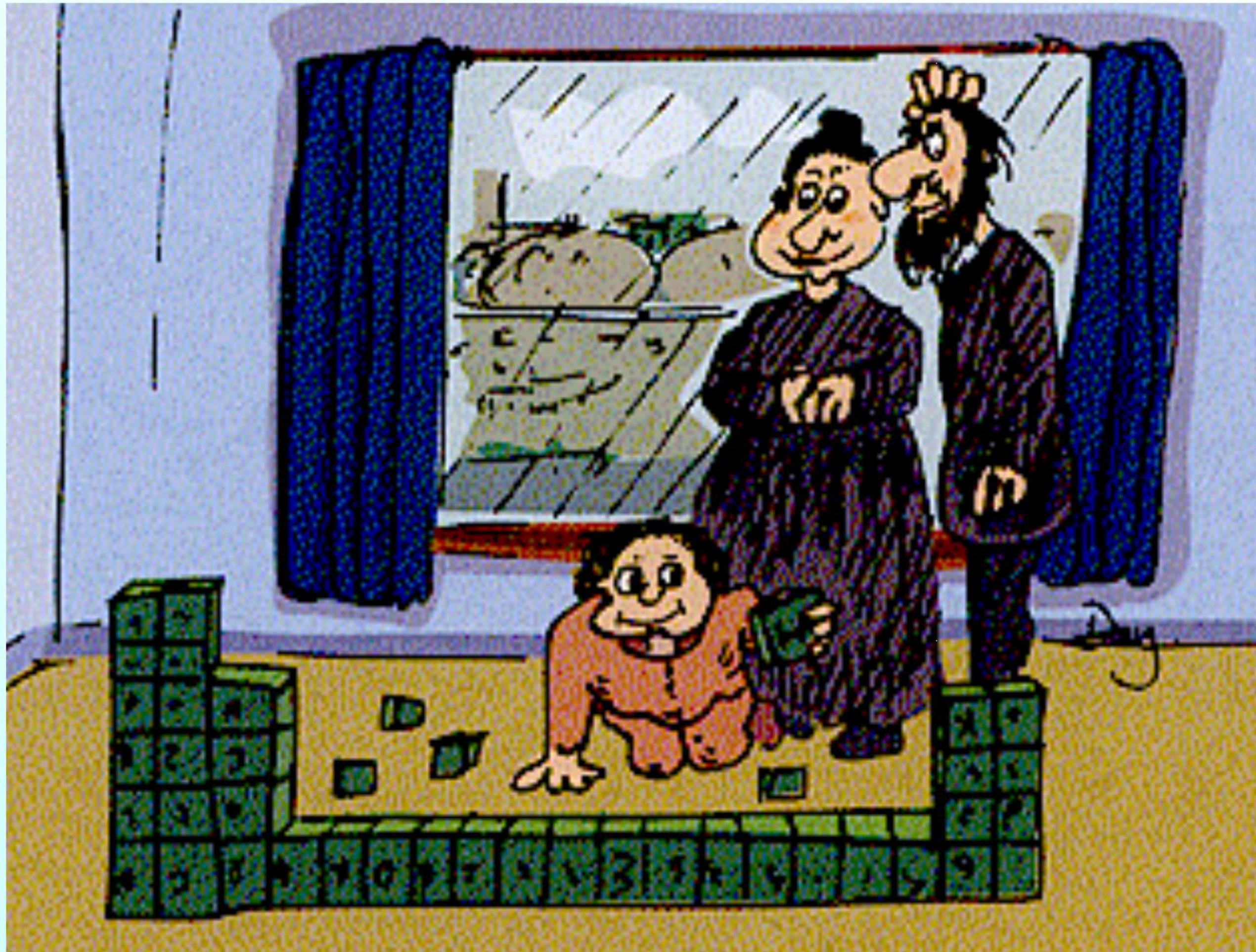
C. Dimitri Mendeleev

- **1869 came up with the eight – column table**
- **Because of blank spots on the table Mendelevium was able to predict properties and masses of several elements not yet discovered**
- **Table arranged in order of increasing atomic mass and the properties of the elements repeated in an orderly way (periodic law)**

As you know, the elements are arranged in the Periodic Table.

The elements were first arranged in this way by Dmitri Mendeleev, a professor at St. Petersburg University, in 1869. His arrangement was based on atomic mass.

When Mendeleev was setting out the table, only 63 elements had been discovered. His big idea was to leave gaps for yet to be discovered elements. He was able to predict the properties of some of these elements, including silicon and boron. When his predictions were shown to be accurate his table became accepted, and it is the basis of the one we use today.

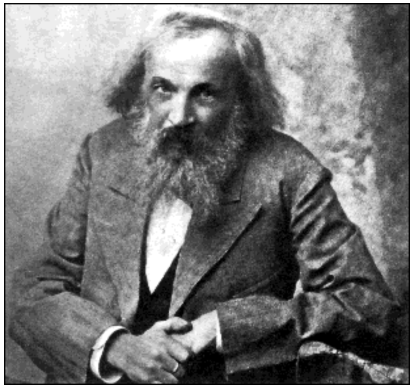


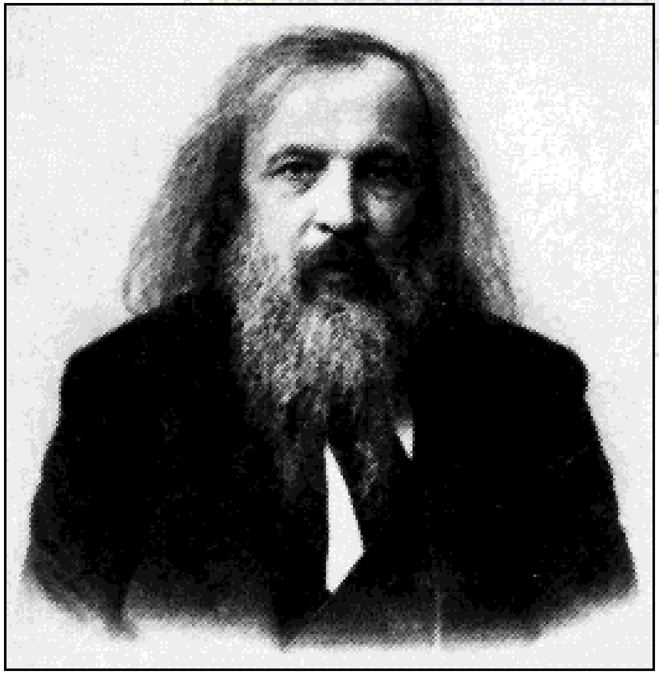
'Maybe one day we'll understand why Dmitri always lays out his blocks this way!'

Mike Turner, Jan. 2004

Dmitri Mendeleev

In 1869 he published a table of the elements organized by increasing atomic mass.

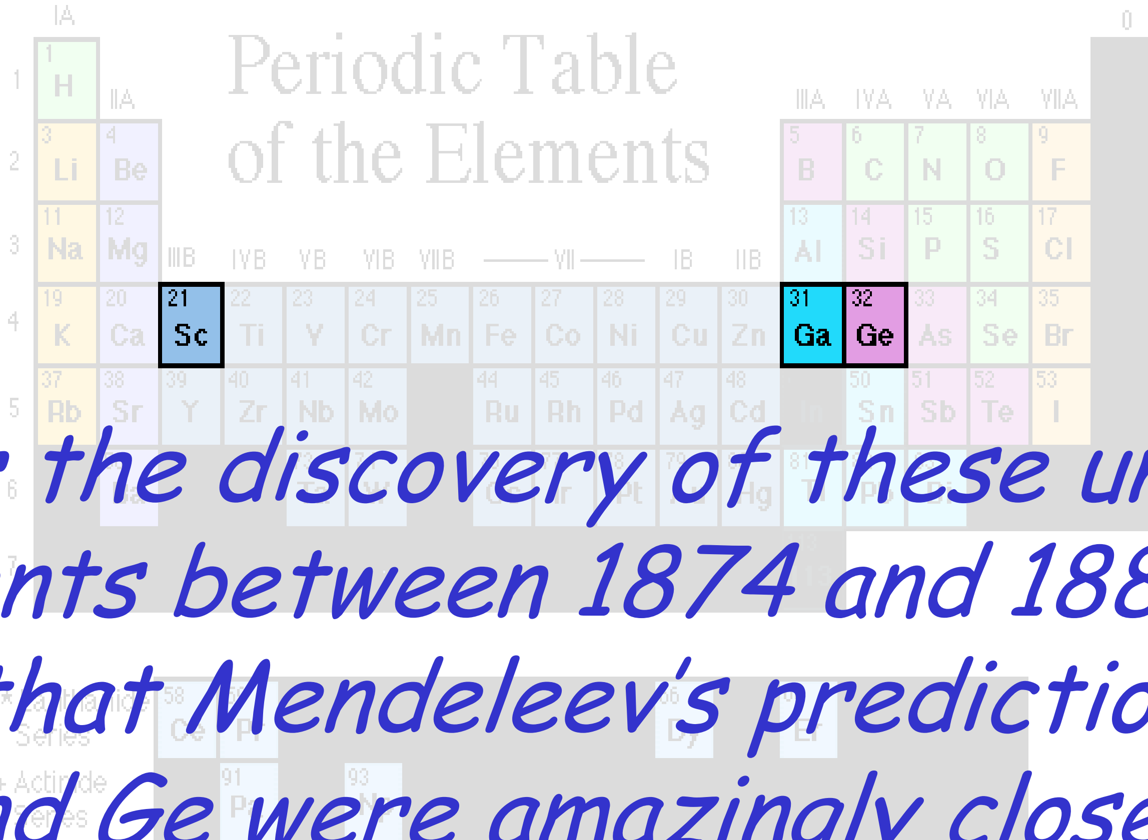




1834 - 1907

Mendeleev…

- stated that if the atomic weight of an element caused it to be placed in the wrong group, then the weight must be wrong. (He corrected the atomic masses of Be, In, and U)
- was so confident in his table that he used it to predict the physical properties of three elements that were yet unknown.



After the discovery of these unknown elements between 1874 and 1885, and the fact that Mendeleev's predictions for Sc, Ga, and Ge were amazingly close to the actual values, his table was generally accepted.

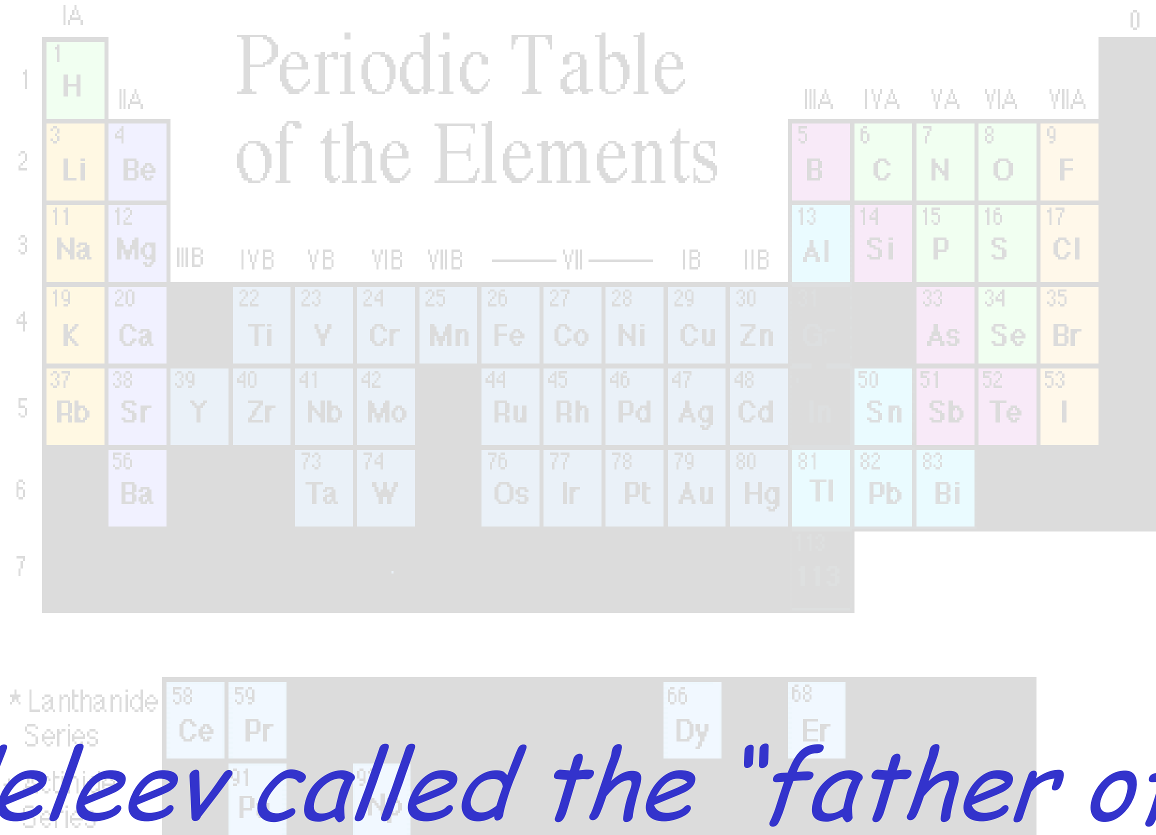
However, in spite of Mendeleev's great achievement, problems arose when new elements were discovered and more accurate atomic weights determined. By looking at our modern periodic table, can you identify what problems might have caused chemists a headache?

Ar and K

Co and Ni

Te and I

Th and Pa



Mendeleev called the "father of the modern periodic table"

Henry Moseley

In 1913, through his work with X-rays, he determined the actual nuclear charge (atomic number) of the elements*. He rearranged the elements in order of increasing atomic number.

*"There is in the atom a fundamental quantity which increases by regular steps as we pass from each element to the next. This quantity can only be the charge on the central positive nucleus."

1887 - 1915

Henry Moseley

His research was halted when the British government sent him to serve as a foot soldier in WWI. He was killed in the fighting in Gallipoli by a sniper's bullet, at the age of 28. Because of this loss, the British government later restricted its scientists to noncombatant duties during WWII.

Glenn T. Seaborg

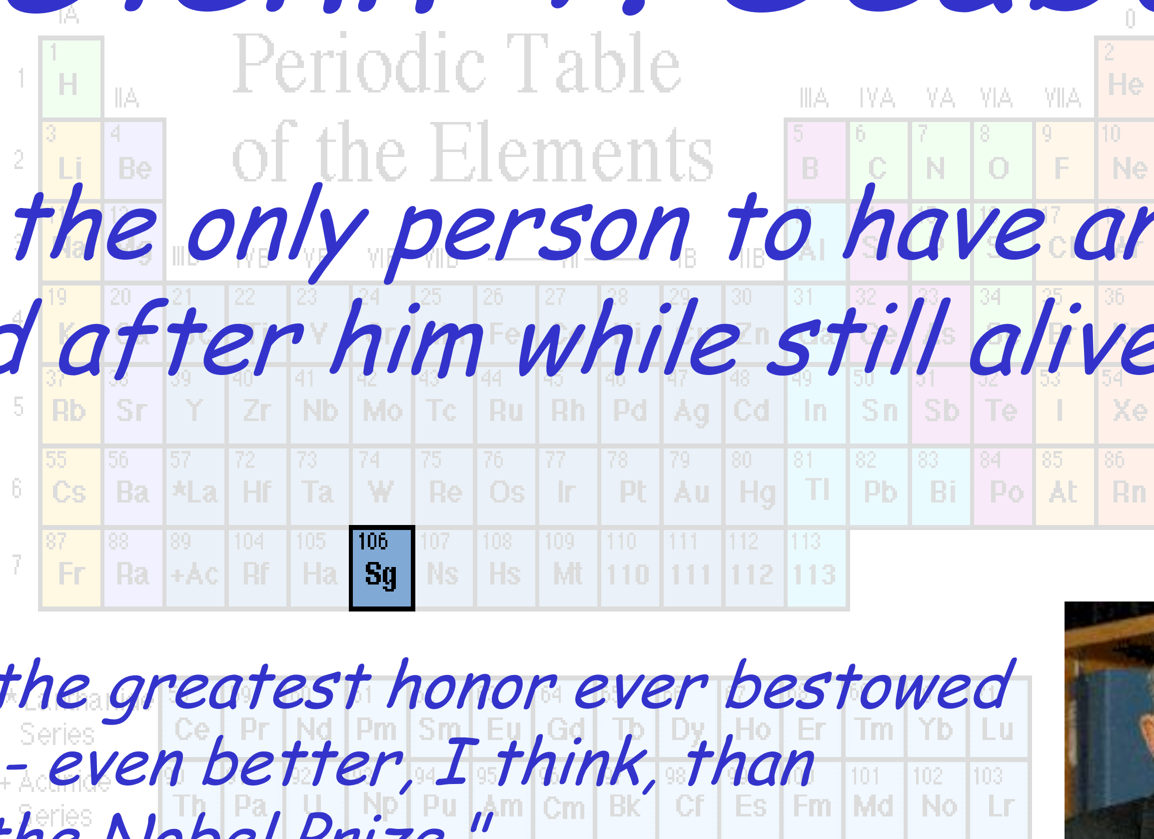
After co-discovering 10 new elements, in 1944 he moved 14 elements out of the main body of the periodic table to their current location below the Lanthanide series. These became known as the Actinide series.

+ Actinide Series

90	91	92	93	94	95	96	97	98	99	100	101	102	103
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

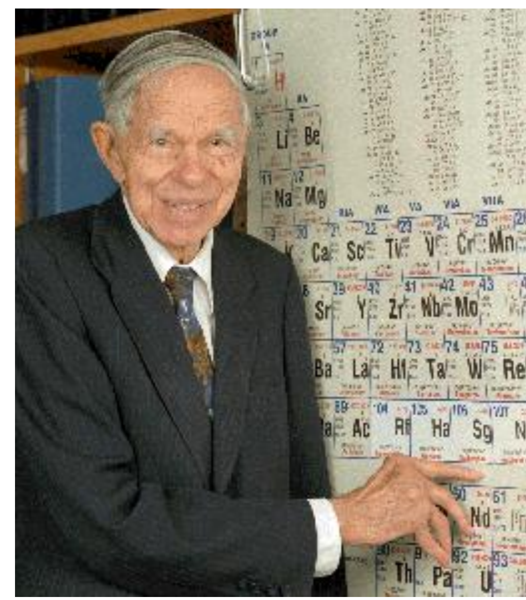
1912 - 1999

Glenn T. Seaborg



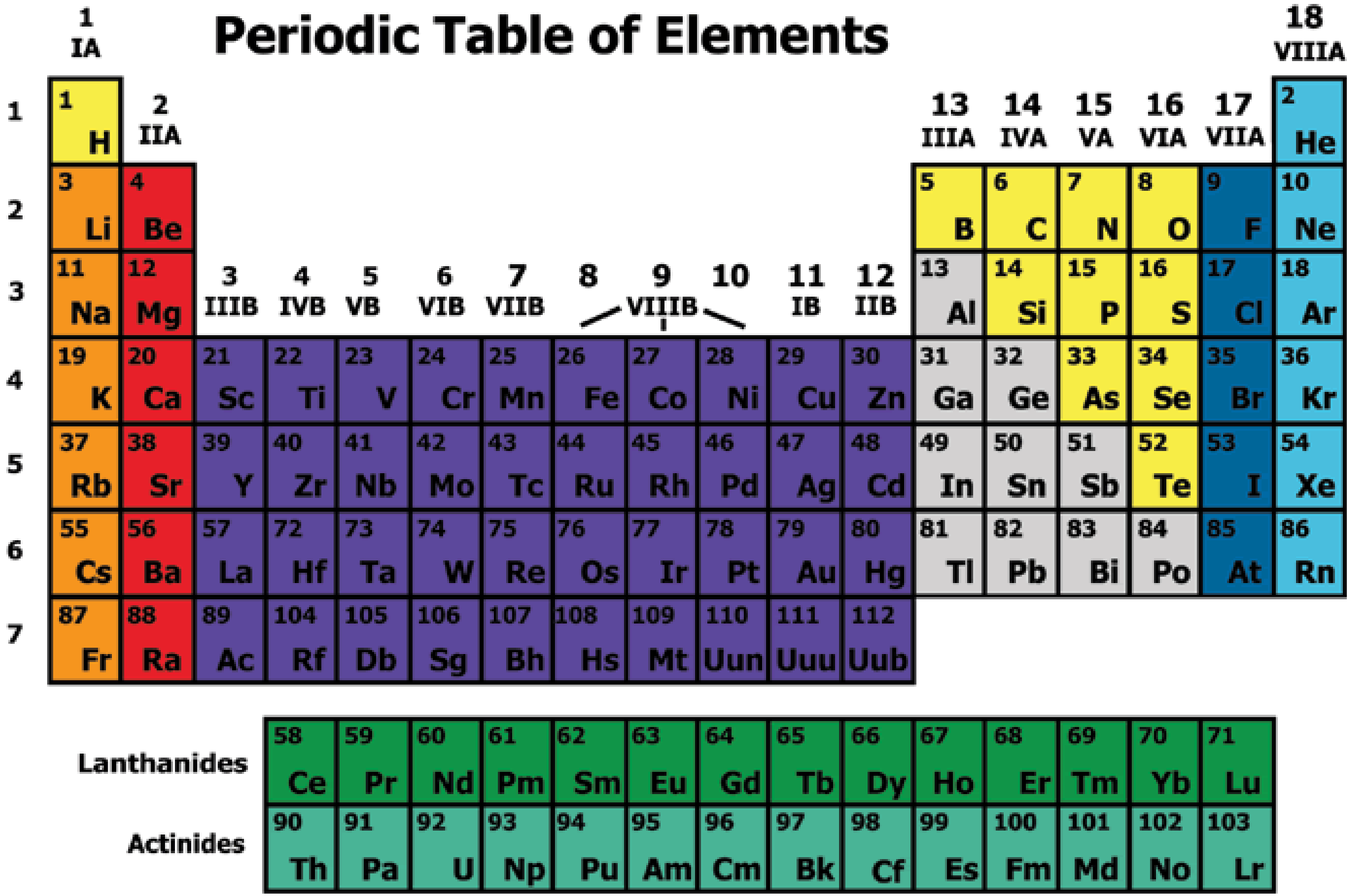
He is the only person to have an element named after him while still alive.

"This is the greatest honor ever bestowed upon me - even better, I think, than winning the Nobel Prize."



1912 - 1999

Periodic Table Geography



Periodic Law

When elements are arranged in order of increasing atomic number, there is a periodic pattern in their physical and chemical properties.

Modern periodic table –

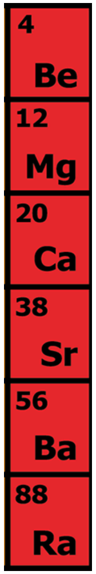
- **where certain electron configurations are periodically repeated**

Periodic properties

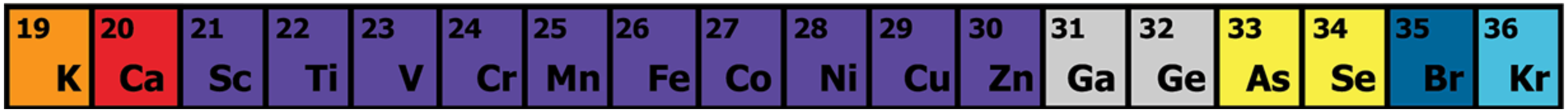
- Both the position and the properties arise from the electron configurations of the atom

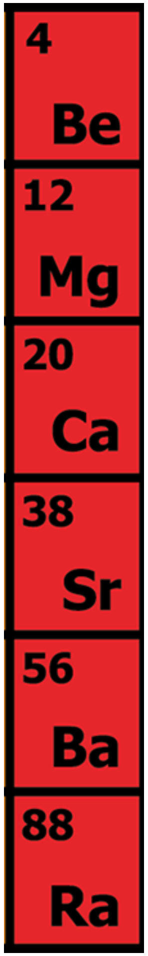
Elements are arranged:

Vertically into **Groups**



Horizontally Into **Periods**



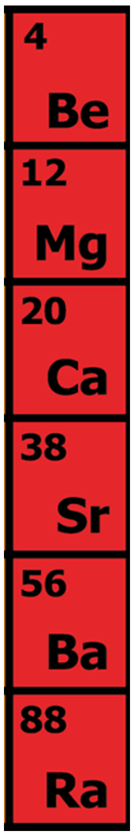


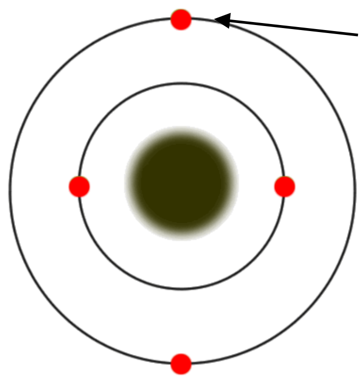
If you looked at one atom of every element in a **group** (family) you would see…

Each atom has the same number of **electrons** in it's outermost shell and sublevel.

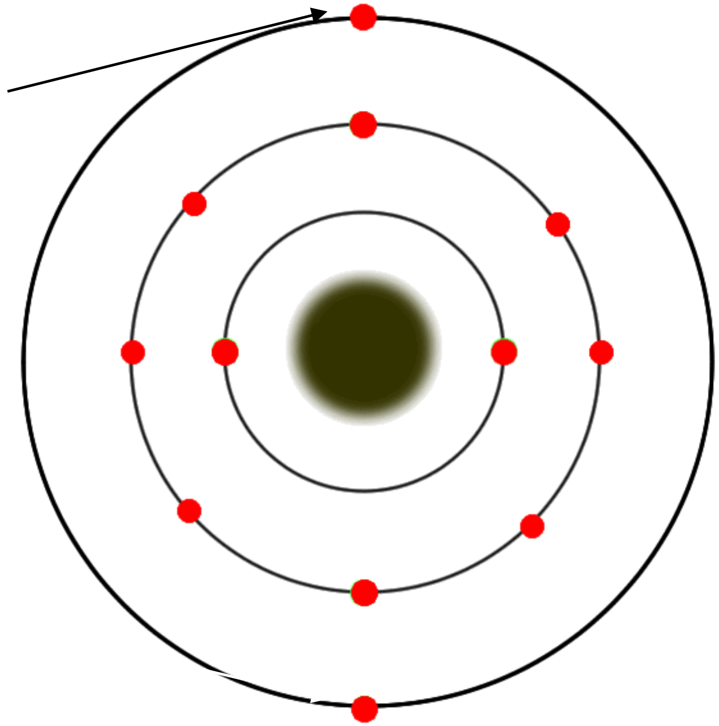
- An example…

The **group 2** atoms all have **2 electrons** in their outer shells





Be (Beryllium) Atom



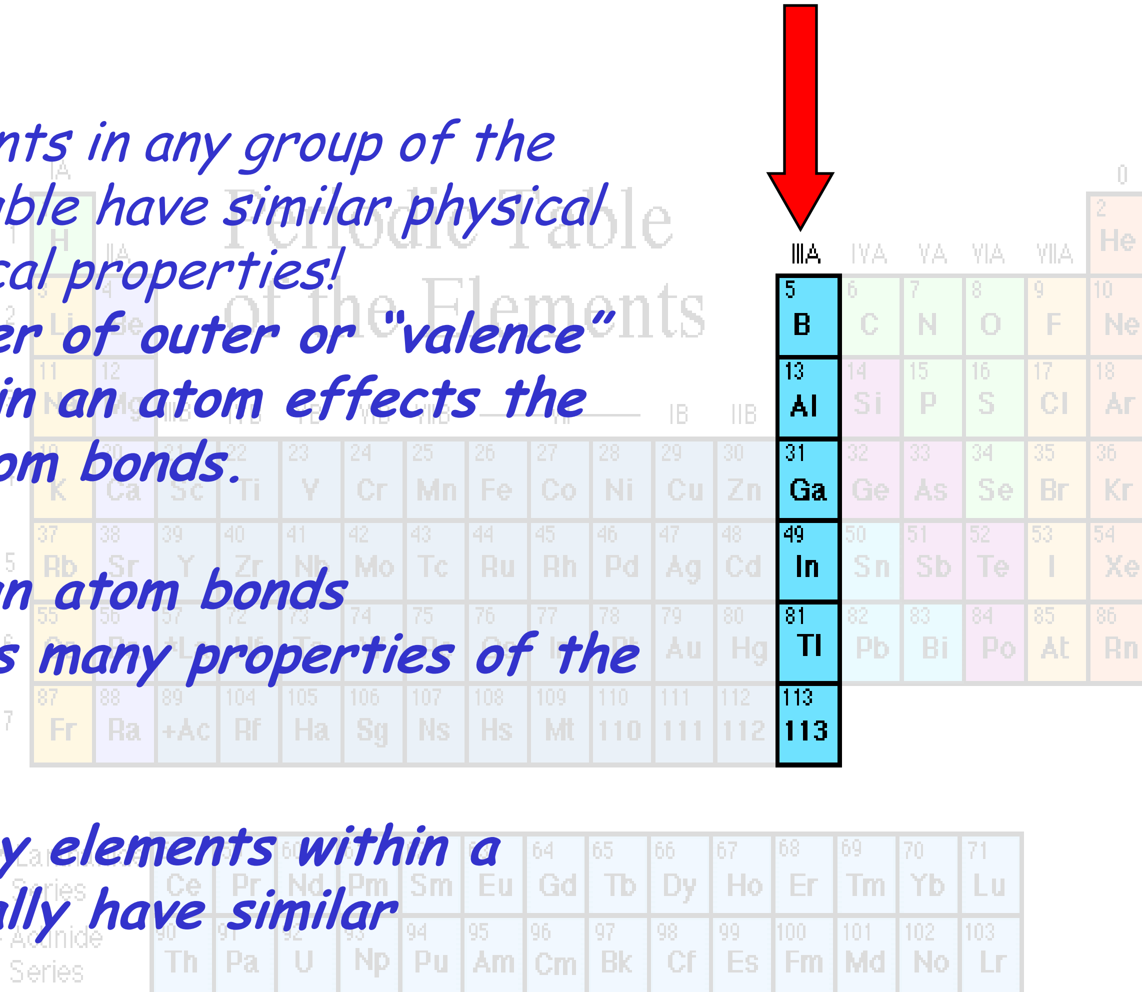
Mg (Magnesium) Atom

The elements in any group of the periodic table have similar physical and chemical properties!

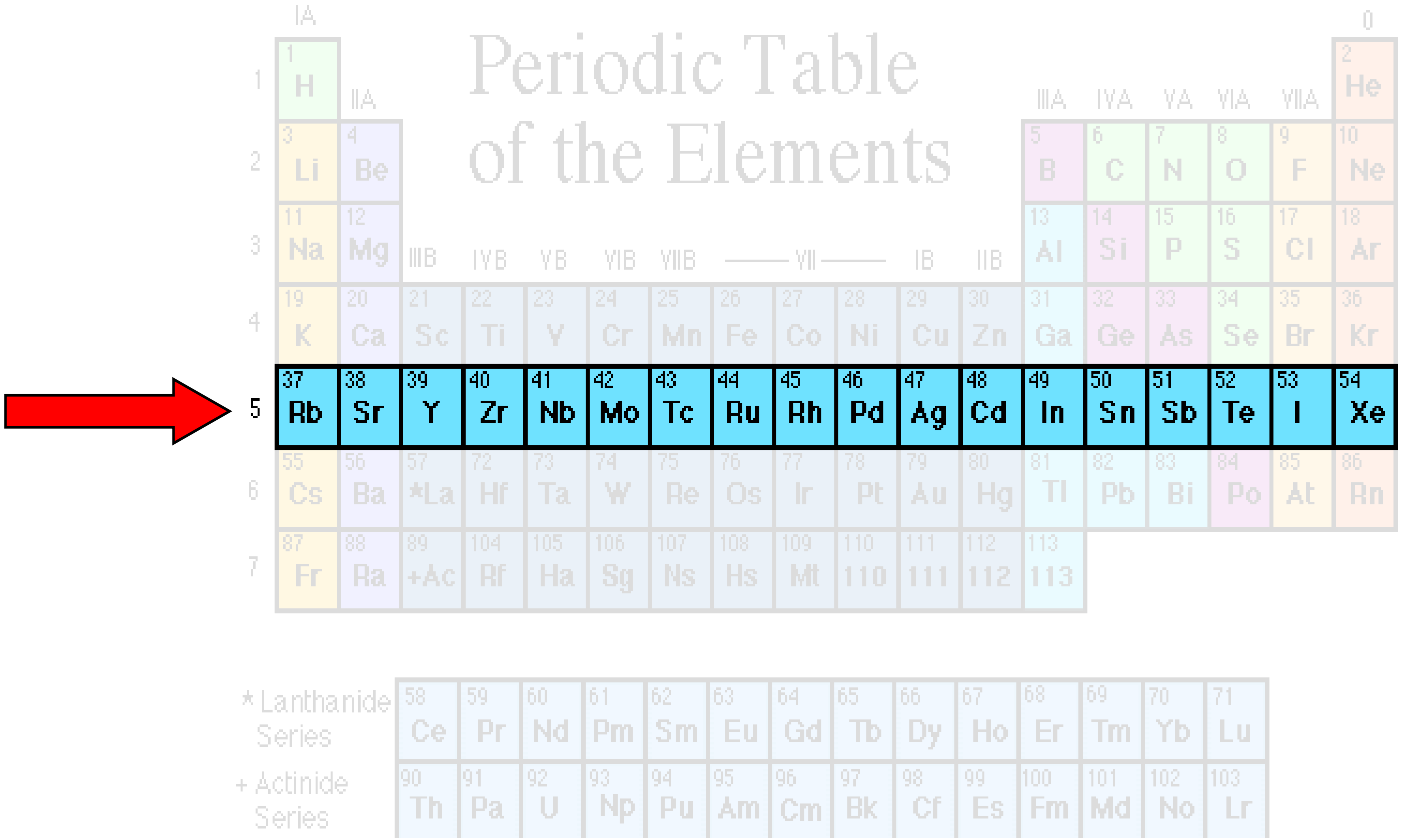
The number of outer or "valence" electrons in an atom effects the way an atom bonds.

The way an atom bonds determines many properties of the element.

This is why elements within a group usually have similar properties



The vertical columns of the periodic table are called GROUPS, or FAMILIES.



The horizontal rows of the periodic table are called PERIODS.

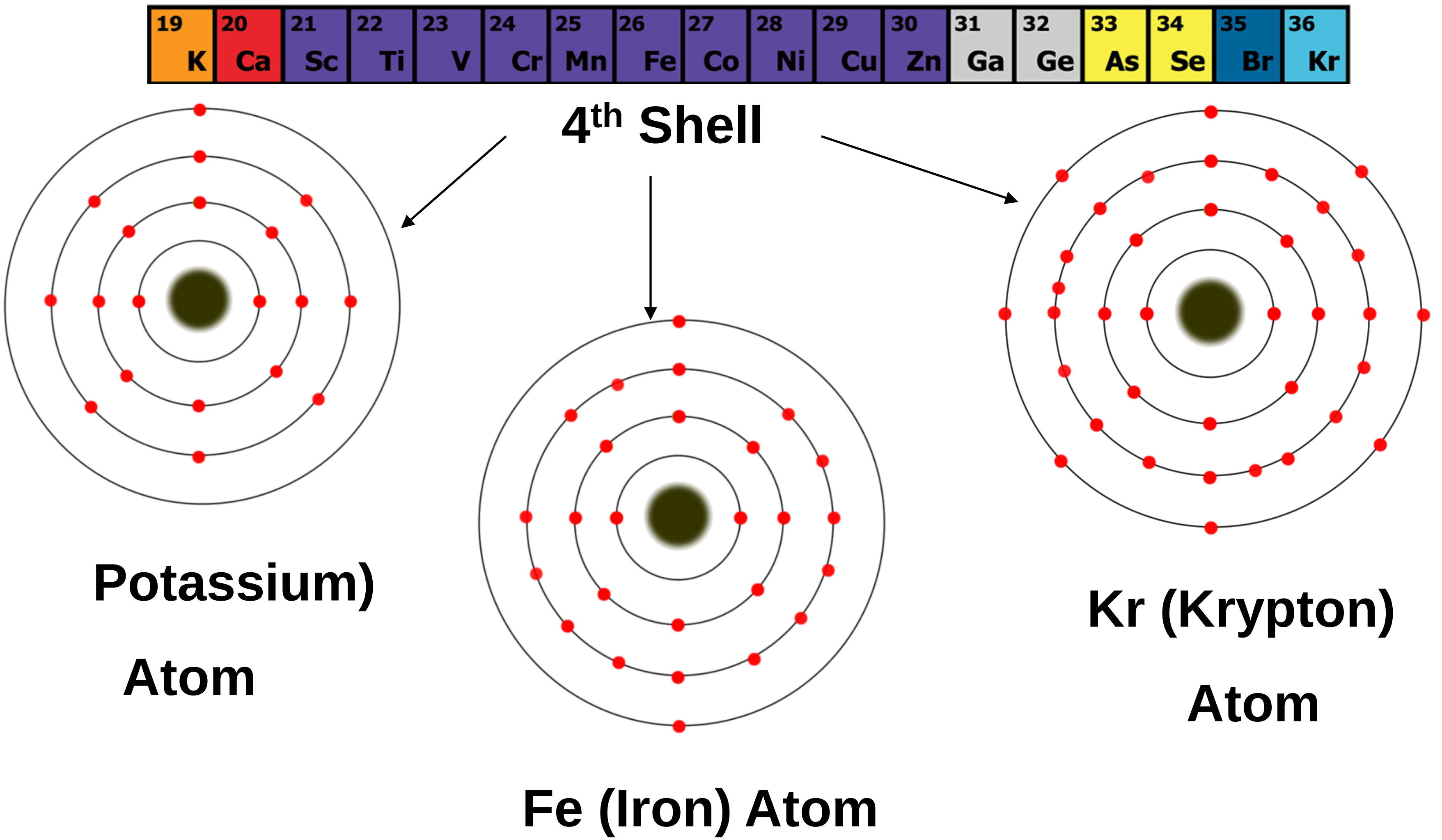
If you looked at an atom from each element in a ***period***

19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr

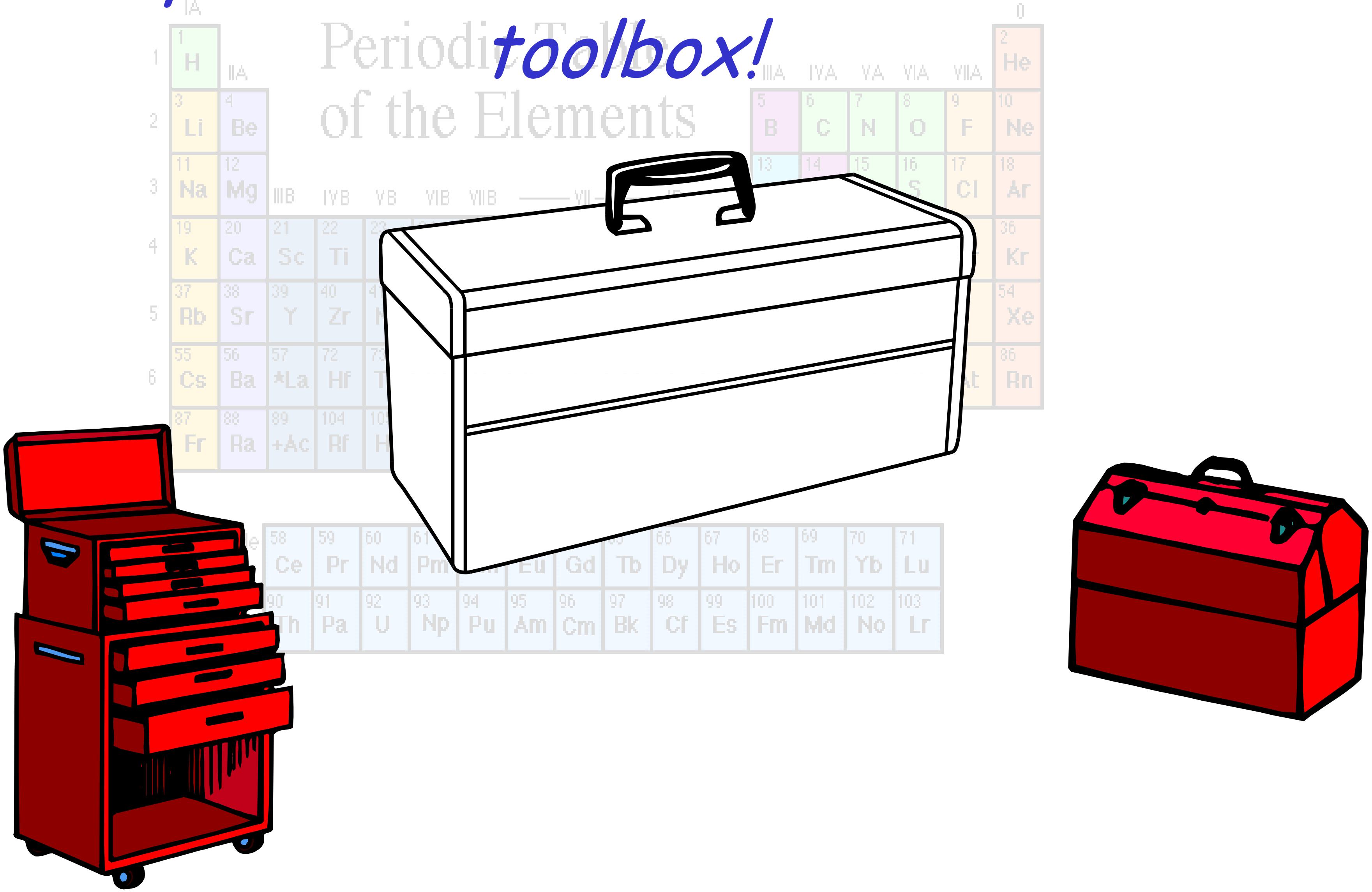
you would see…

Each atom has the same number of electron holding **shells**.

The period 4 atoms each have 4 electron containing shells



The periodic table is the most important tool in the chemist's toolbox!



F. Surveying the table

- Columns 1 and 2 have electrons filling the s sublevel
- Columns 3-12 have electrons filling the d sublevel
- Columns 13-18 have electrons filling the p sublevel

Periodic Table and Electron Configurations

- Build-up order given by position on periodic table; row by row.
- Elements in same column will have the same outer shell electron configuration.

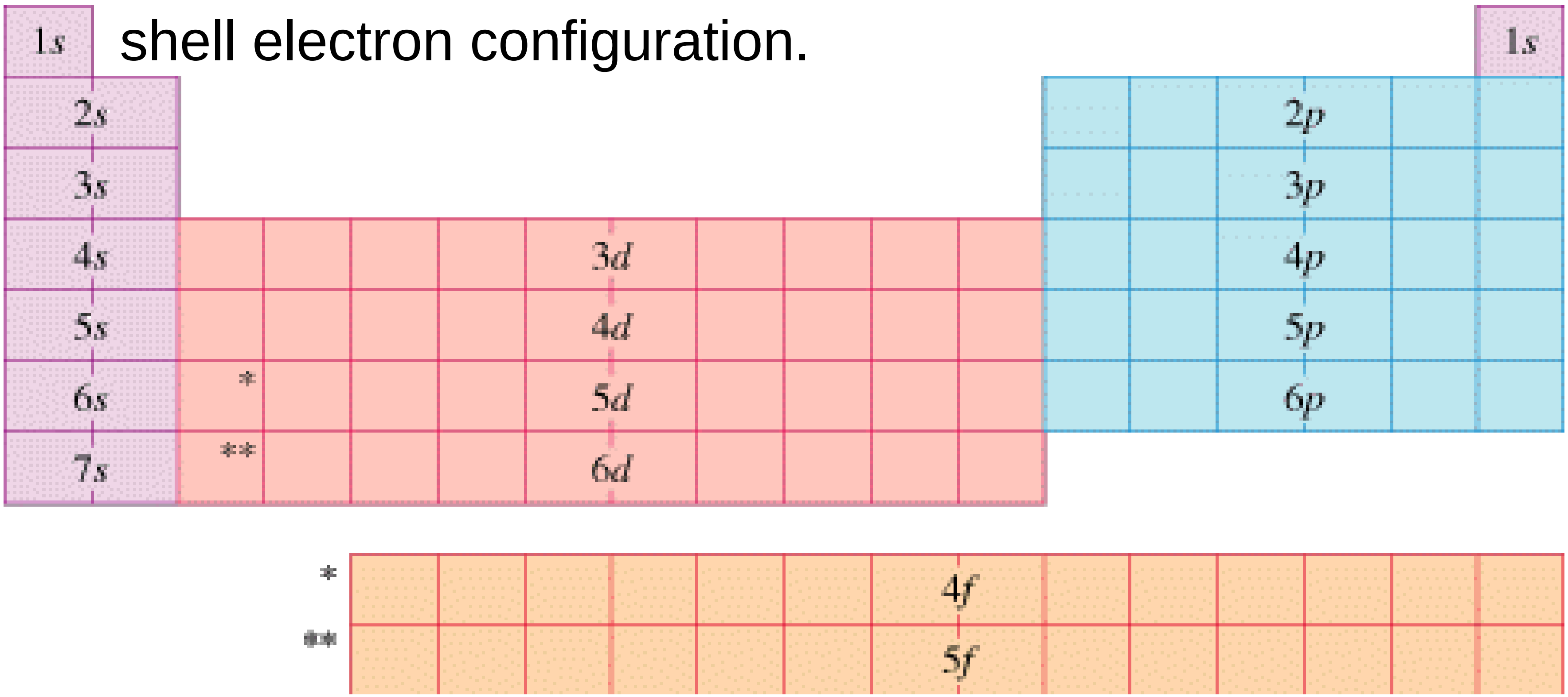
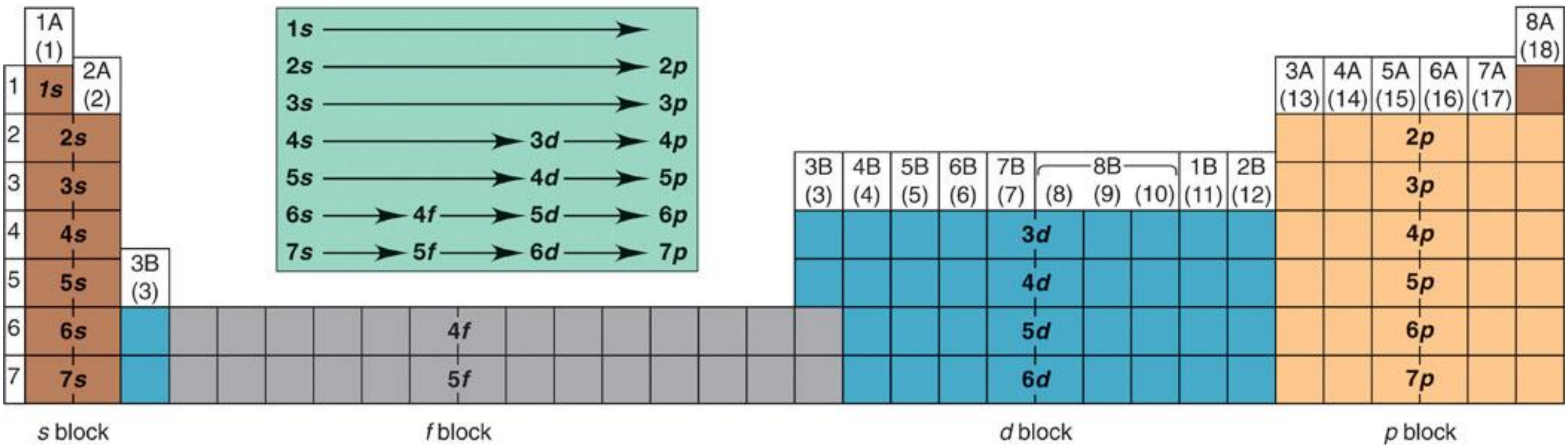


Figure 8.13 **The relation between orbital filling and the periodic table**



Octet rule

- in chemistry an atom that has 8 electrons in its outer level is particularly stable
- Helium is the duet rule

Metals

- **To the left side of the stair step**
  - **Conductors and are malleable and ductile**
  - **Hard and shiny**
  - **3 or less electrons in the outer energy level**
  - **Do not hold outer electrons tightly**

Nonmetals

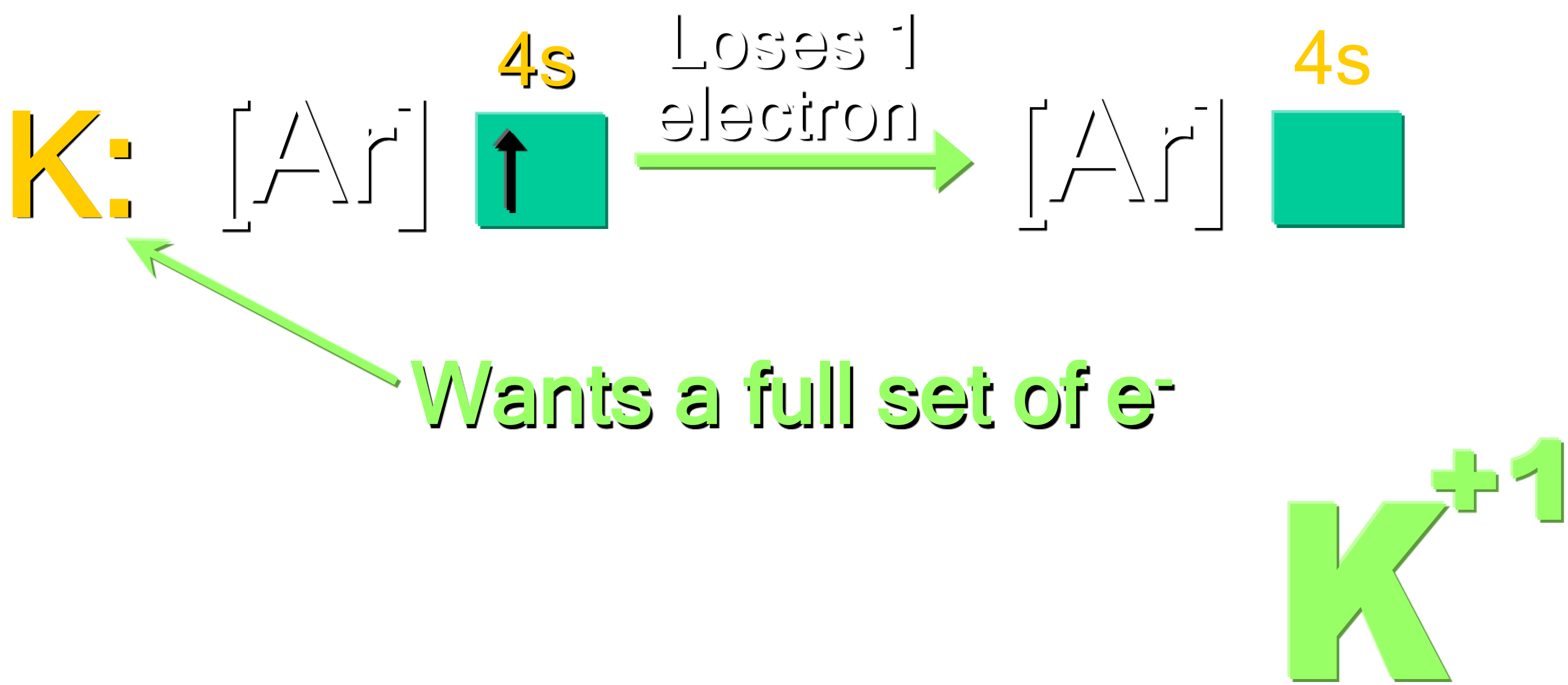
- **To the right of the stair step**
- **Do hold the outer electrons tightly**
- **Brittle**
- **Solids or gases**
- **Five or more electrons in the outer energy level**

Metalloids

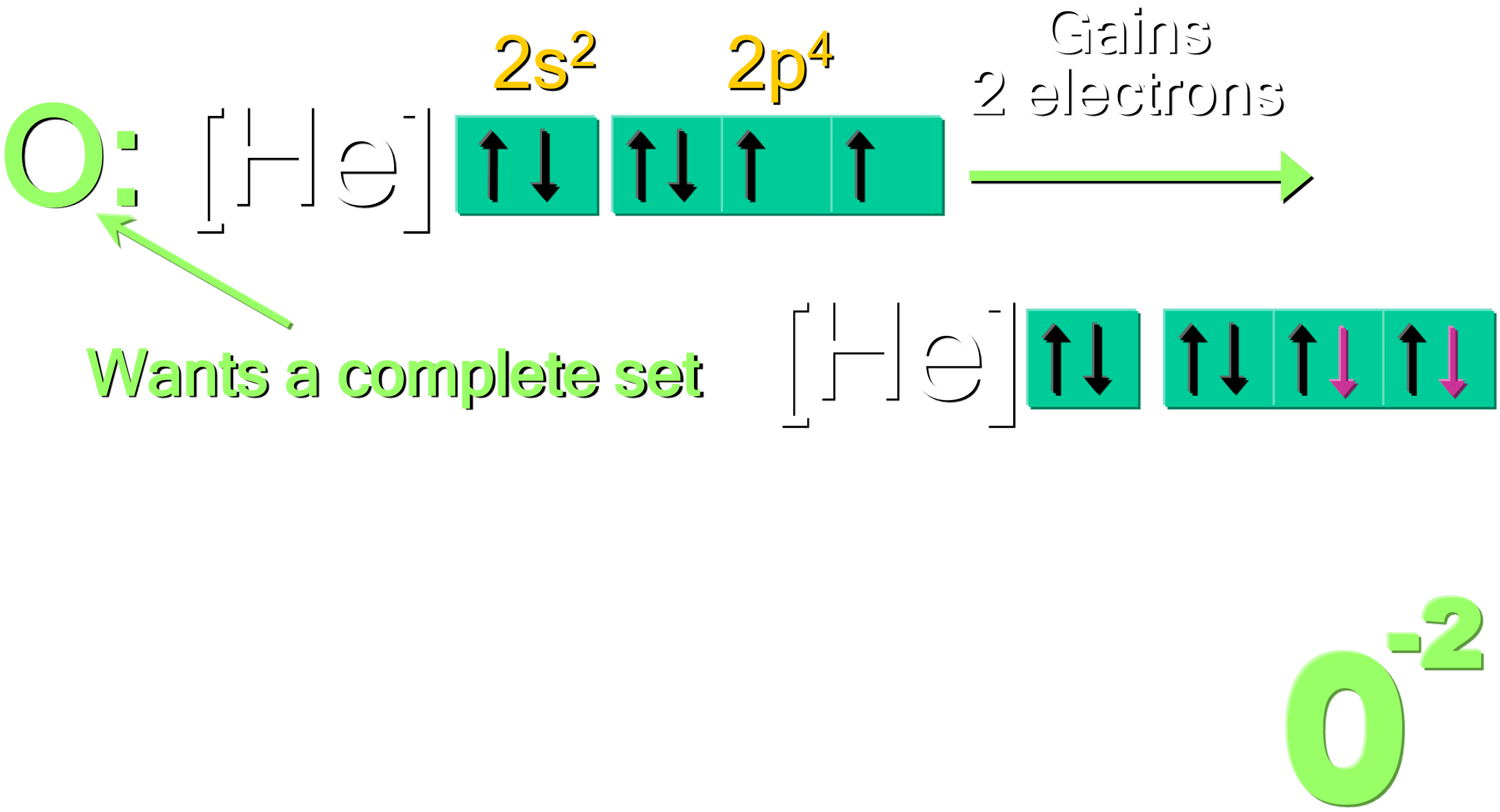
- **have properties of both metals and non metals**

Periodic Properties

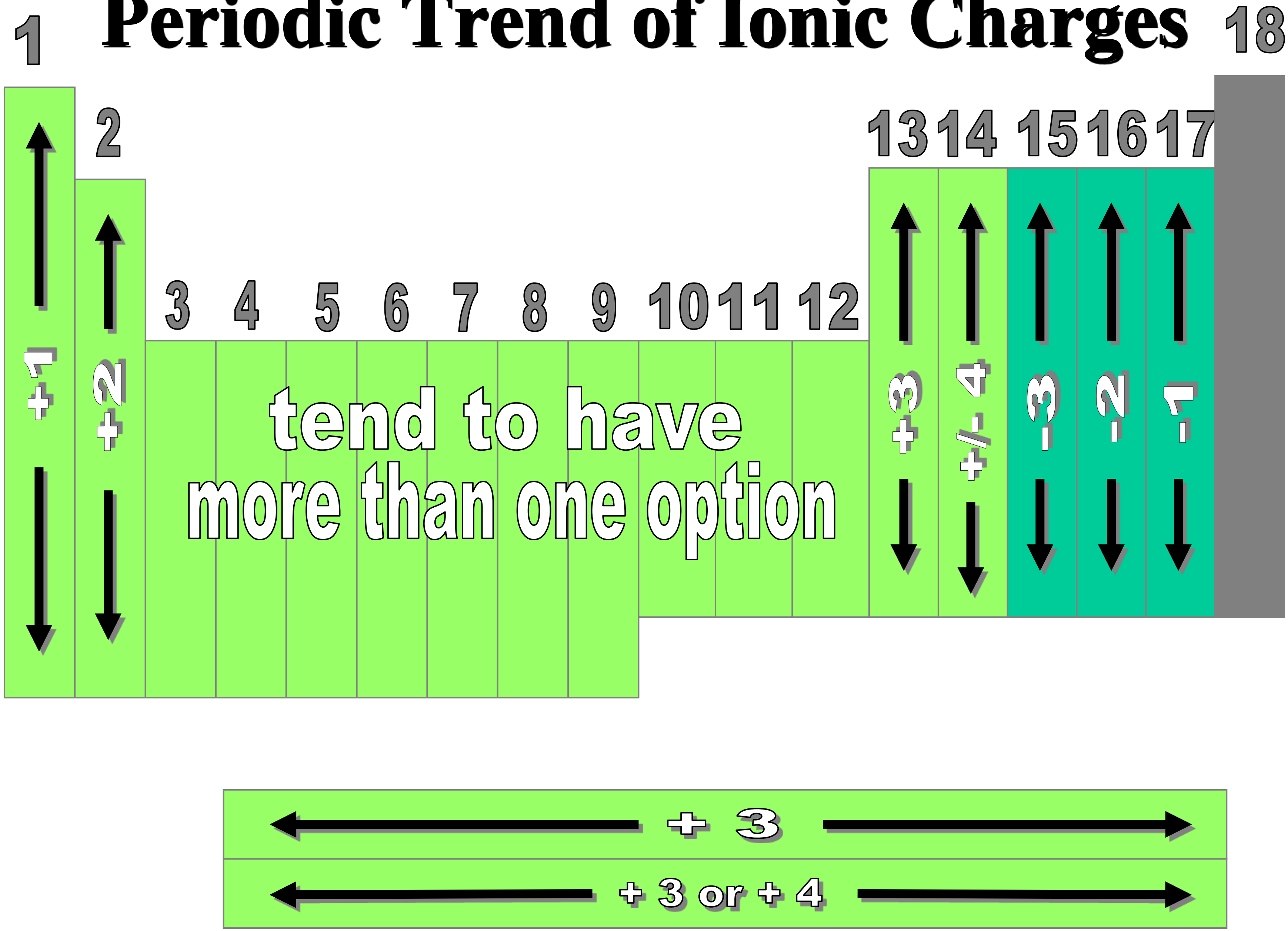
- Elements in a group tend to form ions of the same charge.
  - Modeled by electron configurations.



Periodic Properties



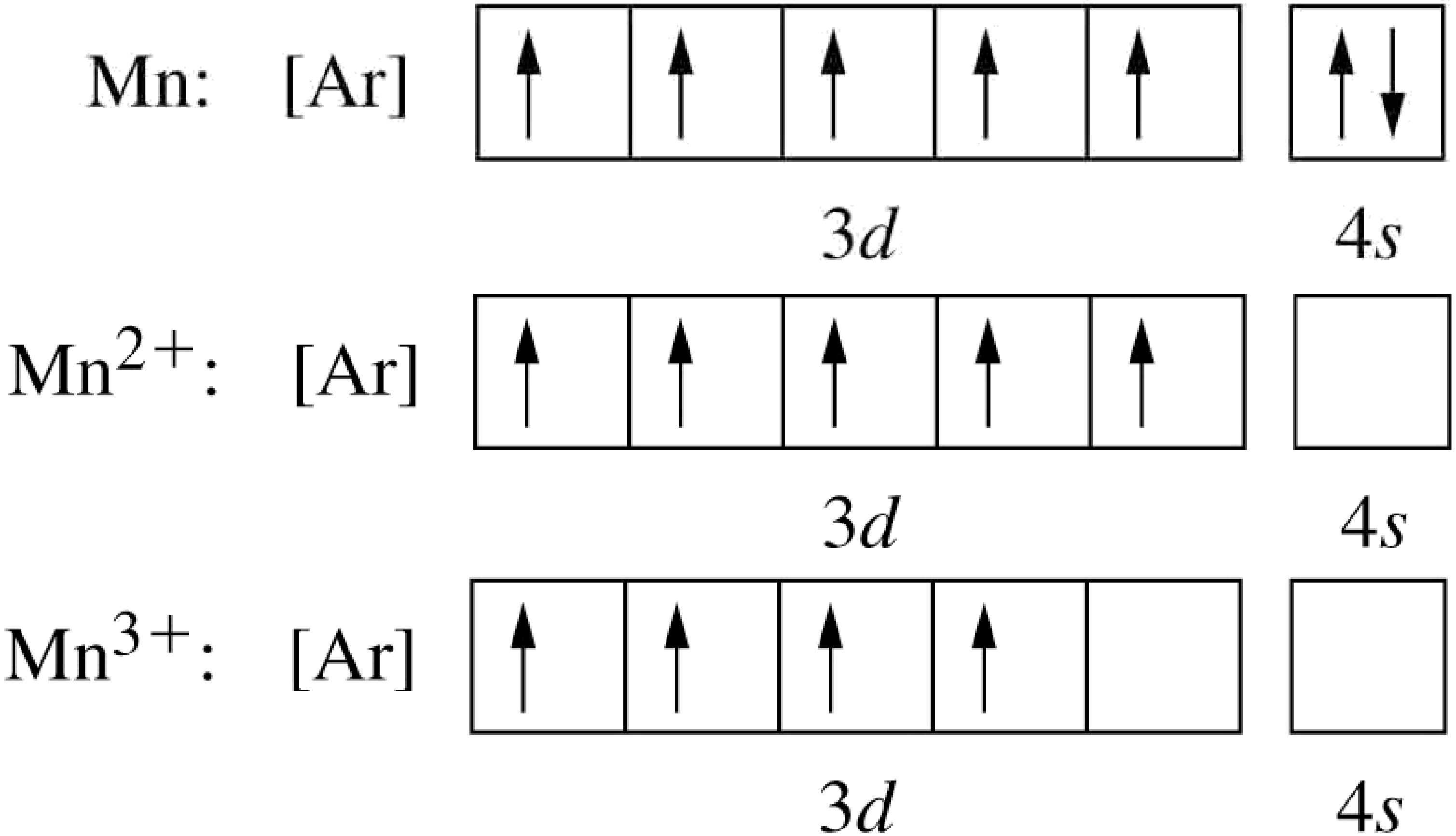
Periodic Trend of Ionic Charges

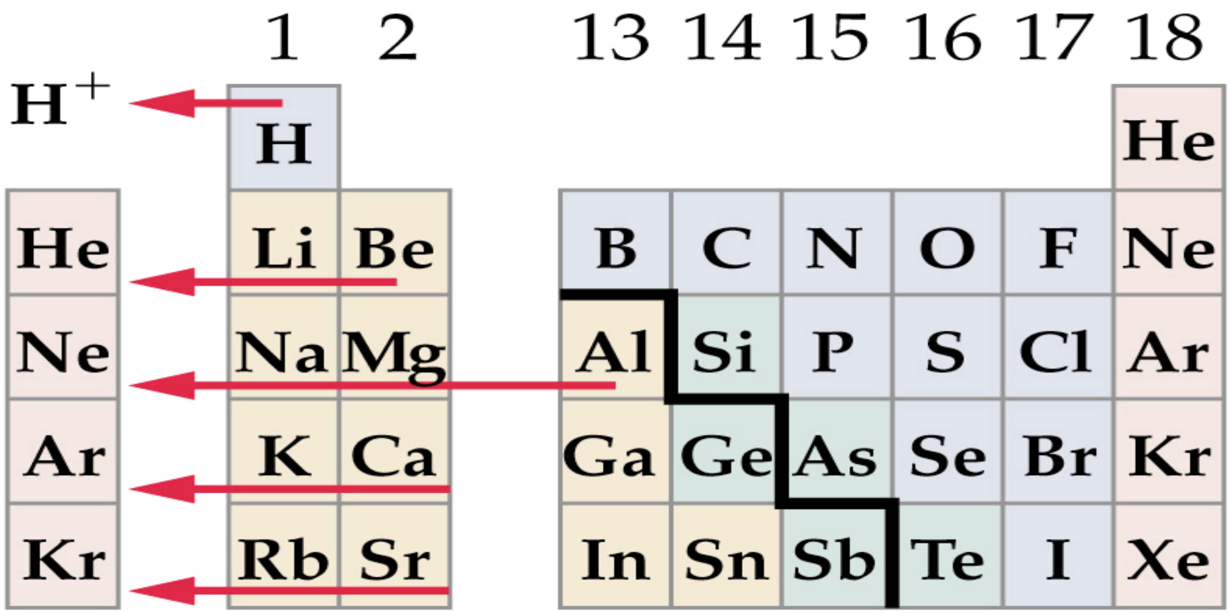


Predicting oxidation numbers

- **+1. . . . . . . s^1**
- **0 . . . . p^6**
- **+2. . . . . . . s^2**
- **+3 . . . . p^1**
- **-4+4 . . . . . p^2**
- **-3 . . . . p3**
- **-2. . . . . . . .p^4**
- **-1 . . . . p5**
  Tend to have more than one oxidation state
  d sublevel series

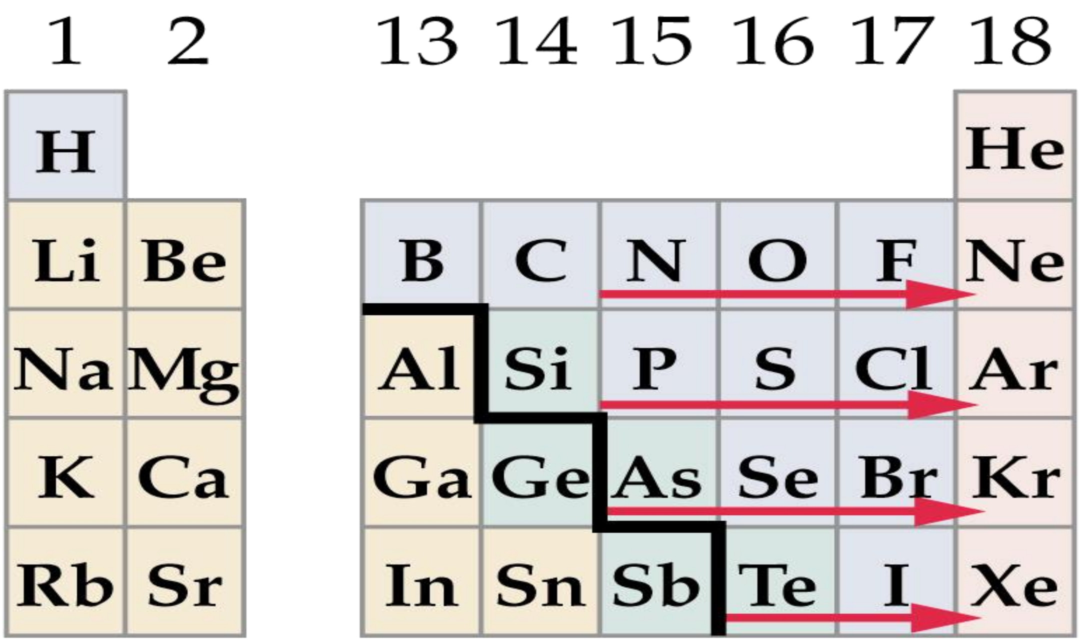
The Transition Elements are almost unpredictable, and sometimes have more than one possible charge -- due to d orbitals --





Tend to lose electrons to become positive

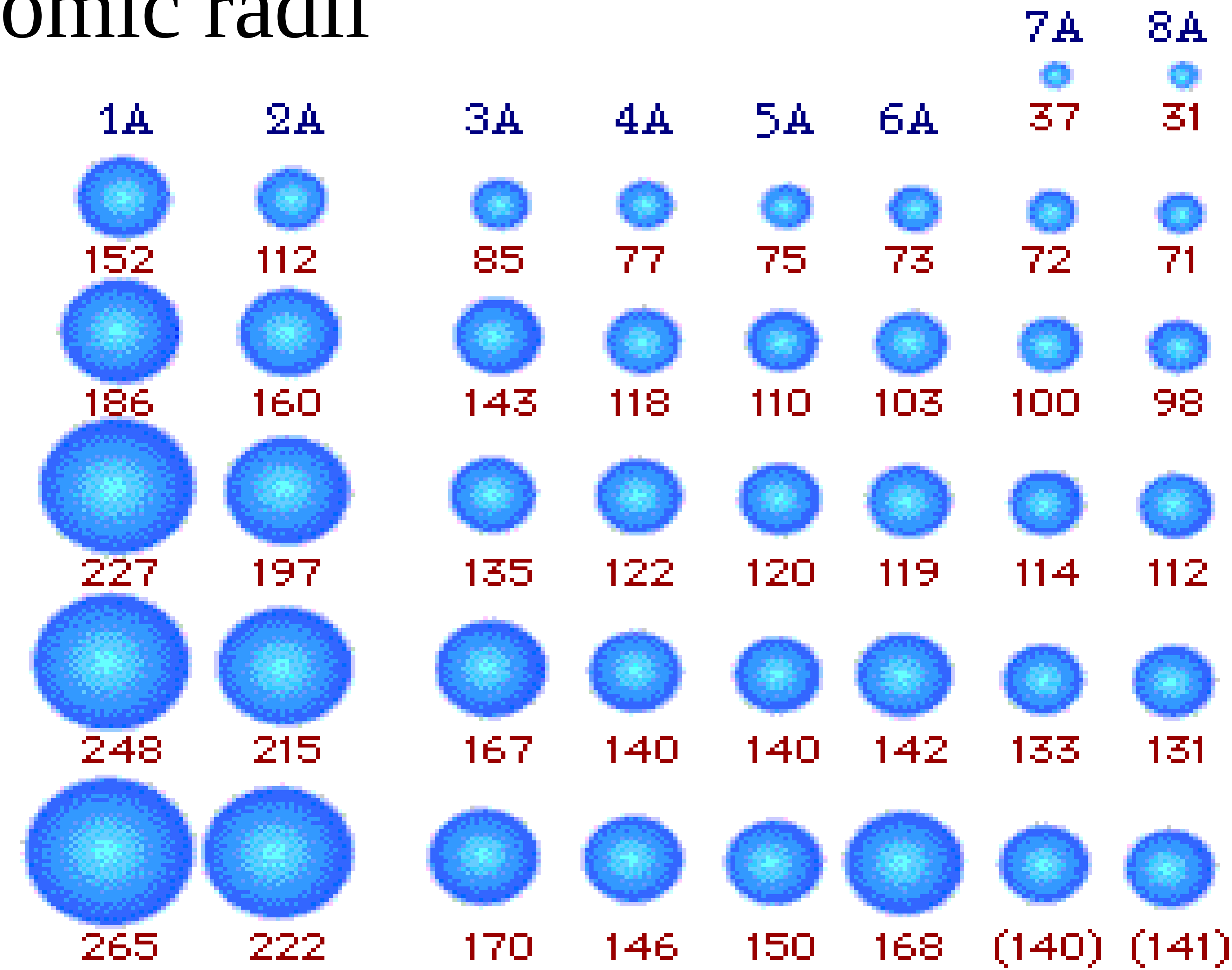
Tend to gain electrons to become negative

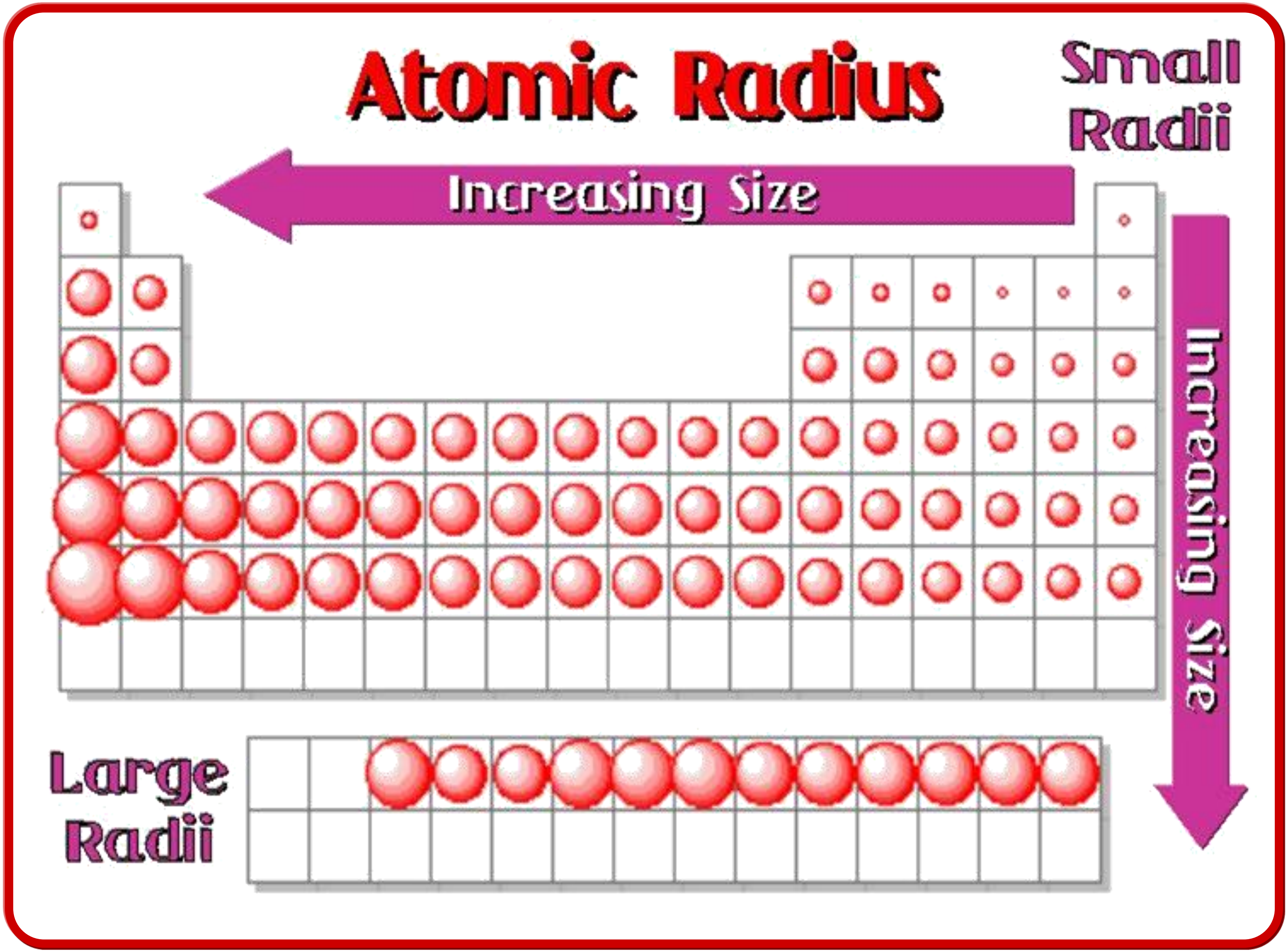


Atomic Radii

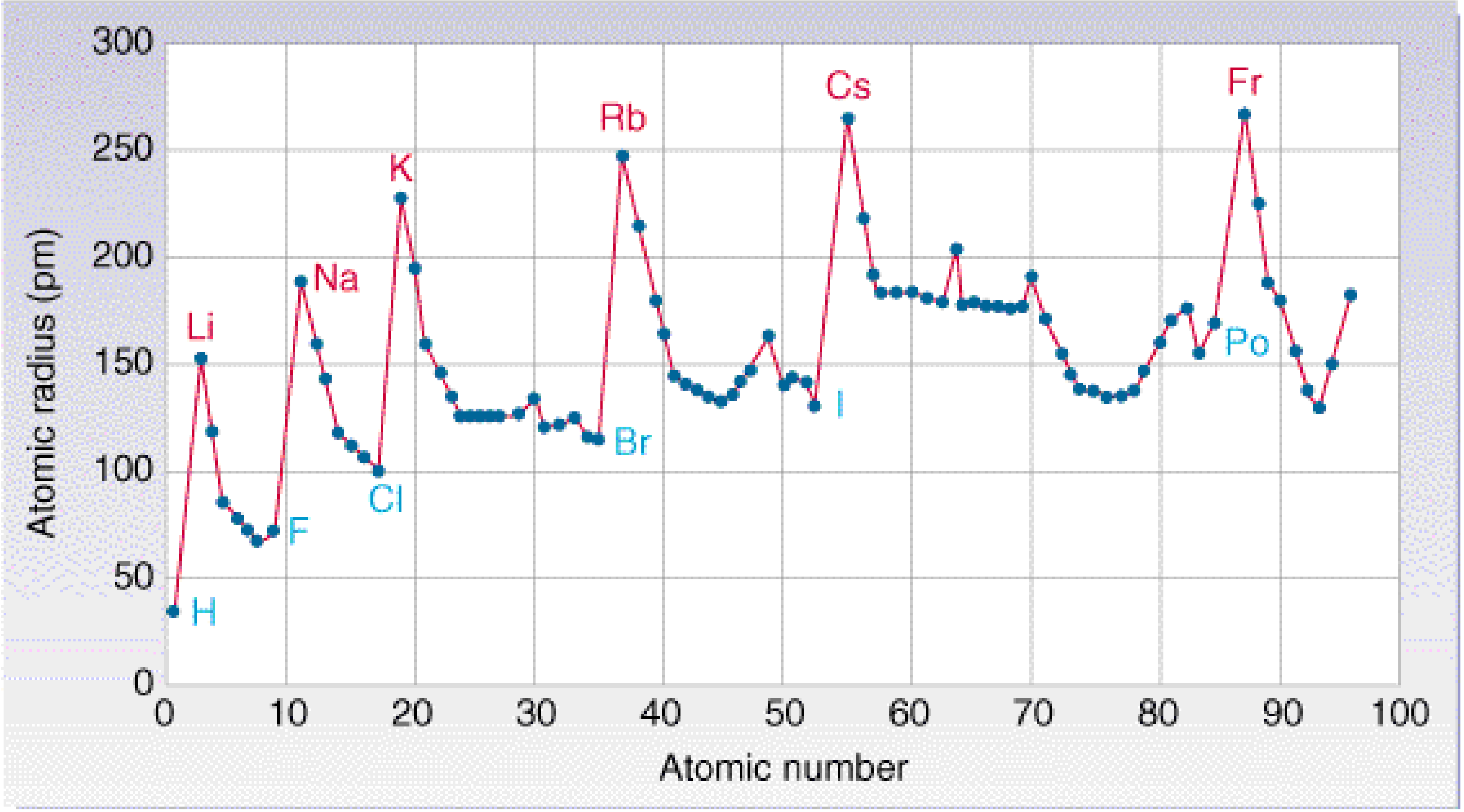
- As the principal quantum number increases the size of the electron cloud increases
- Atomic Radii increases from top to bottom and from right to left
  - As you move to the right the number of protons increase in that particular energy level
  - As you move down the number of electrons increase with less ability to hold the electrons

Atomic radii



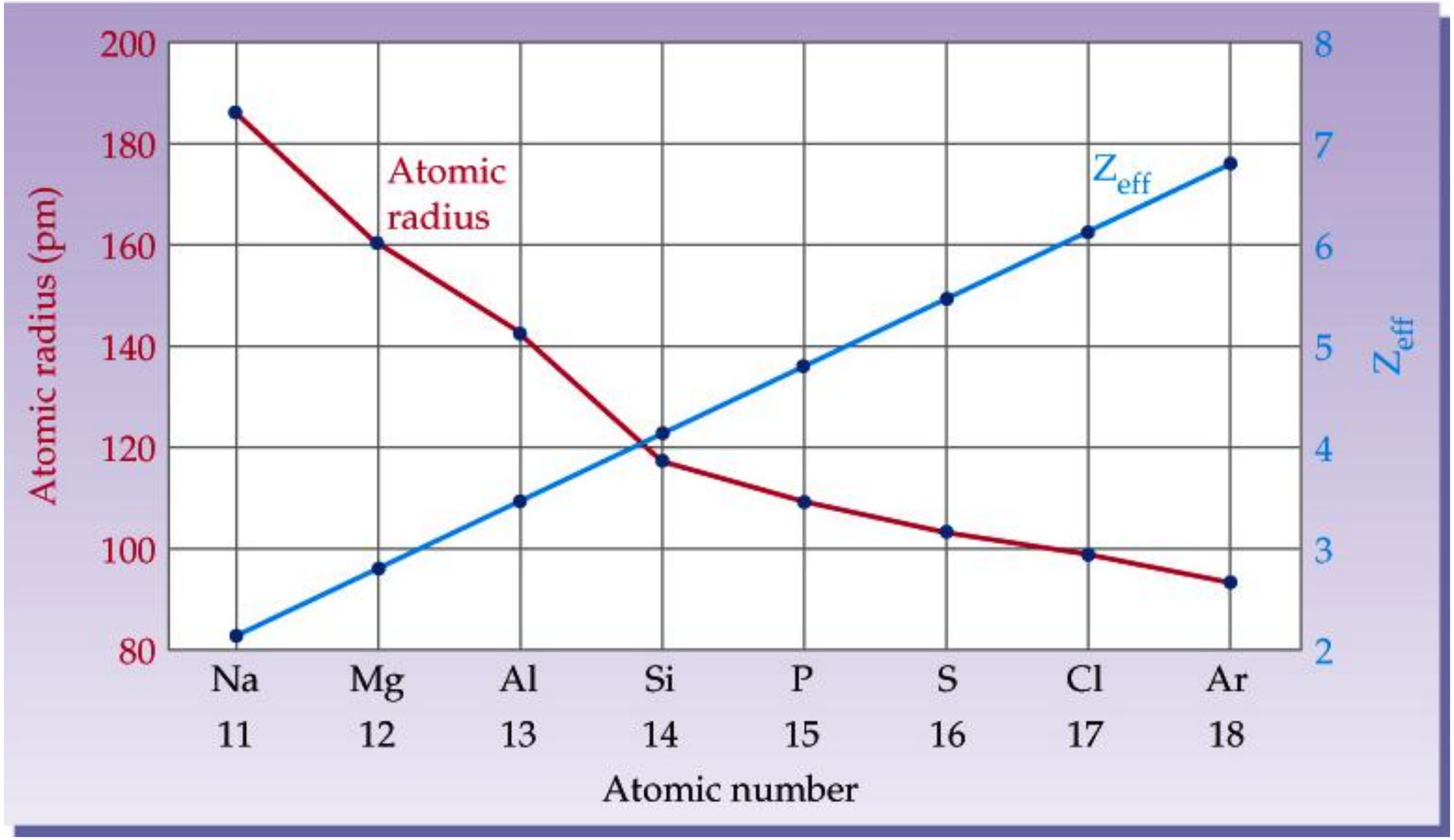

Atomic Radius
Small Radii
Increasing Size
Increasing Size
Large Radii

The Periodic Table 02



Periodic Properties 02

Atomic Radii 02:



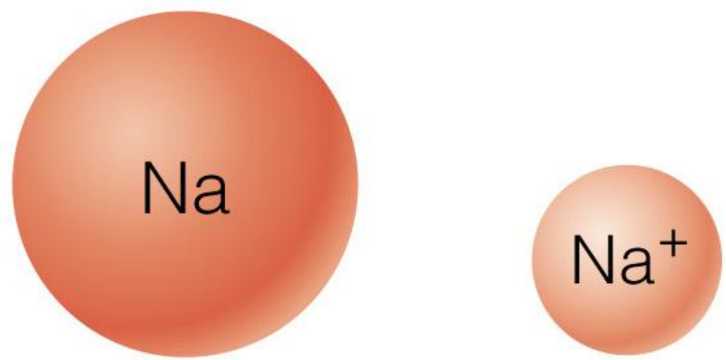
Ion Radii

- All atoms electrons are arranged to made it as stable as possible
- Stable atoms are noble gases
- If sodium loses one electron it has a noble gas configuration and a charge of +1
- Chlorine gains an electron and then has a noble gas configuration and a – 1 charge
- Metallic ions on the left and in the center of the chart are formed by the loss of electrons making them smaller in radii
- Non metallic ions on the right are formed by gaining electrons

Periodic Properties

- How does the size of an atom change when electrons are added or removed?

As an Atom loses 1 or more electrons (becomes positive), it loses a layer therefore, its radius decreases.

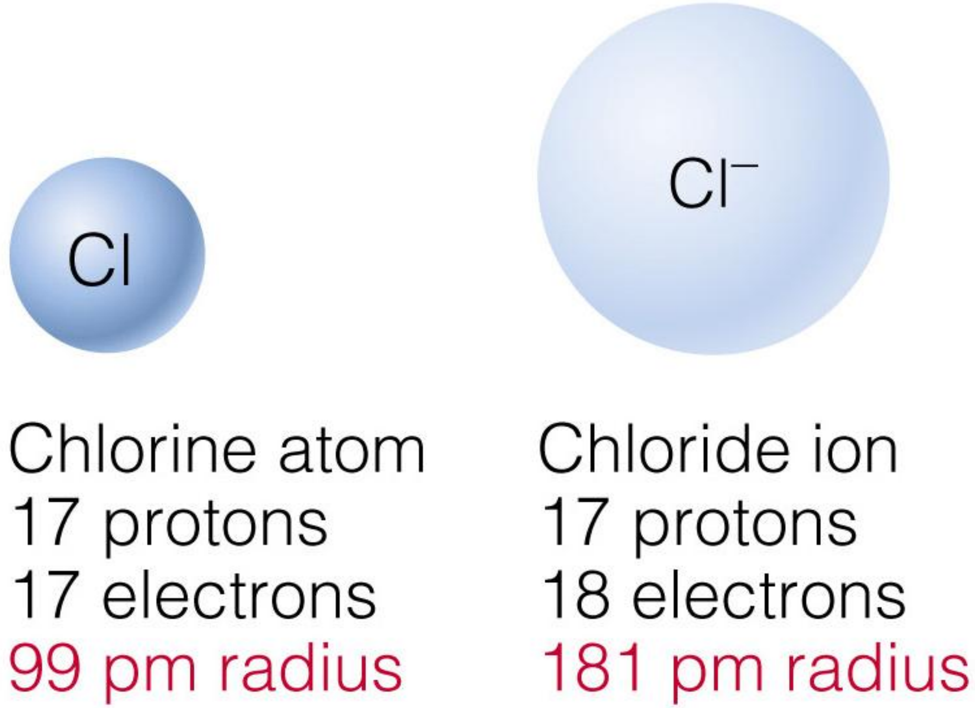


Sodium atom	Sodium ion
11 protons	11 protons
11 electrons	10 electrons
186 pm radius	95 pm radius

Periodic Properties

- How does the size of an atom change when electrons are added or removed?

As an Atom gains 1 or more electrons (negative), it fills its valence layer, therefore, its radius increases.

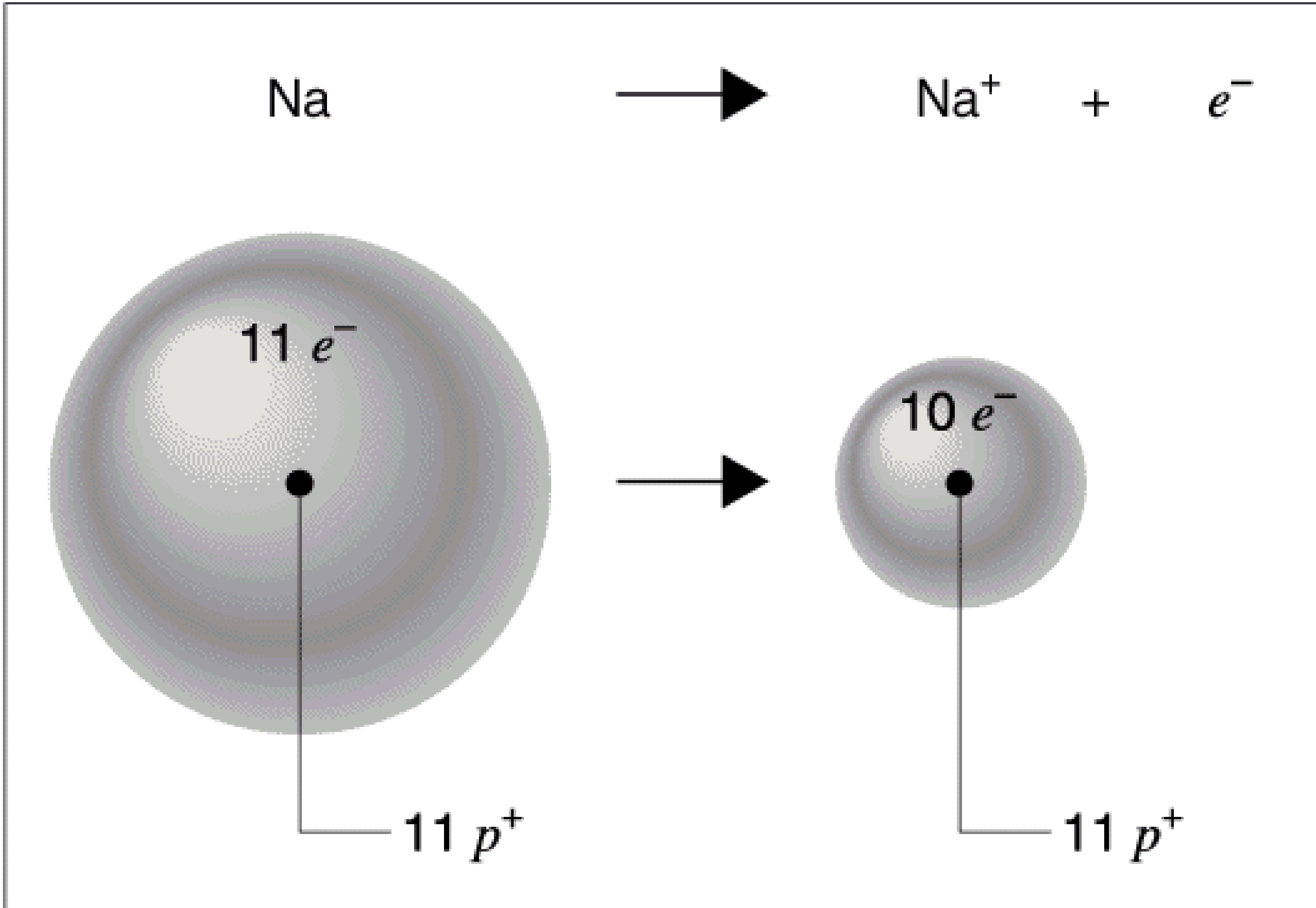


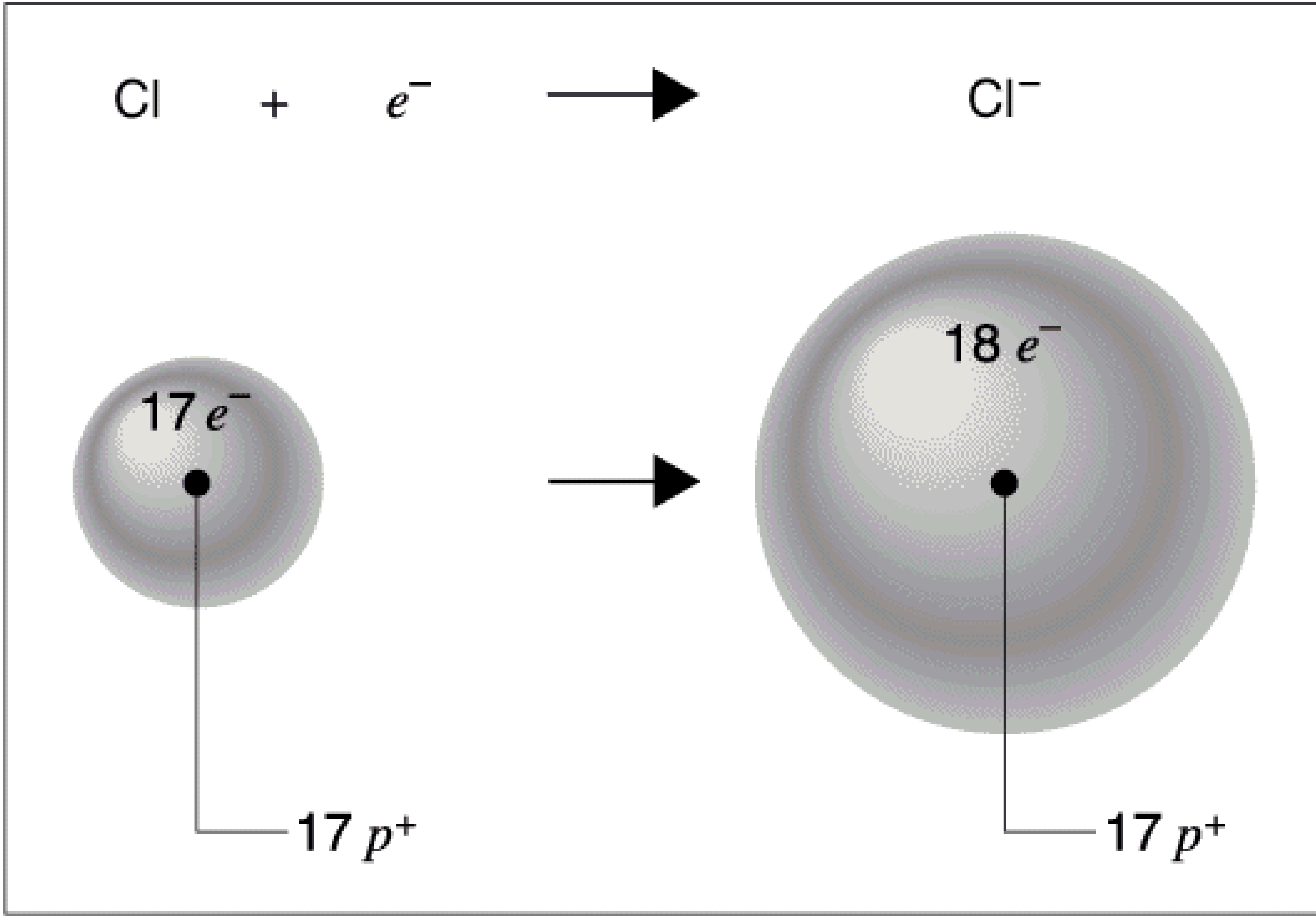
Ions and Ionic Radii02

TABLE 6.1 Some Common Main-Group Ions and Their Noble Gas Electron Configurations

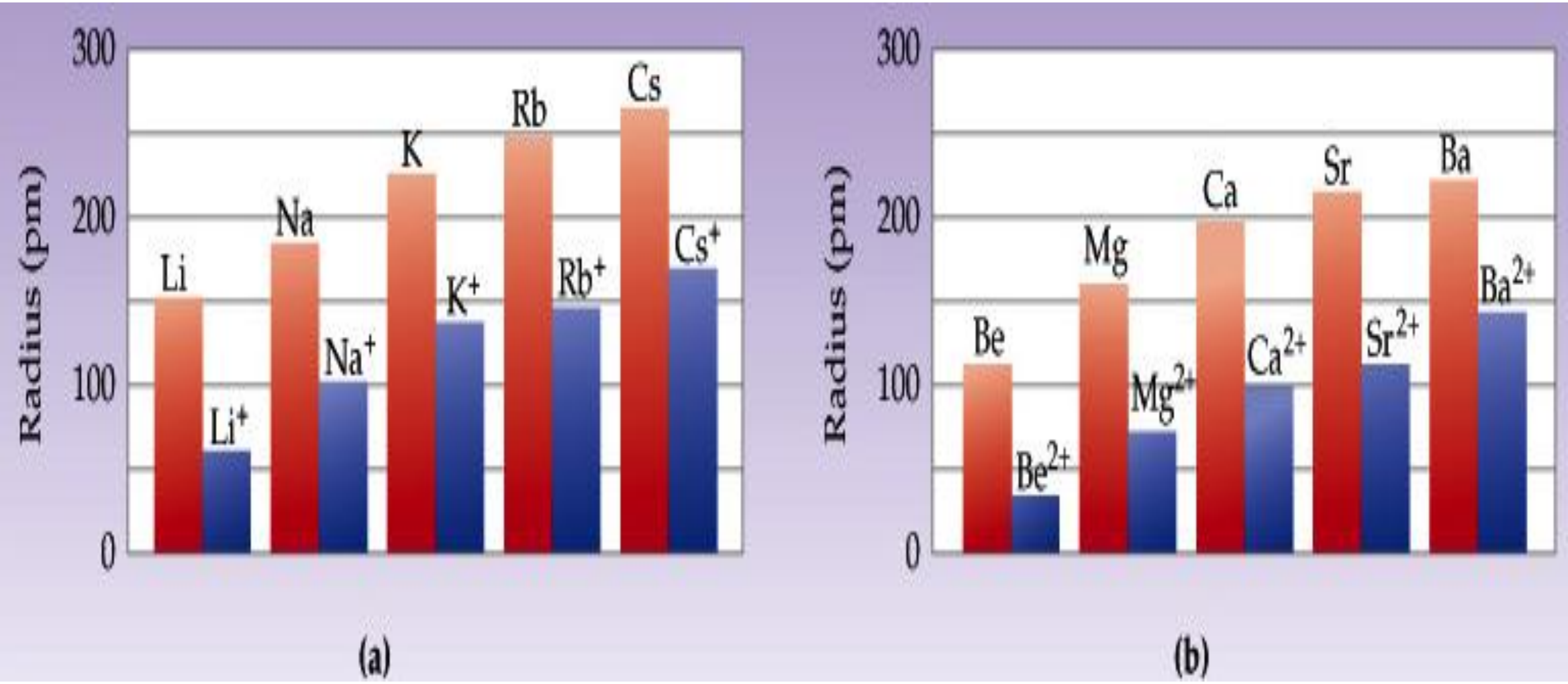
Group 1A	Group 2A	Group 3A	Group 6A	Group 7A	Electron Configuration
H^+					[None]
H^-					[He]
Li^+	Be^{2+}				[He]
Na^+	Mg^{2+}	Al^{3+}	O^{2-}	F^-	[Ne]
K^+	Ca^{2+}	*Ga^{3+}	S^{2-}	Cl^-	[Ar]
Rb^+	Sr^{2+}	*In^{3+}	Se^{2-}	Br^-	[Kr]
Cs^+	Ba^{2+}	*Tl^{3+}	Te^{2-}	I^-	[Xe]

* These ions do not have a true noble gas electron configuration because they have an additional filled *d* subshell.

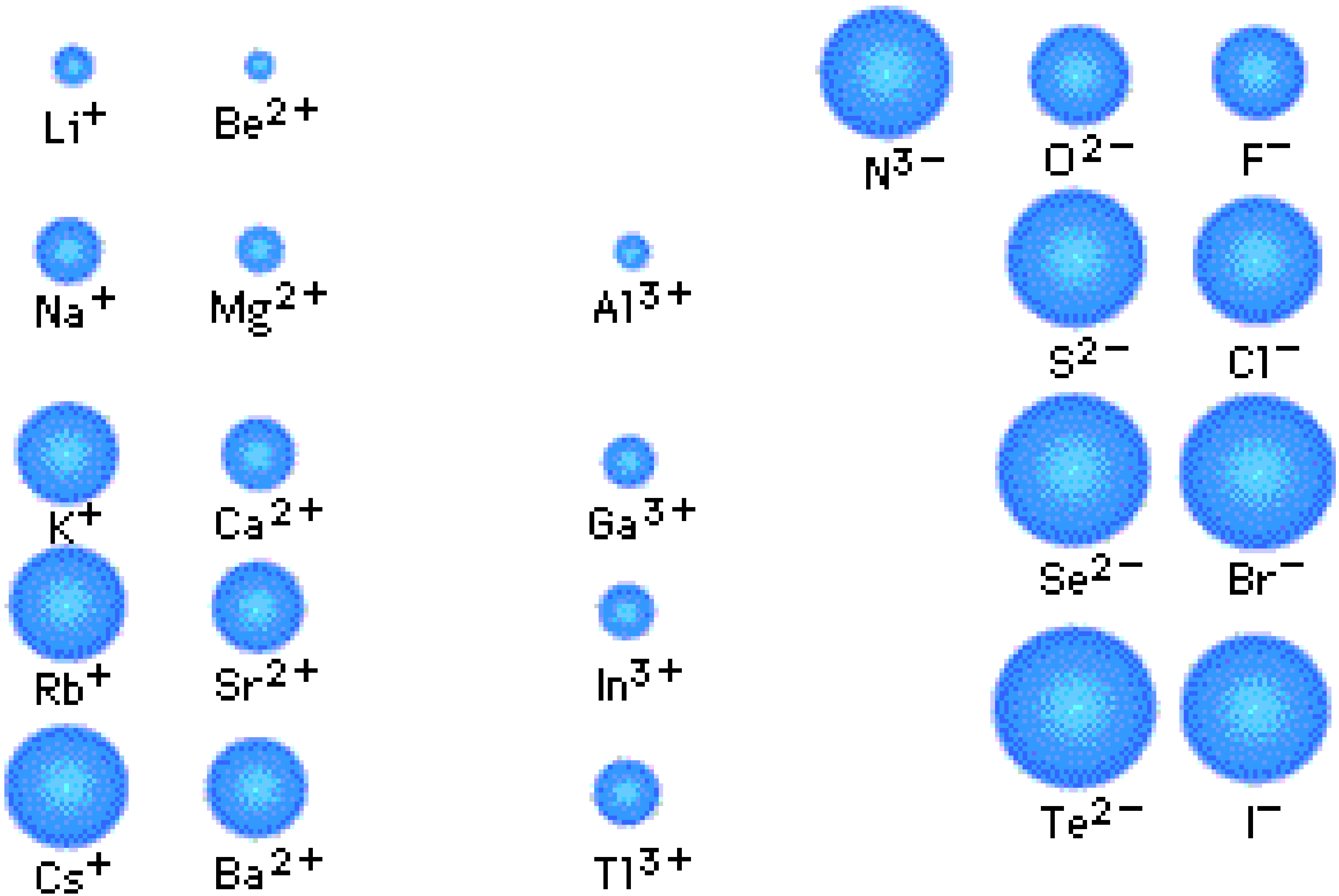

Na
Na+
+
e−
11 e−
10 e−
11 p+
11 p+


$Cl + e^- \rightarrow Cl^-$
$17\ e^-$
$17\ p^+$
$18\ e^-$
$17\ p^+$

Ions and Ionic Radii



Ionic Radii



Li^{+}
Be^{2+}
N^{3-}
O^{2-}
F^{-}
Na^{+}
Mg^{2+}
Al^{3+}
S^{2-}
Cl^{-}
K^{+}
Ca^{2+}
Ga^{3+}
Se^{2-}
Br^{-}
Rb^{+}
Sr^{2+}
In^{3+}
Te^{2-}
I^{-}
Cs^{+}
Ba^{2+}
Tl^{3+}

Trend of ionic radii

The cation of an atom decreases in size while the anion of an atom increases in size.

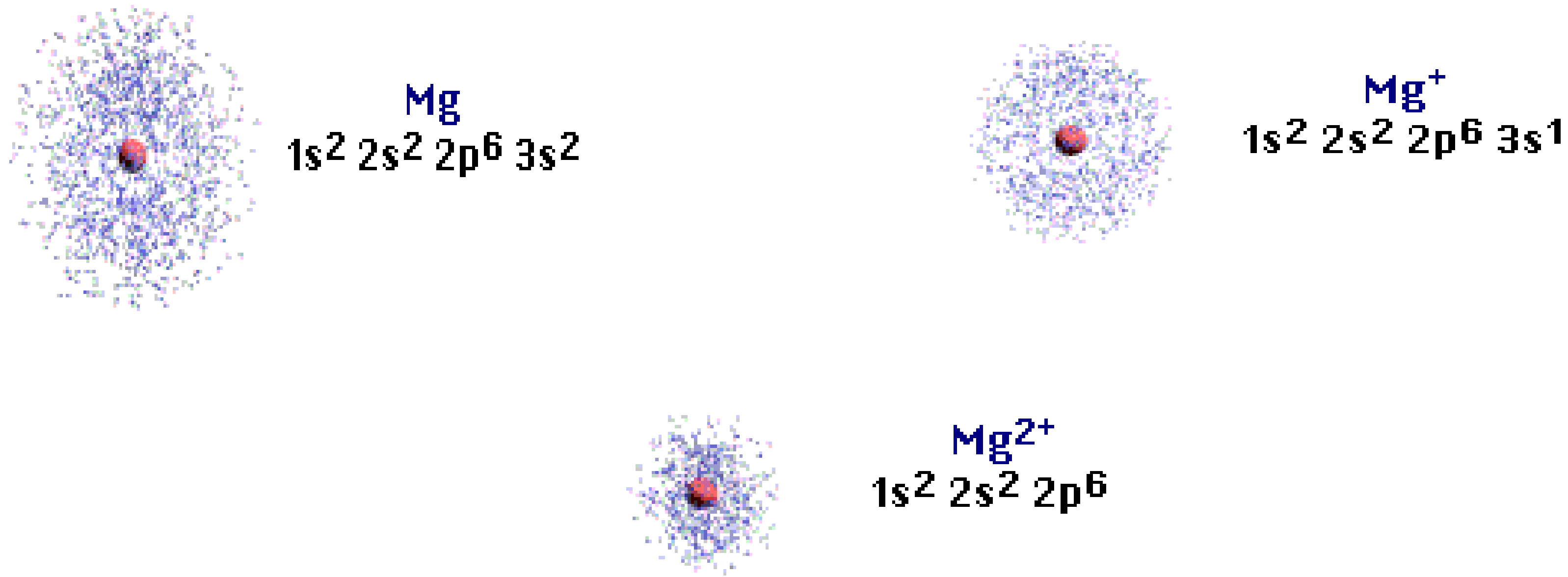
The trend can not be made according to the periodic table, but by the isoelectronic series.

The more positive an ion is the smaller it is because Z_{eff} increases, while the more negative an ion, the larger it is because Z_{eff} decreases.

First ionization Energy

- energy needed to remove the most loosely held electron from the atom (Measured in kilojoules/mole)

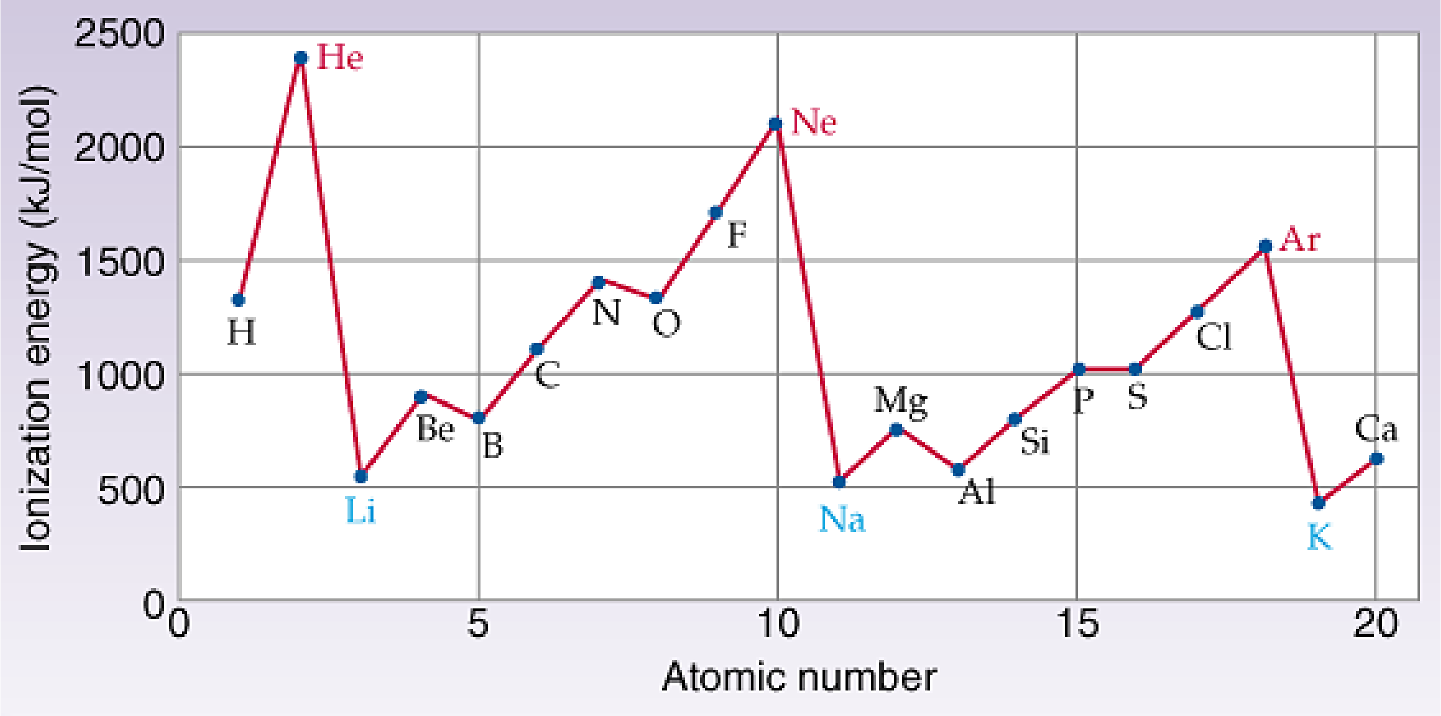
successive ionization energy



- $Mg(g) \longrightarrow Mg^{+}(g) + e^{-}$ IE(1) = 738 kJ/mol
- $Mg^{+}(g) \longrightarrow Mg^{2+}(g) + e^{-}$ IE(2) = 1451 kJ/mol
- $Mg^{2+}(g) \longrightarrow Mg^{3+}(g) + e^{-}$ IE(3) = 7733 kJ/mol

Ionization Energy

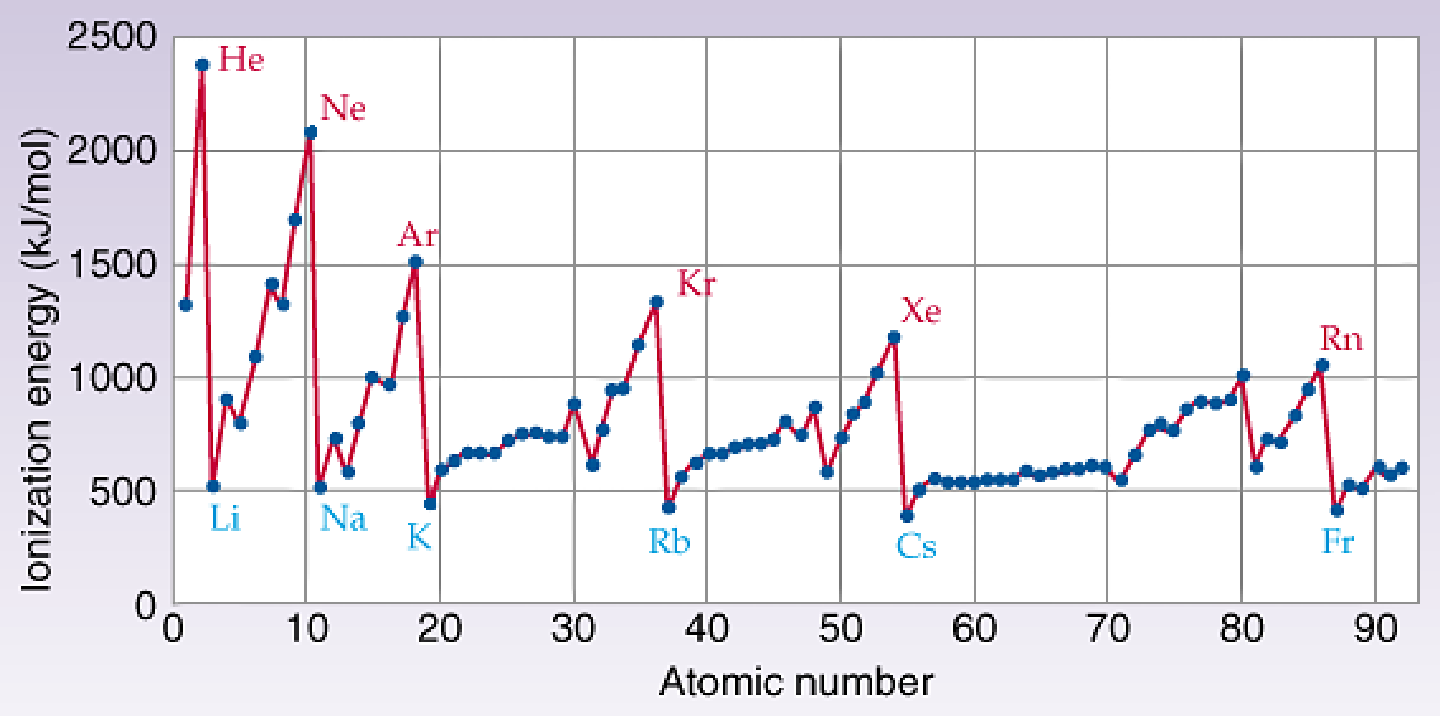
Minor irregularities in the E_i values are explained by looking at the electron configurations.



IE Trend

Ionization Energy

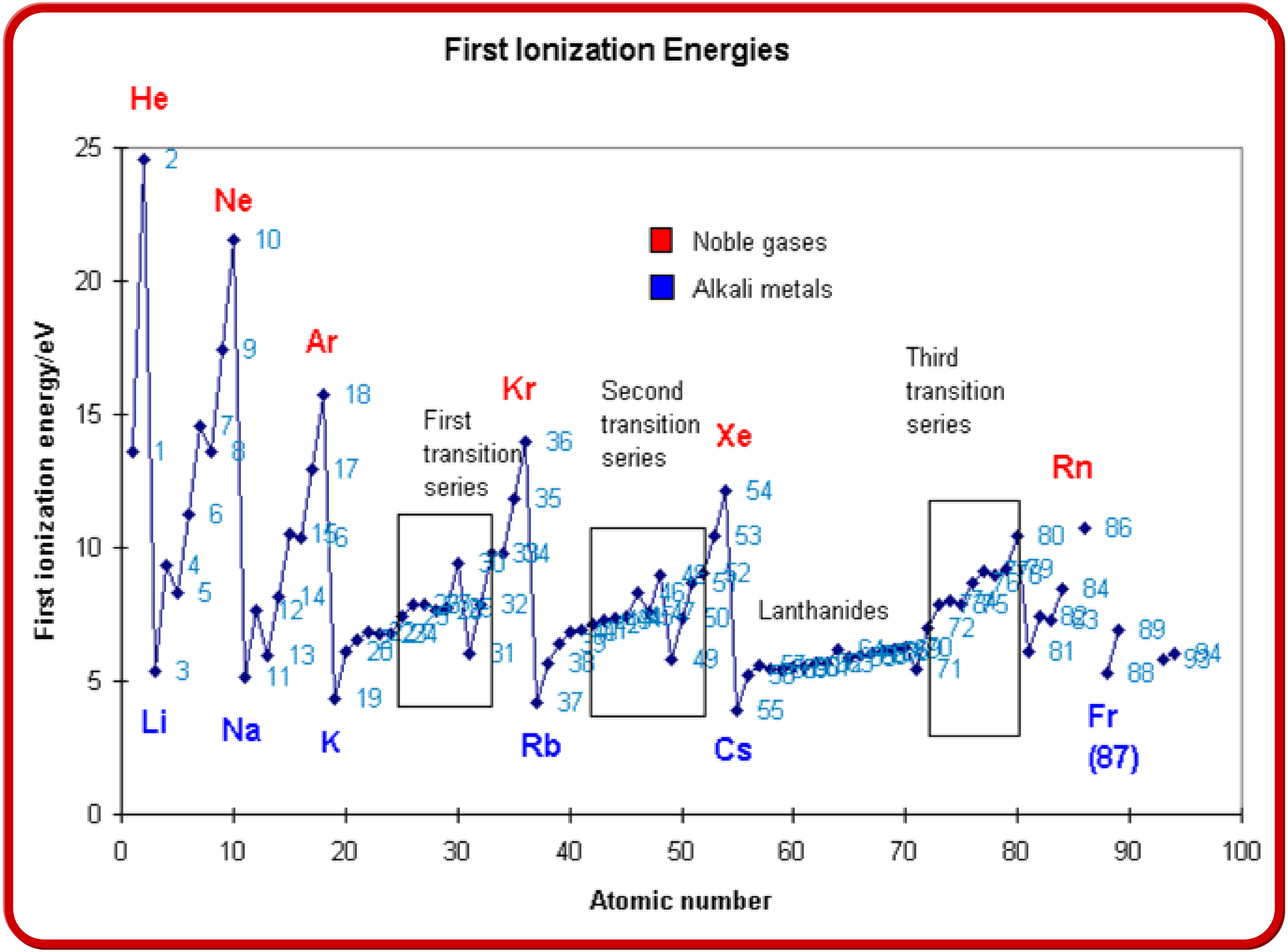
Ionization energies vary periodically, which is explained by the changes in Z_{eff} .



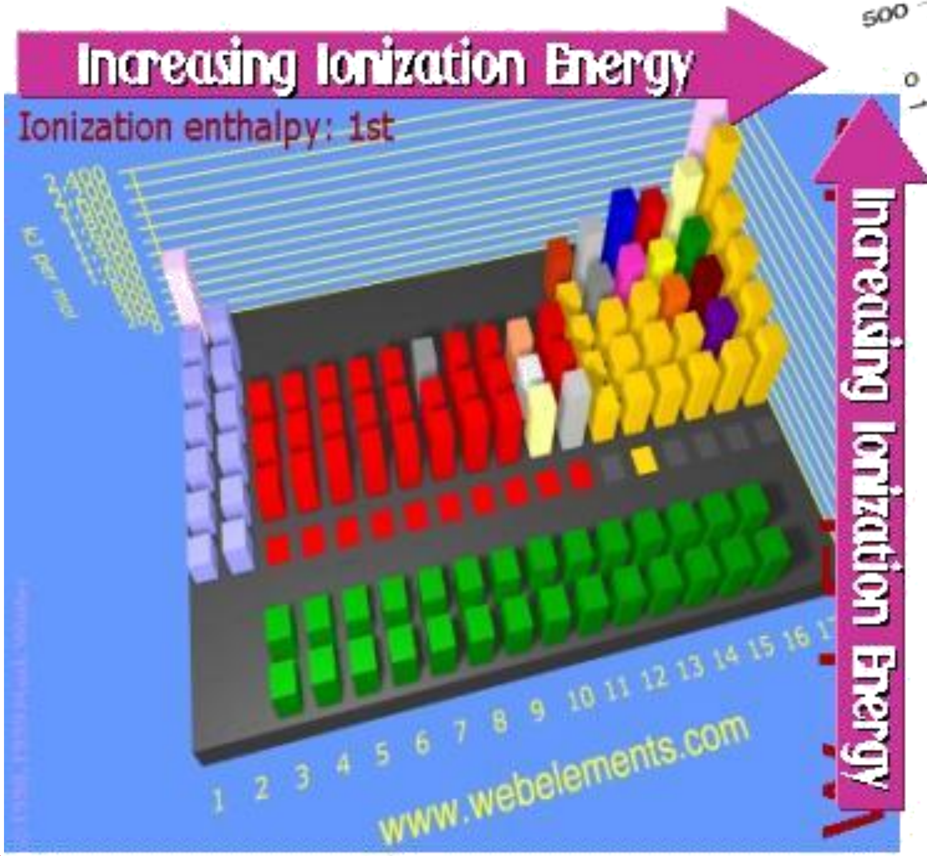
Z_{eff}

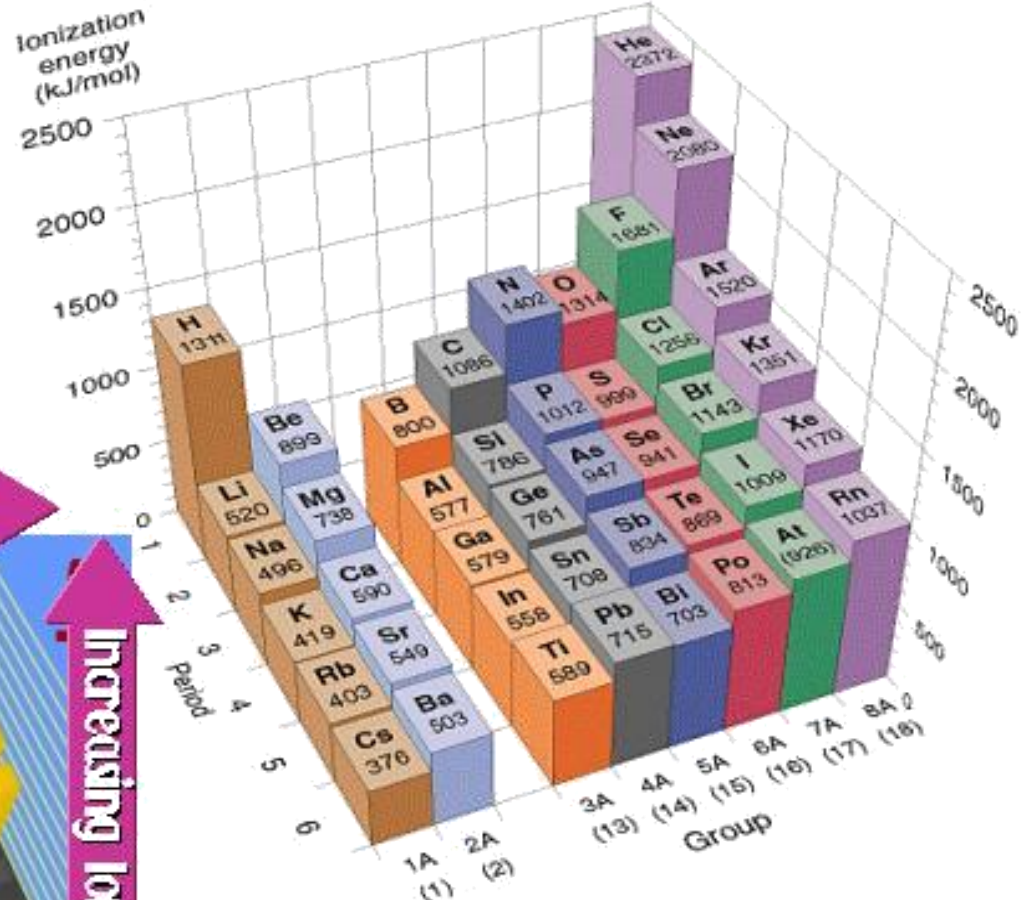
Ionization properties

- **The ionization energies are periodic**
- **The ionization energy tends to increase as atomic number increases in the period**
- **In the column there is a decrease of ionization energy as the atomic # increases**
- **The shielding effect causes the inner electrons to shield the outer electrons from the pull of the protons**
- **Nuclear Charge, Shielding effect, Radius and sublevel all effect ionization energy**


First Ionization Energies
He
Ne
Ar
Kr
Xe
Rn
Li
Na
K
Rb
Cs
Fr
(87)
Noble gases
Alkali metals
First transition series
Second transition series
Third transition series
Lanthanides
First ionization energy/eV
Atomic number
0
5
10
15
20
25
0
10
20
30
40
50
60
70
80
90
100

Ionization Energy Trend





Succesive Values of Ionization Energies, /, for the Elements Sodium Through Argon (kJ/mol)

Na	496	4560	(Inner-shell electrons)				
Mg	738	1450	7730				
Al	578	1820	2750	11,600			
Si	786	1580	3230	4360	16,000		
P	1012	1900	2910	4960	6270	22,200	
S	1000	2250	3360	4560	7010	8500	27,100
Cl	1251	2300	3820	5160	6540	9460	11,100
Ar	1521	2670	3930	5770	7240	8780	12,000

the amount of energy required to remove a 2p e^- (an e^- in a full sublevel) from a Na ion is almost 10 times greater than that required to remove the sole 3s e^-

Periodic Properties

- Another periodic trend dealing with an e- is electron affinity
  - Which is a measure of the ability of an atom to attract or gain an electron.

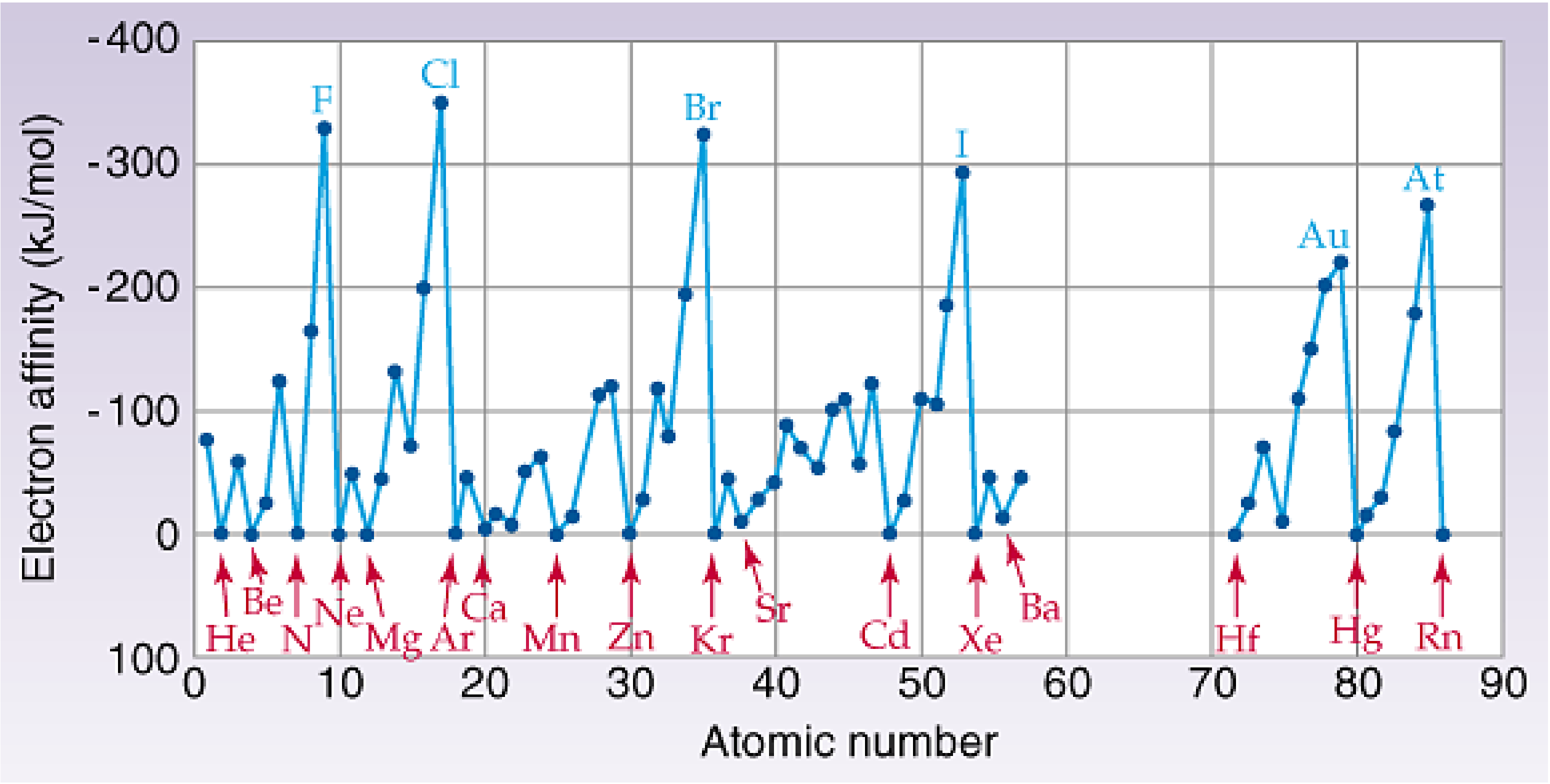
Electron Affinity 01

Electron Energy: Energy change that occurs when an electron is added to an isolated atom in the gaseous state.

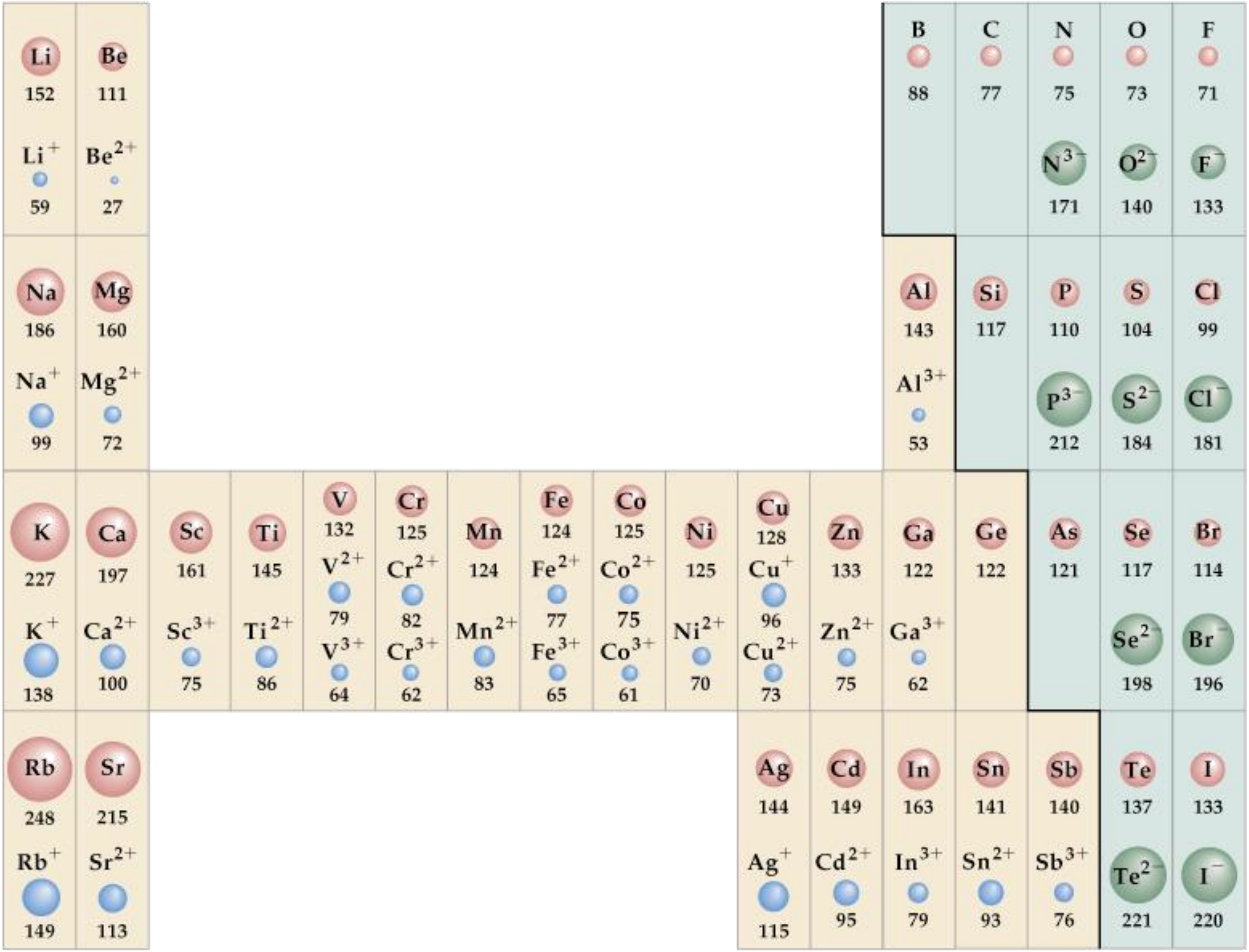
Abbreviation is E_{ea}, it has units of kJ/mol. Values are generally negative because energy is released.

Value of E_{ea} results from interplay of *nucleus electron attraction*, and *electron–electron repulsion*.

Electron Affinity 02



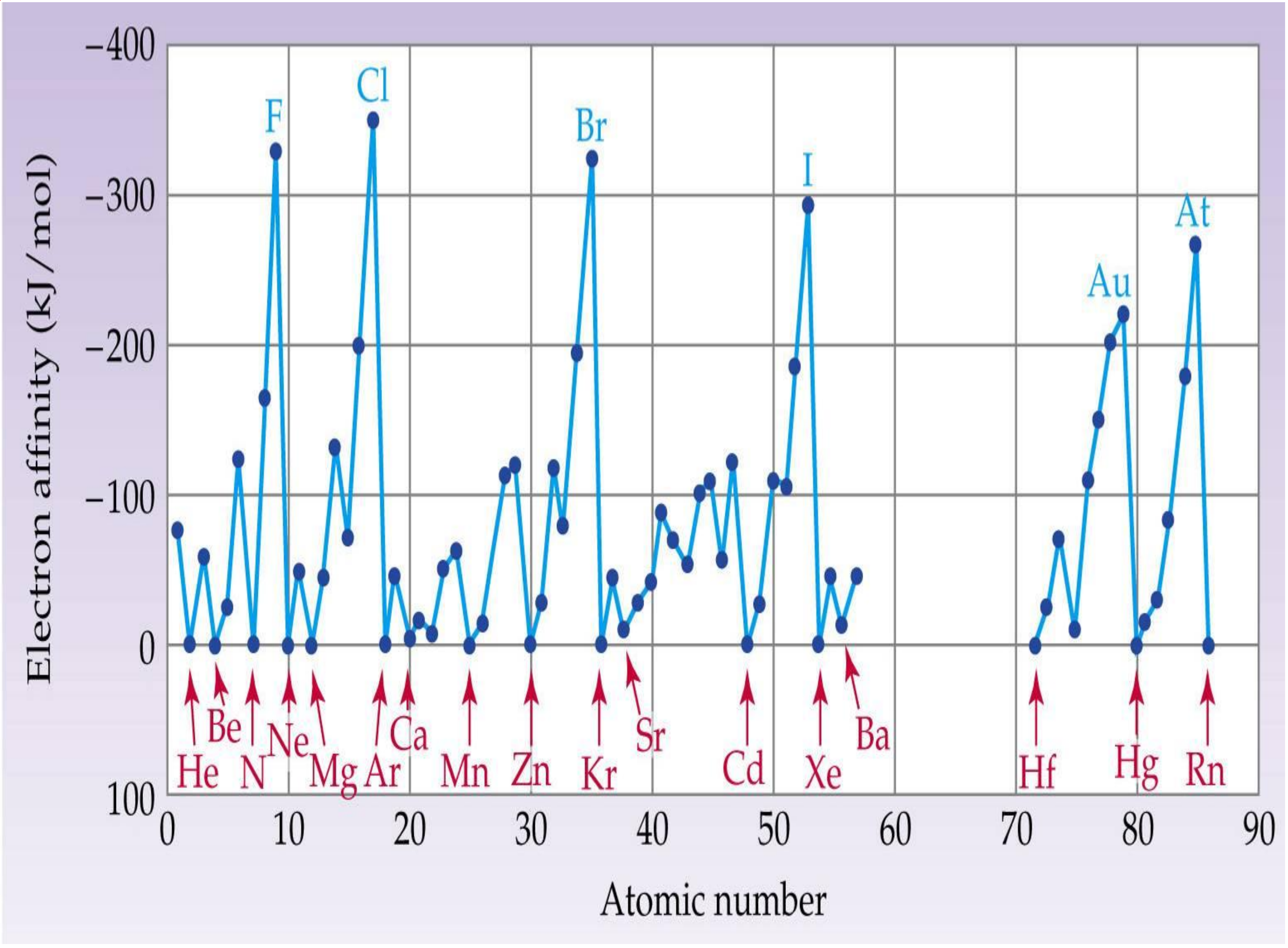
EA Trend



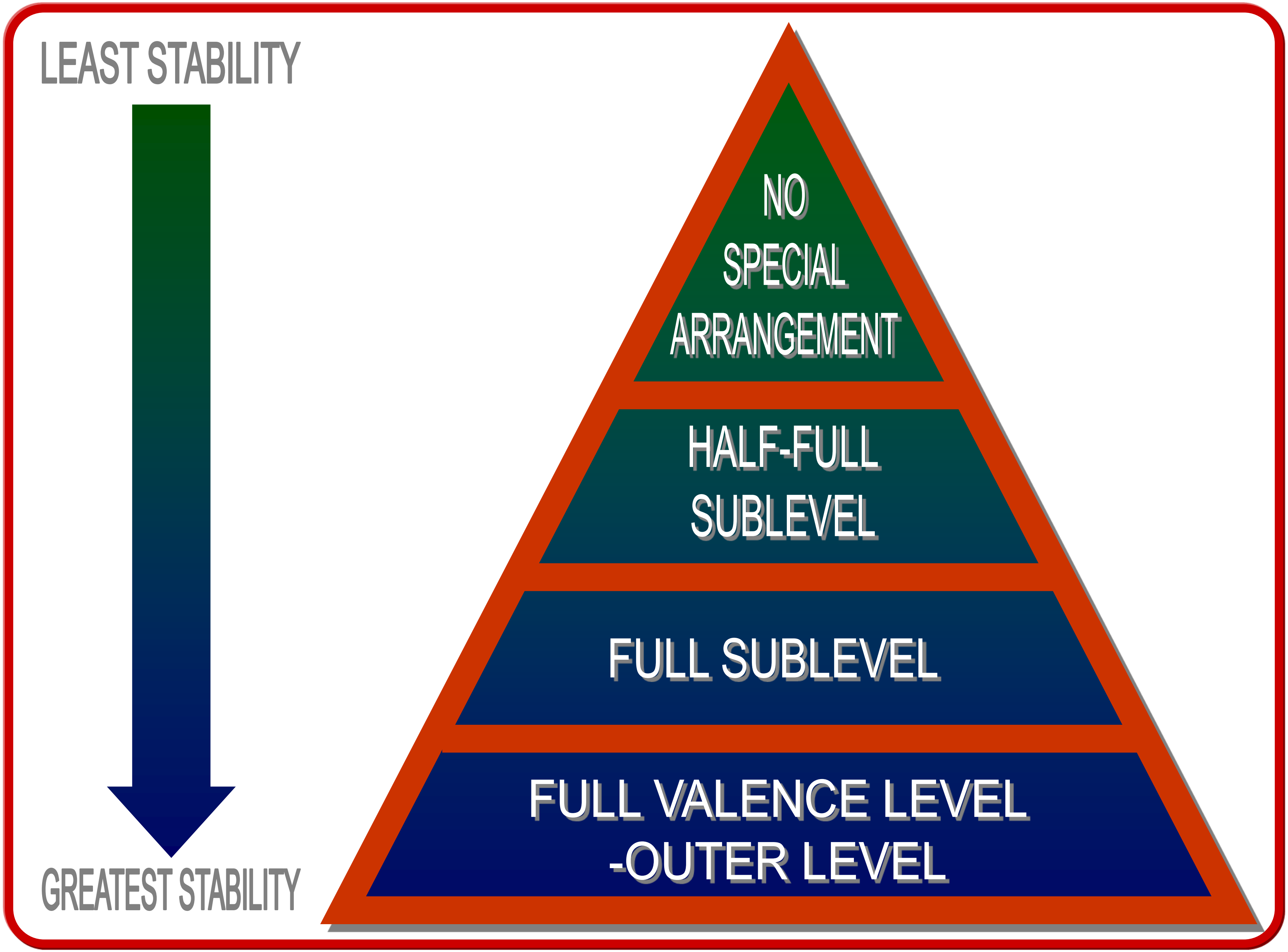
Li
152
Li+
59
Be
111
Be2+
27
B
88
C
77
N
75
N3−
171
O
73
O2−
140
F
71
F−
133
Na
186
Na+
99
Mg
160
Mg2+
72
Al
143
Al3+
53
Si
117
P
110
P3−
212
S
104
S2−
184
Cl
99
Cl−
181
K
227
K+
138
Ca
197
Ca2+
100
Sc
161
Sc3+
75
Ti
145
Ti2+
86
V
132
V2+
79
V3+
64
Cr
125
Cr2+
82
Cr3+
62
Mn
124
Mn2+
83
Fe
124
Fe2+
77
Fe3+
65
Co
125
Co2+
75
Co3+
61
Ni
125
Ni2+
70
Cu
128
Cu+
96
Cu2+
73
Zn
133
Zn2+
75
Ga
122
Ga3+
62
Ge
122
As
121
Se
117
Se2−
198
Br
114
Br−
196
Rb
248
Rb+
149
Sr
215
Sr2+
113
Ag
144
Ag+
115
Cd
149
Cd2+
95
In
163
In3+
79
Sn
141
Sn2+
93
Sb
140
Sb3+
76
Te
137
Te2−
221
I
133
I−
220

Periodic Properties

- **Atoms that tend to accept an e^- are those that tend to give a neg. charge.**
  - The closer to a full outer shell an atom has, the **higher the affinity** (more neg. the measurement)

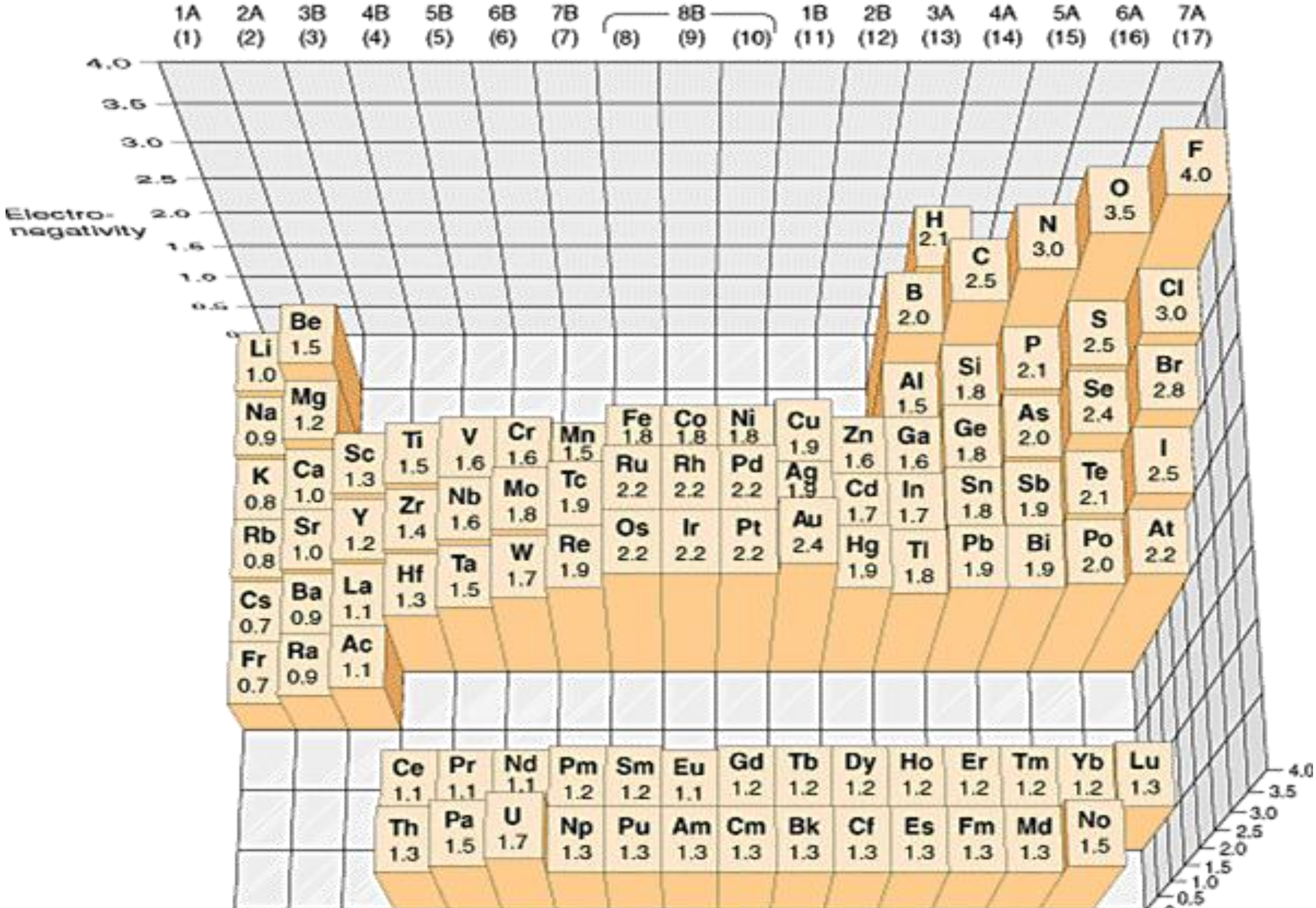


Electron affinity (kJ / mol)
−400
−300
−200
−100
0
100
F
Cl
Br
I
Au
At
He
Be
N
Ne
Mg
Ar
Ca
Mn
Zn
Kr
Sr
Cd
Xe
Ba
Hf
Hg
Rn
0
10
20
30
40
50
60
70
80
90
Atomic number

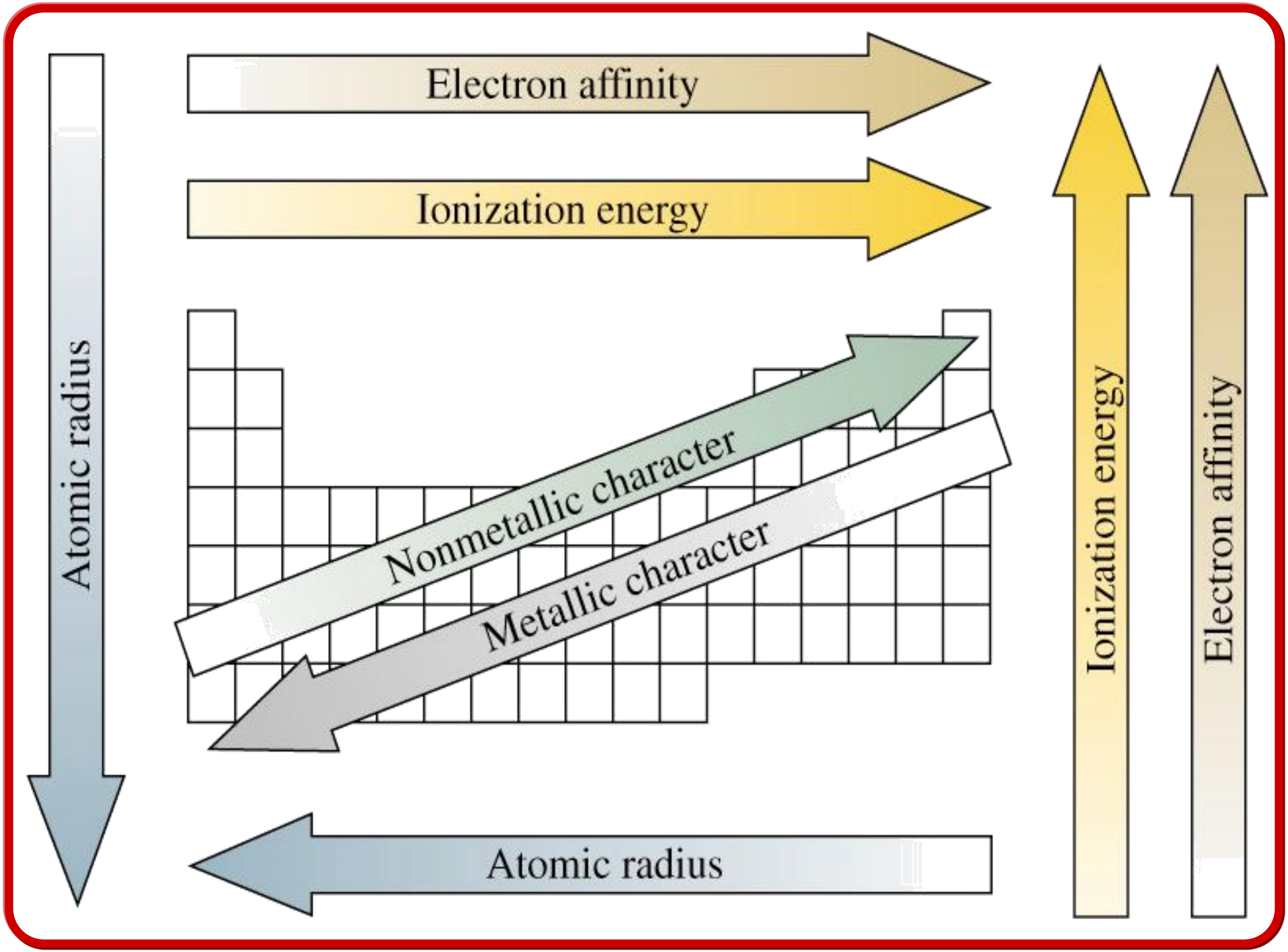

LEAST STABILITY
GREATEST STABILITY
NO SPECIAL ARRANGEMENT
HALF-FULL SUBLEVEL
FULL SUBLEVEL
FULL VALENCE LEVEL -OUTER LEVEL

Periodic Properties

- Electronegativity is a key trend.
  - It reflects the ability of an atom to attract electrons in a chemical bond.
  - F is the most electronegative element and it decreases moving away from F.
- Electronegativity correlates to an atom's ionization energy and electron affinity



1A (1)
2A (2)
3B (3)
4B (4)
5B (5)
6B (6)
7B (7)
8B (8) (9) (10)
1B (11)
2B (12)
3A (13)
4A (14)
5A (15)
6A (16)
7A (17)
Electro-negativity
4.0
3.5
3.0
2.5
2.0
1.5
1.0
0.5
0
H 2.1
Li 1.0
Be 1.5
B 2.0
C 2.5
N 3.0
O 3.5
F 4.0
Na 0.9
Mg 1.2
Al 1.5
Si 1.8
P 2.1
S 2.5
Cl 3.0
K 0.8
Ca 1.0
Sc 1.3
Ti 1.5
V 1.6
Cr 1.6
Mn 1.5
Fe 1.8
Co 1.8
Ni 1.8
Cu 1.9
Zn 1.6
Ga 1.6
Ge 1.8
As 2.0
Se 2.4
Br 2.8
Rb 0.8
Sr 1.0
Y 1.2
Zr 1.4
Nb 1.6
Mo 1.8
Tc 1.9
Ru 2.2
Rh 2.2
Pd 2.2
Ag 1.9
Cd 1.7
In 1.7
Sn 1.8
Sb 1.9
Te 2.1
I 2.5
Cs 0.7
Ba 0.9
La 1.1
Hf 1.3
Ta 1.5
W 1.7
Re 1.9
Os 2.2
Ir 2.2
Pt 2.2
Au 2.4
Hg 1.9
Tl 1.8
Pb 1.9
Bi 1.9
Po 2.0
At 2.2
Fr 0.7
Ra 0.9
Ac 1.1
Ce 1.1
Pr 1.1
Nd 1.1
Pm 1.2
Sm 1.2
Eu 1.1
Gd 1.2
Tb 1.2
Dy 1.2
Ho 1.2
Er 1.2
Tm 1.2
Yb 1.2
Lu 1.3
Th 1.3
Pa 1.5
U 1.7
Np 1.3
Pu 1.3
Am 1.3
Cm 1.3
Bk 1.3
Cf 1.3
Es 1.3
Fm 1.3
Md 1.3
No 1.5



Electron affinity
Ionization energy
Atomic radius
Nonmetallic character
Metallic character
Ionization energy
Electron affinity
Atomic radius