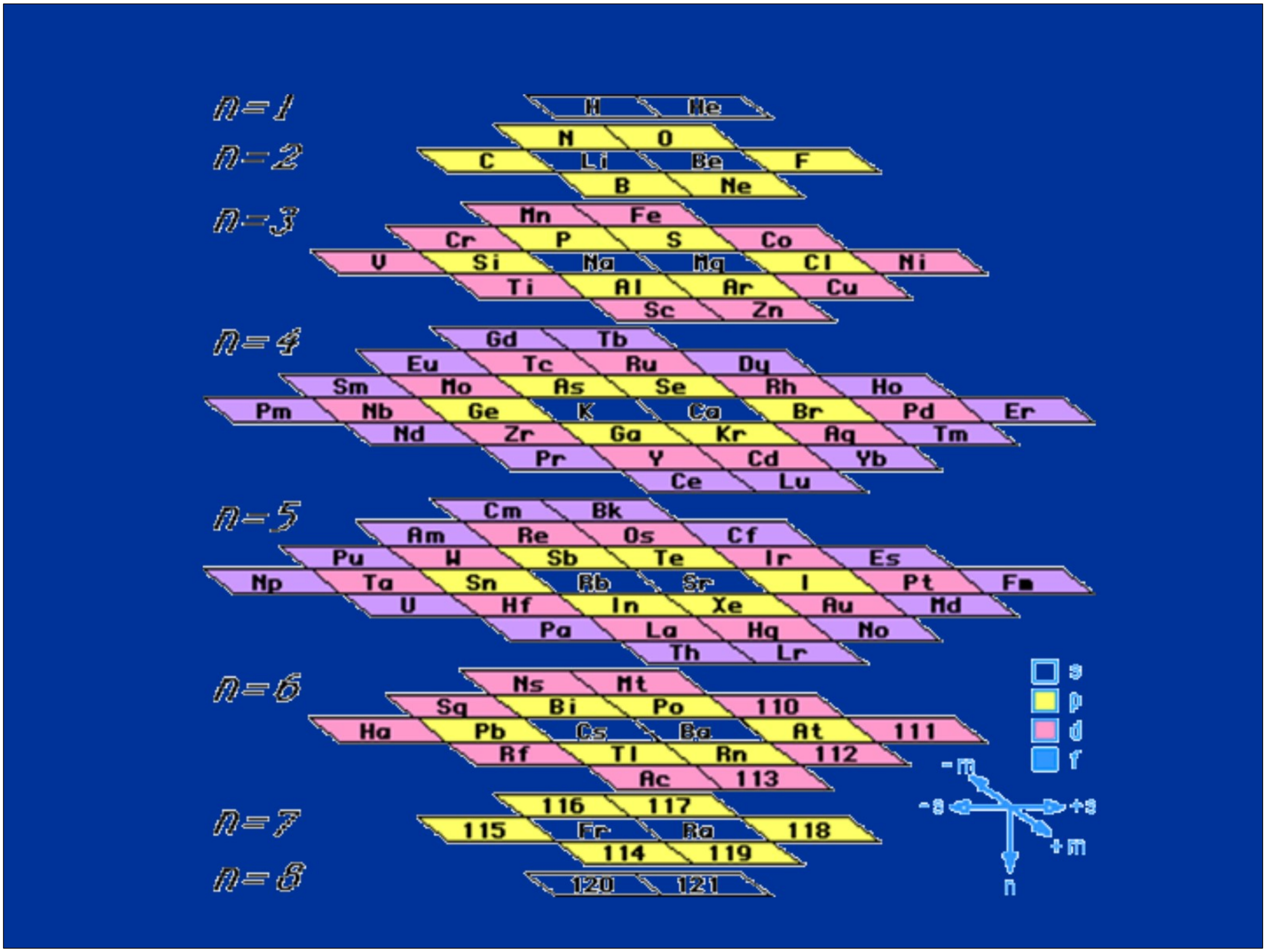
The Periodic Table

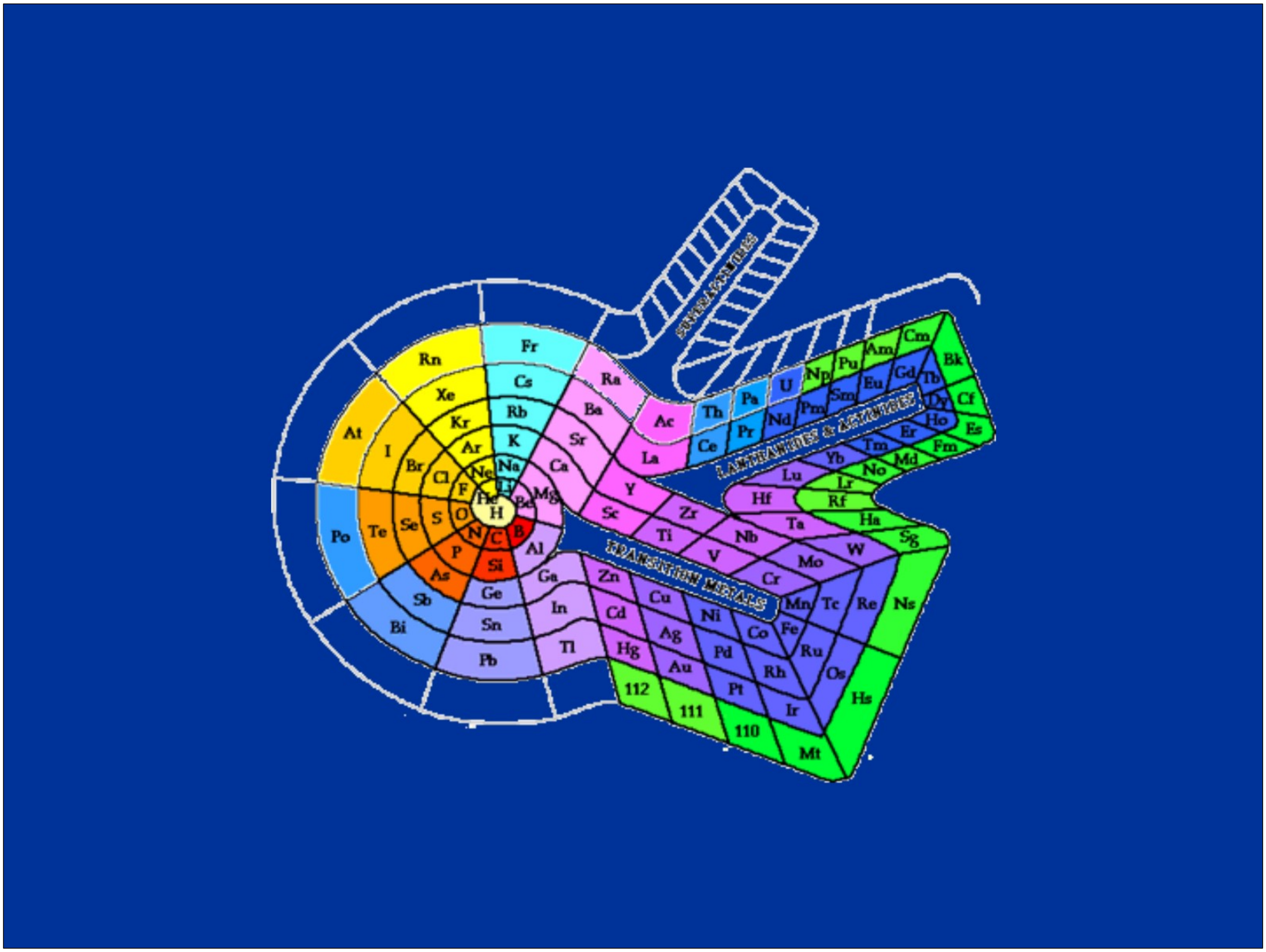
Chapter 5

Dmitri Mendeleev

- Dmitri Mendeleev developed the periodic table in 1869
  1. He grouped elements with similar chemical and physical properties together
  2. He arranged the elements in increasing atomic mass.
  3. He predicted the properties of elements not yet discovered.



n=1
H He
N O
n=2
C Li Be F
B Ne
n=3
Mn Fe
Cr P S Co
V Si Na Mg Cl Ni
Ti Al Ar Cu
Sc Zn
n=4
Gd Tb
Eu Tc Ru Dy
Sm Mo As Se Rh Ho
Pm Nb Ge K Ca Br Pd Er
Nd Zr Ga Kr Ag Tm
Pr Y Cd Yb
Ce Lu
n=5
Cm Bk
Am Re Os Cf
Pu W Sb Te Ir Es
Np Ta Sn Rb Sr I Pt Fm
U Hf In Xe Au Md
Pa La Hg No
Th Lr
n=6
Ns Mt
Sg Bi Po 110
Ha Pb Cs Ba At 111
Rf Tl Rn 112
Ac 113
116 117
n=7
115 Fr Ra 118
114 119
n=8
120 121
s
p
d
f
-m
-s
+s
+m
n


SUPERACTINIDES
LANTHANIDES & ACTINIDES
TRANSITION METALS
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
Ac
La
Y
Sc
Ti
Zr
Hf
V
Nb
Ta
Cr
Mo
W
Mn
Tc
Re
Fe
Ru
Os
Co
Rh
Ir
Ni
Pd
Pt
Cu
Ag
Au
Zn
Cd
Hg
112
111
110
Mt
Hs
Ns
Sg
Ha
Rf
Lr
Lu
Th
Pa
U
Np
Pu
Am
Cm
Bk
Cf
Es
Fm
Md
No
Ce
Pr
Nd
Pm
Sm
Eu
Gd
Tb
Dy
Ho
Er
Tm
Yb

Periodic Table

- Today's periodic table, the elements are arranged in <u>increasing atomic number</u> so elements with similar properties <u>fall in the same group</u>.

  Note: Ar and K both in according to atomic number, not atomic mass

Periodic Law

- Periodic Law: When elements are arranged in increasing atomic numbers, there is a repeating (periodic) pattern to the properties. See tables p. 142, 144, 148, 152.

Group Names

- Group 1 elements – Alkali Metals –
  - Very, very reactive metals
- Group 2 elements – Alkaline Earth Metals –
  - Very reactive metals
- Group 17 elements – Halogens –
  - Very reactive nonmetals
- Group 18 elements – Noble gases –
  - Nonreactive elements

Transition Metals

- Group 3 to 12 – Transition Metals
  - Typical metals

- Inner Transition Elements : lower two rows detached from main table
  - Lanthanides – atomic numbers from 58 to 71
  - Actinides – atomic numbers from 90 to 103

- Write electron configuration of the following elements. Then answer these questions. How many valence electrons does each have? Which group are these elements found in?

Li
Na
K
Rb

Be
Mg
Ca
Sr

F
Cl
Br

He
Ne
Ar

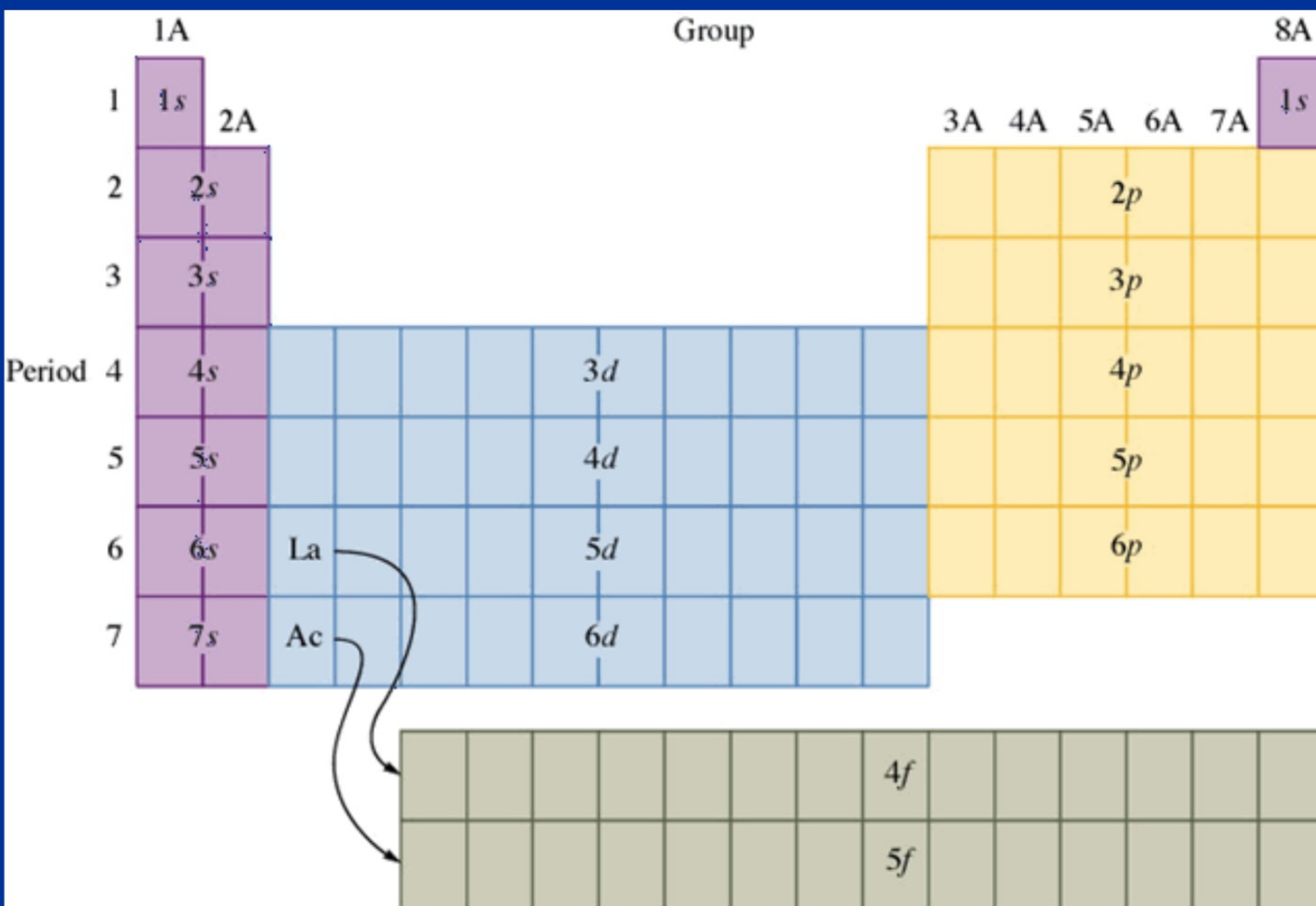
Blocks of Elements

S-block elements- valence electrons are filling the **s** orbitals- **Groups 1-2**

D-block elements-valence electrons are filling the **d** orbtials- **Groups 3-12**

P-block elements-valence electrons are filling the **p** orbitals- **Groups 13-18 except He**

F-block elements-valence electrons are filling the **f** orbitals- **Lanthanide and Actinide**



1A
Group
8A
2A
3A
4A
5A
6A
7A
Period
1
2
3
4
5
6
7
1s
2s
3s
4s
5s
6s
7s
La
Ac
3d
4d
5d
6d
2p
3p
4p
5p
6p
1s
4f
5f

Practice

- p.133, 136, 138, 139 and section review p139

- Practice reading off the electron configuration directly from the periodic table.

Coulomb's Law

$$F = \frac{kq_1q_2}{d^2}$$

F - force of attraction
k - a constant
q - charge on particle
d - distance between partiles

Distance

Attractive forces dissipate with increased distance.

Shielding

Electrons in the “core” effectively shield the nucleus’ attractive force for the valence electrons.

-

Use this ONLY when going up and down the table, NOT across.

Effective Nuclear Charge, Z_{eff}

The net positive charge attracting the electron. (Essentially equal to the group #).

Think of the 1's having a Z_{eff} of one while the 17's have a Z_{eff} of 17!

The idea is that the higher the Z_{eff}, the more attractive force there is emanating from the nucleus, drawing electrons in or holding them in place.

What is atomic Radii?

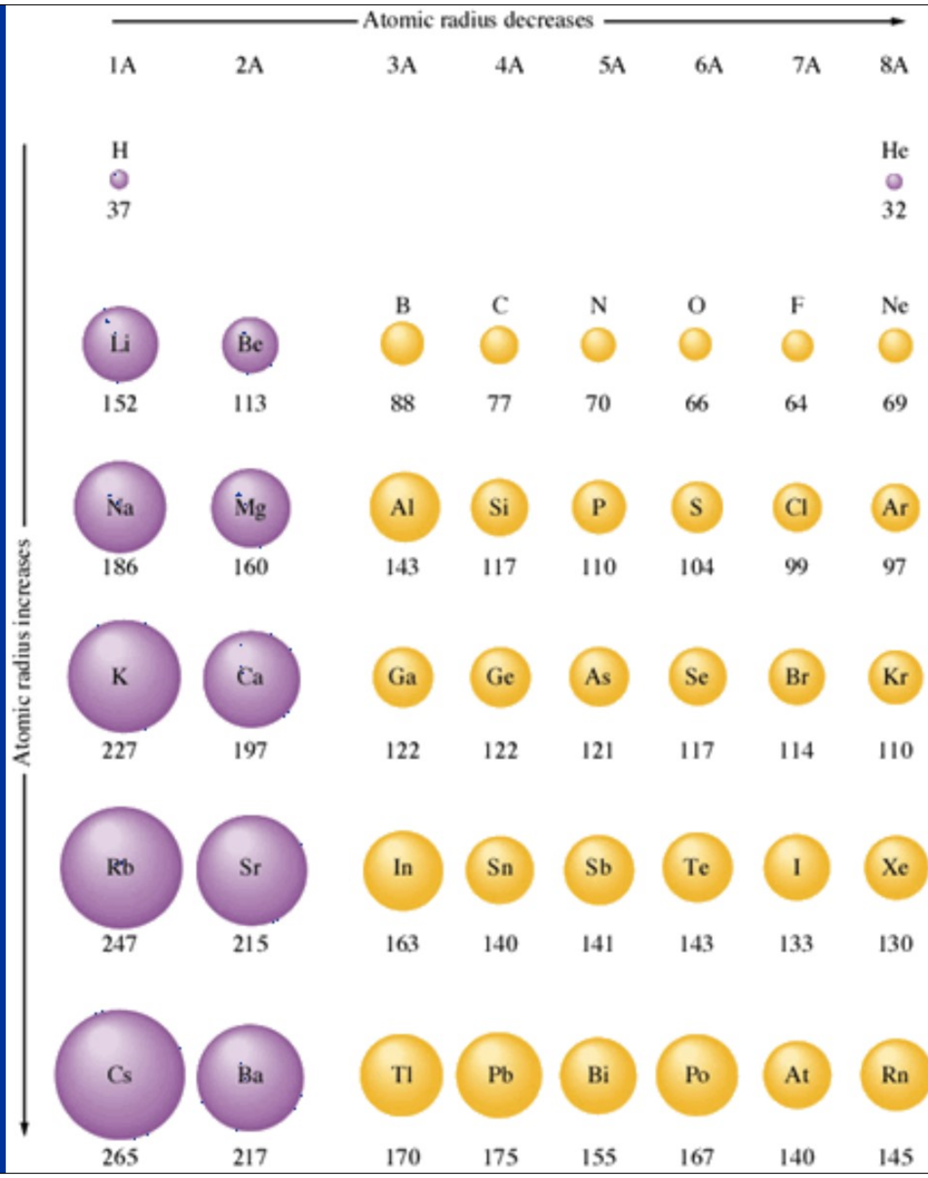
- **Atomic Radii** – **½ the distance between the nuclei of 2 of the same atoms bonded together.**

What happens to the atomic radii as you go across a period?

- atoms get smaller - because the Zeff gets larger, attracting the electron cloud closer to the nucleus.

What happens to atomic radii as you go down a group?

- atoms get larger - because energy levels are being added


Atomic radius decreases
Atomic radius increases
1A
2A
3A
4A
5A
6A
7A
8A
H
37
He
32
Li
152
Be
113
B
88
C
77
N
70
O
66
F
64
Ne
69
Na
186
Mg
160
Al
143
Si
117
P
110
S
104
Cl
99
Ar
97
K
227
Ca
197
Ga
122
Ge
122
As
121
Se
117
Br
114
Kr
110
Rb
247
Sr
215
In
163
Sn
140
Sb
141
Te
143
I
133
Xe
130
Cs
265
Ba
217
Tl
170
Pb
175
Bi
155
Po
167
At
140
Rn
145

Ions

- Ion is an atom with a <u>positive or negative charge</u>.

- Cation formed by <u>loss of electron(s)</u> and has a <u>positive</u> charge.

- Anion formed by a gain of electron(s) and has a negative charge.

- Illustrate example of atom, cation, and anion.

Ionic Radii

- Ionic Radii – <u>positive ions are smaller than neutral atom due to loss of electron</u> –
- <u>negative ions are larger than neutral atom due to gain of electron</u>

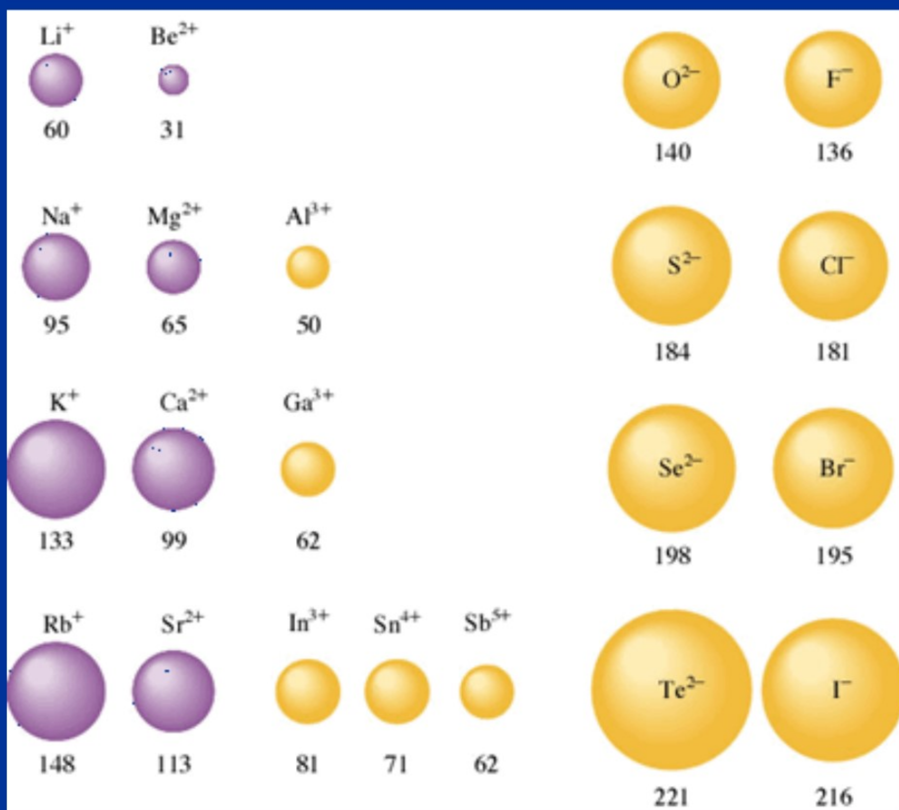
- Ionic radii trends:
- Going across a period – <u>atoms tend to get smaller -</u>
- Going down a group – <u>atoms tend to get larger</u>

Cations

shrink big time since the nucleus is now attracting fewer electrons

Anions

expand since the nucleus is now attracting MORE electrons



Li+
60
Be2+
31
O2–
140
F–
136
Na+
95
Mg2+
65
Al3+
50
S2–
184
Cl–
181
K+
133
Ca2+
99
Ga3+
62
Se2–
198
Br–
195
Rb+
148
Sr2+
113
In3+
81
Sn4+
71
Sb5+
62
Te2–
221
I–
216

Which is larger?

O^{-2} or F^{-1}

Explain.

What is ionization energy?

<u>Energy it takes to remove an electron from a neutral atom</u>

What happens to the IE as you go across a period?

IE increases - because of an increase in the Zeff. The valence electron is held more tightly, therefore is harder to remove.

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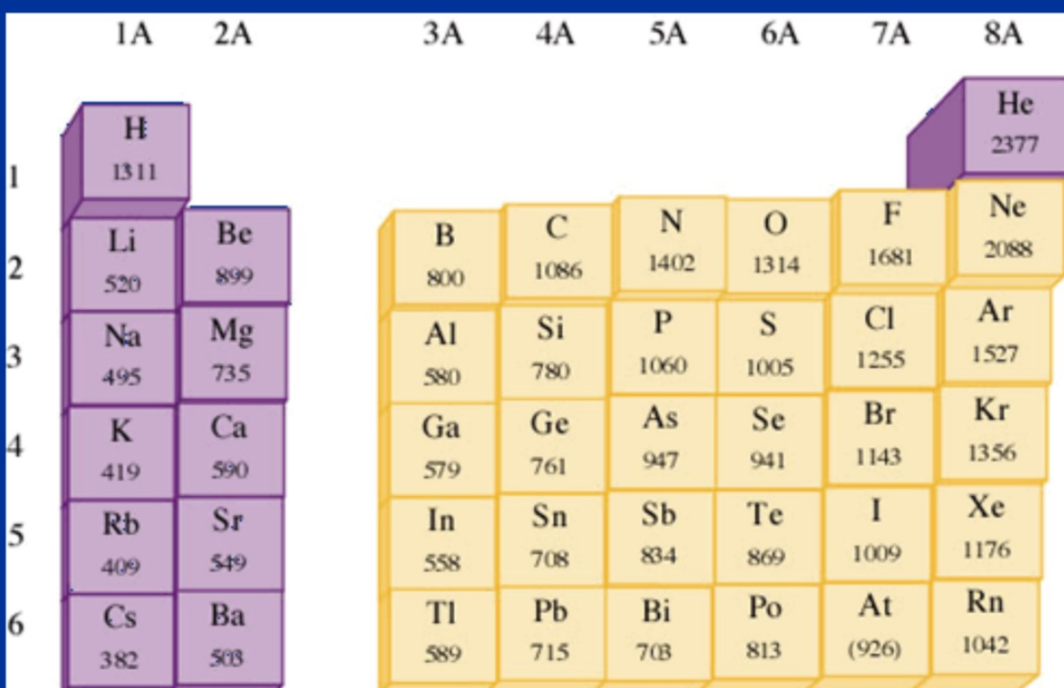
What happens to the IE as you go down a group?

- IE decreases - because the electron is farther from the nucleus & the attractive force between the nucleus and valence electron dissipates.


1A 2A 3A 4A 5A 6A 7A 8A
1 2 3 4 5 6
H 1311
He 2377
Li 520
Be 899
B 800
C 1086
N 1402
O 1314
F 1681
Ne 2088
Na 495
Mg 735
Al 580
Si 780
P 1060
S 1005
Cl 1255
Ar 1527
K 419
Ca 590
Ga 579
Ge 761
As 947
Se 941
Br 1143
Kr 1356
Rb 409
Sr 549
In 558
Sn 708
Sb 834
Te 869
I 1009
Xe 1176
Cs 382
Ba 503
Tl 589
Pb 715
Bi 703
Po 813
At (926)
Rn 1042

Electronegativity

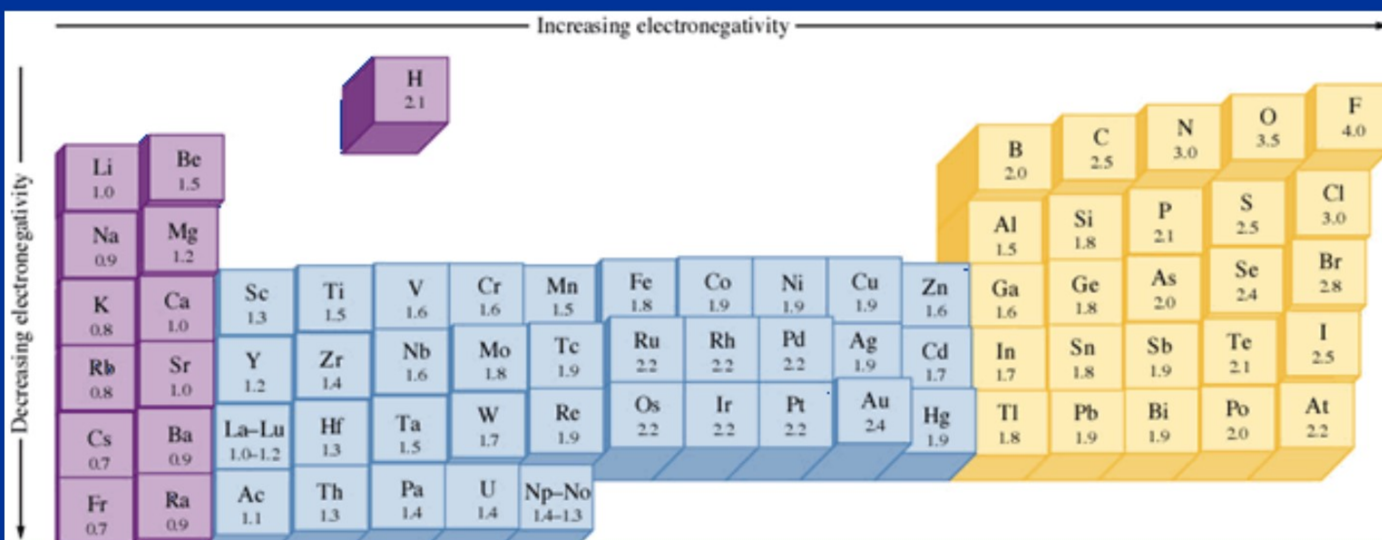
- Electronegativity – <u>measure of the ability of an atom to attract electrons to itself</u>

- Electronegativity trends:
- **Going across a period – increases – the nucleus attracts the electron more since the Zeff increases.**

- **Going down a group – <u>decreases</u> – <u>b/c the nuclear attraction to the valence electron decreases with distance from nucleus.</u>**

Prac P.152 and section review p.154

Pauling's Scale



Fluorine is the most E_n

Francium is the least E_n

Electron Affinity

An atom's liking for electrons – energy associated with the addition of an electron to an atom

Going down a group

EA decreases - The nucleus is farther from the valence level and there is more shielding. So the EA decreases down a group.

Going across a period

EA increases because the Zeff increases which draws the electron.

Why is F the most?

Highest Z_{eff} and smallest so that the nucleus is closest to the "action".

Why is Fr the least?

Lowest Z_{eff} and largest so that the nucleus is farthest from the "action".