Reminder & Announcement

Starting the week of October 15th Group 1 is doing the tutorial on Strong Acids and Bases and Redox. Group 2 is doing the Acid/Base lab.

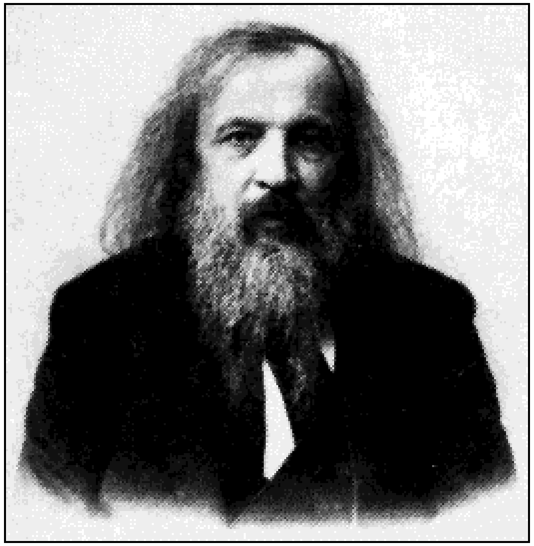
I am away next Monday thru Wednesday. Professor Martin will substitute for me in class and tutorial.

The Periodic Table

When the elements are grouped together on the basis of their physical and chemical properties, they fall into groups illustrated by the **periodic table**

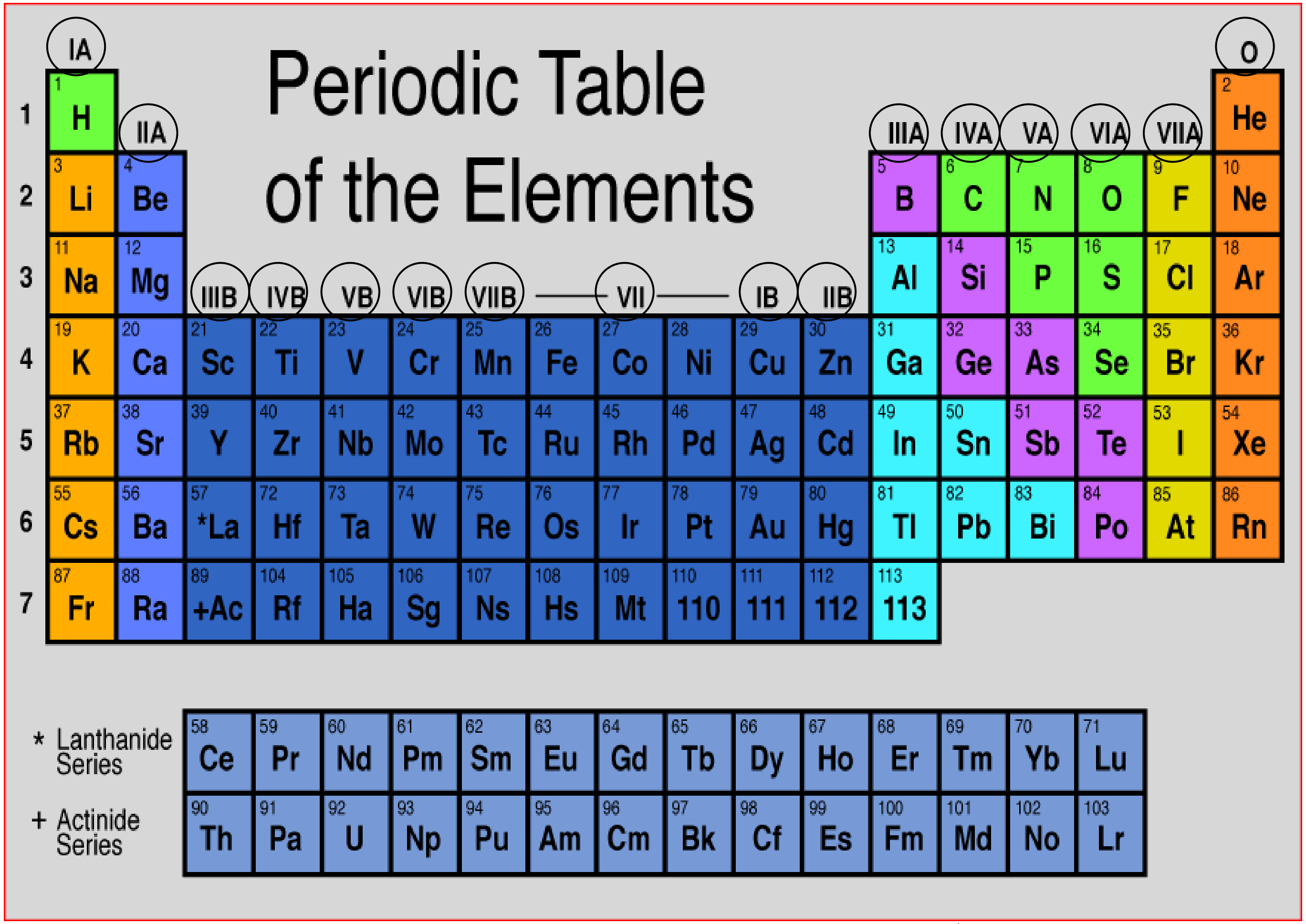
First devised by D. Mendeleev long before the advent of quantum theory

Who was Mendeleev?



- Demitri Mendeleev was a Russian scientist.
- He was born 1834 and died 1907.
- Although not the only chemist of his day to make tables of elements, he was the first to use his table to predict the existence of other elements such as Ge, Ga and Sc.
- As Director of the Russian Bureau of Weights and Measures, he decided that Russian vodka should be 80 proof (40% ethanol).

1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18



Groups

Periods (Horizontal Rows)

20th century chemistry has shown that these groups have **similar electronic configurations**

This demonstrating that the behavior of the elements is a consequence of the **number and configuration of their electrons only**:

The atomic nucleus is unimportant chemically

The electrons that are the most important in determining the properties of an element are those in the **highest energy**, **incompletely filled orbitals**:

≡ the **valence electrons**

Main group elements are those with s or p subshells being filled, with other subshells being full or empty

Those with d orbitals being filled are termed **transition metal** elements. Each d subshell holds 10 electrons

Elements in the f-block are those with f-orbitals being filled.
These orbitals hold 14 electrons.
The elements with the 4f orbitals being filled are called the **lanthanides**
Those with the 5f orbitals being filled are called the **actinides**

Since elements in any one group have the same number of valence electrons, it is no surprise that they have similar chemical properties

Group 1A Alkali metals valence config. = ns^1

Li, Na, K, Rb, Cs, Fr Old English for base

Group 2A Alkaline Earths valence config. = ns^2

Be, Mg, Ca, Sr, Ba, Ra Earth = old English for nonmetallic substance insoluble in water

Group 3A (13) valence config. = ns^2np^1

B, Al, Ga, In, Tl

Group 4A (14) valence config. = ns^2np^2

C, Si, Ge, Sn, Pb

Group 5A (15) valence config. = ns^2np^3

N, P, As, Sb, Bi

Group 6A (16) Chalcogens valence config. = ns^2np^4

O, S, Se, Te, Po (Greek for "ore former")

Group 7A (17) Halogens valence config. = ns^2np^5

F, Cl, Br, I, At (French and Greek Halos = salt; genes = production)

Group 8A (18) Noble gases valence config. = ns^2np^6

He($1s^2$), Ne, Ar, Kr, Xe, Rn

Noble gases are extremely stable and not very reactive due to their full valence shell. (They can undergo some chemical reactions).

Other elements can lose or gain electron to attain the same electronic configuration as these noble gases.

For example:

K tends to give away one electron and Cl to accept one electron to attain the Ar configuration

Periodic Trends in Atomic Properties

1. Atomic Size

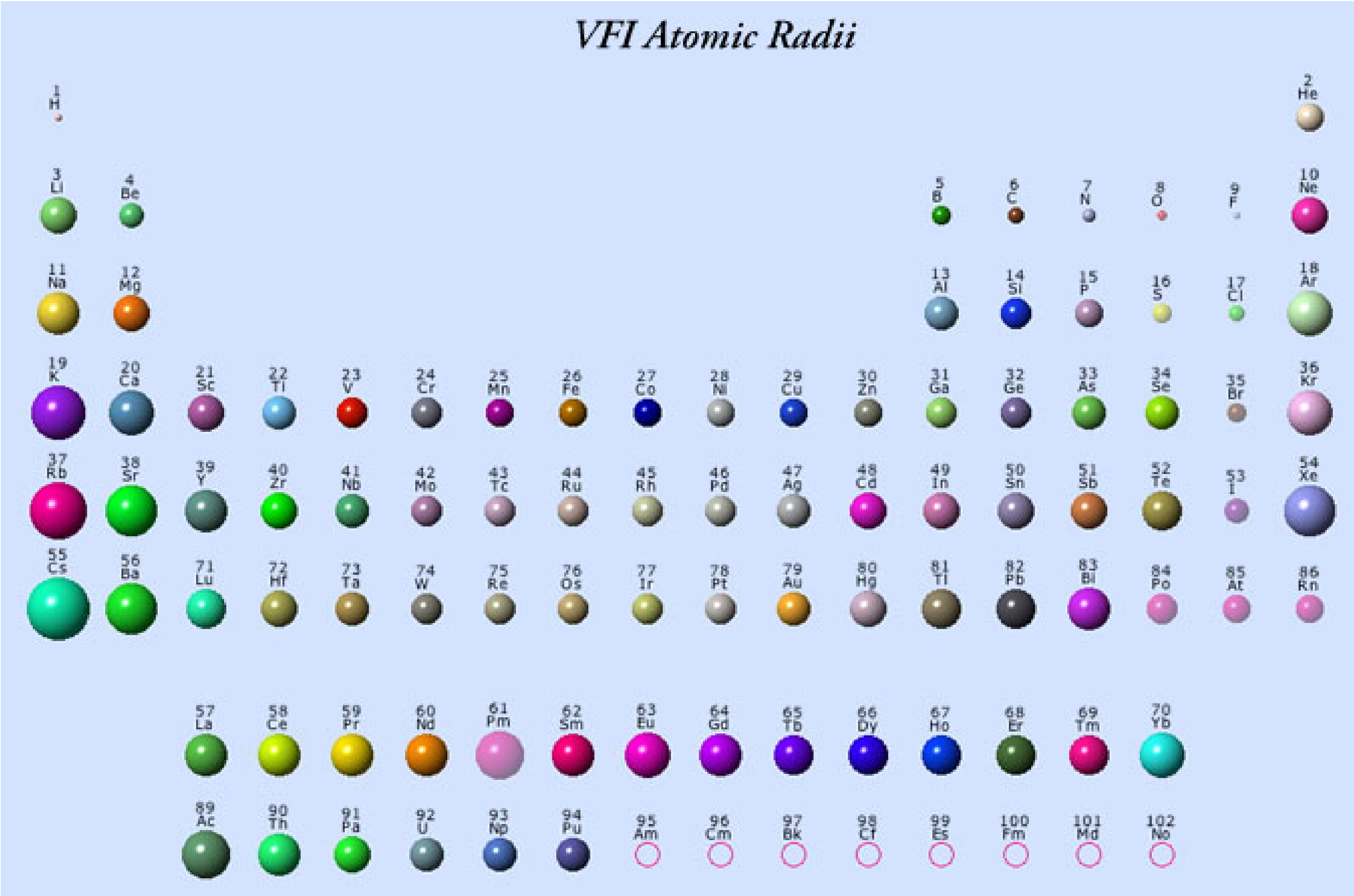
The atomic radius reflects the distance from the nucleus to the highest occupied electronic orbital.
Therefore we expect the following:

i) Atomic radii **increase** from top to bottom within a group.
For example: r(Li) < r(Na) < r(K) < r(Rb) < r(Cs)

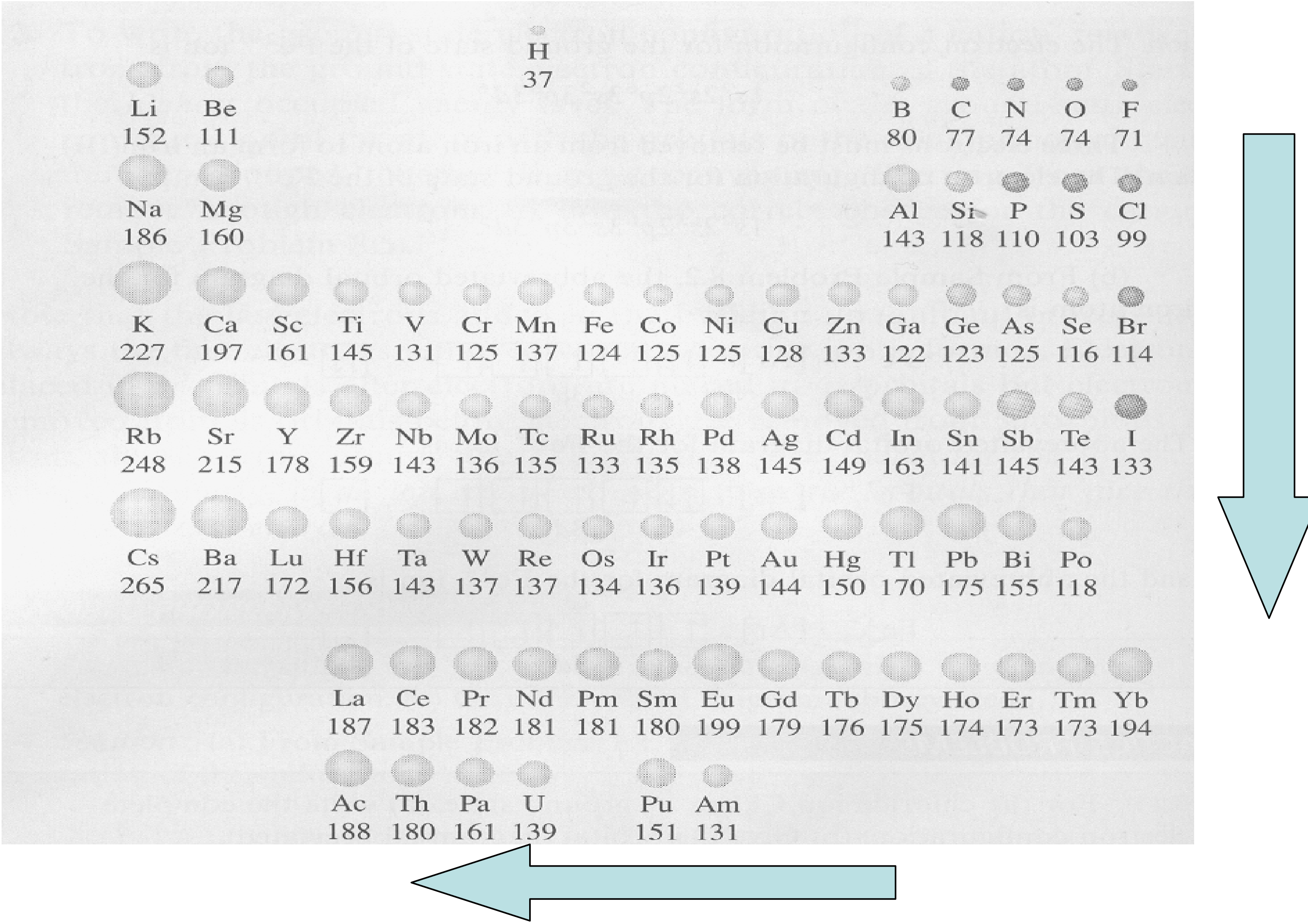
However
ii) Atomic radii **decrease** from left to right across a period
For example: r(Li) > r(Be) > r(B) > r(C) > r(N) >r(O) > r(F)

This is because the electrons in the same n-orbital (here n = 2) are more strongly attracted to the greater number of protons in the nuclei of the heavier elements

iii) Cations are smaller than the neutral parent atom. For example r(Li^+) < r(Li)
iv) Anions are larger than the neutral parent atom. For example r(Cl^-) > r(Cl)


VFI Atomic Radii
1 H
2 He
3 Li
4 Be
5 B
6 C
7 N
8 O
9 F
10 Ne
11 Na
12 Mg
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
19 K
20 Ca
21 Sc
22 Ti
23 V
24 Cr
25 Mn
26 Fe
27 Co
28 Ni
29 Cu
30 Zn
31 Ga
32 Ge
33 As
34 Se
35 Br
36 Kr
37 Rb
38 Sr
39 Y
40 Zr
41 Nb
42 Mo
43 Tc
44 Ru
45 Rh
46 Pd
47 Ag
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
55 Cs
56 Ba
71 Lu
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
57 La
58 Ce
59 Pr
60 Nd
61 Pm
62 Sm
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tm
70 Yb
89 Ac
90 Th
91 Pa
92 U
93 Np
94 Pu
95 Am
96 Cm
97 Bk
98 Cf
99 Es
100 Fm
101 Md
102 No

Relative atomic radii in picometers

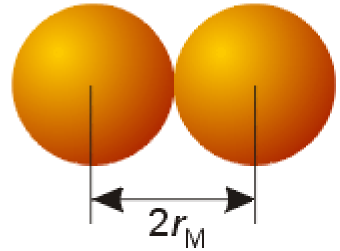


Atomic Sizes and Radii (An aside: not responsible for this)

There is no one way to define the 'size' of an atom, like we can define the size of a billiard ball. This is a limitation of quantum theory.

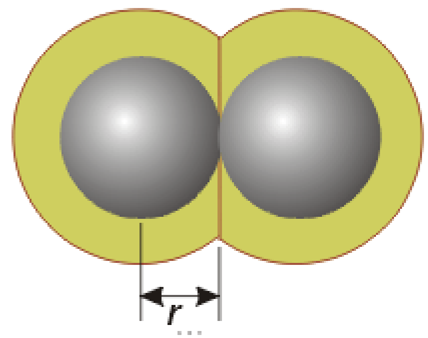
Common approaches:

- Single-bond covalent radii - radii assigned within a covalent bonding situation.
- Ionic radii - radii assigned to ions of the elements in predominantly ionic compounds
- Metallic or van der Waals radii - non-bonded contact distances.

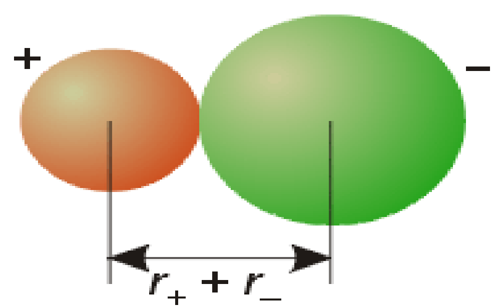


1 Metallic radius

Could also be taken as van der Waals radius



2 Covalent radius



3 Ionic radius

2.Ionization Energies

Ionization energy, IE, is the energy required to remove one electron from an atom or ion;
an **endothermic** process

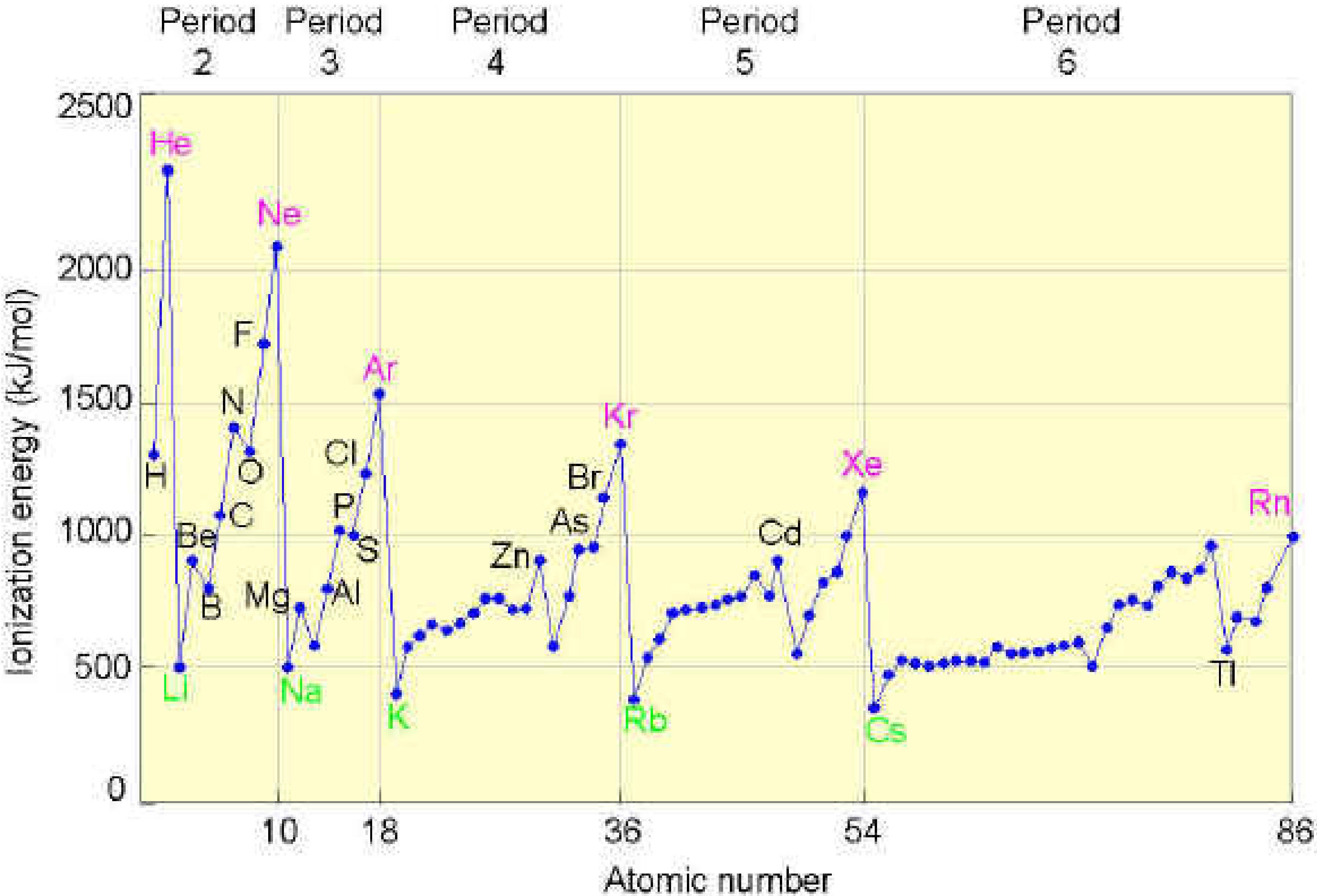
that is, $A \rightarrow A^{+} + 1\ e^{-} \quad \Delta H = +ve$

The energy, in kJ mol^{-1}, required to remove 1 electron from the neutral atom (1st ionization limit) depends on the orbital in which the electron resides

Two general trends:

1.) IEs **decrease down a group**
(an electron removed from an orbital more distant from the nucleus is less tightly bound)

2.) IEs **increase from left to right** across a period
(electrons being removed from orbitals of equal n are subject to increasing nuclear attraction)



Period 2
Period 3
Period 4
Period 5
Period 6
Ionization energy (kJ/mol)
2500
2000
1500
1000
500
0
He
Ne
F
Ar
N
Kr
Cl
O
H
Br
Xe
C
P
Be
As
Cd
Rn
S
Zn
B
Mg
Al
Li
Na
K
Rb
Cs
Tl
10
18
36
54
86
Atomic number