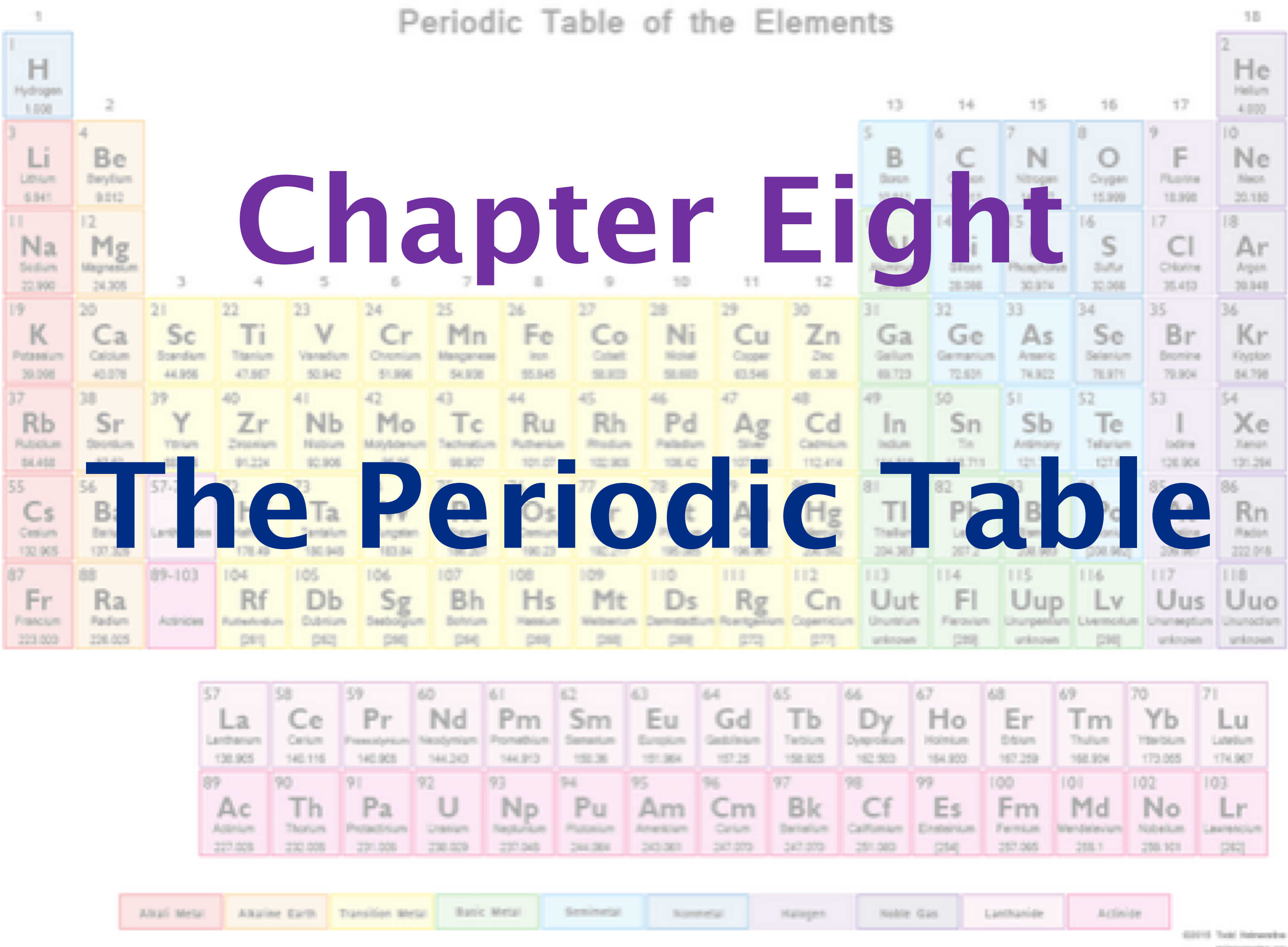
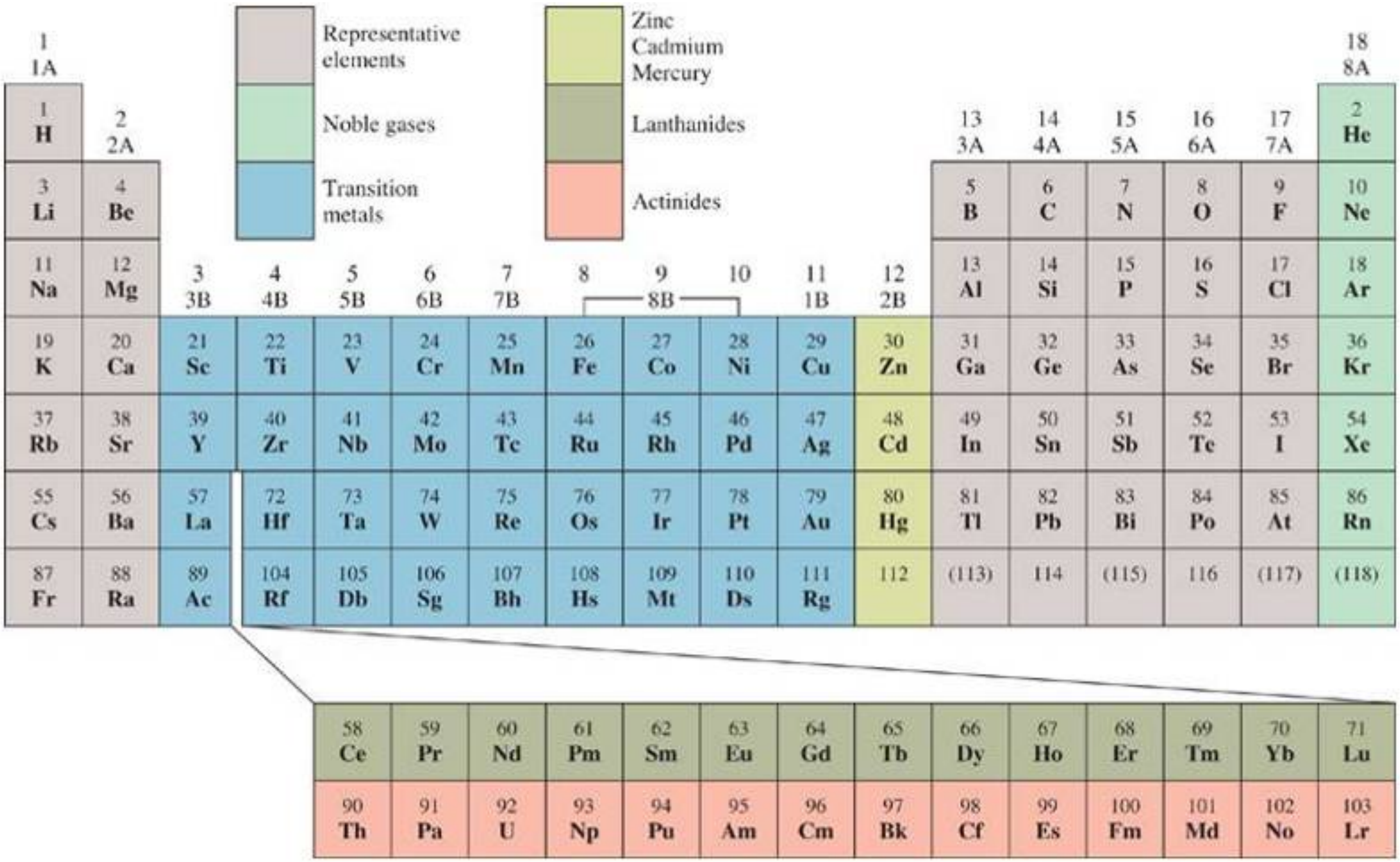


Chapter Eight

The Periodic Table

The Modern Periodic Table

- Representative Elements: (main group elements)
  - Incomplete s or p shell determine elemental properties
- Transition metals – d orbitals also play a role in bonding



Some Groups in the Periodic Table

Alkali Metals

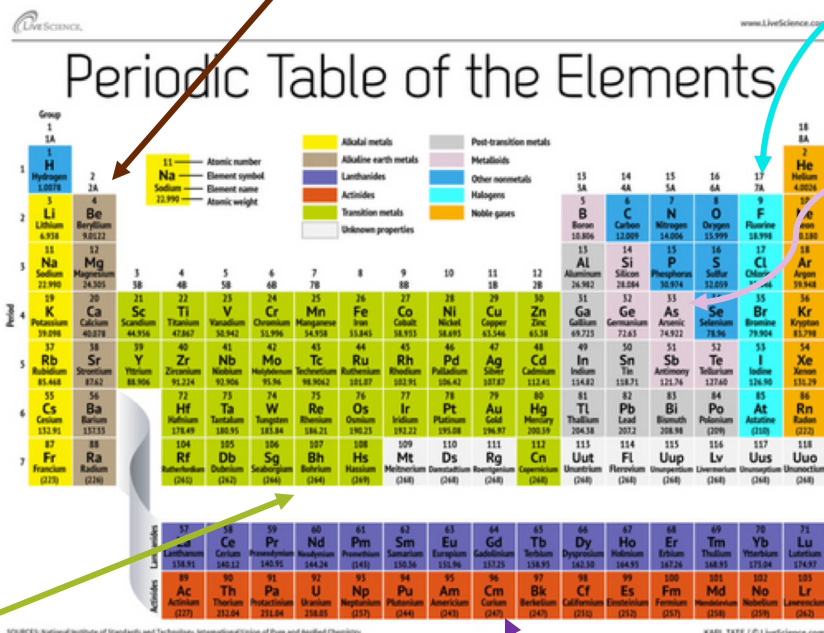
- Group 1A
- +1 charge
- Highly reactive

Alkali Earth Metals

- Group 2A
- +2 charge
- Reactive

Halogens

- Group 7A
- -1 charge
- Highly reactive if single atoms
- Diatomic molecules



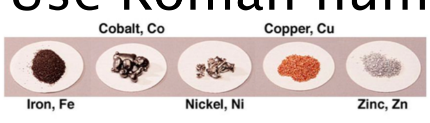
Metalloids

- Some characteristics of metals, some of nonmetals
- semiconductors

Silicon (Si)

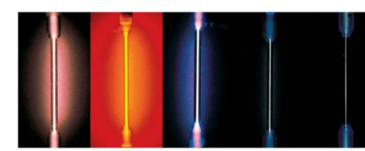
Transition metals

- Center of table
- Varying (+) charge
- Use Roman numerals



Lanthanides & Actinides

- Bottom of table
- Very reactive
- + charge
- Often radioactive



Noble Gases

- Group 8A
- + charge (if charged)
- Inert (least reactive)

Valence and Core Electrons

Valence electrons:

- Highest energy shell (largest principle quantum #, n)
- Furthest from nucleus
- Outermost electrons
- Available for bonding
  - Determine the behavior of the atom

TABLE 8.1

Electron Configurations of Group 1A and Group 2A Elements

Group 1A		Group 2A	
Li	$[He]2s^1$	Be	$[He]2s^2$
Na	$[Ne]3s^1$	Mg	$[Ne]3s^2$
K	$[Ar]4s^1$	Ca	$[Ar]4s^2$
Rb	$[Kr]5s^1$	Sr	$[Kr]5s^2$
Cs	$[Xe]6s^1$	Ba	$[Xe]6s^2$
Fr	$[Rn]7s^1$	Ra	$[Rn]7s^2$

Core electrons

- Located on the inside in inner shells.
- Principal quantum number is lower

Example

Oxygen, O Z = 8

valence electrons e- = 6

Core electrons e- = 2

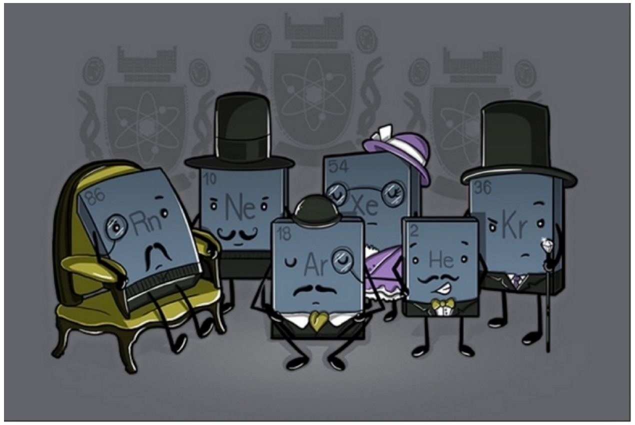
Valence e^-

$1s^2 2s^2 2p^4$

Effect of Valence Electrons on Elements

Octet Rule:

- Elements most stable with 8 valence electrons (2s + 6p)
- Noble gases have 8 valence electrons
  - No e^- want to be added or removed
  - Why they are so unreactive
- Main group elements form ions to become isoelectronic with the noble gases
  - Same electron configuration
- He & H follow duet rule
  - 2 e^-; too small for 8e^-



Main Group elements – gain or lose s & p e^- to get 8

Transition metals – all form cations – remove e^- from s orbital before d orbital (ie 4s e^- lost before 3d e^-)

Periodic Properties in Main Group Elements

Table of Pauling Electronegativity Values

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA
1	1 H 2.1																	2 He
2	3 Li 1.0	4 Be 1.5											5 B 2.0	6 C 2.5	7 N 3.0	8 O 3.5	9 F 4.0	10 Ne
3	11 Na 0.9	12 Mg 1.2											13 Al 1.5	14 Si 1.8	15 P 2.1	16 S 2.5	17 Cl 3.0	18 Ar
4	19 K 0.8	20 Ca 1.0	21 Sc 1.3	22 Ti 1.5	23 V 1.6	24 Cr 1.6	25 Mn 1.5	26 Fe 1.8	27 Co 1.8	28 Ni 1.8	29 Cu 1.9	30 Zn 1.6	31 Ga 1.6	32 Ge 1.8	33 As 2.0	34 Se 2.4	35 Br 2.8	36 Kr
5	37 Rb 0.8	38 Sr 1.0	39 Y 1.2	40 Zr 1.4	41 Nb 1.6	42 Mo 1.8	43 Tc 1.9	44 Ru 2.2	45 Rh 2.2	46 Pd 2.2	47 Ag 1.9	48 Cd 1.8	49 In 1.8	50 Sn 1.8	51 Sb 1.9	52 Te 2.1	53 I 2.5	54 Xe
6	55 Cs 0.7	56 Ba 0.9	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl 1.8	82 Pb 1.9	83 Bi 1.9	84 Po 2.0	85 At 2.2	86 Rn
7	87 Fr 0.7	88 Ra 0.9	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110	111	112		114		116		

Lanthanides	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	71 Lu
Actinides	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	103 Lr

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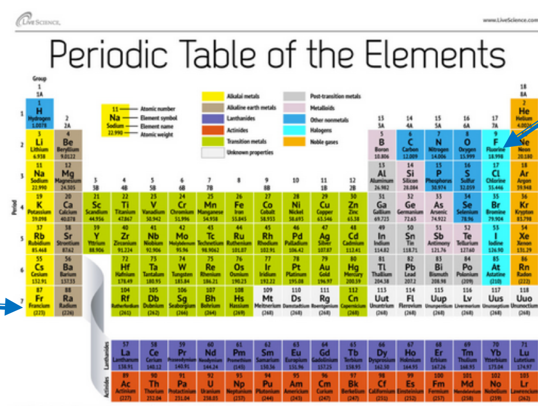
Trends in the Periodic Table

Effective Nuclear Charge (Z_{eff})

- The attractive force felt by an electron in an atom
- Takes into account two things:
  - The actual nuclear charge (Z)
  - The repulsive effects of the other electrons (referred to as shielding effects)
    - Most shielding is due to core electrons
- Depends on size of nucleus & energy level



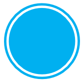
Fluorine: Valence electrons experience high Z_{eff}

Francium: Valence electrons experience low Z_{eff}

Trends in the Periodic Table

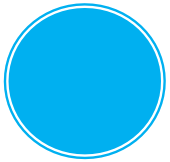
When looking at trends, consider 3 things:

- Amount of positive charge in nucleus (Z)
- Distance of the electron from the nucleus (Energy level)
- Number of other electrons between the electron in question and the nucleus (Shielding)



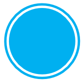
Lithium
- 3 protons
- 2^{nd} energy level
- 2 core electrons

Fluorine
- 9 protons
- 2^{nd} energy level
- 2 core electrons



Rubidium
- 37 protons
- 5^{th} energy level
- 36 core electrons

Iodine
- 53 protons
- 5^{th} energy level
- 46 core electrons

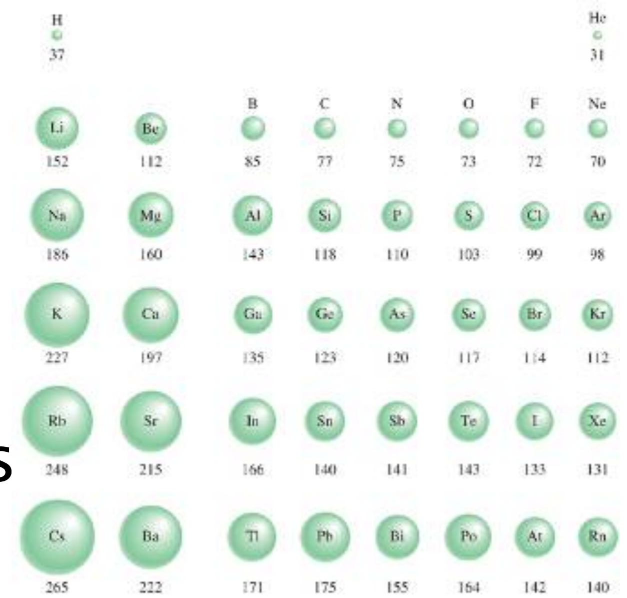


Sizes not exactly to scale

Atomic Radius

Atomic radius increases from top to bottom in a group/column

- Electrons are shielded from nucleus
- Previous shells blocks attraction
- Effective nuclear charge decreases
- Large size difference between shells



Atomic radius decreases from left to right across a row/period

- Little shielding as all electrons in same shell
- Effective nuclear charge higher as protons added
- Electrons pulled closer to nucleus

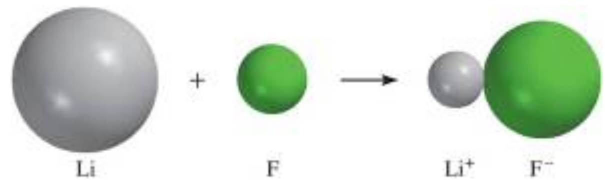
Ionic Radius

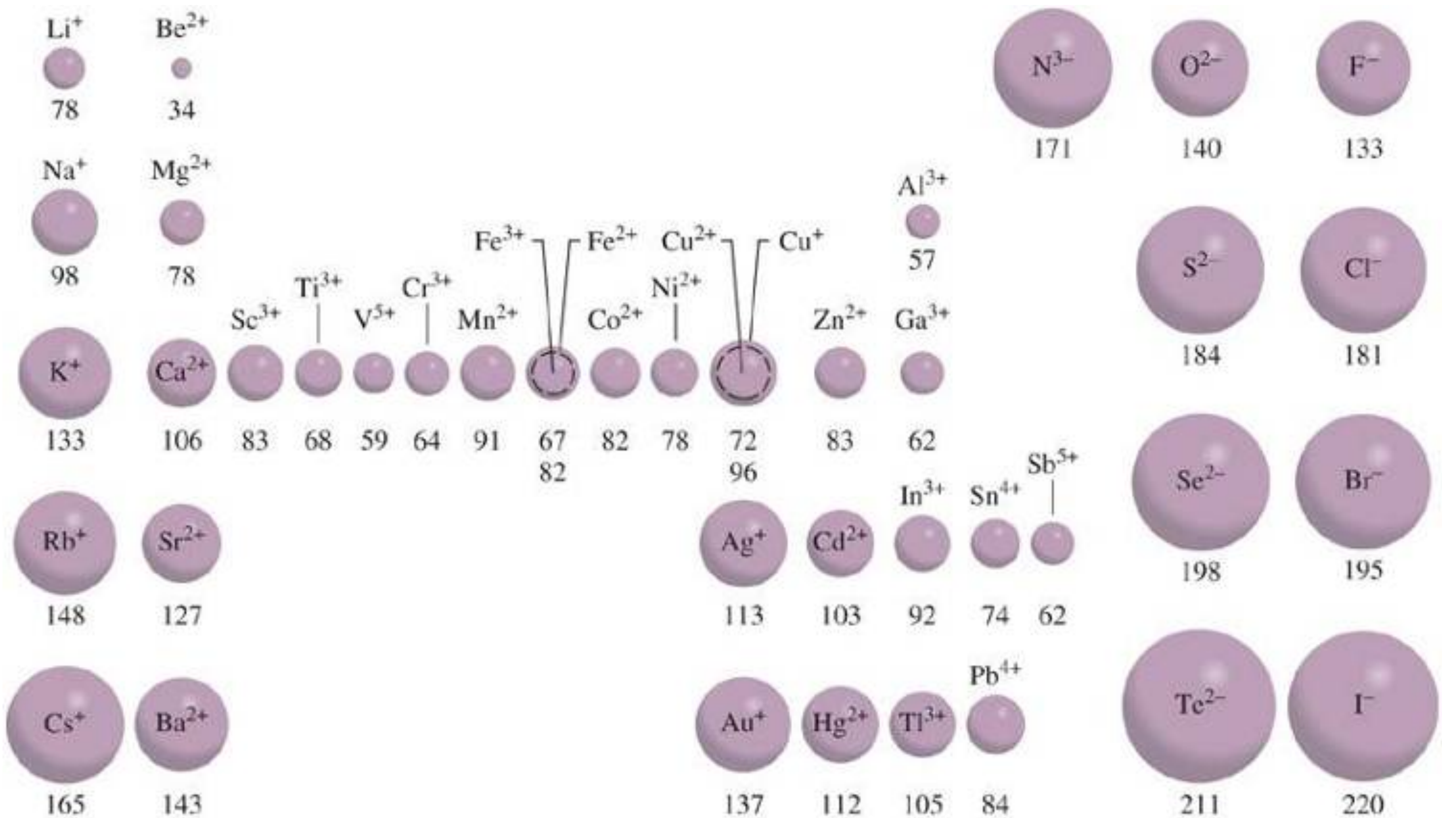
Anions larger than atoms

- Low effective nuclear charge
- More electrons
- More repulsion

Cations smaller than atoms

- High effective nuclear charge
- Fewer electrons
- Less repulsion





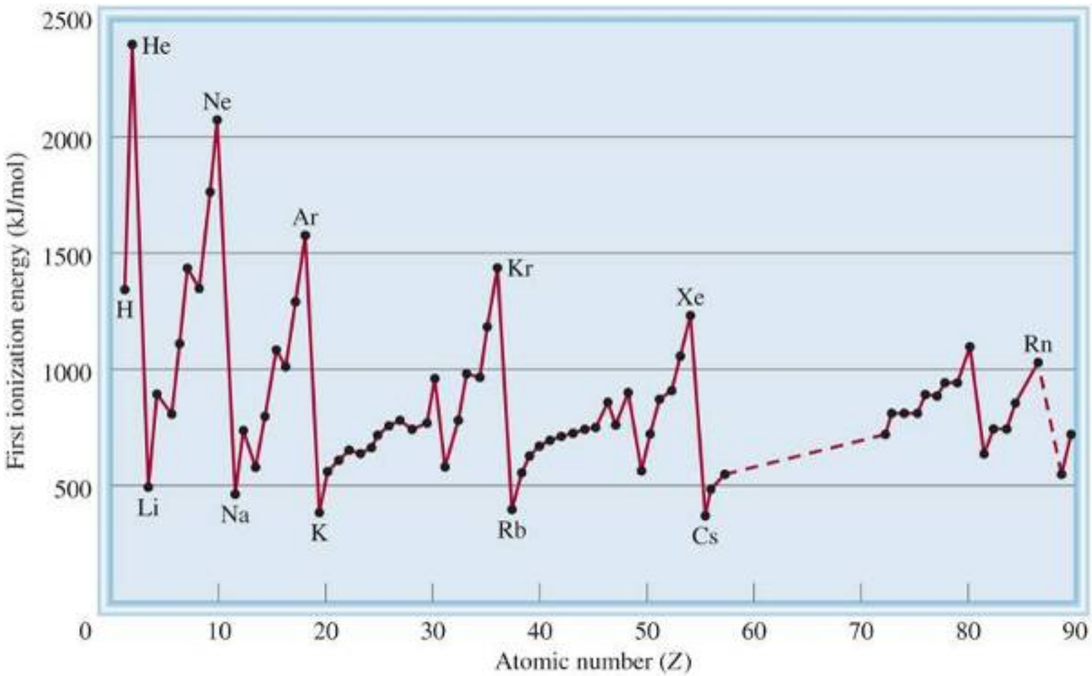
Ionization Energy

Energy needed to remove an e- from a gaseous atom or ion

$X(g) \rightarrow X^{+}(g)$ + 1e- **Endothermic Process**

Decreases top to bottom: Bigger atom = more shielding

Increases from left to right: Atoms want to gain electrons



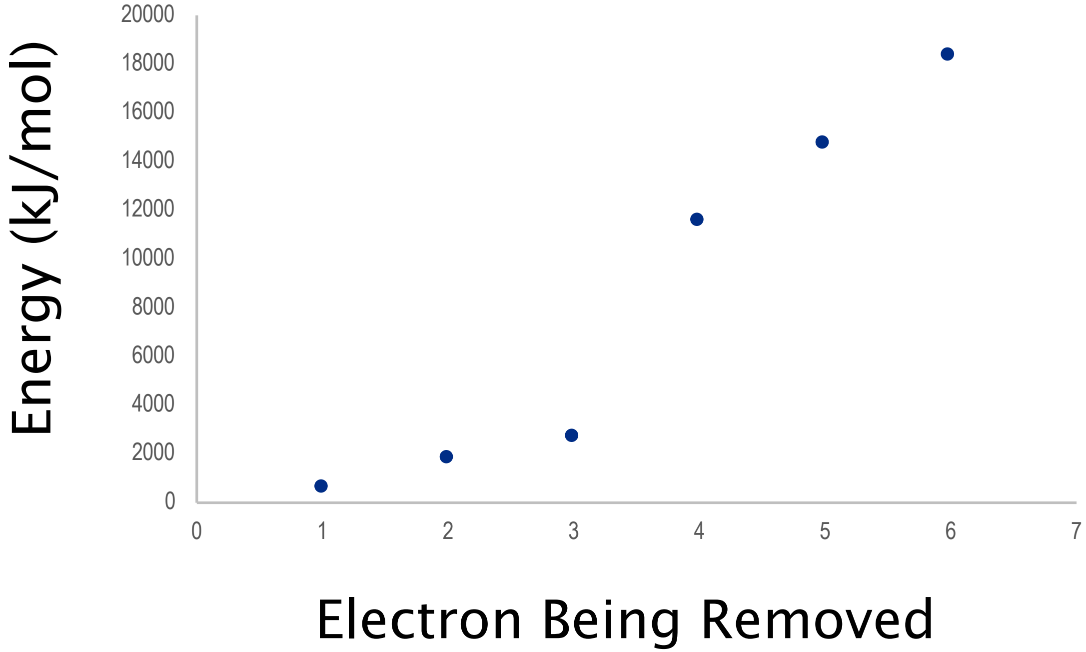
Ionization Energy con't

3rd ionization energy > 2nd > 1st:

- takes less energy to remove the first electron
- 2nd , 3rd, 4th, etc. electrons are held more strongly

Very large jump once all valence e^- have been removed

Ionization Energies of Aluminum



Elemental Ionization Energies

TABLE 8.2 The Ionization Energies (kJ/mol) of the First 20 Elements

Z	Element	First	Second	Third	Fourth	Fifth	Sixth
1	H	1,312					
2	He	2,373	5,251				
3	Li	520	7,300	11,815			
4	Be	899	1,757	14,850	21,005		
5	B	801	2,430	3,660	25,000	32,820	
6	C	1,086	2,350	4,620	6,220	38,000	47,261
7	N	1,400	2,860	4,580	7,500	9,400	53,000
8	O	1,314	3,390	5,300	7,470	11,000	13,000
9	F	1,680	3,370	6,050	8,400	11,000	15,200
10	Ne	2,080	3,950	6,120	9,370	12,200	15,000
11	Na	495.9	4,560	6,900	9,540	13,400	16,600
12	Mg	738.1	1,450	7,730	10,500	13,600	18,000
13	Al	577.9	1,820	2,750	11,600	14,800	18,400
14	Si	786.3	1,580	3,230	4,360	16,000	20,000
15	P	1,012	1,904	2,910	4,960	6,240	21,000
16	S	999.5	2,250	3,360	4,660	6,990	8,500
17	Cl	1,251	2,297	3,820	5,160	6,540	9,300
18	Ar	1,521	2,666	3,900	5,770	7,240	8,800
19	K	418.7	3,052	4,410	5,900	8,000	9,600
20	Ca	589.5	1,145	4,900	6,500	8,100	11,000

Electron Affinity

Energy released when an e^- is added to a gaseous atom

$X(g) + 1e^- \rightarrow X^-(g)$

Exothermic Process

- Decreases top to bottom
- Increases left to right
- Fluorine at top right
  - small atom
  - limited shielding
  - nucleus relatively large compared to overall size

TABLE 8.3 Electron Affinities (kJ/mol) of Some Representative Elements and the Noble Gases*

1A	2A	3A	4A	5A	6A	7A	8A
H							He
73							< 0
Li	Be	B	C	N	O	F	Ne
60	≤ 0	27	122	0	141	328	< 0
Na	Mg	Al	Si	P	S	Cl	Ar
53	≤ 0	44	134	72	200	349	< 0
K	Ca	Ga	Ge	As	Se	Br	Kr
48	2.4	29	118	77	195	325	< 0
Rb	Sr	In	Sn	Sb	Te	I	Xe
47	4.7	29	121	101	190	295	< 0
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
45	14	30	110	110	?	?	< 0

2nd electron affinities lower: Ion is already negative – doesn't want to add more negative charges

Electronegativity: measure of attraction for e^- in a chemical bond – follows similar trend; F has greatest electronegativity

Trends in the Periodic Table

1.) Which has the highest ionization energy: nitrogen, phosphorus, arsenic, or antimony?

2.) Which atom is smaller, potassium, calcium, iron, or arsenic?

3.) Which is the largest ion, K^+, Ca^{2+}, Se^{2-}, Br^-?

4.) Which has the highest electronegativity, fluorine, chlorine, bromine, or iodine?