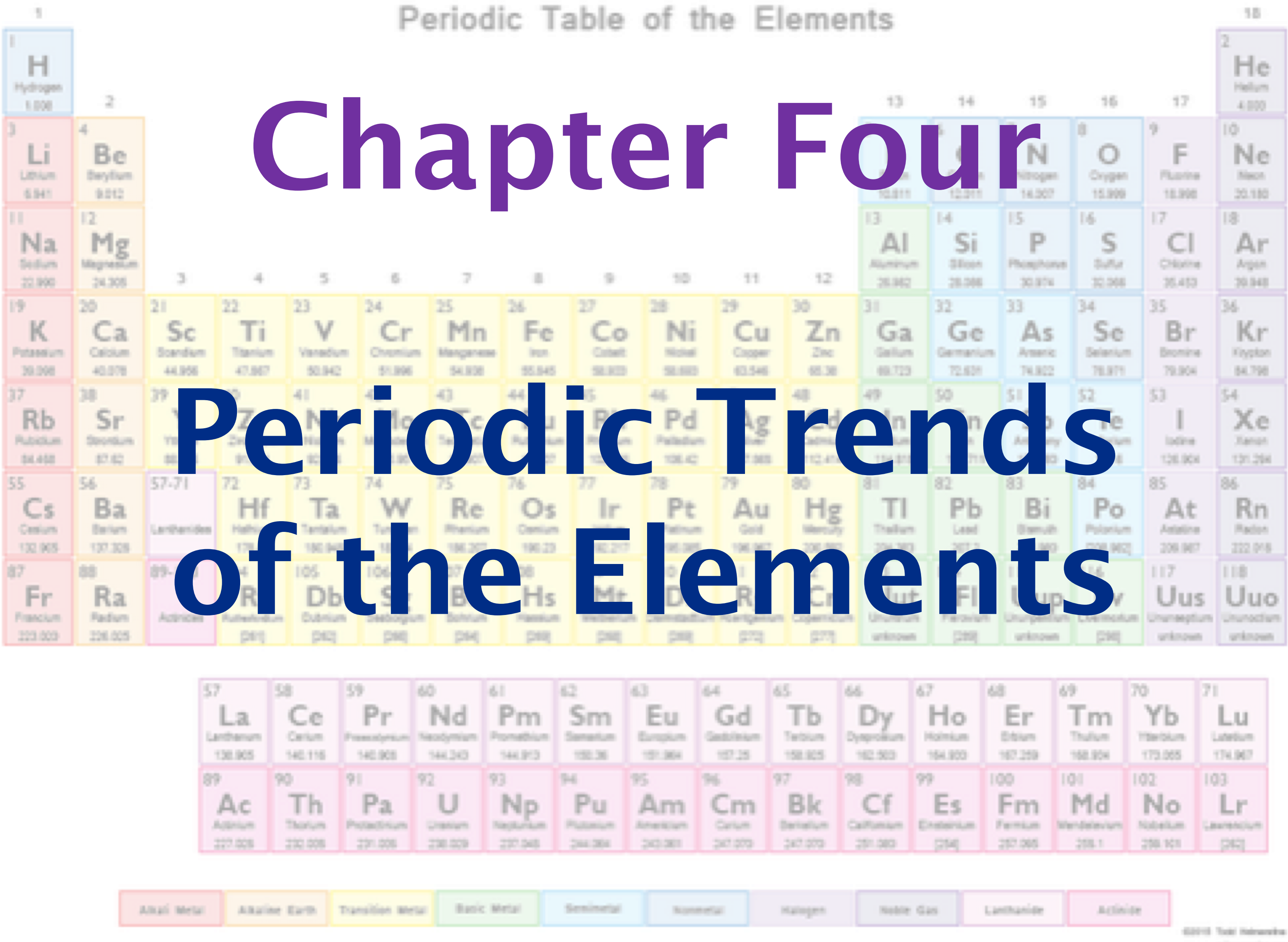
Chapter Four

Periodic Trends of the Elements



Mendeleev’s Periodic Table 1869

Known elements arranged in order of increasing **atomic mass** from left to right and from top to bottom in groups.

Elements with similar properties are placed in same column.

	H 1						
He 4	Li 7	Be 9	B 11	C 12	N 14	O 16	F 19
Ne 20	Na 23	Mg 24	Al 27	Si 28	P 31	S 32	Cl 35
Ar 40	K 39	Ca 40		Ge	As 75	Se 79	Br 80
Kr 84	Rb 85	Sr 88	In 115	Sn 119	Sb 122	Te 128	I 127
Xe 131	Cs 133	Ba 137	Tl 204	Pb 207	Bi 209		
Rn (222)							

Used table to predict properties of undiscovered elements!

Eka-Silicon →	Germanium
MM: 72	MM: 72.6
Density: 5.5g/mL	Density: 5.47g/mL
Color: dirty gray	Color: grayish white

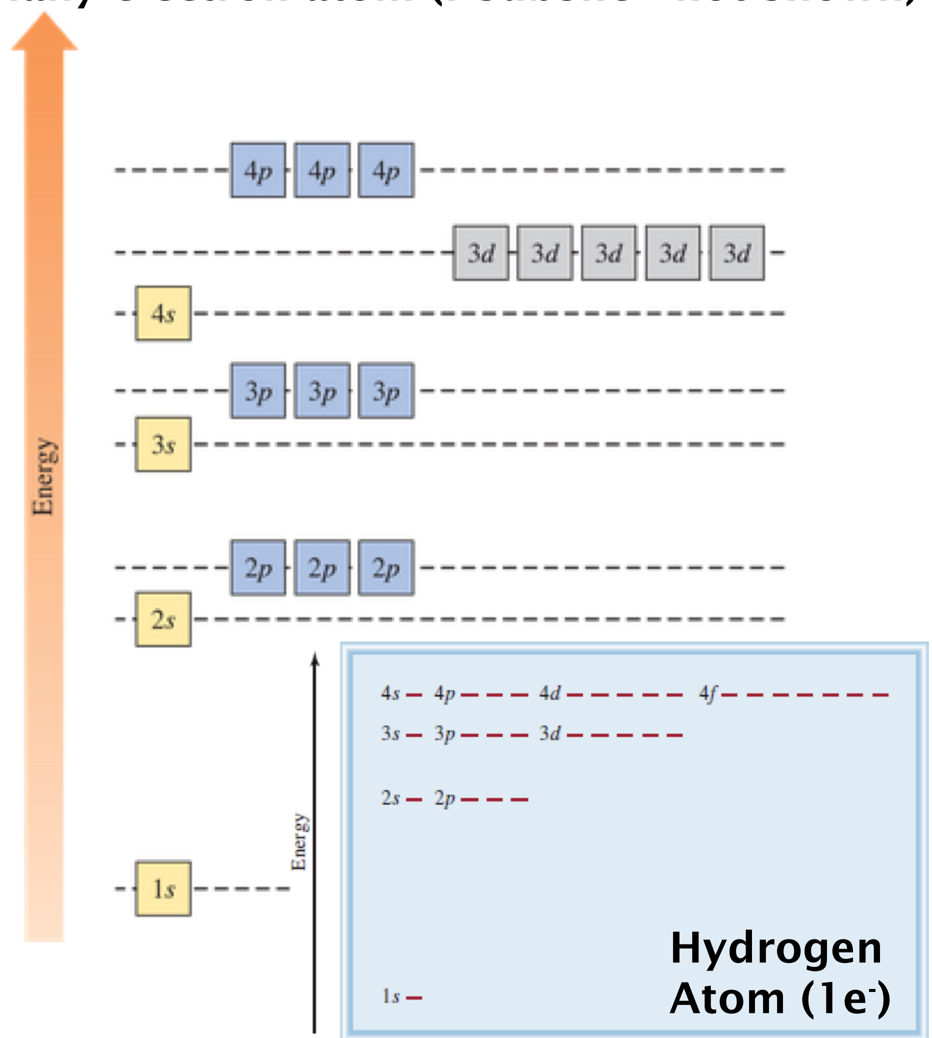
Early 1900’s: Henry Moseley discovered connection between **atomic number** & periodic trends. Today’s Periodic Table is arranged by atomic number.

Electron Configuration:
Finding a home for each electron

The energy of an electron is defined by both n & ℓ

- Principle shells (size)
  - n = 1,2, 3, 4 or 5
- Subshells (shape)
  - ℓ = 0, 1, 2, or 3
  - n determines number of subshells
  - s, p, d, f orbitals
  - <u>Shielding</u> impacts relative energies

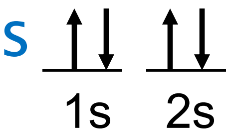
Many-electron atom (f subshell not shown)



Rules & Principles Governing e^- Configurations

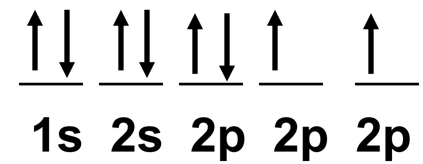
Pauli Exclusion Principle:

- No 2 e^- in an atom can have the same set of 4 quantum #s
  - If in the same orbital, e^- must have opposite spins



Hund's rule:

- Electrons in the same subshell occupy degenerate orbitals singly, before pairing
  - Degenerate = same energy



Ex: Oxygen, O   Z = 8

The Aufbau Principle:

- In general, each successive electron added to an atom occupies the lowest energy orbital available
  - There are some exceptions

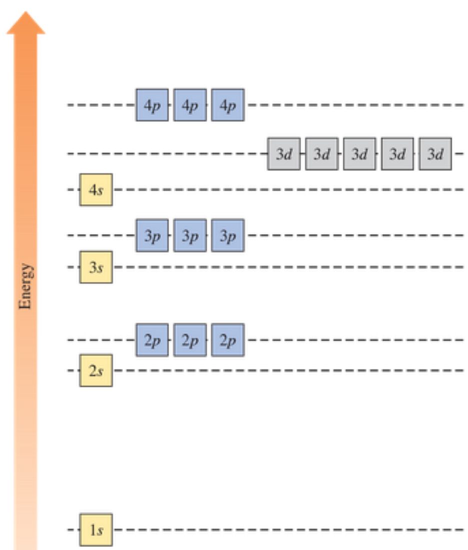
(Z = 1) H $1s^1$
(Z = 2) He $1s^2$
(Z = 3) Li $1s^2 2s^1$

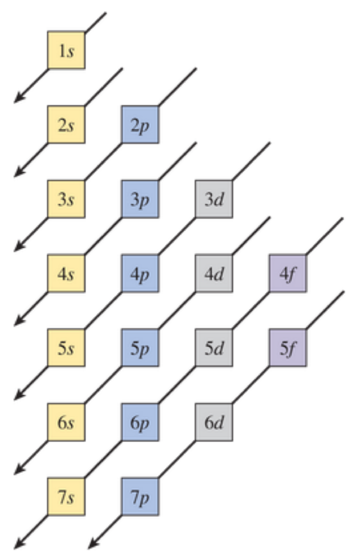
Orbital Filling in Multi-electron Atoms

- Fill low to high energy
- 2 electrons per orbital
- Shielding impacts the energy of orbitals
- Use chart to account for overlap of n values

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

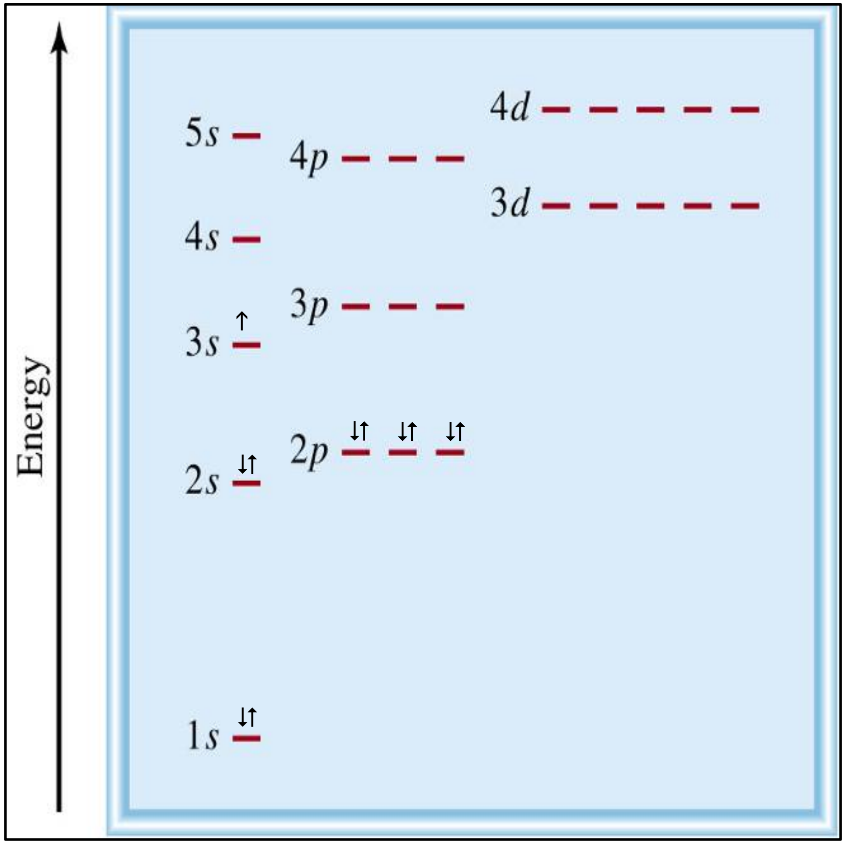
- Format: spdf or orbital notation
- Ends when a home is found for each electron





Electron Configuration con't

- Defines the orbital ("home") for each electron
- # electrons = atomic number (Z) of atom (if neutral)
- Max 2 electrons per orbital



- Energy increases from bottom to top
  - Higher energy levels at top
- Boxes or lines represent orbitals
  - # lines at one level = # degenerate orbitals
- Arrows (⇂↾) represent e^-
- 2 e^- allowed per orbital
  - one arrow up & one down to show the different spins

Generally do not want to draw a box, so use other formats

Formats for Electron Configurations

Electron Configuration notation

- Front number = energy level
- Letter = type of orbital (s, p, d, or f)
  - Degenerate orbitals are combined together
- Superscript = # electrons in that type of orbital
  - Degenerate orbitals are combined, so the superscript can be more than 2 if it is a p, d, or f orbital (p max 6, d max 10, f max 14)

Examples:

Ne: Z = 10

$1s^2 2s^2 2p^6$

Na: Z = 11

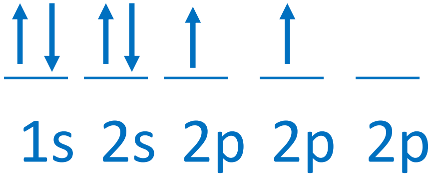
$1s^2 2s^2 2p^6 3s^1$

Orbital Notation (diagram)

- Number = energy level
- Letter = type of orbital
  - Degenerate orbitals are NOT combined
- Arrows = electrons
  - Put one e^- in each degenerate orbital before pairing
  - If 2 e^- in one orbital, one arrow must be up, the other down

Example:

C: Z = 6



Writing Electron Configurations

Sulfur:

Vanadium:

Noble Gas Configuration

- Abbreviation of Electron Configuration (ex: Na: [Ne]$3s^1$)
- Noble gas symbol replaces the portion of the e^- config. that is identical to the e^- config. of the noble gas.
- Always use the largest noble gas that is smaller than the element
- Can use for either spdf or orbital notation
- Will always start with an s orbital – just not 1s.
- ex: Arsenic

  regular configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

- ex: Strontium

Exceptions To The Aufbau Principle

Half filled & filled subshells provide additional stability

Cr and Cu ½ fill/fill their 3d shell before the 4s shell.

Elements in same columns as Cr & Cu behave in same way.

Similar behavior seen in p block.

Cr & Cu are the only exceptions you need to know.

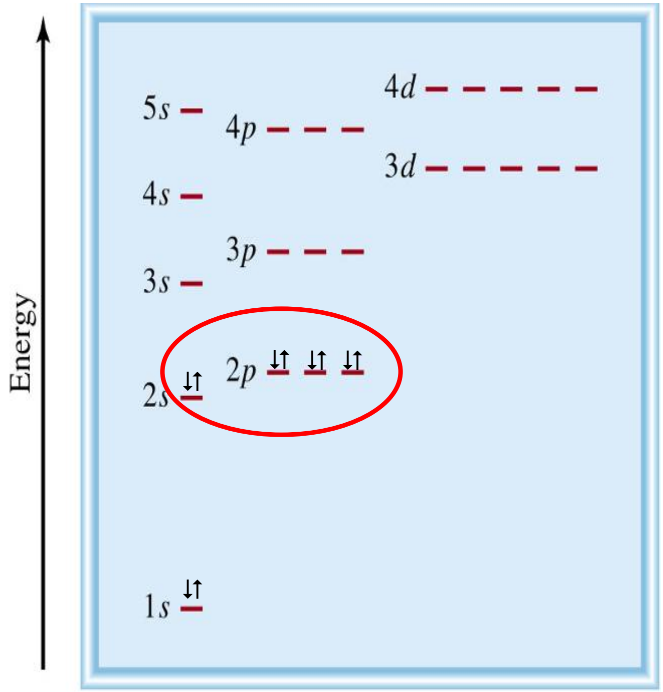
Cr: 3d ½ full & 4s ½ full

Cu: 3d full & 4s ½ full

		3*d*					4*s*	
Sc	[Ar]	↑					↑↓	$[Ar]3d^{1}4s^{2}$
Ti	[Ar]	↑	↑				↑↓	$[Ar]3d^{2}4s^{2}$
V	[Ar]	↑	↑	↑			↑↓	$[Ar]3d^{3}4s^{2}$
Cr	[Ar]	↑	↑	↑	↑	↑	↑	$[Ar]3d^{5}4s^{1}$
Mn	[Ar]	↑	↑	↑	↑	↑	↑↓	$[Ar]3d^{5}4s^{2}$
Fe	[Ar]	↑↓	↑	↑	↑	↑	↑↓	$[Ar]3d^{6}4s^{2}$
Co	[Ar]	↑↓	↑↓	↑	↑	↑	↑↓	$[Ar]3d^{7}4s^{2}$
Ni	[Ar]	↑↓	↑↓	↑↓	↑	↑	↑↓	$[Ar]3d^{8}4s^{2}$
Cu	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑	$[Ar]3d^{10}4s^{1}$
Zn	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	$[Ar]3d^{10}4s^{2}$

Magnetism in Multi-electron Atoms

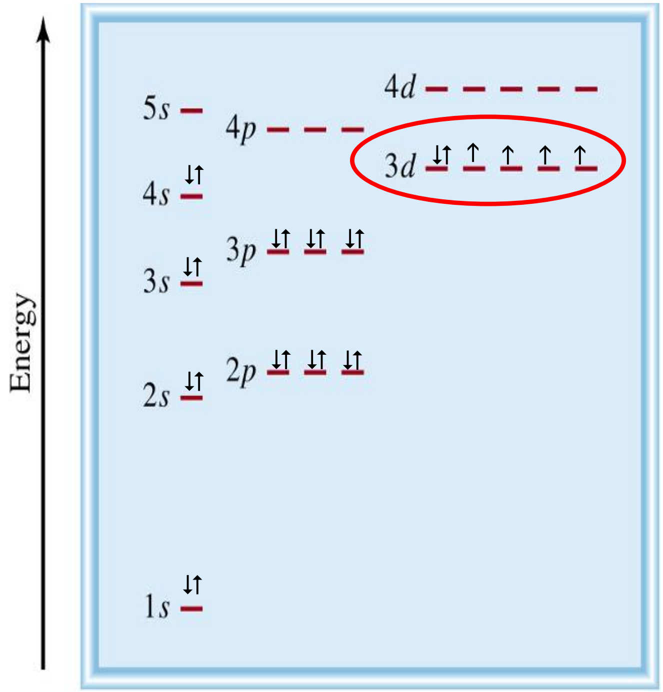
- +1/2 & -1/2 spins will cancel if electrons paired
- No magnetic properties without spin present
  - \# unpaired electrons proportional to magnetic properties



Diamagnetic

All electrons paired

Ne: $1s^2 2s^2 2p^6$



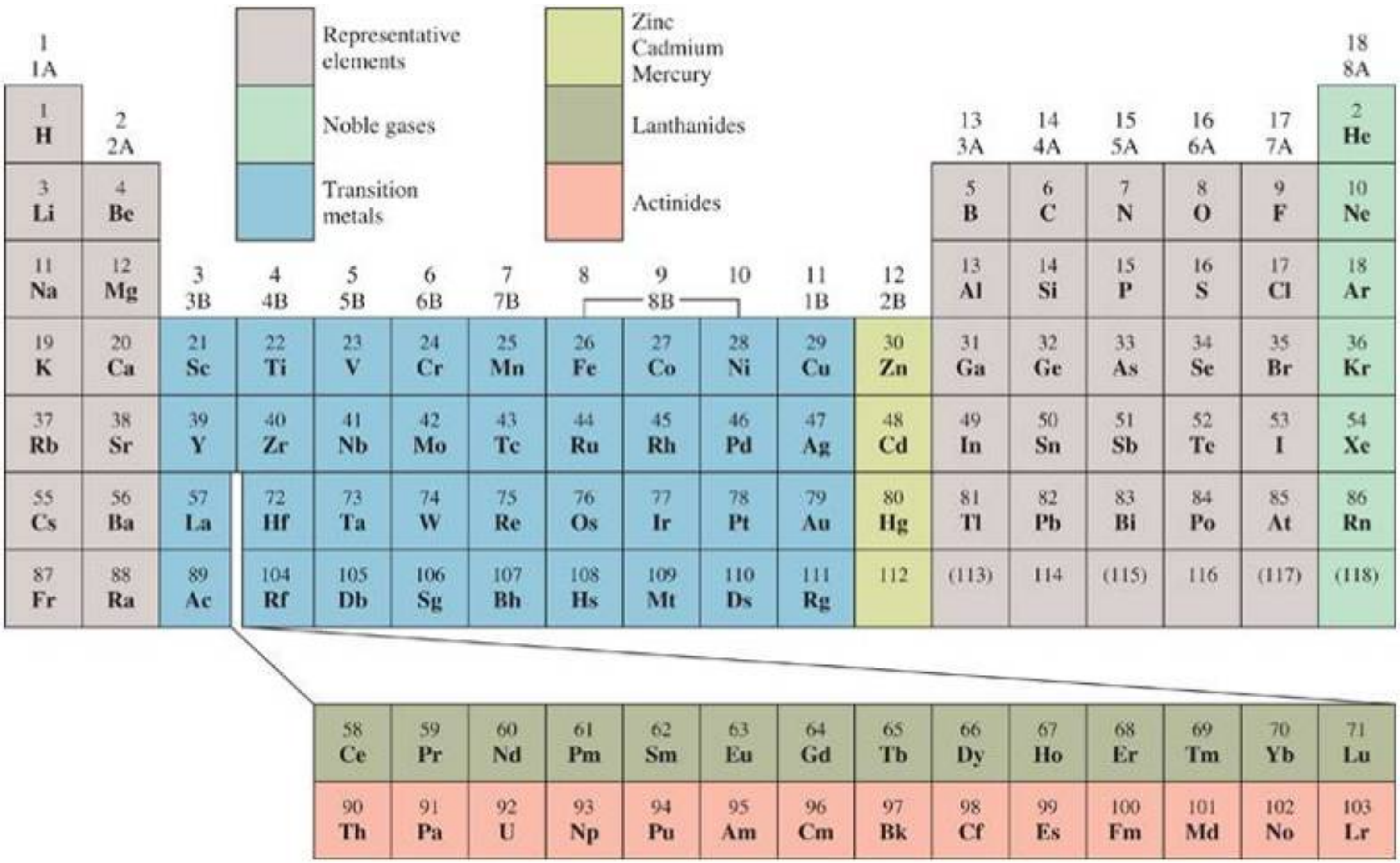
Paramagnetic

At least 1 unpaired electron

Fe: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

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

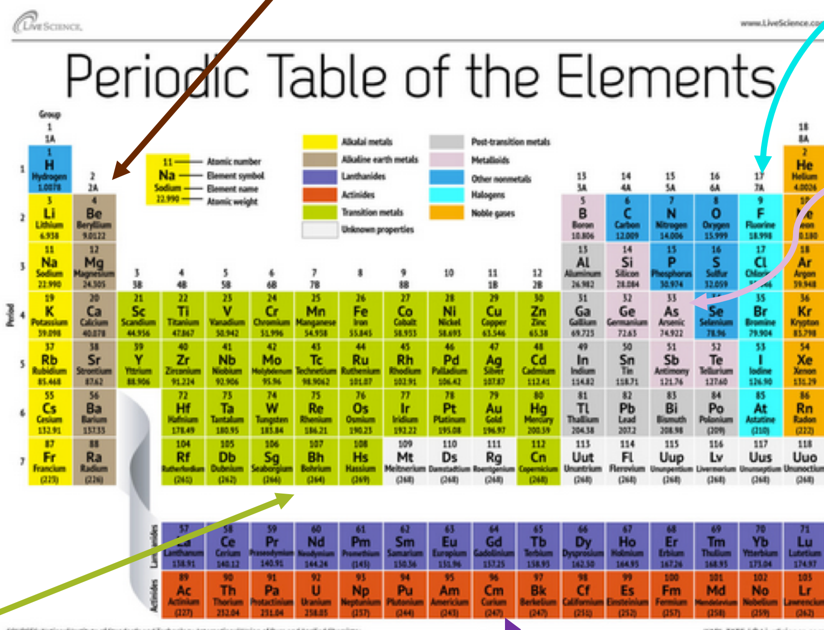
- Group 1A
- +1 charge
- Highly reactive

Alkali Earth Metals

- Group 2A
- +2 charge
- Reactive

Halogens

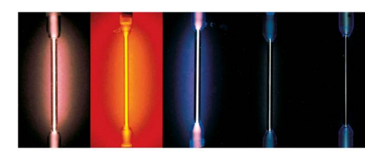
- Group 7A
- −1 charge
- Highly reactive if single atoms
- Diatomic molecules



Metalloids

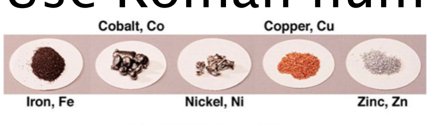
- Some characteristics of metals, some of nonmetals
- semiconductors

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

Valence e^-

Oxygen, O — $Z = 8$

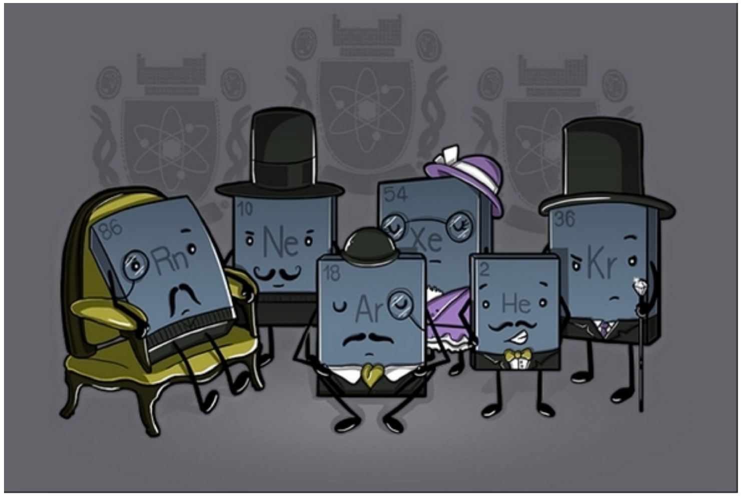
valence electrons — e- = 6

Core electrons — e- = 2 — $1s^2 2s^2 2p^4$

Effect of Valence Electrons on Elements

Octet Rule:

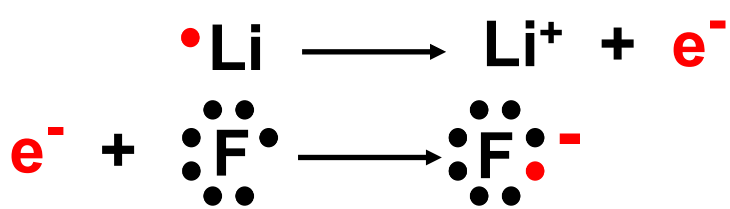
- Elements most stable with 8 valence electrons (2s + 6p)
- Noble gases have 8 valence electrons
  - No e^- want to be added or removed
  - Why they are so unreactive
- Main group elements form ions to become <u>isoelectronic</u> with the noble gases
  - Same electron configuration
- He & H follow duet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



Ionization

Atoms gain or lose electrons to obtain an octet

- Electrons are negatively charged
- Gaining electrons increases negative particles in atom
  - More electrons than protons
  - Forms negative ions
  - Called anions
- Losing electrons decreases negative particles in atom
  - More protons than electrons
  - Forms positive ions
  - Called cations
- **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^-)



$$1s^2 2s^1 \rightarrow 1s^2$$

$$1s^2 2s^2 2p^5 \rightarrow 1s^2 2s^2 2p^6$$

Writing Electron Configurations for Ions

Remove Electrons from Highest Energy Level First

Magnesium ion (Mg^{2+}):

Magnesium atom: $1s^2 2s^2 2p^6 3s^2$

Magnesium ion:

Fluorine ion (F^-):

Fluorine atom: $1s^2 2s^2 2p^5$

Fluorine ion:

Manganese (II) ion (Mn^{2+}):

Manganese atom: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$

Manganese ion:

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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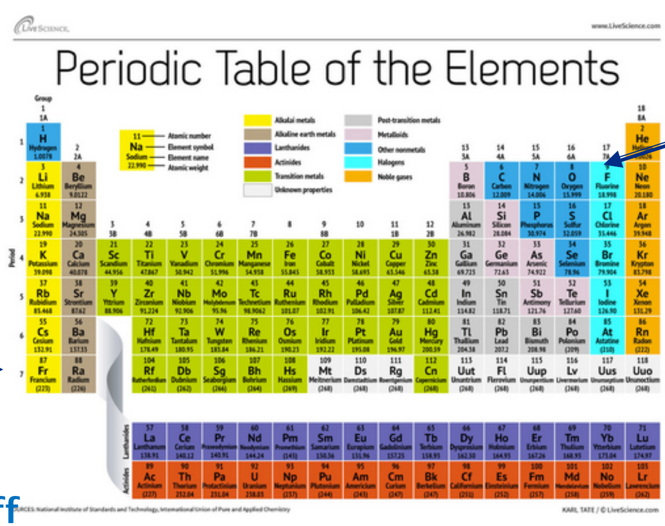
Trends in the Periodic Table

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}

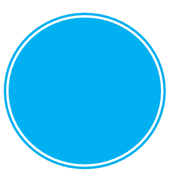
Trends in the Periodic Table

When looking at trends, consider 3 things:

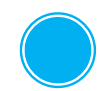
- Amount of positive charge in nucleus (Z)
- Distance of the electron from the nucleus (Energy level)
- Number of other electrons between the electron in question and the nucleus (Shielding)

Lithium
- 3 protons
- 2nd energy level
- 2 core electrons

Fluorine
- 9 protons
- 2nd energy level
- 2 core electrons



Rubidium
- 37 protons
- 5th energy level
- 36 core electrons



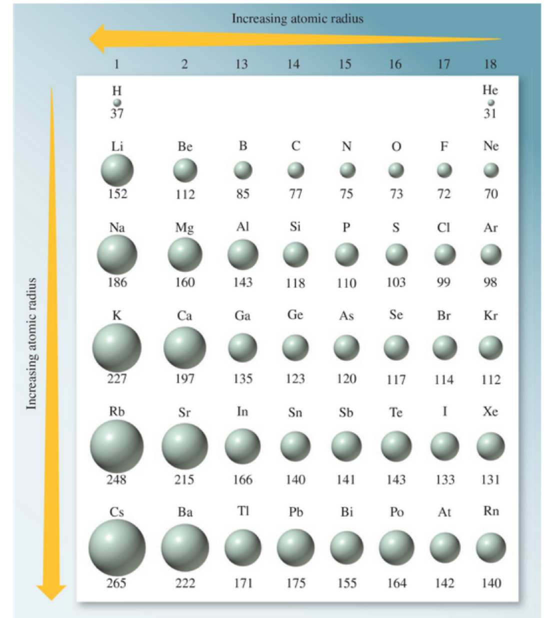
Iodine
- 53 protons
- 5th energy level
- 46 core electrons

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

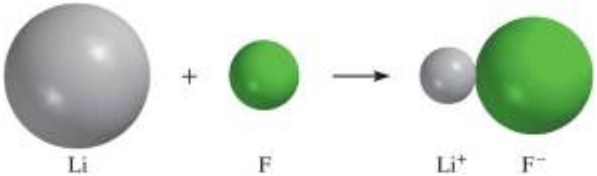
Ionic Radius

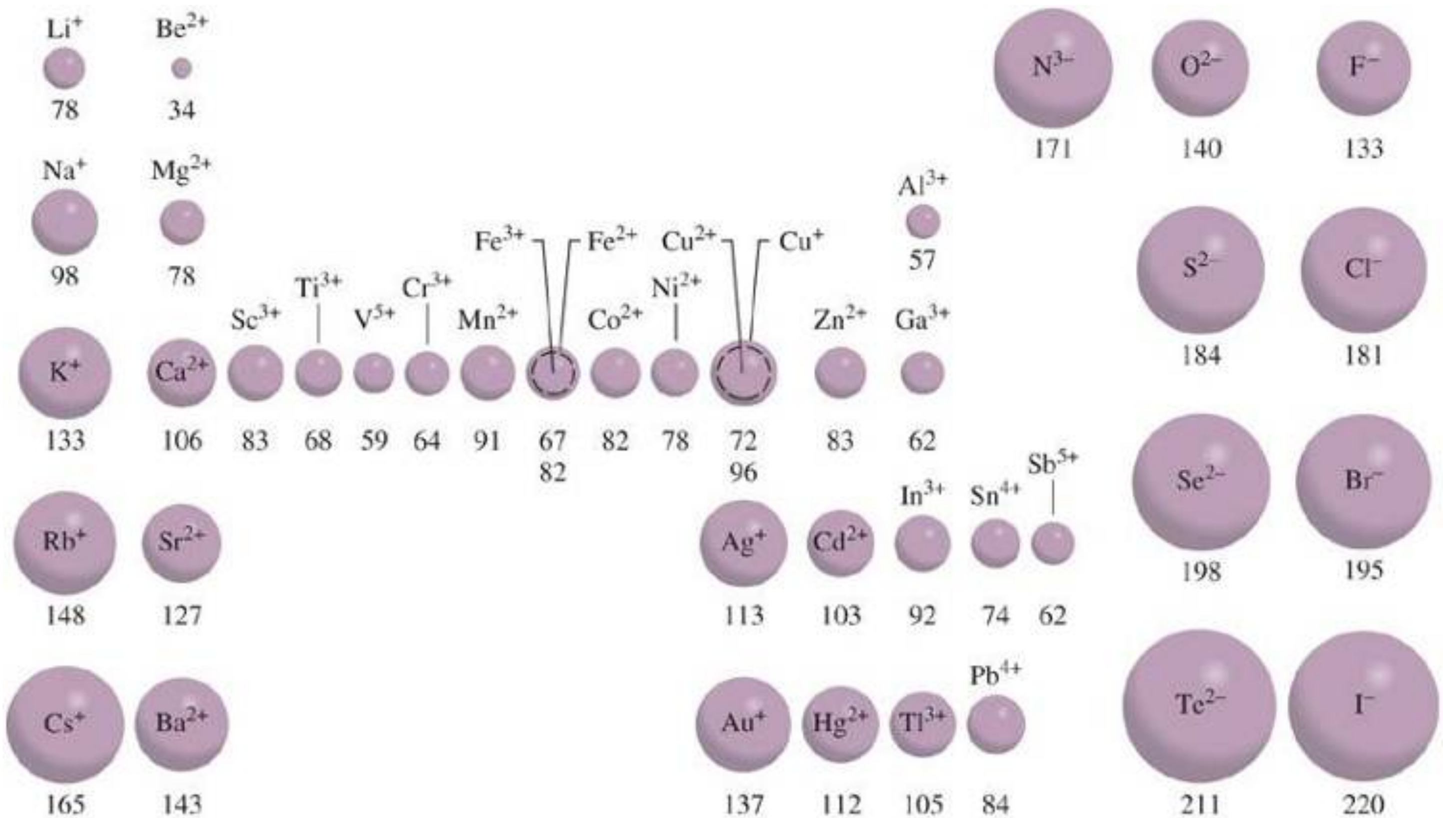
Anions larger than atoms

- Low effective nuclear charge
- More electrons
- More repulsion

Cations smaller than atoms

- High effective nuclear charge
- Fewer electrons
- Less repulsion



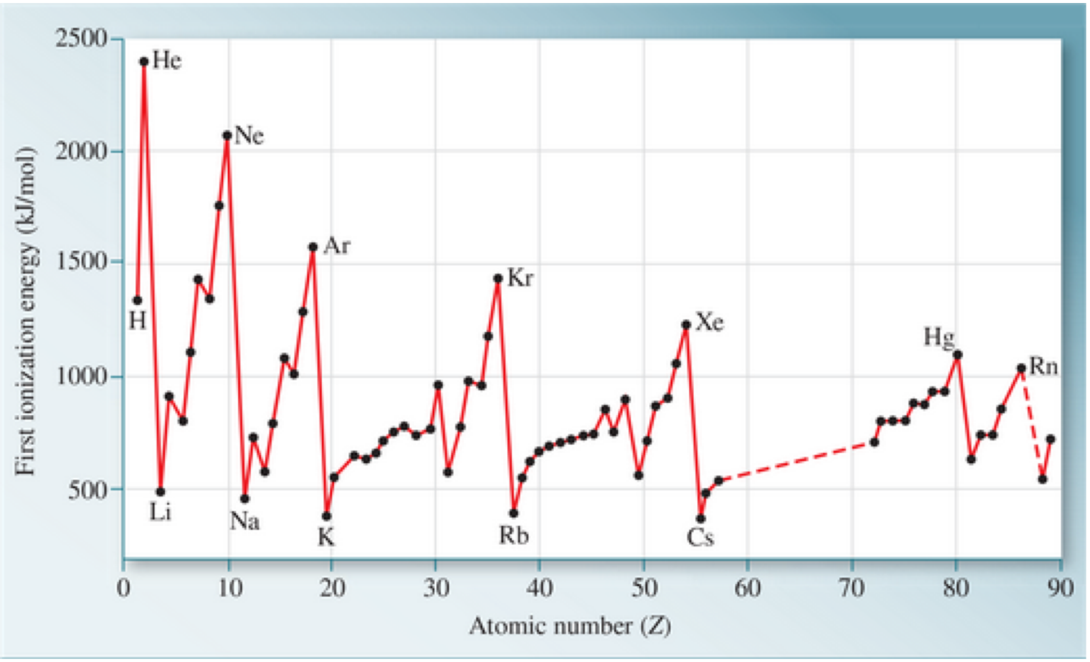


Ionization Energy

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

X(g)→ X⁺(g) + 1e- Endothermic Process (requires energy)

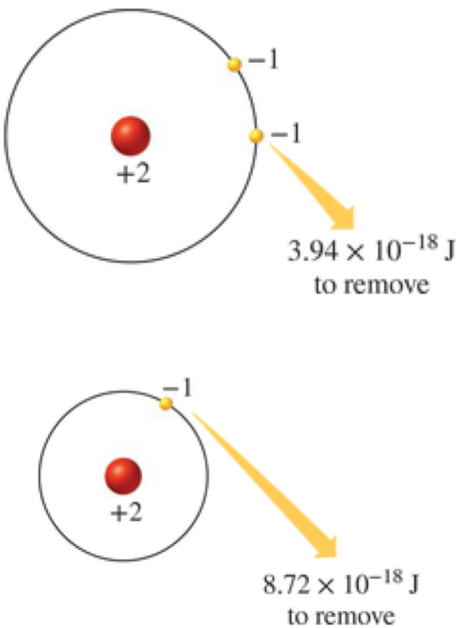
Decreases top to bottom: Bigger atom = more shielding
Increases from left to right: Atoms want to gain electrons



Ionization Energy con't

3[rd] ionization energy > 2[nd] > 1[st]

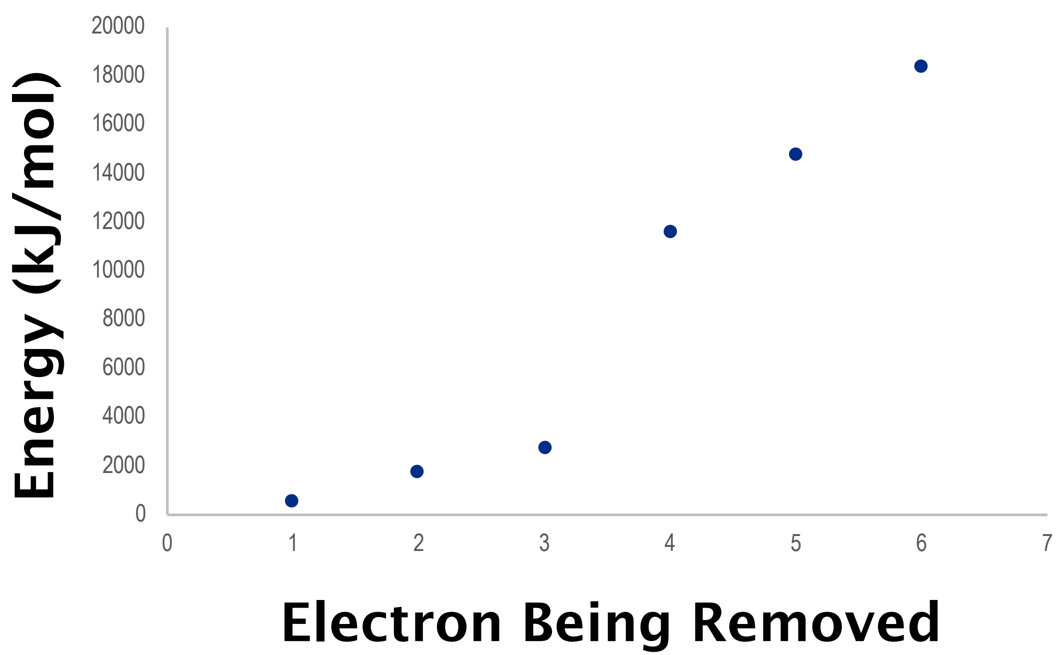
- takes less energy to remove the first electron
- 2[nd] , 3[rd], 4[th], etc. electrons are held more strongly



Helium: removing 2[nd] e^- requires more than 2X as much energy

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

Ionization Energies of Aluminum



Elemental Ionization Energies

TABLE 4.2 **Ionization Energies (in kJ/mol) for Elements 3 through 11***

	Z	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
Li	3	520	7,298	11,815							
Be	4	899	1,757	14,848	21,007						
B	5	800	2,427	3,660	25,026	32,827					
C	6	1,086	2,353	4,621	6,223	37,831	47,277				
N	7	1,402	2,856	4,578	7,475	9,445	53,267	64,360			
O	8	1,314	3,388	5,301	7,469	10,990	13,327	71,330	84,078		
F	9	1,681	3,374	6,050	8,408	11,023	15,164	17,868	92,038	106,434	
Ne	10	2,080	3,952	6,122	9,371	12,177	15,238	19,999	23,069	115,380	131,432
Na	11	496	4,562	6,910	9,543	13,354	16,613	20,117	25,496	28,932	141,362

Electron Affinity

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

$X(g) + 1e^- \rightarrow X^-(g)$ Exothermic Process (releases energy)

- Decreases top to bottom
- Increases left to right
    - Except column 18
- Fluorine at top right
    - small atom
    - limited shielding
    - nucleus relatively large compared to overall size

1	2	13	14	15	16	17	18
H +72.8							He (0.0)
Li +59.6	Be ≤0	B +26.7	C +122	N −7	O +141	F +328	Ne (−29)
Na +52.9	Mg ≤0	Al +42.5	Si +134	P +72.0	S +200	Cl +349	Ar (−35)
K +48.4	Ca +2.37	Ga +28.9	Ge +119	As +78.2	Se +195	Br +325	Kr (−39)
Rb +46.9	Sr +5.03	In +28.9	Sn +107	Sb +103	Te +190	I +295	Xe (−41)
Cs +45.5	Ba +13.95	Tl +19.3	Pb +35.1	Bi +91.3	Po +183	At +270	Rn (−41)

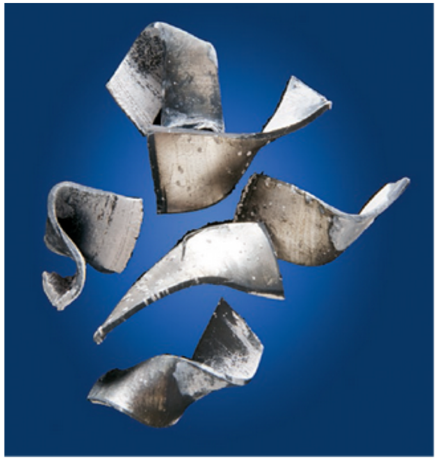
2nd electron affinities lower: Ion is already negative – doesn't want to add more negative charges

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

Metallic Character

Metals tend to be:

- Shiny & lustrous
- Malleable
- Ductile
- Good conductors of heat & electricity
- Low ionization energy
  - form cations



Lithium

Solid Nonmetals tend to be:

- Dull with variable colors
- Brittle
- Poor conductors
- High ionization energy
  - form anions

Metallic Character:

- Increases from top to bottom
- Decreases from left to right

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 lowest electronegativity, fluorine, chlorine, bromine, or iodine?