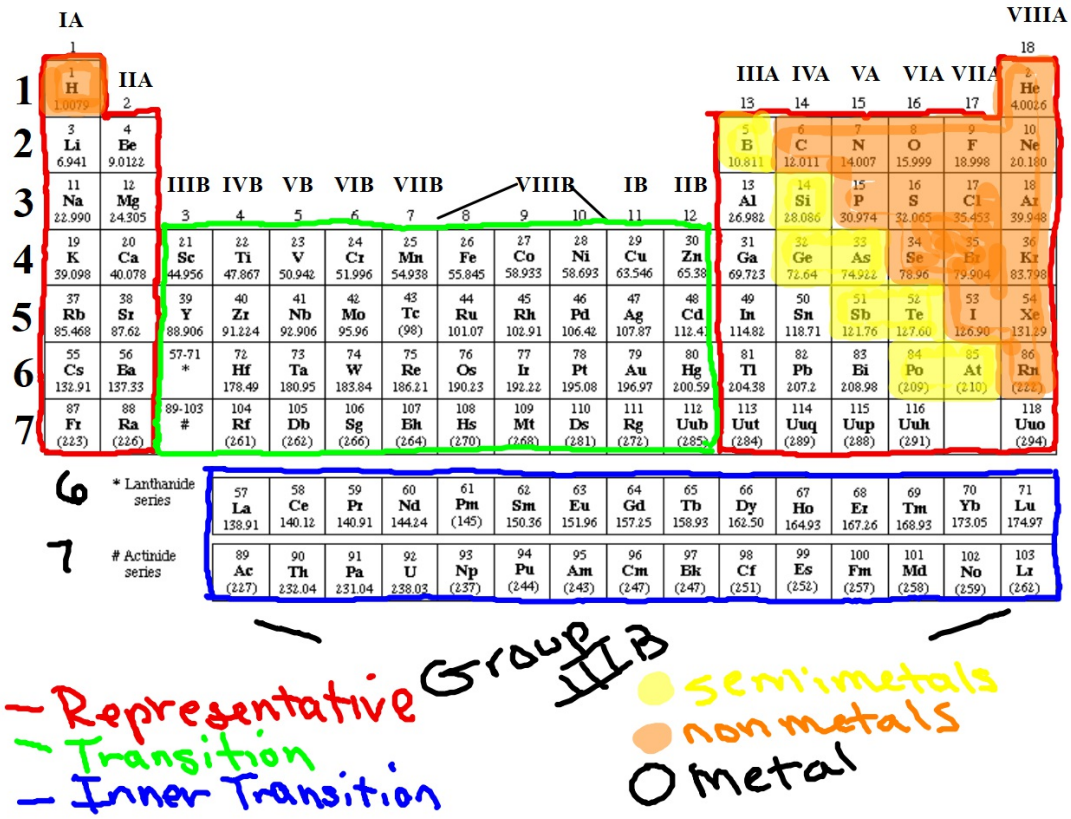
IA
VIIIA
IIA
IIIA IVA VA VIA VIIA
IIIB IVB VB VIB VIIB VIIIB IB IIB
1 H 1.0079
2 He 4.0026
3 Li 6.941
4 Be 9.0122
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6 C 12.011
7 N 14.007
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102 No (259)
103 Lr (262)
Group IIIB
Representative
Transition
Inner Transition
semimetals
nonmetals
Metal

The Periodic Table

Groups- vertical columns of the periodic table (labeled IA to VIIIA and IB to VIIIB)- elements in the same group have similar properties

Periods- Horizontal rows of the periodic table (labeled 1 through 7

Representative elements- Group A elements
- **Alkali metals- Group IA**
- **Alkaline Earth metals- Group IIA**
- **Halogens- Group VIIA**
- **Noble Gases- Group VIIIA**

Transition elements (metals)- Group B elements

Inner transition elements (metals)- Group IIIB, periods 6 and 7

Metals- elements with a high luster, high electrical conductivity (all are solids except for Hg)

Nonmetals- nonlustrous, poor electrical conductivity (all gases are nonmetals, some are brittle solids, one liquids- Br)

Semimetals (metalloids)- have the properties of both metals and nonmetals (all solids- B, Si, Ge, As, Sb, Te, Po, At)

IONS- Atom or group of atoms with a positive or negative charge due to loss or gain of electrons

CATIONS- Ion with a positive charge (loses electrons)
(to get electrons, SUBTRACTcharge from # of protons) – end in ion
ANIONS- Ion with a negative charge (gains electrons) – end in -ide
(to get electrons, ADD charge to #~~o~~f protons) +

Ion	Atomic #	Mass #	p^+	n^0	e^-	Name
Na^{+1}	11	23	11	12	10	sodium ion
Ca^{+2}	20	40	20	20	18	calcium ion
I^{-1}	53	127	53	74	54	iodide
N^{-3}	7	14	7	7	10	nitride
O^{-2}	8	16	8	8	10	oxide

ATOMIC MASS- **The weighted average mass of all isotopes for a given element**

A weighted average is calculated using masses of all isotopes for an element and the relative abundance in which they occur in nature

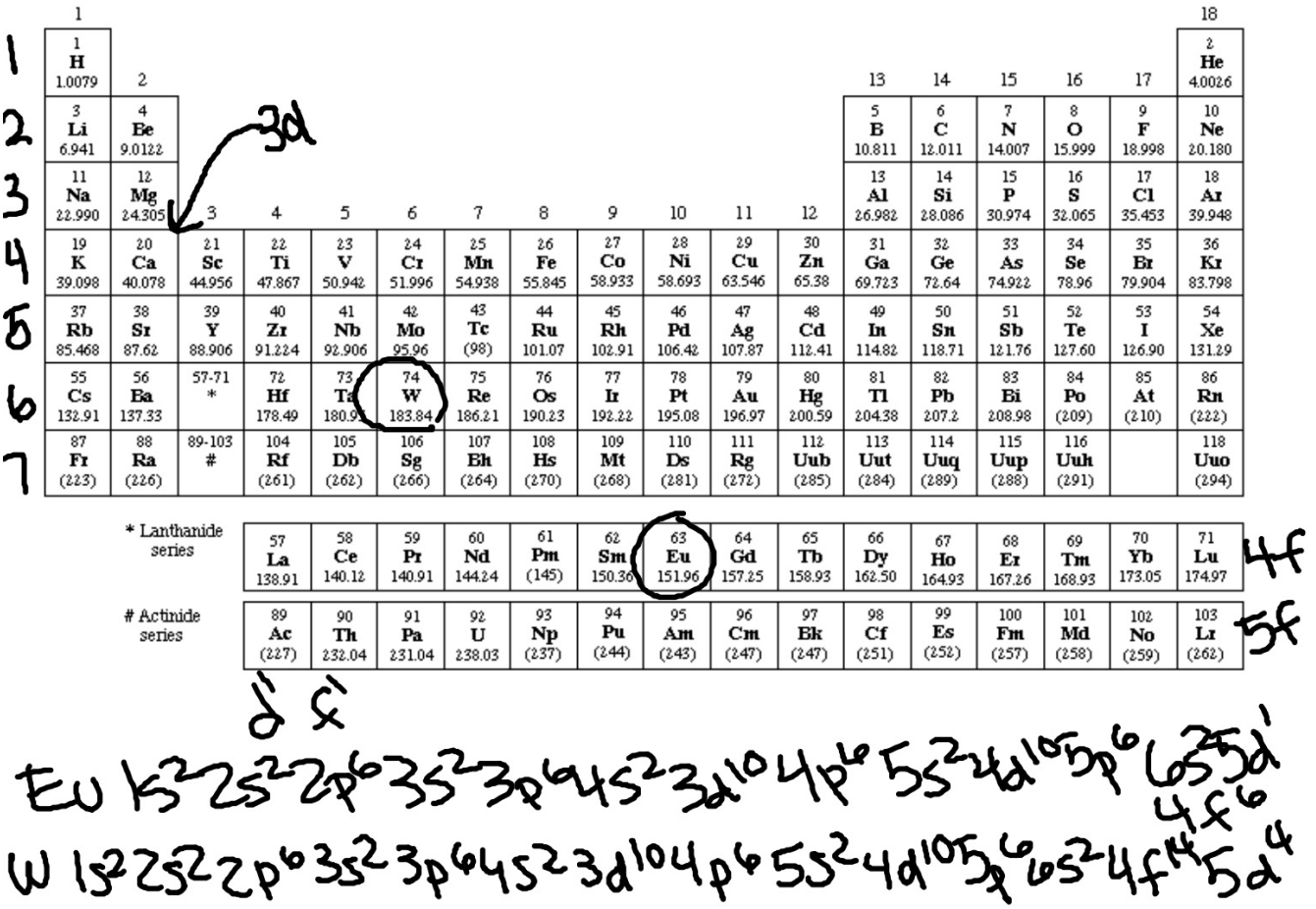
Example:

Name	Symbol	Natural % abundance	Average atomic mass
Oxygen	^{16}O	99.759	(.99759)(16)
	^{17}O	0.037	(.00037)(17)
	^{18}O	0.204	+(.00204)(18)
			16.00

To calculate average atomic mass:

Values can be compared with atomic mass values on the periodic table

Name	Symbol	Natural % abundance	Average atomic mass
Sulfur	^{32}S	95.00	(.95)(32)
	^{33}S	0.76	(.0076)(33)
	^{34}S	4.22	(.0422)(34)
	^{36}S	0.014	+(.00014)(36)
			32.09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18
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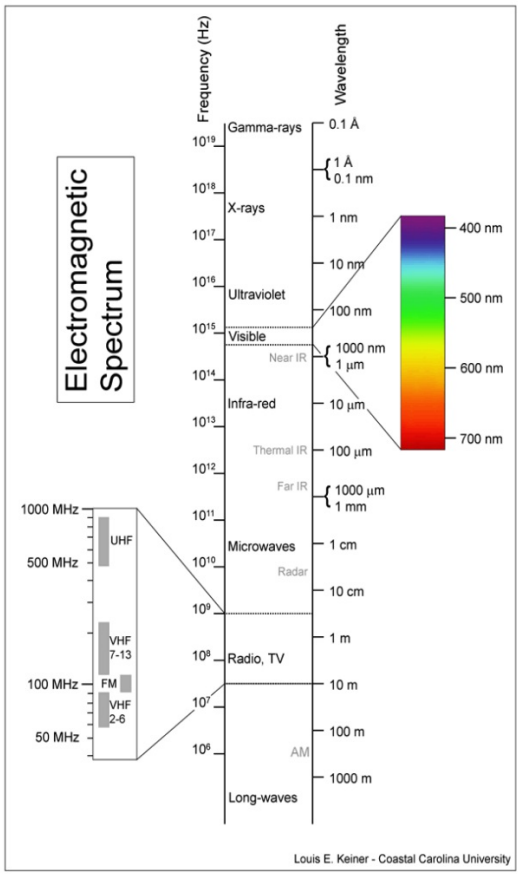
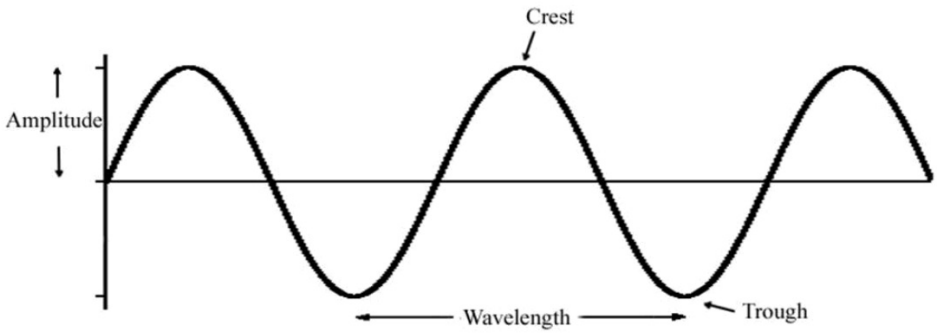
103 Lr (262)
3d
4f
5f
d
f
Eu 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶5s²4d¹⁰5p⁶6s²4f⁷
W 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶5s²4d¹⁰5p⁶6s²4f¹⁴5d⁴

Electromagnetic Radiation- any form of energy that exhibits wave-like behavior as it travels through space



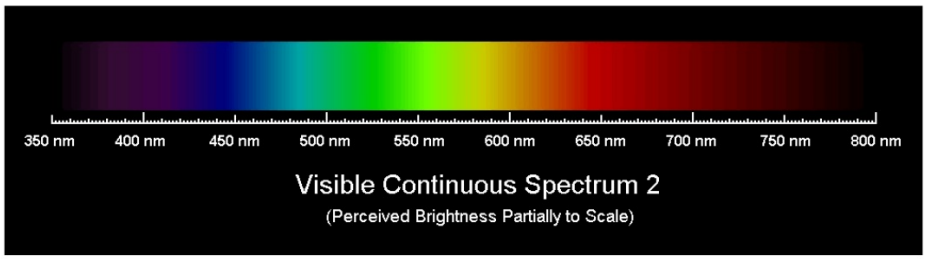


Amplitude- vertical distance from the middle of the wave to the crest or trough

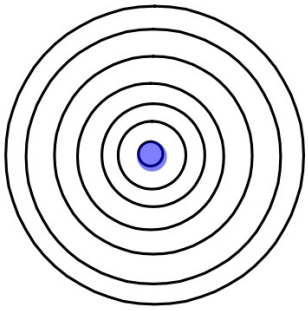
Wavelength- (symbolized λ- Greek letter lambda)- The distance between identical points on two identical waves- measured in units of distance

Frequency (symbolized ν- Greek letter nu)- The number of waves that pass a given point per unit time- measured in Hertz 1 Hz. = 1 cycle/sec.

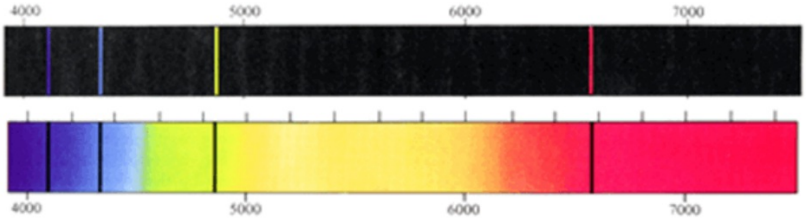
Continuous spectrum- spectrum in which all wavelengths in a given range are included



Energy level- region around the nucleus of an atom where an electron is likely to be moving- atoms have <u>quantized</u> energy levels (set amounts of energy)



<u>Atomic emission spectrum</u>- line spectra emitted by substances when energy is added to an atom and electrons move from higher energy levels back to their ground states- discrete amounts of energy released- each element has a unique pattern like a fingerprint



The Hydrogen Spectrum
Above: Emission
Below: Absorption

Hydrogen spectrum:
Lyman series- ultraviolet part of spectrum
Balmer series- visible parts of spectrum
Paschen series- infrared parts of spectrum

Arrangement of electrons in atoms

Energy levels- designated by principal quantum numbers n=1,2,3,4,5,6... (smaller numbers indicate closer proximity to nucleus)

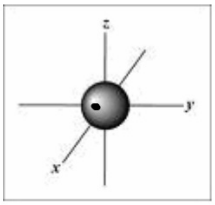
Energy sublevels- located within each energy level- 4 types based on the type of cloud shapes present (designated by 4 letters- s,p,d,f)

(Electron cloud)

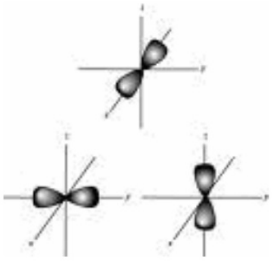
Atomic orbitals- region in space where there is a high probability of finding an electron- Each orbital can have a maximum of 2 electrons

Atomic orbitals

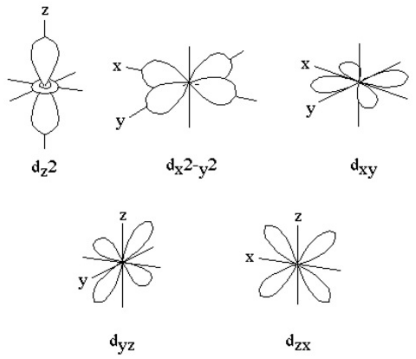
s orbital- spherical shape- 1 orbital (max 2 electrons)



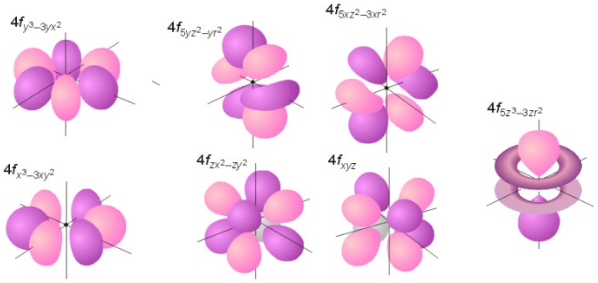
p orbitals- dumbbell shape- 3 orbitals (max 6 electrons)



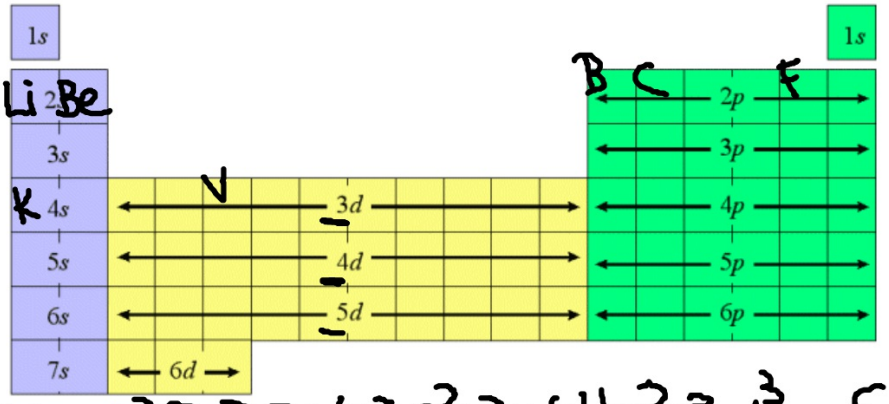
d orbitals- more complex shape- 5 orbitals (max 10 electrons)



f orbitals- most complex- 7 orbitals (max 14 electrons)



Energy level	# of sublevels	Type of sublevel	# of elec.
n=1	1	1s	2
n=2	2	2s, 2p	8
n=3	3	3s, 3p, 3d	18
n=4	4	4s, 4p, 4d, 4f	32



V $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

C $1s^2 2s^2 2p^2$

F $1s^2 2s^2 2p^5$

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

S $1s^2 2s^2 2p^6 3s^2 3p^4$

He $1s^2$

Li $1s^2 2s^1$

Be $1s^2 2s^2$

B $1s^2 2s^2 2p^1$

K $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

V

W

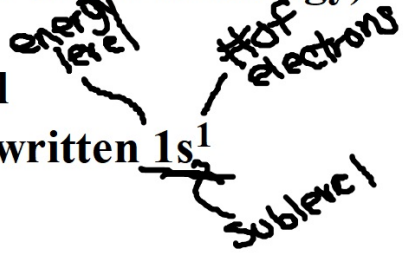
<u>Electron configuration</u>- the distribution of electrons among the various orbitals in an atom or molecule

Rules for filling atomic orbitals:

<u>The Aufbau Principle</u>: Electrons enter orbitals of lowest energy first (see periodic table diagram for order of energy)

Ex. Hydrogen- 1 electron in the 1s orbital

Electron configuration is written $1s^1$



He

Li

Be

B

K

V

W

Shortcut electron configurations

Use the noble gas (Group VIIIA element) that comes just before in brackets [ ]. Write only the additional orbits that come after the noble gas.

Example: Mg [Ne] $3s^2$

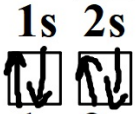
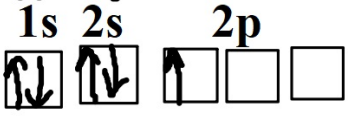
V

W

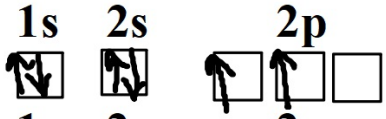
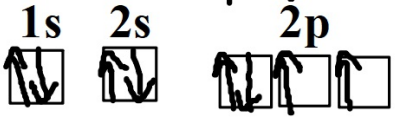
<u>Pauli Exclusion Principle</u>- An atomic orbital may describe at most two electrons. To occupy the same orbital, two electrons must have opposite spin.

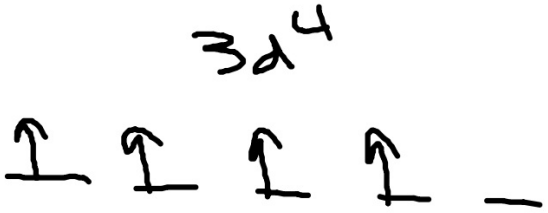
Orbital diagrams- use boxes or lines to represent each orbital, arrows to represent each electron

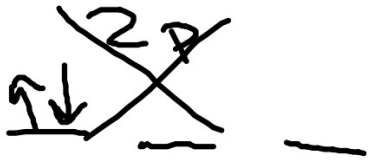
s-1 p-3 d-5 f-7

Element	Elec. configuration	Orbital diagram
Hydrogen	$1s^1$	
Helium	$1s^2$	
Beryllium	$1s^2 2s^2$	
Boron	$1s^2 2s^2 2p^1$	

Hund's Rule- when electrons occupy orbitals of equal energy, one electron enters each orbital until all the orbitals contain one electron with spins parallel

Element	Elec. configuration	Orbital diagram
Carbon	$1s^2\ 2s^2\ 2p^2$	
Oxygen	$1s^2\ 2s^2\ 2p^4$	





Exceptional electron configurations

Chromium
Expected
Actual

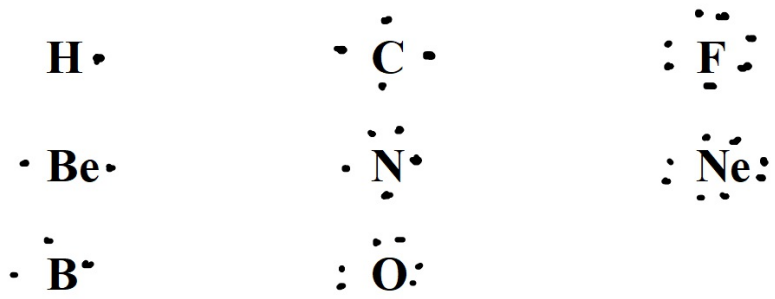
Copper
Expected
Actual

Why?
Filled energy levels are more stable than partially filled sublevels.

<u>Valence electrons</u>- the electrons in the highest occupied energy level of an atom; those involved in bonding

For representative elements- the number of valence electrons is the same as the group number (Roman numeral)

<u>Electron dot structure</u>- representation that depicts valence electrons as dots around the element's symbols

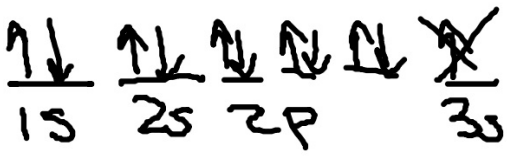


Stable electron configurations for cations or anions

<u>Octet rule</u>- atoms tend to form ions so that they have eight electrons in their outer shell (same as a noble gas)

metals- lose electrons

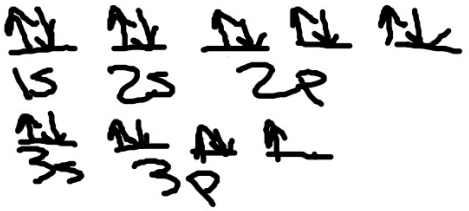
Sodium atom **Sodium ion** Na^{+1}



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

same as Ne

Chlorine atom **Chlorine ion** Cl^{-1}



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

same as Ar

Periodicity

Periodic table- organizes elements according to similar properties

History

1860s- first organization by Mendeleev (Mendeleev's grouping was amazingly similar to present groupings

http://corrosion-doctors.org/Periodic/Periodic-Mendeleev.htm

1913- Henry Mosely determined atomic numbers of elements and arranged atoms in order of increasing atomic number

Periodic Law- when the elements are arranged in order of increasing atomic number, there is a periodic pattern in their physical and chemical properties

Trends in Atomic Size

covalent atomic radius- half the distance between the nuclei of two atoms in a homonuclear diatomic molecule

-- *This is one way to estimate the size of an atom due to the fact that the radius cannot be determined directly because an atom does not have a sharply defined boundary*

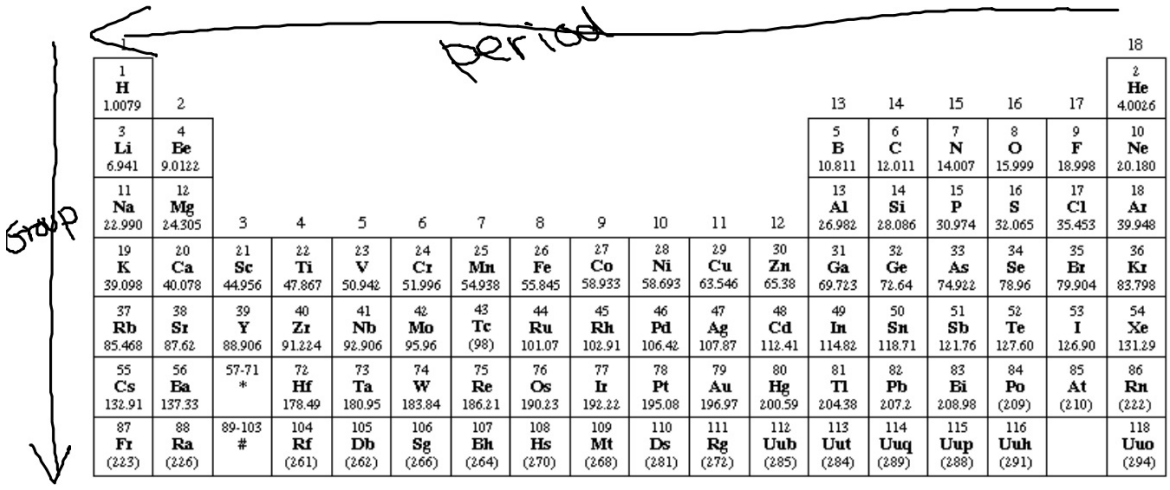
Group Trends

Atomic size increases as we move down a group because electrons are added to higher energy levels as we move down

Periodic Trends

Atomic size decreases from left to right across a period. Electrons and protons are added as you move across a period. The effect of increasing nuclear charge on the outermost electrons is to pull them closer to the nucleus.

Trend in Atomic Size



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Increasing trend

Ionization energy- the energy to remove and electron from a gaseous atom

$$Na\,(g) \rightarrow Na^{+1} + e^{-}$$

First ionization energy- energy to remove first electron

Group IA- greatest increase from 1st to 2nd ion. en.

Second ionization energy- energy to remove 2nd electron

Group IIA- greatest increase from 2nd to 3rd ion en.

Group trend

Decreases down a group because size increases as you move down so outermost electrons are farther away from nucleus

Periodic trend

Increases from left to right because nuclear attraction increases as you go from left to right- more difficult to remove elec.

Electronegativity- the tendency of an atom to attract electrons to itself when chemically combined with another element (closely tied to ionization energy and electron affinity)

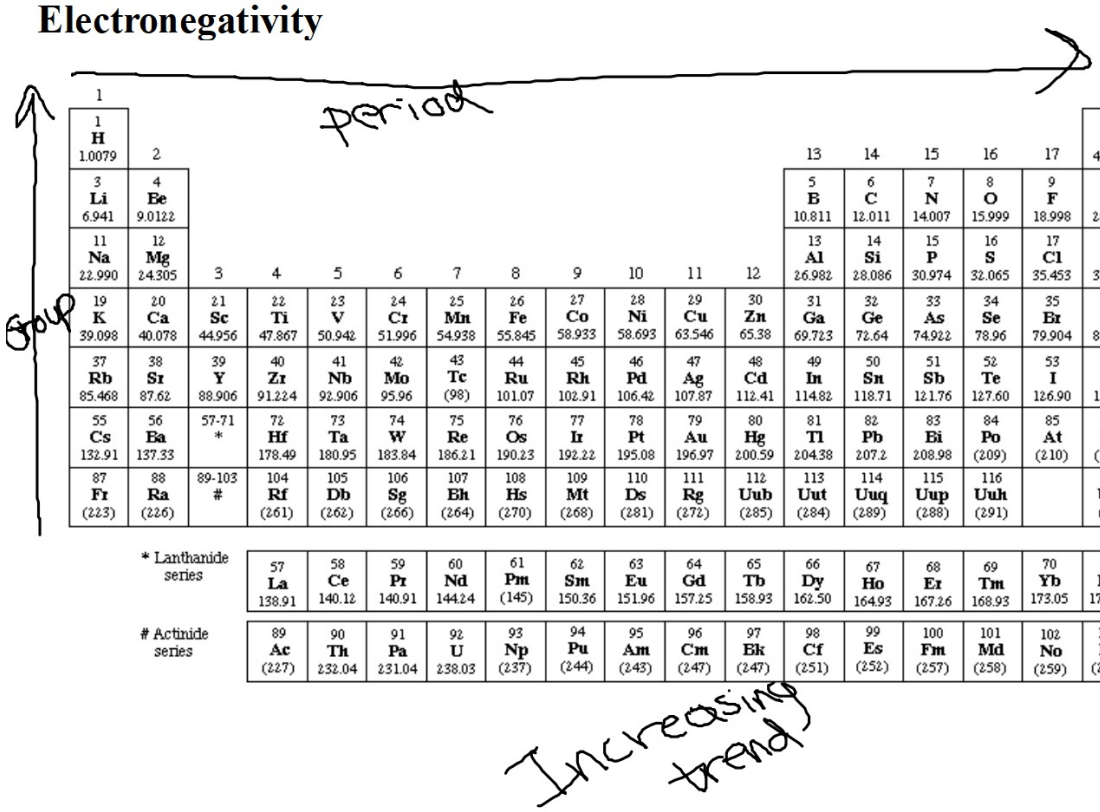
Group trend

Decrease down a group

Periodic trend

Increases left to right

Trend for Ionization energy, ~~Electron affinity~~, and Electronegativity



1																	18
1 H 1.0079	2											13	14	15	16	17	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305	3	4	5	6	7	8	9	10	11	12	13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.065	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.38	31 Ga 69.723	32 Ge 72.64	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.798
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.96	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 *	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89-103 #	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (270)	109 Mt (268)	110 Ds (281)	111 Rg (272)	112 Uub (285)	113 Uut (284)	114 Uuq (289)	115 Uup (288)	116 Uuh (291)		118 Uuo (294)

* Lanthanide series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
# Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

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

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

Br $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

Ba $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

Np $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^{10} 4f^{14} 6p^6 7s^2 6d^1 5f^4$

Co $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$

Ag $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^9$

Te $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^4$

Ra $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2$

Lr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 6d^1 5f^{14}$