Periodic Law
Topic#6
AMSAT Chem 1H

Student Edition

History

Objectives

- **Explain** the roles of Mendeleev and Moseley in the development of the periodic table.
- **Describe** the modern periodic table.
- **Explain** how the periodic law can be used to predict the physical and chemical properties of elements.
- **Describe** how the elements belonging to a group of the periodic table are interrelated in terms of atomic number.

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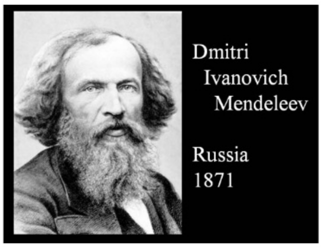
History

- **Mendeleev**
  - arranged the known elements by properties and then looked for trends (patterns)
  - when arranged by increasing atomic mass, similar properties appeared at regular intervals
    - this repeating interval is considered *periodic*
  - created a table where elements with similar properties were grouped together
    - a few elements grouped according to properties had an atomic mass less than the element that proceeded it.
      - For example, I was placed after Te even though I has a lower atomic mass than Te)
        - Why was I placed after Te?
        - Other examples include Ni and Co,
  - considered the father of the modern periodic table
  - procedure left empty spaces where he predicted the existence of new, undiscovered elements and some of their properties

ekasilicon (after silicon)

- predicted properties:	72amu,	$5.5 g/cm^3$	high mp, gray
- real:	72.59amu	$5.36 g/cm^3$	high mp, gray



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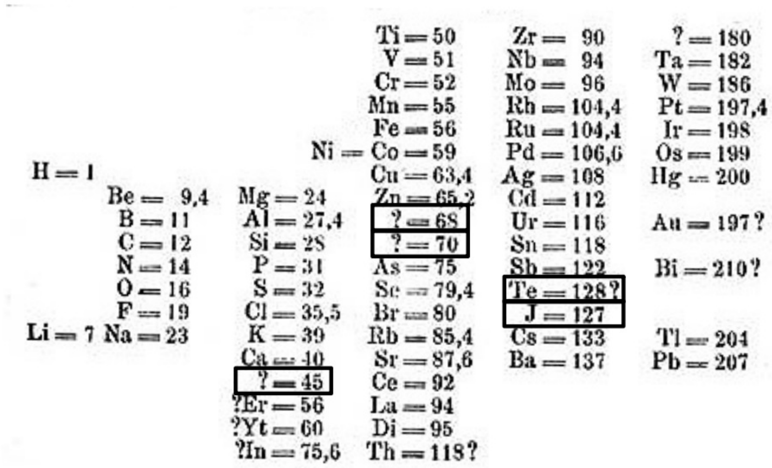

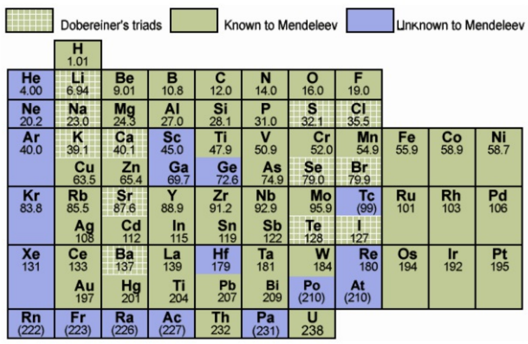

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Properties	Ekaaluminum (gallium, discovered 1875)		Ekaboron (scandium, discovered 1877)		Ekasilicon (germanium, discovered 1886)	
	Predicted	Observed	Predicted	Observed	Predicted	Observed
Density	6.0 g/cm^3	5.96 g/cm^3	3.5 g/cm^3	3.5 g/cm^3	5.5 g/cm^3	5.47 g/cm^3
Melting point	low	30°C	*	*	high	900°C
Formula of oxide	Ea_2O_3	Ga_2O_3	Eb_2O_3	Sc_2O_3	EsO_2	GeO_2
Solubility of oxide	*	*	dissolves in acid	dissolves in acid	*	*
Density of oxide	*	*	*	*	4.7 g/cm^3	4.70 g/cm^3
Formula of chloride	*	*	*	*	$EsCl_4$	$GeCl_4$
Color of metal	*	*	*	*	dark gray	grayish white

Predicted elements	Element and year discovered	Properties	Predicted properties	Observed properties
Ekaaluminum	gallium 1875	density of metal	6.0 g/mL	5.96 g/mL
		melting point	low	30°C
		oxide formula	Ea_2O_3	Ga_2O_3
Ekaboron	scandium 1877	density of metal	3.5 g/mL	3.86 g/mL
		oxide formula	Eb_2O_3	Sc_2O_3
		solubility of oxide	dissolves in acid	dissolves in acid
Ekasilicon	germanium 1886	melting point	high	900°C
		density of metal	5.5 g/mL	5.47 g/mL
		color of metal	dark gray	grayish white
		oxide formula	EsO_2	GeO_2
		density of oxide	4.7 g/mL	4.70 g/mL
		chloride formula	$EsCl_4$	$GeCl_4$

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History

The Modern Periodic Table

- **Mosely**
  - worked with Rutherford, discovered atomic number, arranged periodic table according to atomic number, confirmed the broken pattern (Te/I and Co/Ni) of Mendeleev, and confirmed the ordering of elements by properties.
- **periodic law**
  - The physical and chemical properties of elements are periodic functions of their atomic numbers.
- **Noble Gases (Group 18)**
  - Argon was discovered by English physicist John William Strutt (Lord Rayleigh) and Scottish chemist Sir William Ramsey.
  - Soon afterward He was discovered
  - Ramsey proposed that a new group needed to be added between the fluorine family (group 17) and the lithium family (group 1) thus creating group 18, The Noble gases.
- **Lanthanides (4*f*'s)**
  - 14 elements from cerium, Ce (58) and lutetium, Lu (71)
  - Incredibly similar in physical and chemical properties, found together in the earth and very hard to separate from each other.
- **Actinides (5*f*'s)**
  - 14 elements from thorium, Th (90) and lawrencium, Lr (103)
  - All are radioactive and most are man made (only thorium, protactinium, uranium are found in nature with the others being solely synthetic)

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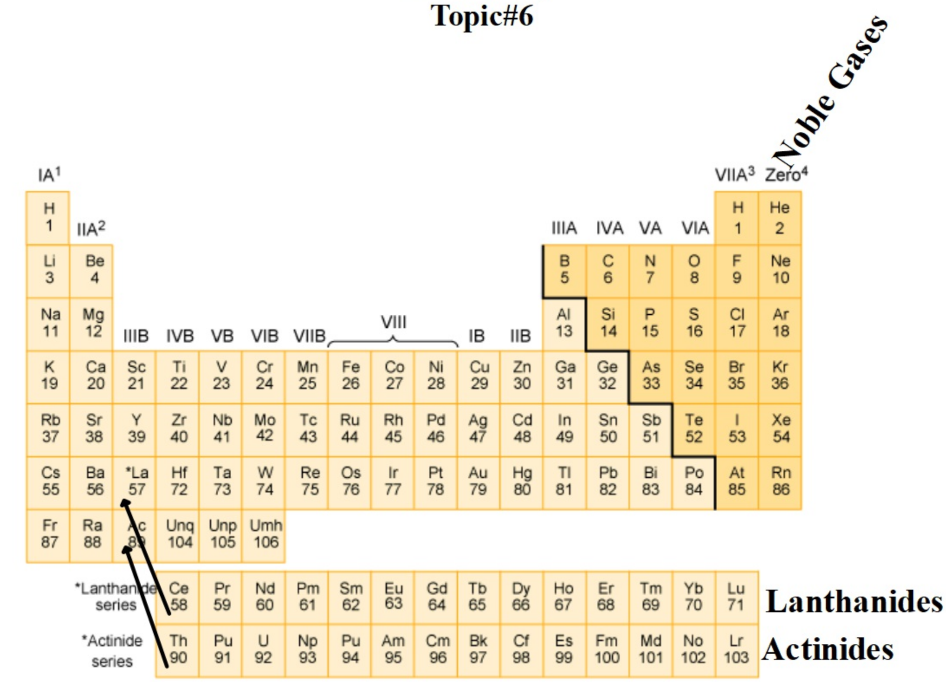
History



[1]Group IA (excluding hydrogen) compromises the alkali metals.
[2]Group IIA compromises the alkaline-earth metals.
[3]Group VIIA (excluding hydrogen) compromises the halogens.
[4]Group zero compromises the noble gases.

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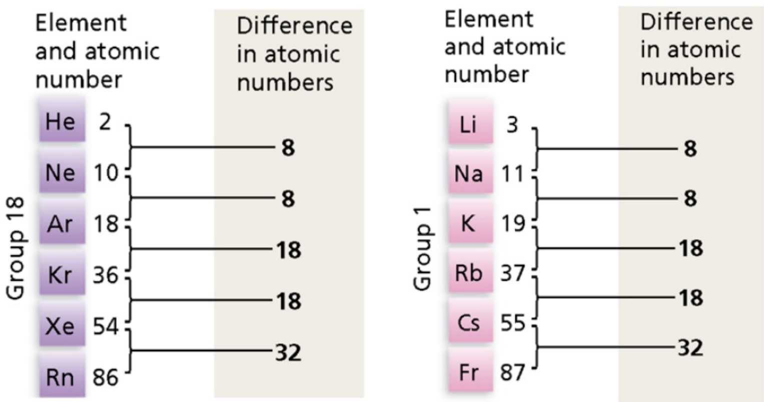
History

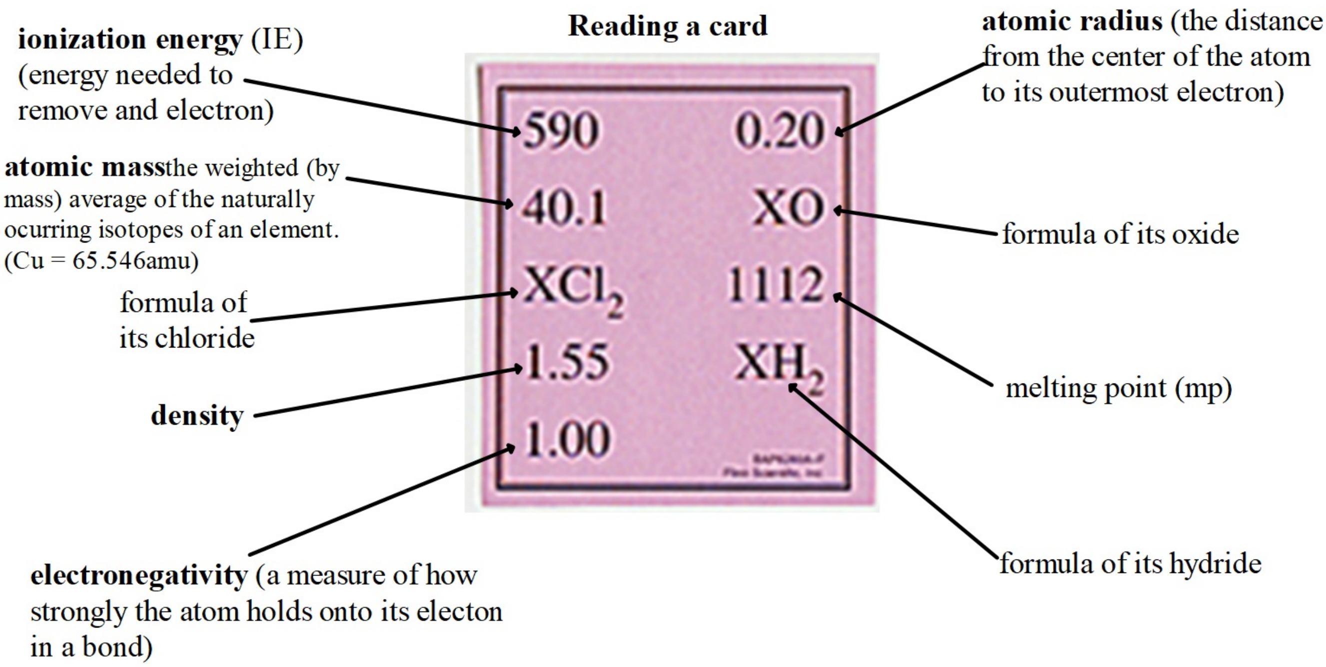
Periodicity

- With respect to atomic number can be found in any group.



Lab: It's in the Cards

Reading a card



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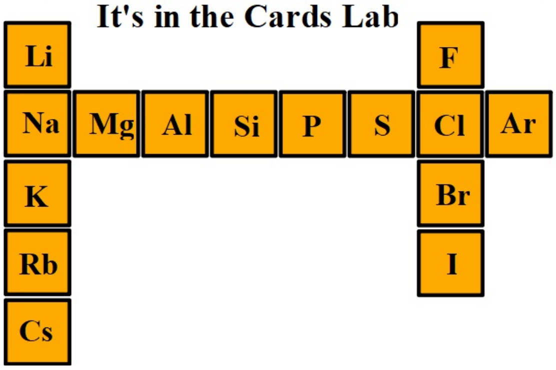
Lab: It's in the Cards

Property	Na	Mg	Al	Si	P	S	Cl	Ar
Electronegativity	1.0	1.3	1.6	1.9	2.2	2.6	3.2	--
Atomic Mass (amu)	23	24	27	28	31	32	35	40
Atomic Radius (Å)	2.23	1.72	1.82	1.46	1.23	1.09	0.97	0.88
mp(K)	371	923	933	1687	317	388	172	84
I_E(kJ/mol)	496	738	578	787	1012	1000	1251	1521
Density	0.97g/cm^3	1.7g/cm^3	2.7g/cm^3	2.3g/cm^3	1.8g/cm^3	2.1g/cm^3	3.2g/L	1.8g/L

Element	At Mass (amu)	At Radius (Å)	MP(K)	I_E(kJ/mol)	EN	Density
Li/F	7/19	2.05/0.57	454/54	520/1681	1.0/4/0	0.53(g/cm^3)/1.70(g/L)
Na/Cl	23/35	2.23/0.97	371/172	496/1251	1.0/3.2	0.97(g/cm^3)/3.2(g/L)
K/Br	39/80	2.77/1.12	337/266	419/1140	0.8/3.0	0.86(g/cm^3)/3.1(g/cm^3)
Rb/I	85/127	2.98/1.32	312/387	403/1008	0.8/2.7	1.5(g/cm^3)/4.9(g/cm^3)
Cs	133	3.34	302	376	0.7	1.9(g/cm^3)

It's in the Cards Lab



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Electron Configuration and Periodic Table

Objectives

•**Describ**e the relationship between electrons in sublevels and the length of each period of the periodic table.

•**Locate** and name the four blocks of the periodic table. Explain the reasons for these names.

•**Discuss** the relationship between group configurations and group numbers.

•**Describe** the locations in the periodic table and the general properties of the alkali metals, alkaline earth metals, halogens, and noble gases.

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Electron Configuration and Periodic Table

Alkali Metals

- Group 1, most reactive metals, form 1+ ions, ns^1, valance 1

Alkaline Earth metals

- Group 2, reactive metals, form 2+ ions, ns^2, valence 2

Hydrogen

- 1 valence, nonmetal, forms 1+ and 1- ions, ns^1

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Electron Configuration and Periodic Table

Sample WS#1 - Using Electron Configuration *s*-block Elements

1. Without looking at the periodic table, give the group (___), period (___), and block (___) in which the element with the electron configuration [Xe]6s^2 is located.

2. Without looking at the periodic, write the electron configuration for the group 1 element in the third period. Is this element likely to be more active or less active than the element described in question 1?

3. Without looking at the periodic table, give the group (___), period (___), and block (___) in which element with the electron configuration [Kr]5s^1 is located.

4. Without looking at the periodic table, write the complete electron configuration for the group 2 element in the fourth period.

5. Without looking at the periodic table, give the group configuration for the group 2 elements.

Electron Configuration and Periodic Table

Transition Elements (all metals)

- the *d*-block elements: Groups 3-12
- group configuration for these elements is $(n-1)d^{1-10}ns^2$
  - 4th row transition metals: 4-1 = 3 so $3d^{1-10}4s^2$
    - i.e. Ti is $3d^2 4s^2$
- good conductors of heat and electricity
- some are not reactive: Pt, Au, and Pd
- a few can exist in nature as free elements (Au and Pt)
- coinage metals: Cu, Ag, and Au

Electron Configuration and Periodic Table

d-block Elements

6. An element has the electron configuration [Kr]$4d^5 5s^1$. Without looking at the periodic table, identify the period (___), block (___), and group (___) in which the element is located. Then consult the periodic table to identify this element and the ohthers in its group.

7. Without looking at the periodic table, identify the period (___), block (___), and group (___) in which this element with the electron configuration [Ar]$3d^8 4s^2$ is located.

8. Without looking at the periodic table, write the outer electron configuration for the group 12 element in the fifth period.

9. Refer to the periodic table to identify the element described in question 8 and write the element's Noble-gas notation.

Electron Configuration and Periodic Table

The *p*-block elements (metals, metalloids, and nonmetals)

- outer electron config - $ns^2np^{1\text{-}6}$
- valence 3-8
  - bromine in group 17 : val = 17-10 = 7
- includes all of the nonmetals except He and H
- includes all of the metalloids: B, Si, Ge, Te, As, and Sb
- *p*-block metals can lose their *s* and *p* electrons to form ions
  - Sn has an outer electron config of $5s^24d^{10}5p^2$
    - tin ions are Sn^{2+} (2 *p* electrons lost) and Sn^{4+} (2 *p* and 2 *s* electrons lost)
- Group 17
  - halogens, ns^2np^5, valence of 7, form (1-) ions, most reactive nonmetals, fluorine is the most reactive nonmetal
- Group 18
  - Noble gases, inert (do not react: Kr, Xe, and Rn can react with Cl, F, and O at high temperatures and pressures),

Main group elements (representative elements) - *s* and *p*-block elements

Valence electrons:	1	2	3	4	5	6	7	8
Bonding:	1	2	3	4	3	2	1	0

Electron Configuration and Periodic Table

p-block Elements

10. Without looking at the periodic table, write the outer electron configuration for the group 17 element in the third period.

11. Name the element described in question 10, and identify it as a metal, metalloid, or nonmetal.

12. Without looking at the periodic table, identify the period (___), block (___), and group (___) of an element with the electron configuration [Ar]$3d^{10}4s^24p^3$.

13. Name the element described in question 12, and identify it as a metal, metalloid, or nonmetal.

Electron Configuration and Periodic Table

f-block Elements

- lanthanides and actinides
- all actinides are radioactive
- all very similar in properties, so they are very hard to separate when mined
- similar to group 2 in properties
- actinides first three are naturally occuring others are person-made

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Electron Configuration and Periodic Table

Review: The Element Blocks of the Periodic Table

14. The electron configuration of atoms of four elements are written below. For each element, name the block and group in which it is located. Then name the element by consulting the periodic table. Identify each element as a metal, metalloid, or nonmetal. Finally, describe it as likely to be of high reactivity or of low reactivity.

	Block	Period	Group	Group name	Element	Type	Reactivity
a. $[Xe]4f^{14}5d^{9}6s^{1}$	d	6	10	transition metals	Pt	metal	low
b. $[Ne]3s^{2}3p^{5}$							
c. $[Ne]3s^{2}3p^{6}$							
d. $[Ar]3d^{10}4s^{2}\ 4p^{2}$							
e. $[He]2s^{2}2p^{5}$							
f. $[Ar]3d^{10}4s^{1}$							
g. $[Kr]5s^{2}$							

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Periodic Trends

Atomic Radius (A_r) - distance from center of atom to outer most electron.

Ionic Radius (I_r) - half the distance between two ionically bonded atoms (ions, transfer of electrons from metal to nonmetal)

Ionization Energy (IE) - the amount of energy needed to remove an electron from a gaseous atom.

Electron Affinity (A_e) - the energy change when a gaseous atom accepts an electron

Electronegativity (EN) - measures the ability of an atom to hold on to its electron when in a bond.

Metallic Character (M_c) - the level of reactivity of a metal

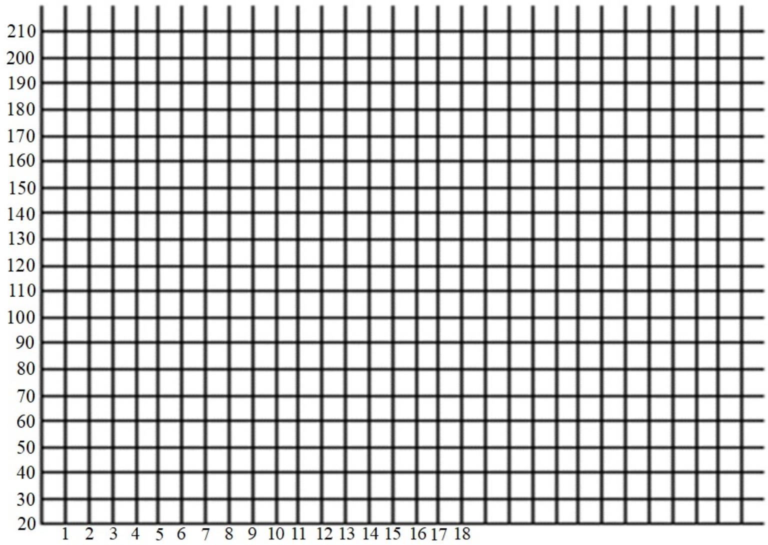
<u>Periodic Trends - Atomic Radius (A_r)</u>

Periodic Trends Graphing Activity/Study Aide

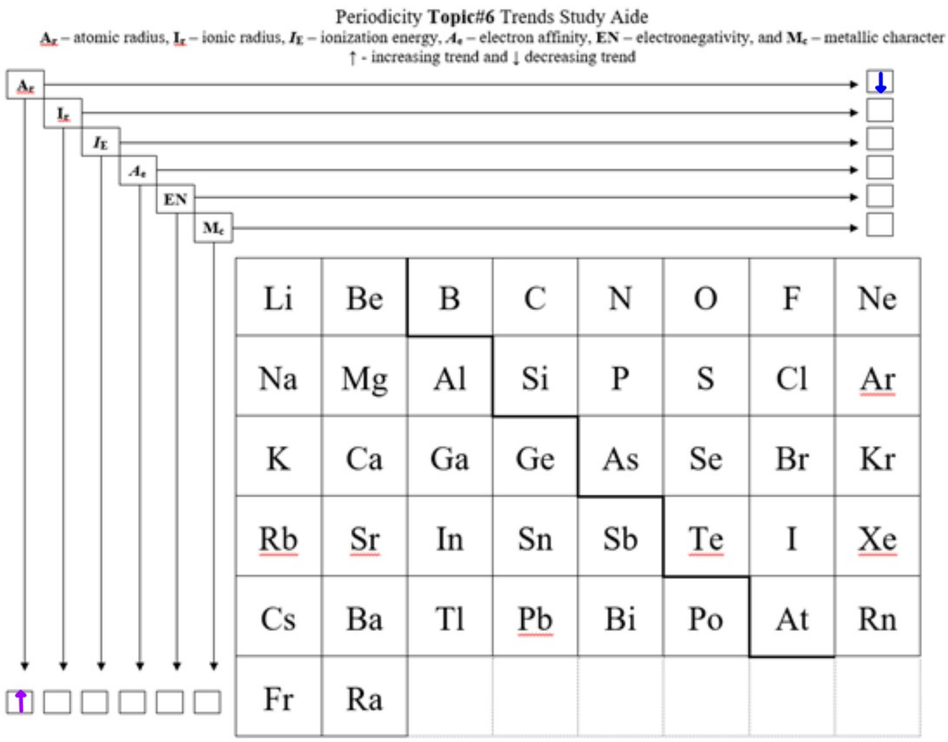
Need: 2 pieces of double sided graph paper and Periodic Trends Study Sheet

Element	Atomic Radius*	Element	Atomic Radius*
H	53	Ne	38
He	31	Na	190
Li	167	Mg	145
Be	112	Al	118
B	87	Si	111
C	67	P	98
N	56	S	88
O	48	Cl	79
F	42	Ar	71

*Atomic radius is measured in picometers (pm)



Periodic Trends - Atomic Radius (A_r)



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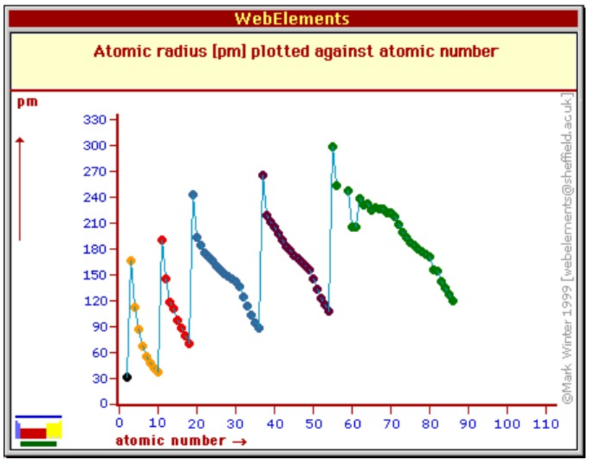
Periodic Trends - Atomic Radius (A_r)

Trend for atomic radius

- decreases across a period
  - Why? electrons move into the same energy level with more protons in nucleus pulling them towards the center of the atom (an increase in coulombic force due to increase in nuclear charge decreases distance between nucleus and outer electron(s))
- increases down a group
  - Why? Each move down increases the number of energy levels the atom has (more shells) (decrease in coulombic force due to an increase in distance between nucleus and outer electron(s))



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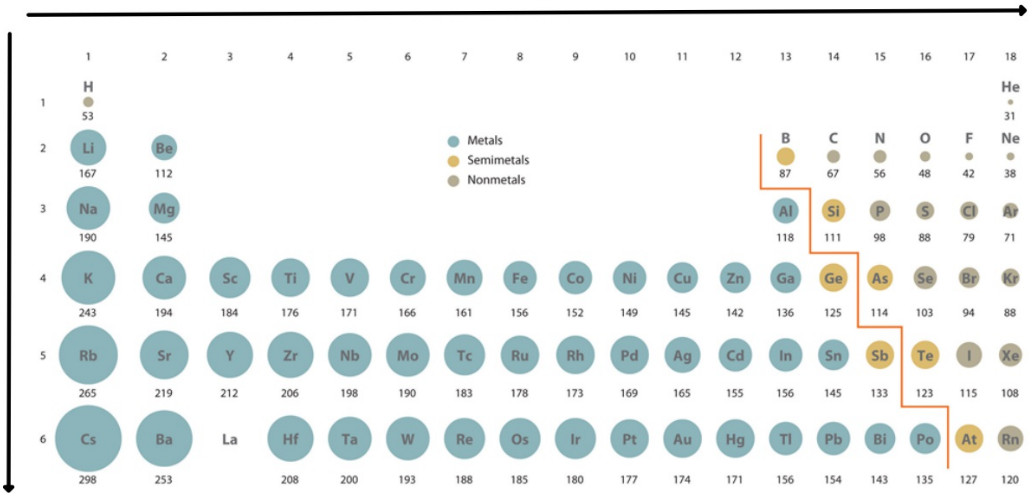
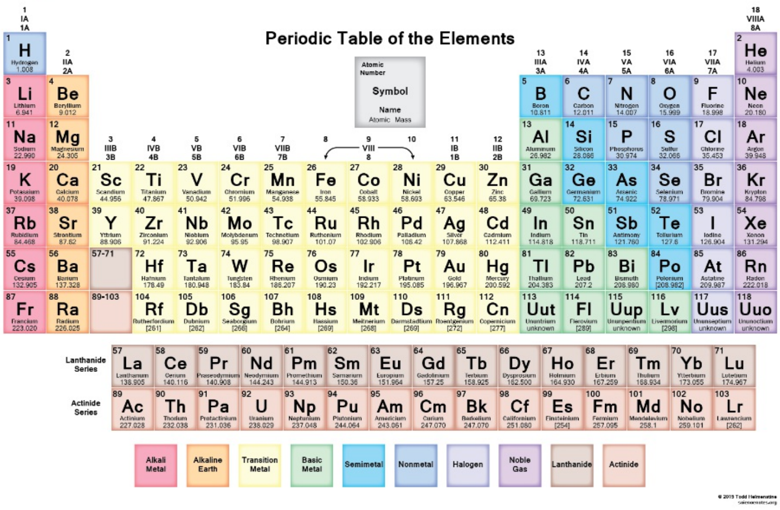
Periodic Trends - Atomic Radius (A_r)



Periodic Trends - Atomic Radius (A_r)

Sample WS#2 - Periodic Trends

1. Of the elements Mg, Cl, Na, and P, which has the largest atomic radius? Smallest? Explain your answer in terms of trends in the periodic table.
2. Of the elements Ca, Be, Ba, and Sr, which has the largest atomic radius? Smallest? Explain your answer in terms of trends in the periodic table

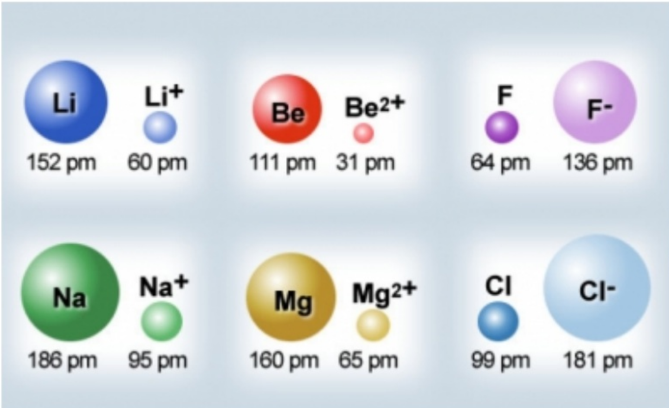


Periodic Trends - Ionic Radius (I_r)

What occurs to the radius of an atom as it gains or loses electrons (becomes an ion)?

- Loses electrons - becomes a cation ($A \rightarrow A^+ + e^-$).
  - the protons out number the electrons, so the electron-electron repulsion forces decrease allowing the protons to pull the outer electrons closer to the nucleus, thus decreasing the radius of the ion.
- Gains electrons - becomes an anion ($A + e^- \rightarrow A^-$).
  - the electrons out number the protons, so the electron-electron repulsion forces increase forcing the electrons to move farther apart while the protons have less pull on the electrons. Since the protons have less pull on the outer electrons, the radius of the ion increases.

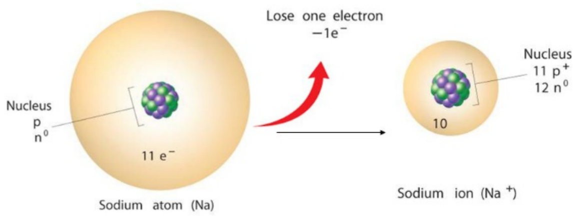
Question: Which ion is larger Fe^{3+} or Fe^{2+}? Why?



Periodic Trends - Ionic Radius (I_r)

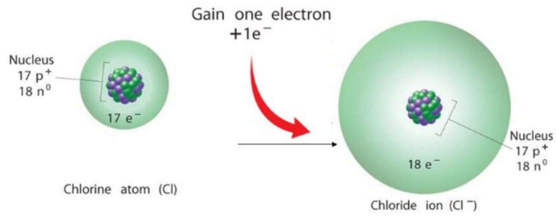
Trends in Ionic Size

Positive and negative ions form when electrons are transferred between atoms.



Trends in Ionic Size

Positive and negative ions form when electrons are transferred between atoms.



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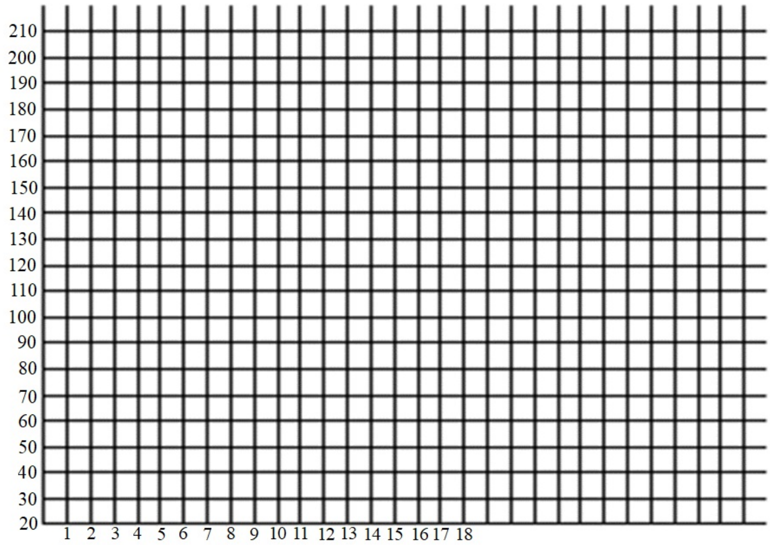
Periodic Trends - Ionic Radius (I_r)

Element	Ionic Radius*	Element	Ionic Radius*
Li^{+}	68	Na^{+}	97
Be^{2+}	31	Mg^{2+}	66
B^{3+}	23	Al^{3+}	51
C	-	Si	-
N^{3-}	171	P^{3-}	212
O^{2-}	140	S^{2-}	184
F^{1-}	133	Cl^{1-}	181

*Ionic radius is measured in picometers (pm)



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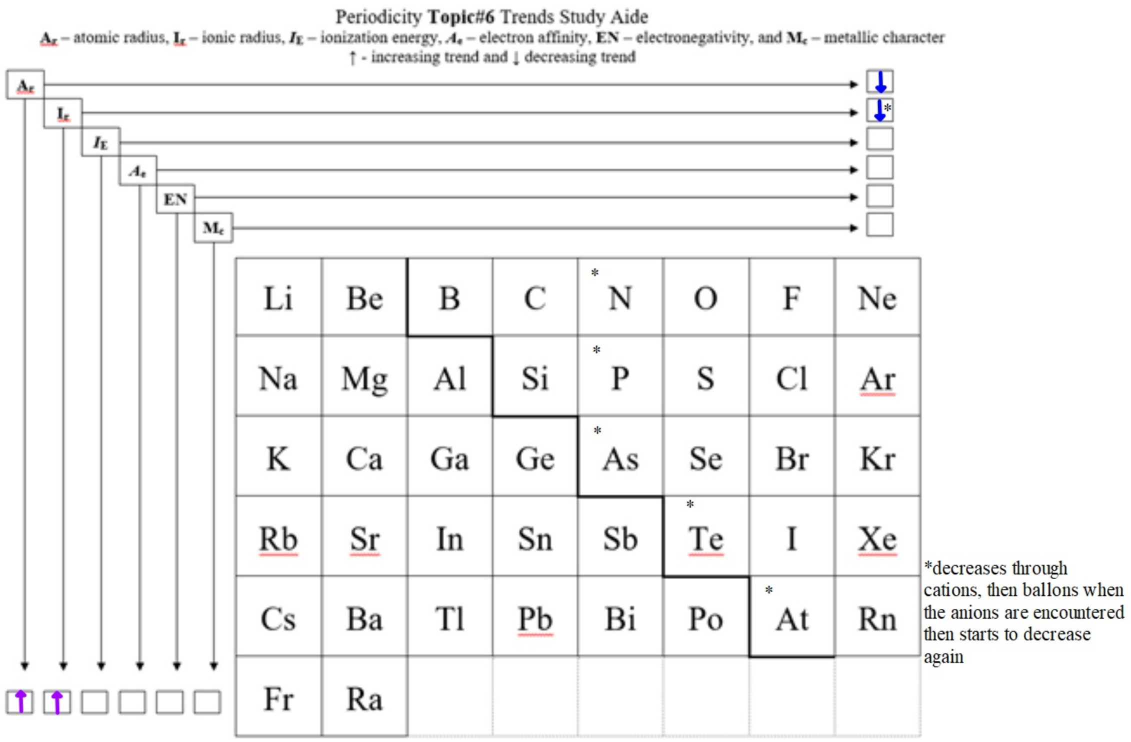
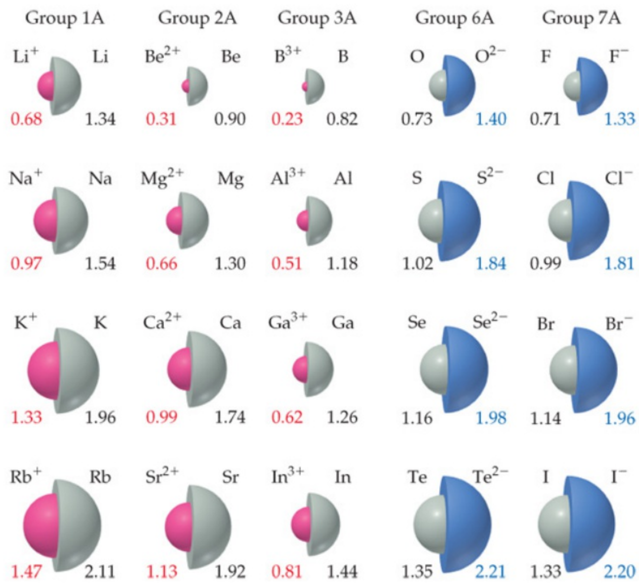
Periodic Trends - Ionic Radius (I_r)



Periodic Trends - Ionic Radius (I_r)

- decreases across a period, but jumps in radius when moving through the metalloids into the nonmetals.
  - Why? Same as above. (an increase in coulombic force due to increase in nuclear charge decreases distance between nucleus and outer electron(s))
  - decreases through cations in a period then expands greatly with the anions in a period then starts to decrease as one moves to the end of the period.
  - Metals form cations (positive ions) which are smaller than their atoms.
  - Nonmetals form anions (negative ions) which are larger than their atoms.
- increases down a group
  - Why? Same as above. (decrease in coulombic force due to an increase in distance between nucleus and outer electron(s))



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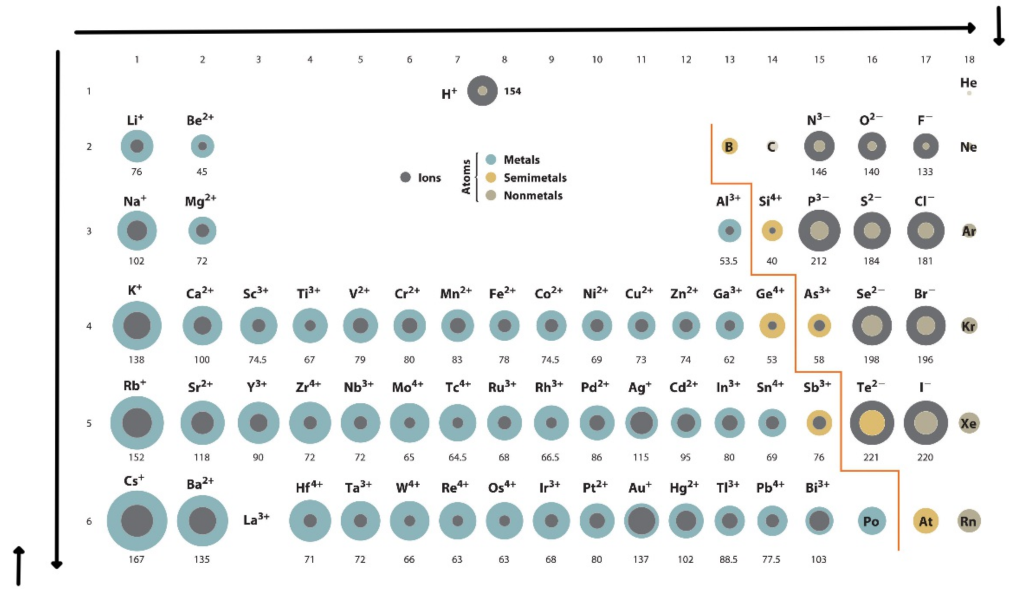
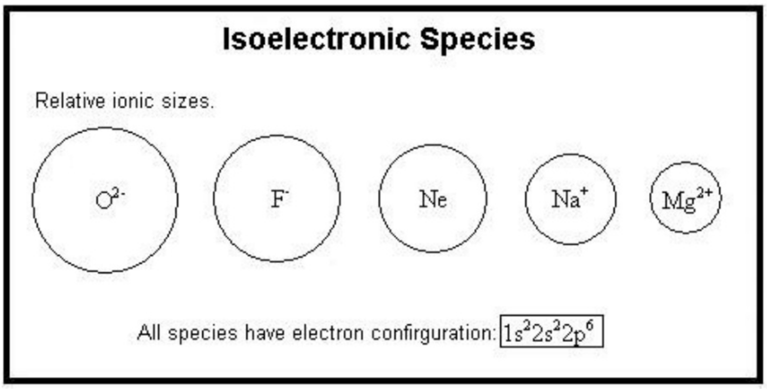
Periodic Trends - Ionic Radius (I_r)



$Fe^{3+} = 23e^-$
$Fe^{2+} = 24e^-$
$Fe^{1+} = 25e^-$
$Fe^{0} = 26e^-$

$Fe^{3+} < Fe^{2+} < Fe^{1+} < Fe$

Periodic Trends - Ionic Radius (I_r)



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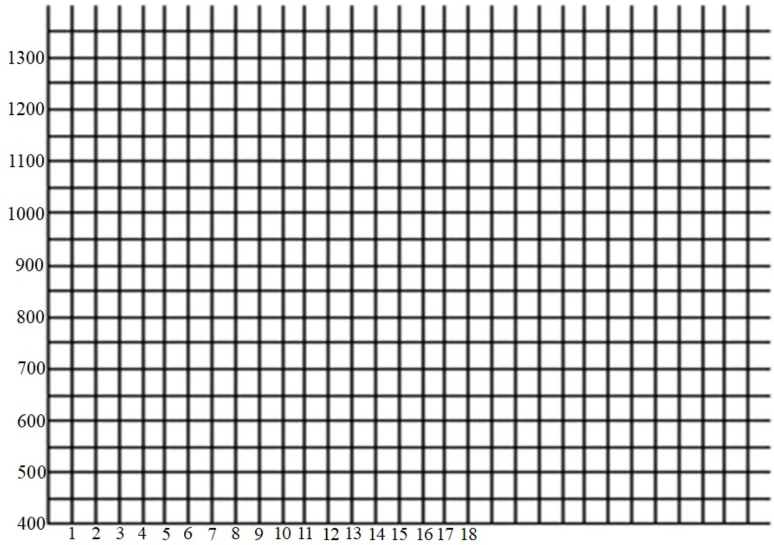
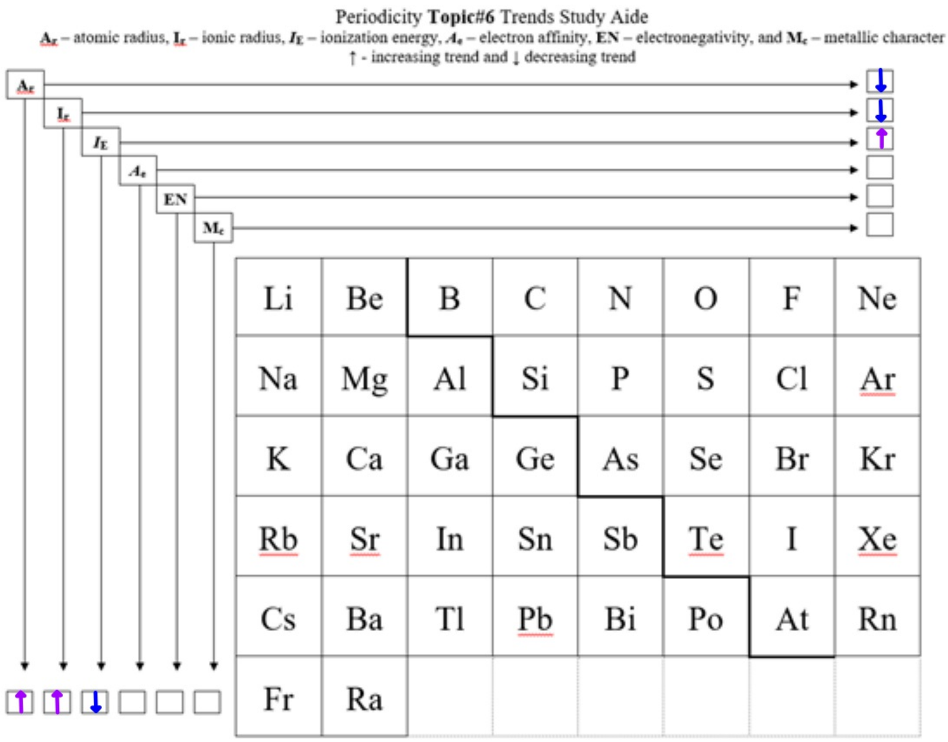
Periodic Trends - Ionization Energy (*IE*)

Element	IE_1*	Element	IE_1*
Li	520	Na	496
Be	899	Mg	738
B	801	Al	578
C	1086	Si	787
N	1420	P	1012
O	1314	S	1000
F	1681	Cl	1251

*Ionization energy (IE) is in kJ/mol



Periodic Trends - Ionization Energy (*IE*)



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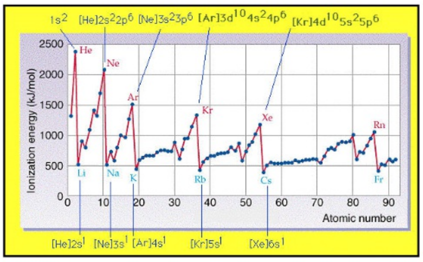
Periodic Trends - Ionization Energy (*IE*)



- The E required to remove one electron from a neutral atom
- A + energy → $A^+ + e^-$ (cation formation)
  - $IE_1 < IE_2 < IE_3 <$ etc.

Trend

- period: ionization energy increases across
  - Why? Pull of nucleus on outer electrons increases, harder to remove. (an increase in coulombic force due to increase in nuclear charge decreases distance between nucleus and outer electron(s))
- group: ioniziation energy decreases down
  - Why? Outer electrons are further from nucleus, easier to remove (decrease in coulombic force due to an increase in distance between nucleus and outer electron(s))

Periodic Trends - Ionization Energy (*IE*)

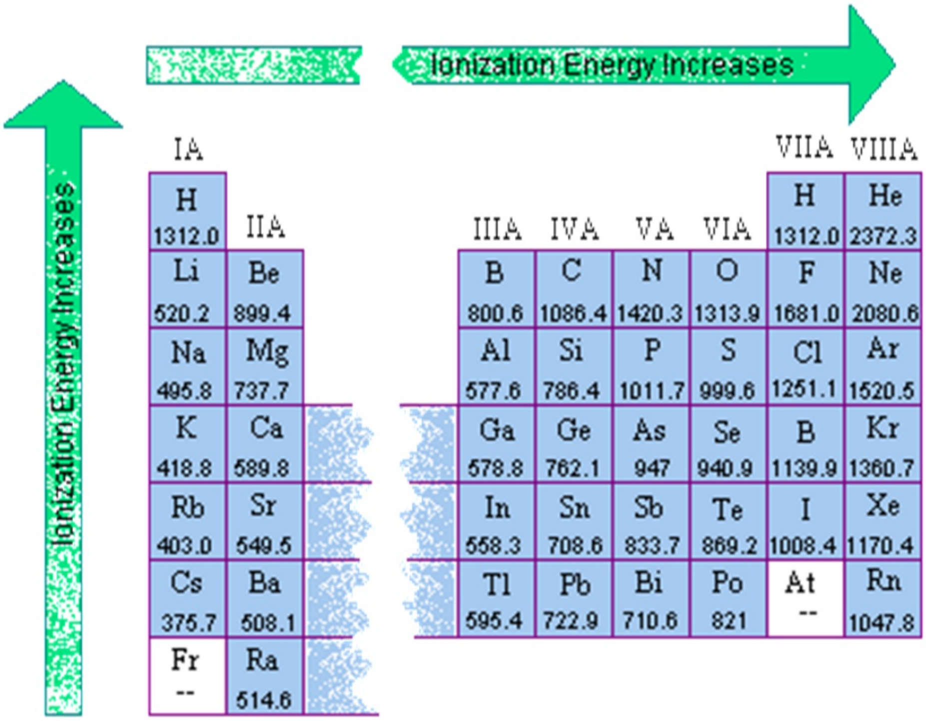
General increase →

General decrease ↓

Successive Ionization Energies for Period 3 Elements

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7
Na	498	4560	6910	9540	13 400	16 600	20 100
Mg	736	1445	7730	10 600	13 600	18 000	21 700
Al	577	1815	2740	11 600	15 000	18 310	23 290
Si	787	1575	3220	4350	16 100	19 800	23 800
P	1063	1890	2905	4950	6270	21 200	25 400
S	1000	2260	3375	4565	6950	8490	27 000
Cl	1255	2295	3850	5160	6560	9360	11 000
Ar	1519	2665	3945	5770	7230	8780	12 000

Periodic Trends - Ionization Energy (*IE*)



Periodic Trends - Ionization Energy (*IE*)

3. Consider the two main-group elements A and B. Element A has a first IE_1 of 419 kJ/mol. Element B has an IE_1 of 1000 kJ/mol. For each element, decide if it is more likely to be in the *s*-block or *p*-block. Which element is more likely to form a positive ion?

Periodic Trends - Ionization Energy (*IE*)

4. Consider the four hypothetical main-group elements Q, R, T, and X with the outer electron configuration indicated below. Then answer the questions that follow.

Q: $3s^2 3p^5$ R: $3s^1$ T: $4d^{10} 5s^2 5p^5$ X: $4d^{10} 5s^2 5p^1$

a. Identify the block locations of each hypothetical main-group element.

b. Which of these elements are in the same period? Same group?

c. Which element would you expect to have the highest IE_1? Which would have the lowest IE_1?

d. Which element would you expect to have the highest IE_2?

e. Which of the elements is most likely to form a 1^+ ion?

Periodic Trends - Electron Affinity (A_e)

- the energy change when a gaseous atom accepts an electron.
- $A + e^- \rightarrow A^{1-}$ + energy (anion formation)
  - usually a negative number (release of energy)
  - if number is zero or positive, energy needed to add electron (ion unstable)

Trend

- energy released (increase in electron affinity) increases across a period
  - Why? Shielding of nucleus by inner electrons decreases allowing for a stronger attraction between added electron and nucleus. Stronger attraction , more energy released. Atoms closer to an octet.
    - An increase in coulombic attraction due to decreased distance allows the added electron to "drop" to a lower energy level releasing more energy.
- decreases down a group
  - Why? Greater shielding of nuclear charge (positive charge) by the inner electrons.
    - A decrease in coulombic attraction due to increased distance. The drop to a lower energy level is not as great due the lower columbic force on the added electron.

Periodic Trends - Electron Affinity (A_e)

Periodicity
Topic#6

Electron affinities in the periodic table [hide]

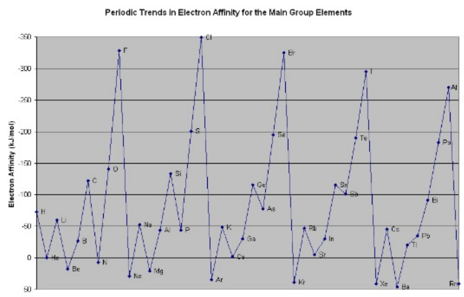
Group → / ↓ Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	H 73																	He *
2	Li 60	Be *											B 27	C 122	N *	O 141	F 328	Ne *
3	Na 53	Mg *											Al 42	Si 134	P 72	S 200	Cl *349*	Ar *
4	K 48	Ca 2	Sc 18	Ti 8	V 51	Cr 65	Mn *	Fe 15	Co 64	Ni 112	Cu 119	Zn *	Ga 41	Ge 119	As 79	Se 195	Br 324	Kr *
5	Rb 47	Sr 5	Y 30	Zr 41	Nb 86	Mo 72	Tc *	Ru 101	Rh 110	Pd 54	Ag 126	Cd *	In 39	Sn 107	Sb 101	Te 190	I *295*	Xe *
6	Cs 46	Ba 14	*	Hf	Ta 31	W 79	Re *	Os 104	Ir 150	Pt 205	Au 223	Hg *	Tl 36	Pb 35	Bi 91	Po	At	Rn *
7	Fr	Ra	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo
* Lanthanides				La 45	Ce 92	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm 99	Yb	Lu 33
** Actinides				Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Legend

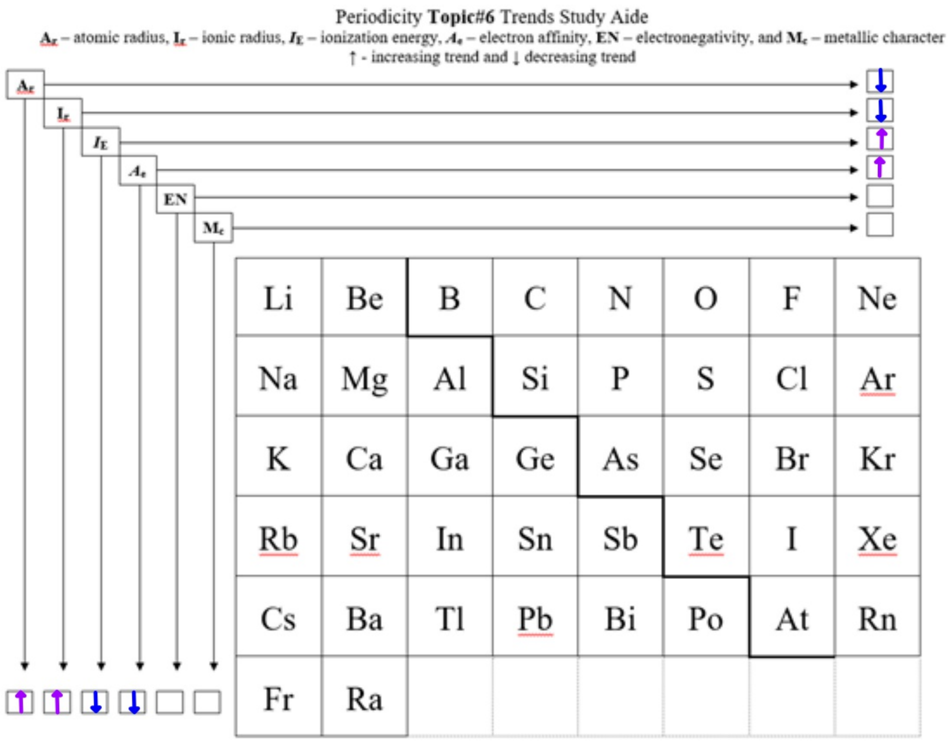
The number mentioned is **Electron affinity** in kJ/mol (rounded)

* Denotes elements that are expected to have electron affinities close to zero on quantum mechanical grounds

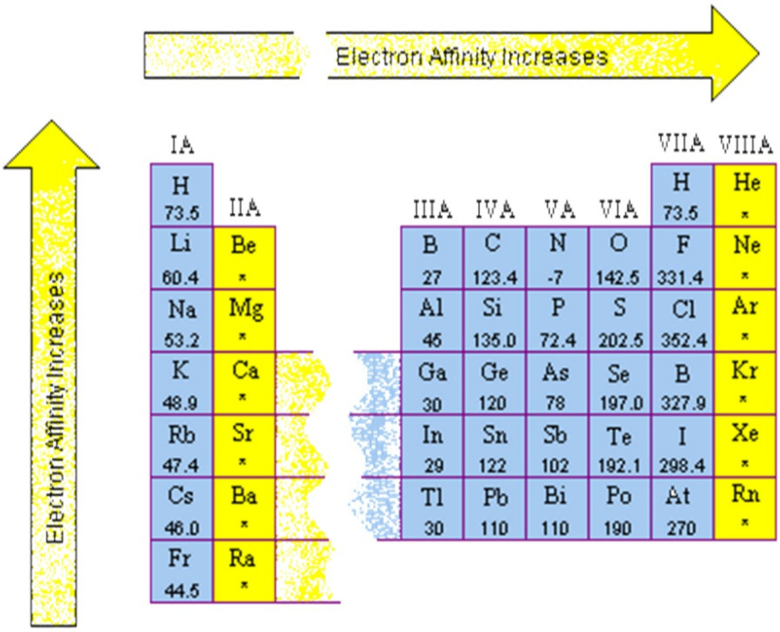
For the equivalent value in eV, see: Electron affinity (data page)



Periodic Trends - Electron Affinity (A_e)



Periodic Trends - Electron Affinity (A_e)



Periodicity
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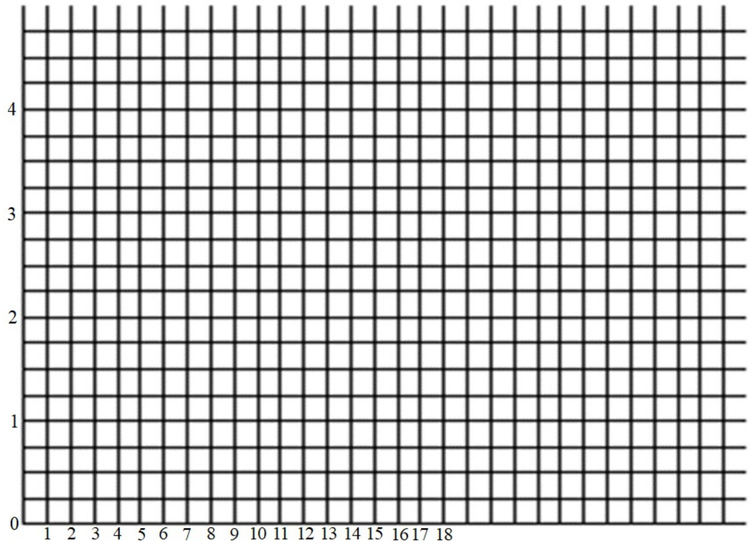
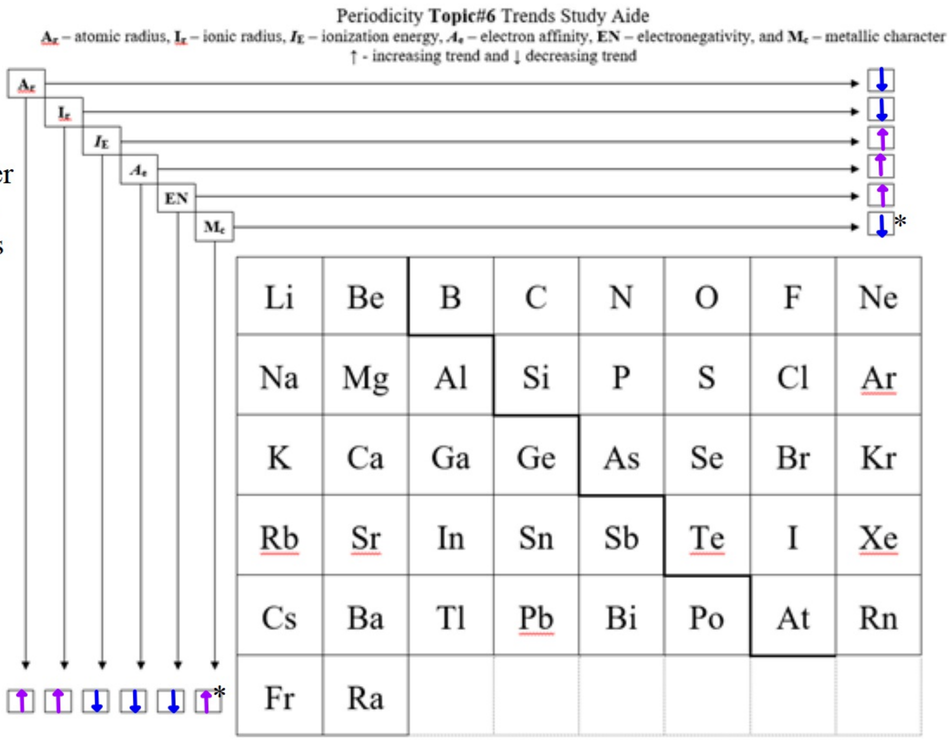
Periodic Trends - Electronegativity (EN)

Element	EN	Element	EN
Li	1.0	Na	0.9
Be	1.5	Mg	1.2
B	2.0	Al	1.5
C	2.5	Si	1.8
N	3.0	P	2.1
O	3.5	S	2.5
F	4.0	Cl	3.0



Periodic Trends - Electronegativity (EN)

*Metallic Character - increases down a group and deceases across a period



Periodic Trends - Electronegativity (EN)

measures the ability of an atom to hold on to its electron involved in a bond.

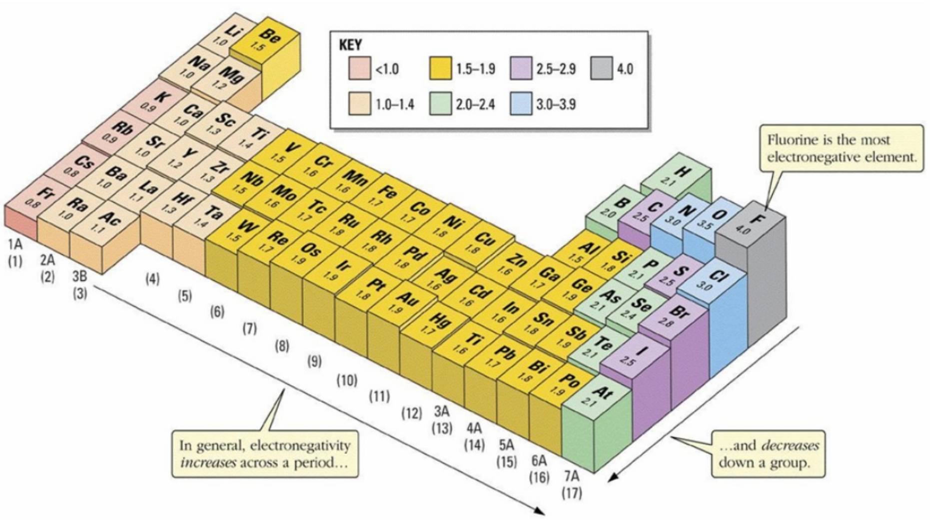
Trend

- increases across a period
  - Why? Increased proton number in nucleus increases the ability to hold on to its own electron in a bond.
- group - decreases down
  - Why? Each additional energy level adds another set of sheilding electrons, so the hold on the outer electron involved in a bond decreases.

H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Ti 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	Ac 1.1	Unq	Unp	Unh	Uns	Uno	Une									

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.2
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

Periodic Trends - Electronegativity (EN)



Electronegativity

5. Among the elements Ga, Br, and Ca, which has the highest electronegativity? Explain why in terms of periodic trends.

Periodic Trends - Electronegativity (EN)

Periodicity
Topic#6

H 2.1																	He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Ti 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra 0.7	Ac 1.1	Unq	Unp	Unh	Uns	Uno	Une									

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.2
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

6. Consider the five hypothetical main-group elements E, G, J, L, and M with the outer electron configurations shown below

$E = 2s^2 2p^5$ $G = 4d^{10} 5s^2 5p^5$ $J = 2s^2 2p^3$ $L = 5d^{10} 6s^2 6p^5$ $M = 2s^2 2p^4$

a. Identify the block location of each element. Then determine which elements are in the same period and which are in the same group.

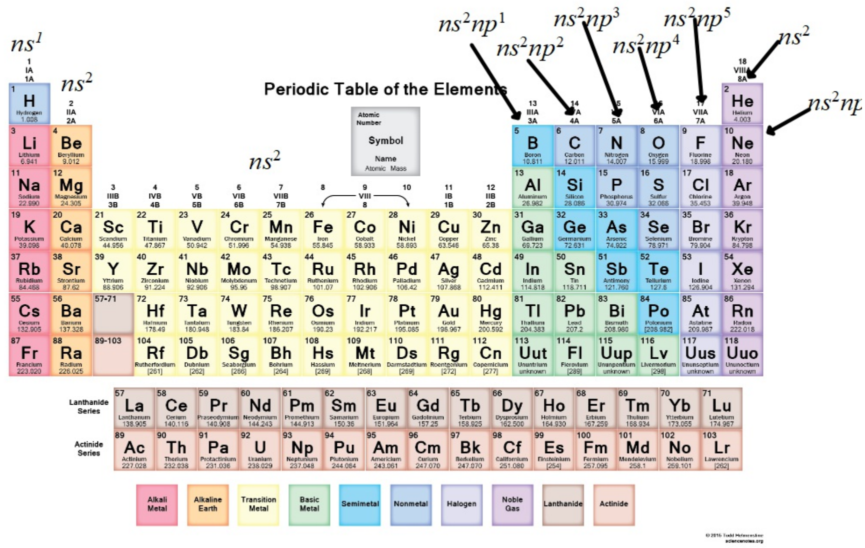
b. Which element would you expect to have the highest electron affinity? To form a 1- ion? Highest electronegativity?

c. Compare the ionic radius of the typical ion formed by the element G with the radius of its neutral atom.

d. Which element(s) contain seven valence electrons?

Valence Electrons

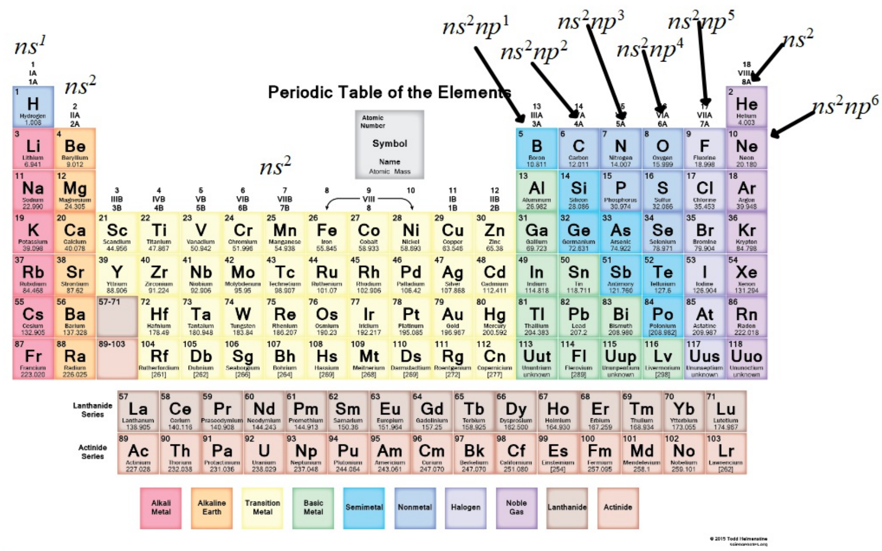
- outer electrons located in ONLY the s and p type orbitals.
- For example,
  - Mg has 2 valence elcectrons, $3s^2$
  - Thallium (Tl) has 3 valence electrons, $6s^26p^1$
- To determine the valance electrons oi an element, go to the row it is in, this i: n, then go to its group, this is the sum of the powers on the s and p. s can onl have a maximum of 2 and p has a maximum of 6.
  - For elements in the d-block, theii valence electrons are ALWAYS ns^2.
    - For example, zirconium (Zr) has 2 valence electrons, $5s^2$.
- Elements in group 18 (except He) have eight valence electrons (ns^2np^6). This i: called the <u>octet rule</u>. This FULL outei energy level lends itself to stability.



Periodicity
Topic#6

Electron Dot Diagrams

- A diagram using the symbol of an atom and its valence electrons (outer *s* and *p* electrons)
- For example, Be (beryllium(has 2 valence electrons ($2s^2$), so its electron dot diagram is Be:



Valence electrons/Electron Dot Diagram

7. Determine the valence electrons and electron dot diagram for the following atoms:

a. K(Group ___)   e. Si (Group ___)

b. In (Group ___)   f. Os (Group ___)

c. Te (Group ___)   g. Xe (Group ___)

d. I (Group ___)