Chapter 8 - Electron Configurations and Periodicity

Section 3 - Periodic Properties

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Introduction – Periodic Properties

Periodicity is being able to predict some trends in the periodic table. While designing the periodic table Mendeleev noticed that groups had similar chemical properties. Initially the periodic table had been arranged by atomic weight but was later changed by Mendeleev to atomic numbers. This was done because of the observation about the chemical properties.

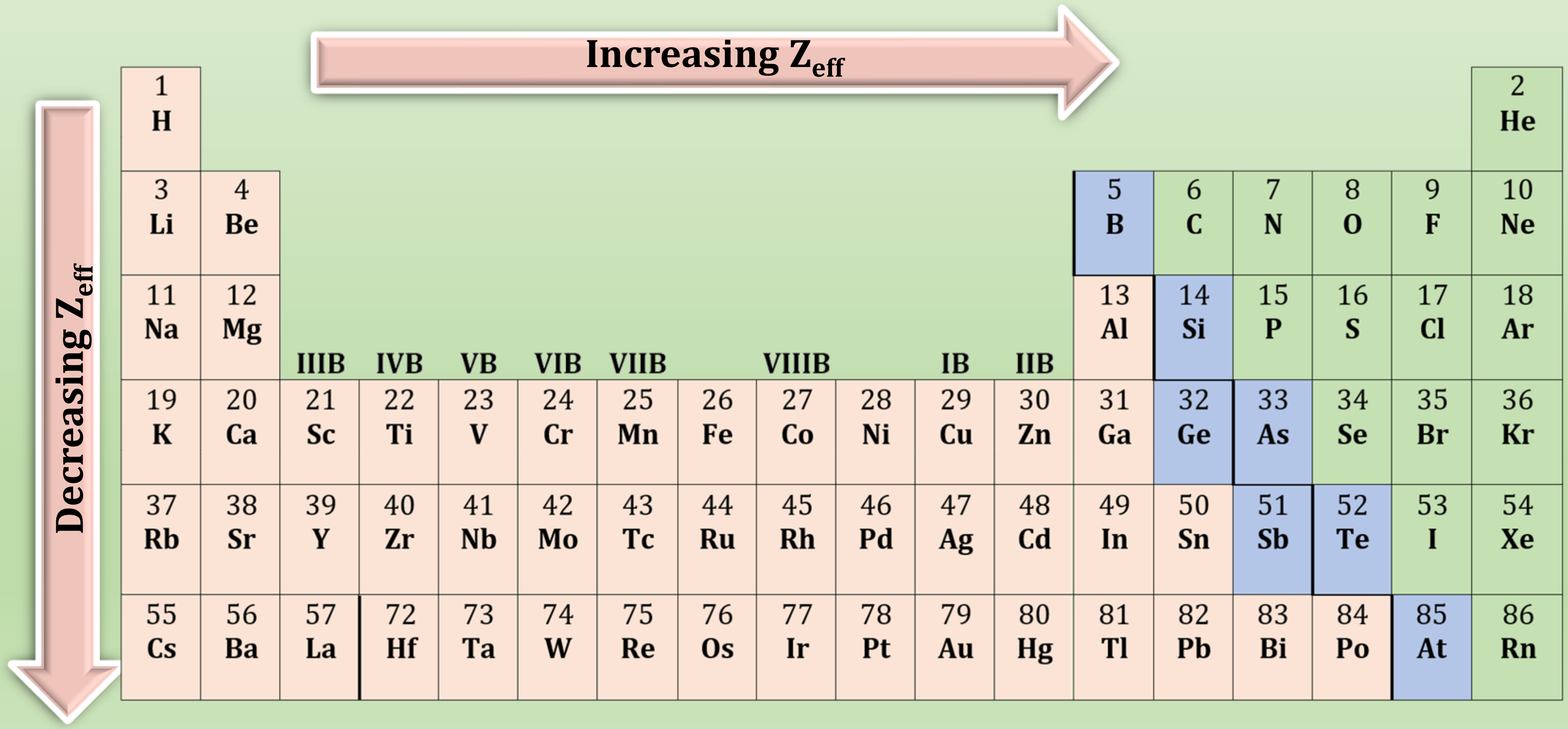
The periodic table, as designed, has a number of trends in the rows and groups. The main trends we will study here are: Atomic radii, Ionization energy and Electron affinity. We will also see how metallic nature of elements changes in the periodic table.

One thing to keep in mind is that all these trends are for the main group elements, groups I – VIII, and does not include transition metals, lanthanides and actinides.

Effective Nuclear Charge (Z_{eff})

All the periodic trends are dependent on Effective Nuclear Charge, which also has a trend in the periodic table.

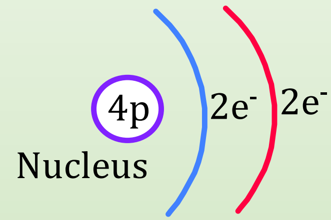
Effective nuclear charge is the attraction of valence electrons (VE) to the positive nucleus. The trend then is that effective nuclear charge increases across the row and decreases going down a group.



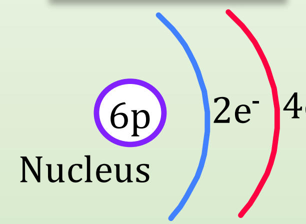
Effective Nuclear Charge (Z_{eff}) - Explanation

Trend in the row: As electrons are added to the same shell the Z_{eff} increases because the number of electrons added are in the same valence shell and the number of protons are also increasing, thus the positive-negative attraction increases. As in the example shown on the right, Be has 2 VE with 4 protons, whereas C has 4 VE with 6 protons, thus attraction of 6p to 4 VE of C is much more than of 4p to 2 VE of Be.

Higher Z_{eff}



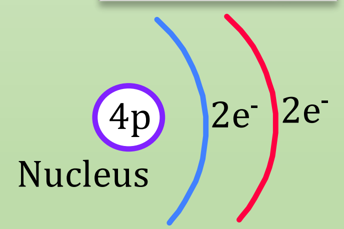
Be – $1s^2$ $\underline{2s^2}$



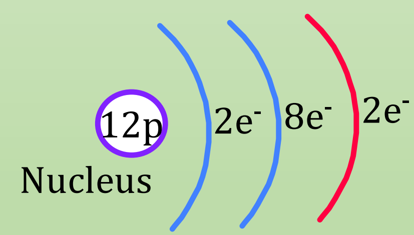
C $1s^2$ $\underline{2s^2\ 2p^2}$

Trend in a group: As electrons are added to a shell farther away from the nucleus the Z_{eff} decreases. In the example on the right, in Mg, the valence shell is farther away from the nucleus thus the VE to proton attraction is not as much as in Be.

Higher Z_{eff}



Be – $1s^2$ $\underline{2s^2}$



Mg $1s^2$ $2s^2$ $2p^6$ $\underline{3s^2}$

The Trends: Atomic Radii; Ionization Energy; Electron Affinity; Metallic Character

- Atomic radius is half the distance between two nuclei next to each other .
- Ionization Energy is the amount of energy required to remove one electron from an atom in a gas phase.
- Electron Affinity is the measure of energy change when an electron is added to a valence shell. A negative energy measurement means an anion is formed and positive energy measurement means cation is formed.
- Metallic Character is how much a metal behaves like a metal i.e., conducts electricity and heat, malleability, reactivity with acids etc.

Atomic Radii

Atomic radius is half the distance between two nuclei next to each other.

The trend is that atomic radii decreases in a row of the periodic table and increases down a group.

Why does this happen? As we go across a row, effective nuclear charge increases, thus the attraction of VE to protons/nucleus increases making the atomic radii smaller.

When going down a group, the shell number of the VE increases making the attraction between the VE and protons/nucleus much smaller and thus making the atom larger.

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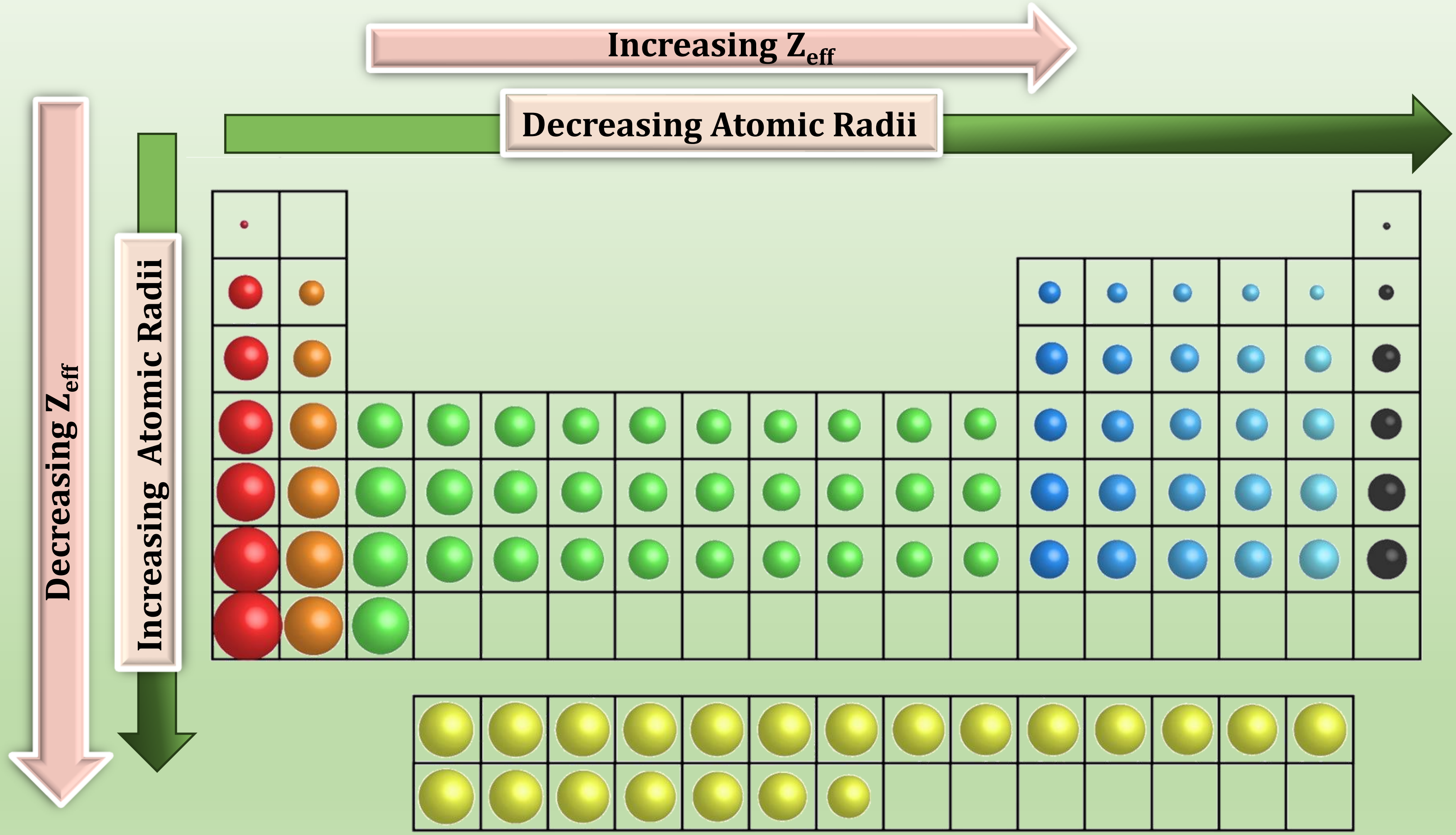
Atomic Radii



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Solved Problem: Arranging in atomic radius

Refer to a periodic table and arrange the following elements in order of increasing atomic radius: Br, Se, Te.

34 Se	35 Br
52 Te	

Te is larger than Se.
Se is larger than Br.

Atomic radius:
Br < Se < Te

Solved Problem: Arranging in atomic radius

Refer to a periodic table and arrange the following elements in order of increasing atomic radius: As, Br, Sb.

33 As		35 Br
51 Sb		

Sb is larger than As.
As is larger than Br.

Atomic size:
Sb > As > Br

Ionization Energy

Ionization Energy (IE) is the amount of energy required to remove one electron from an atom in a gas phase.

The trend is that IE increases in a row of the periodic table and decreases down a group.

Why does this happen? As we go across a row, effective nuclear charge increases, thus the attraction of VE to protons/nucleus increases, making it hard to remove the valence electron.

When going down a group, the VE is in a shell farther away from the nucleus making the attraction between the VE and protons/nucleus much smaller and thus making it easier to remove an electron from the valence shell.

Ionization Energy

Increasing Z_{eff}

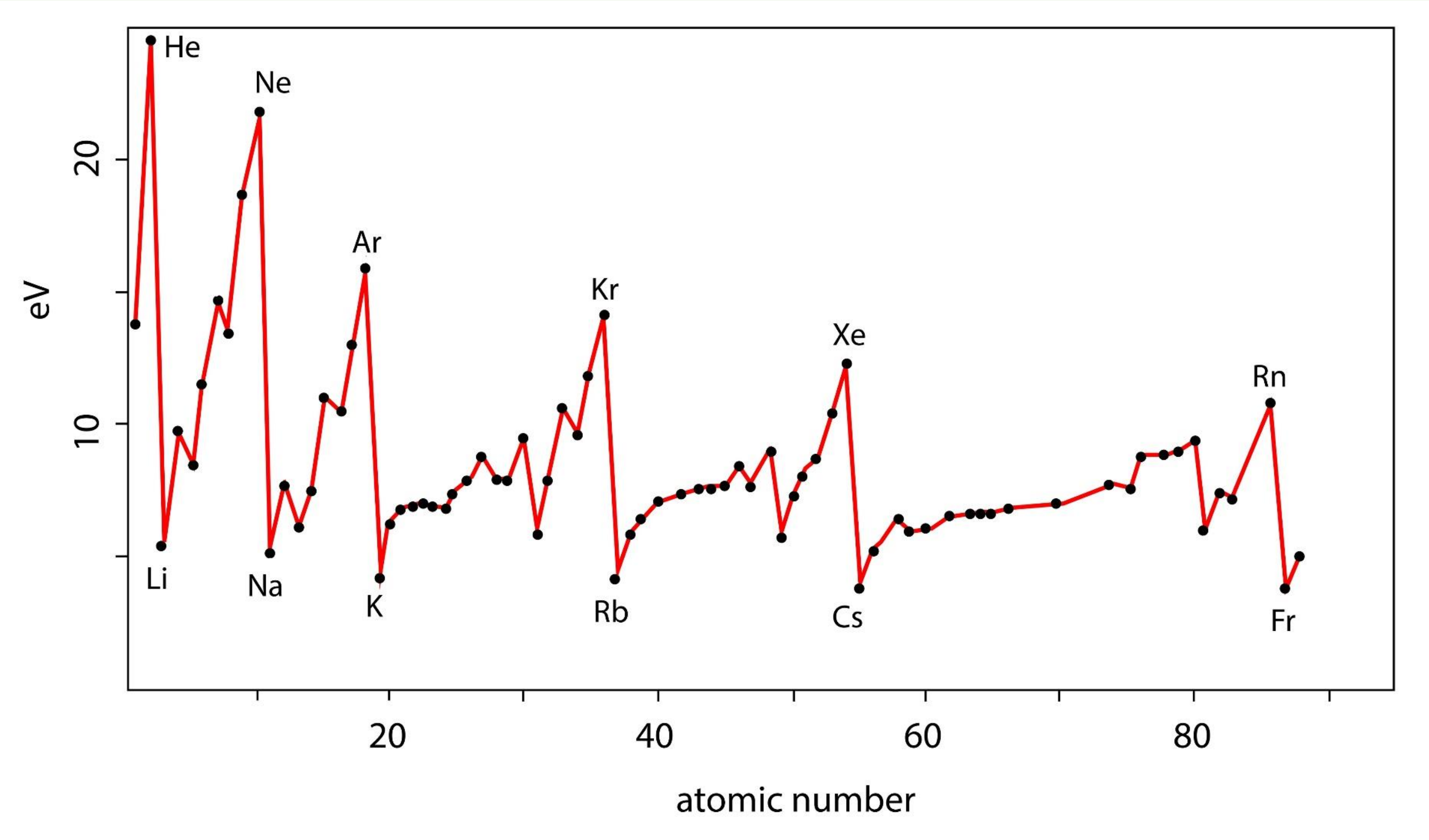
Increasing Ionization Energy

Decreasing Z_{eff}

Decreasing Ionization Energy

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	IIIB	IVB	VB	VIB	VIIB		VIIIB		IB	IIB	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	57 La	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

Ionization Energy



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Ionization Energy – 2nd, 3rd and more

Once the 1st VE is removed, it becomes harder to remove the next electron (this would be the 2nd IE), as the Z_{eff} is much larger because the number of protons attracting the remaining number of electrons is much higher. The 3rd IE is even larger. One can see from the data below that once the entire valence shell is removed, it is extremely hard to remove from the remaining cation, e.g., see the high IE values of the 2nd IE of Li and the 4th IE of B.

Element	First	Second	Third	Fourth	Fifth	Sixth	Seventh
H	1312						
He	2372	5250					
Li	520	7298	11,815				
Be	900	1757	14,848	21,006			
B	801	2427	3660	25,026	32,827		
C	1086	2353	4620	6223	37,831	47,277	
N	1402	2856	4578	7475	9445	53,326	64,360
O	1314	3388	5300	7469	10,990	13,326	71,330
F	1681	3374	6050	8408	11,023	15,164	17,868
Ne	2081	3952	6122	9371	12,177	15,238	19,999

Solved Problem: Arranging in ionization energy

Refer to a periodic table and arrange the following elements in order of increasing ionization energy: Br, Se, Te.

34 Se	35 Br
52 Te	

Te has lower IE than Se.
Se has lower IE than Br.

Ionization energies :
Te < Se < Br

Solved Problem: Arranging in ionization energy

Refer to a periodic table and arrange the following elements in order of increasing ionization energy: As, Kr, K.

19 K		
33 As		36 Kr

K is lower IE than As and Kr.
As lower IE than Kr.

Ionization energies:
K < As < Kr

Electron Affinity

Electron Affinity (EA) is the measure of energy released when an electron is added to a valence shell. A higher energy measurement means more energy is released and lower energy measurement means less energy is released.

Increasing Electron Affinity

Decreasing Electron Affinity

	IA	IIA	IIIA	IVA	VA	VIA	VIIA	VIIIA
Row 1	H							He
	73							**≤ 0**
Row 2	Li	Be	B	C	N	O	F	Ne
	60	**≤ 0**	**27**	**122**	**≤ 0**	**141**	**328**	**≤ 0**
Row 3	Na	Mg	Al	Si	P	S	Cl	Ar
	53	**≤ 0**	**44**	**134**	**72**	**200**	**349**	**≤ 0**
Row 4	K	Ca	Ga	Ge	As	Se	Br	Kr
	48	**2**	**41**	**119**	**78**	**195**	**325**	**≤ 0**
Row 5	Rb	Sr	In	Sn	Sb	Te	I	Xe
	47	**5**	**37**	**107**	**101**	**190**	**295**	**≤ 0**
Row 6	Cs	Ba	Tl	Pb	Bi	Po	At	Rn
	46	**14**	**36**	**35**	**91**	**180**	**270**	**≤ 0**

Metallic Character

Metallic character is when elements follow a lot of the characteristics given in the first column in the table below.

The trend is that metallic character decreases in a row of the periodic table and increases down a group. Some of it has to do with IE and EA trends.

Metallic Character	Nonmetals	Metalloids
•Shiny, lustrous, malleable •Good conductors •Low *IE* (form cations) •Form ionic compounds with chlorine •Form basic, ionic compounds with oxygen	•Vary in color, not shiny •Brittle •Poor conductors •Form acidic, molecular compounds with oxygen •High *EA* (form anions) •Group VII and VIII are all non metals	•Properties both of metals and nonmetals

Solved Problem: Arranging in electron affinity

Refer to a periodic table and arrange the following elements in order of increasing electron affinity: Br, Se, Te.

34 Se	35 Br
52 Te	

Br more than Se.
Te less than Se.

Electron Affinity:
Te < Se < Br

Solved Problem: Arranging in metallic character

Refer to a periodic table and arrange the following elements in order of increasing metallic character: As, Br, Sb.

33 As		35 Br
51 Sb		

Sb is more than As.
As is more than Br.

Ionization energies:
Br < As < Sb

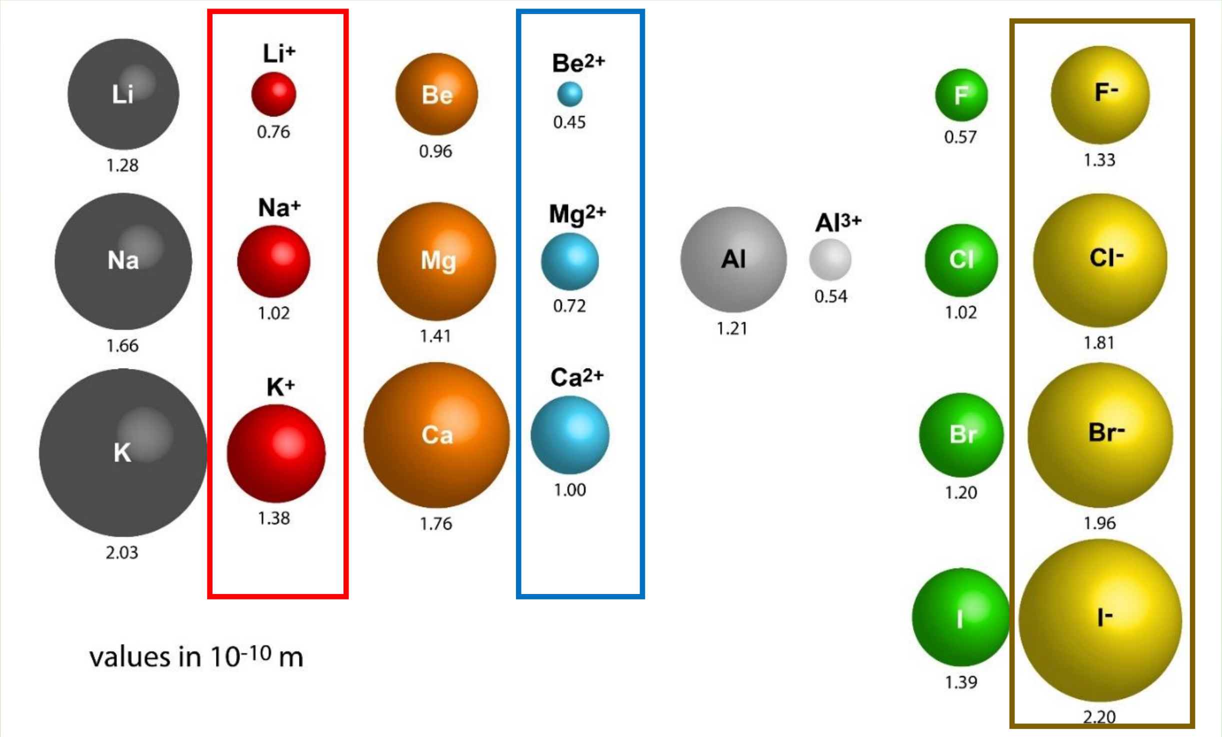
Trends

Trend	Definition	Across a Group	Down a Group
Effective nuclear charge	This is the attraction of negative electrons to the positive nucleus	**Increases** Because electrons are added to the same shell	**Decreases** Because electrons are added to a new shell
Atomic radii	Atomic radius is half the distance between two nuclei	**Decreases** Because electrons are added to the same shell	**Increases** Electrons are added to a new shell
Ionization energy	IE is the amount of energy required to eject one electron from an atom in a gas phase	**Increases** Z increases, so harder to remove an electron	**Decreases** Z decreases so easy to remove an electron
Electron affinity	EA is the measure of energy released when an electron is added to a valence shell	**Increases** easier to gain electrons	**Decreases** harder to add electrons
Metallic Character	how much a metal behaves like a metal e.g. conducts electricity and heat, malleability etc.	**Decreases** Z increases so harder to lose electrons	**Increases** Z decreases so easier to lose electrons

Ionic Radii

Cations have lost electrons hence the number of protons are more than electrons, Z_{eff} is higher and thus the cationic radii gets smaller

Anions have accepted electrons hence the number of protons is less than electrons, Z_{eff} is lower and thus the anionic radii gets larger



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Solved Problem: Arranging in ionic radius

Refer to a periodic table and arrange the following elements in order of increasing ionic radius: Br^-, Se^{2-}, Te^{2-}.

34 Se^{2-}	35 Br^-
52 Te^{2-}	

Te^{2-} is larger than Se^{2-}: because it is below Se in the PT.
Se^{2-} is larger than Br^-: because it is in the same row and accepting more electrons than Br.

Ionic radius:
$Br^- < Se^{2-} < Te^{2-}$

Solved Problem: Arranging in ionic radius

Refer to a periodic table and arrange the following elements in order of increasing ionic radii: K^+, Cl^-, P^{3-}, Ca^{2+}.

15 P^{3-}	17 Cl^-
19 K^+	20 Ca^{2+}

All these elements are isolectric. The best way to figure out size is by their effective nuclear charge – i.e. the number of protons and electrons. If protons are more then Z_{eff} is more and vice versa.
K^+ larger than Ca^{2+} i.e. Z_{eff} is more for Ca^{2+} because Ca^{2+} has 18 e^- and 20 p while K^+ has 18 e^- and 19 p
P^{3-} larger than Cl^-
K^+ smaller than Cl^-

Ionic radii:
$Ca^{2+} < K^+ < Cl^- < P^{3-}$

Decreasing Atomic Radii

Increasing Ionization Energy

Increasing Electron Affinity

Decreasing Metallic Character

Increasing Z_{eff}

Increasing Atomic Radii

Decreasing Ionization Energy

Decreasing Electron Affinity

Increasing Metallic Character

Decreasing Z_{eff}

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**
		IIIB	**IVB**	**VB**	**VIB**	**VIIB**		**VIIIB**		**IB**	**IIB**						
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**	57 **La**	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**

Acidic and Basic Properties of the Groups

- Groups I and II – all metals
- Groups VII and VIII – all nonmetals
- Transition metals – all metals
- Groups III, IV, V and VI – has nonmetals, metalloids and metals.

- Oxides: dissolving oxides in water give bases and acids.
  - Metal oxides give bases
  - Nonmetal oxides give acids
  - Amphoteric oxides can give both acids and bases

Reactions of Oxides

- Metal oxides will produce bases on reaction with water.

$$Na_2O_{(s)} + H_2O_{(l)} \longrightarrow 2NaOH_{(aq)}$$

- Nonmetal oxides will produce acids when dissolved in water.

$$SO_{3(g)} + H_2O_{(l)} \longrightarrow 2H_2SO_{4(aq)}$$

- Amphoteric oxides are located at intermediate positions on the periodic table and can react with both acids and bases.

$$Al_2O_{3(s)} + 6HCl_{(aq)} \longrightarrow 2AlCl_{3(aq)} + 3H_2O_{(l)}$$

$$Al_2O_{3(s)} + 2NaOH_{(aq)} + 3H_2O_{(l)} \longrightarrow 2NaAl(OH)_{4(aq)}$$

Solved Problem: Identifying acidic or basic oxides/elements

Determine if the following elements or oxides will be acidic or basic when dissolved in water.

a) Al

b) NO_2

c) Fe_2O_3

d) Ca

e) C

a) Al metal – should form base in water
b) NO_2 nonmetal – should form acid in water
c) Fe_2O_3 metal – should form base in water
d) Ca metal – should form base in water
e) C nonmetal – should form acid in water

Group Properties

Group I Alkali Metals

3	Li	6.941
11	Na	22.98
19	K	39.1
37	Rb	85.47
55	Cs	132.9
87	Fr	223

Group II Alkaline Earth Metals

4	Be	9.012
12	Mg	24.31
20	Ca	40.08
38	Sr	87.62
56	Ba	137.3
88	Ra	226

Group III

5	B	10.81
13	Al	26.98
31	Ga	69.72
49	In	114.8
81	Tl	204.4
113	Nh	[266]

Group 1

- All elements are metals.
- Never found in nature in elemental state
- All metals very reactive – have to be stored in special containers.
- All metals are very soft (unlike typical metals)

Group II

- All elements are metals.
- Less reactive than group I.
- Some react with acids to form hydrogen gas.

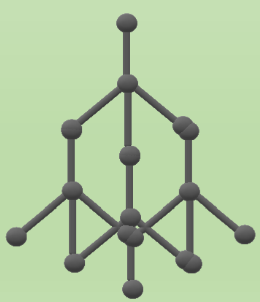
Group III

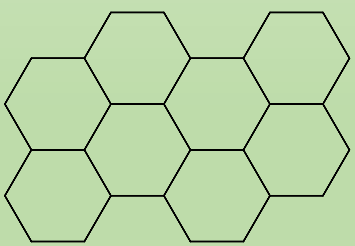
- Metalloid (B) and metals (all others)
- Al forms Al_2O_3 with oxygen and is an amphoteric oxide

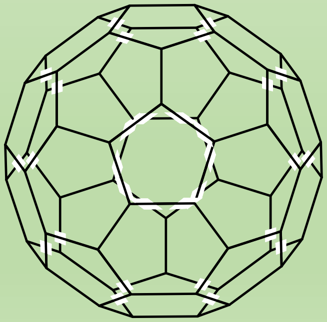
Group IV Properties

Group IV		
6	C	12.01
14	Si	28.09
32	Ge	72.63
50	Sn	118.7
82	Pb	207.2
114	Fl	[289]

- Most interesting group since it has two main elements.
- One that makes organic substances – carbon.
- And one that makes all electronics – silicon.
- This group has nonmetal (C) metalloids (Si, Ge) and metals.
- This group can have acidic oxides (C), amphoteric oxides (Si, Ge) and basic oxides (Sn, Pb).
- Carbon has three main allotropes (same element with different molecular structure): diamond, graphite and buckminsterfullerene. Graphite is a soft amorphous solid, diamond is the hardest substance on earth and buckminsterfullerene is literally out of this world.







Group Properties

Group V	Group VI	Group VII Halogens
7 N 14.01	8 O 15.99	9 F 18.99
15 P 30.97	16 S 32.07	17 Cl 35.45
33 As 74.92	34 Se 78.96	35 Br 79.9
51 Sb 121.8	52 Te 127.6	53 I 126.9
83 Bi 209	84 Po [209]	85 At [210]
115 Mc [289]	116 Lv [293]	117 Ts [294]

Group V

- Nonmetal (N_2, P) metalloid (As,Sb) and metal (Bi).
- Nitrogen, N_2 forms variety of oxides.
- As, Sb, Bi (crystalline).
- Has acidic oxides N, P, As), amphoteric oxides (Sb) and basic oxides (Bi).

Group VI

- Nonmetals (O, S, Se).
- Metalloids (Te, Po).
- Oxides form acids; SO_2, SO_3, H_2S, H_2SO_4.

Group VII

- All nonmetals.
- All elements are diatomic molecules.
- Physical property trend is gas (F_2 and Cl_2), liquid (Br_2), solid (I_2).
- Halogens form acids (HX).

Group VIII Properties

Group VIII Noble Gases

2	He	4.003
10	Ne	20.18
18	Ar	39.95
36	Kr	83.8
54	Xe	131.3
86	Rn	[222]
118	Og	[294]

- All monatomic
- Filled valence shells
- All are gases
- Considered “inert” until 1963 when Xe and Kr were used to form compounds
- No major commercial use

Key Words

- Effective nuclear charge
- Atomic radii
- Ionization energy
- Electron affinity
- Metallic character
- Ionic radii
- Isoelectric elements/ions
- Basic oxide
- Acidic oxide