TOPIC : PERIODIC TABLE & PERIODICITY
EXERCISE # 1

SECTION (A)

1. It is fact.

3. The d-block has 10 columns, because a maximum of 10 electrons can occupy all the orbitals (5) in a d-subshell.

6. Cu, Ag, Au group of element are called coinage metals as these are used in minting coins.

7. For d-block elements group number = number of electrons in (n – 1)d sub-shell + number of electrons in valence shell. So, group number = $(n-1)d^{10} + ns^{1}$

8. All belongs to d-block as differentiating electrons enter in d-subshell.

10. $[Kr]^{36}\ 5s^{1}$
Last electron enters in s-subshell, so it belongs to s-block.
For s-block group number = number of valence electron(s).

12. Yes, hydrogen can be placed in 1st group on the basis of its valency +1 (H^{+}).

13. Z = 118 $[Rn]^{86}\ 5f^{14}\ 6d^{10}\ 7s^{2}\ 7p^{6}$; as last electron enters in p-subshell, it belongs to p-block. Thus its group number will be 10 + 2 + 6 = 18. Hence the element is a noble gas.

SECTION (B)

2. For isoelectronic species, as Z increases, Z_{eff} increases (and vice versa).

3. For isoelectronic species, as Z increases, Z_{eff} increases (and vice versa).

SECTION (C)

2. Ne is bigger than oxygen due to interelectronic repulsion on account of completely filled 2p subshell.

3. Across the period size decreases and down the group size increases. So, Cs is largest ionic radius.

5. All are isoelectronic species and thus Na^{+} has smallest ionic radius because of high effective nuclear charge (i.e., 11 No. of protons as compared to 9, 8 and 7 in F, O and N respectively).

7. O^{+} is smaller than parent atom while anion is bigger than parent atom. O^{2-} and N^{3-} are isoelectronic species. So ionic size $\propto \frac{1}{\text{nuclear charge}}$. Hence the correct order is $O^{+} < O^{2-} < N^{3-}$.

8. Down the group the size of atomic increases as new atomic orbits are added with increasing atomic number.

10. Atomic radii of zero group elements are expressed as their vander Waal's radii.
$r_{\text{van der Waal's}} > r_{\text{covalent}}$.

11. On moving left to right in a period, atomic radii decreases due to increase in Z_{eff} and addition of electrons to the same outermost shell.

12. Among isoelectronic species, size decreases with increases in nuclear charge.

	Mg^{2+}	Na^{+}	Ne	O^{2-}
Number of elecron	10	10	10	10
Number of charge	+12	+11	+10	+8

13. Atomic radius increases on moving top to bottom in a group due to increasing number of shells. However, it decreasing on moving left to right in a period due to increasing Z_{eff} and addition of electrons to the same shell.
Nb (4d) $\approx$ Ta (5d) (due to poor shielding of nuclear charge by 4f electrons).

For isoelectronic species, ionic radius $\propto \frac{1}{\text{nuclear charge}}$. So correct order is $Y^{3+} < Sr^{2+} < Rb^{+}$.

SECTION (D)

1. Orbitals bearing lower value of n will be more closer to the nucleus and thus electrons will experience greater attraction from nucleus and so its removal will be difficult, not easier.

2. Due to stable half filled electronic configuration of outer most shell of N, it has higher ionisation energy than O which has partially filled electron configuration of outer most shell.

3. The IP decreases in a group on moving downward because atomic radius increases hence the correct order is : Be > Mg > Ca

4. Across the period (i.e. 3rd period) the size of atom decreases and nuclear charge increases. So generally the ionisation energy increases. However the ionisation energy of Mg is greater than Al because of more penetration power of 2s sub-shell electrons of Mg as compared to that of the 2p sub-shell electron of Al. Also, Mg has fully filled configuration.

5. Half filled electron configuration has extra stability. Hence the removal of electron from outer most orbit requires higher energy.

6. Completely filled electron configurations and half filled electron configurations are expected to have higher ionisation energies. $ns^2\ np^5$ will have higher first ionisation energy than $ns^2\ np^4$ on account of smaller size of atom and higher nuclear charge.

9. The ionisation energy of Tin (Sn) is less than that of lead (Pb) It is due to the poor sheilding of d- and f-electron in Pb, due to which it feels greater attraction from nucleus.

10. As elements are ionized of the proton to electron ratio increases, so the attraction between valence shell electron and nucleus increases and as a result the size decreases. Therefore, the removal of electron from smaller cation requires higher energy. Hence the second ionisation energy is higher than its first ionisation energy.

11. (1) Down the group size increases and therefore, ionization energies decrease. Hence the order is correct.

(2) Cation is smaller but anion is bigger than its parent atom. As charge on cation increases the ionic radius decreases. Similarly as charge on anion increases the ionic radius increases. $IE \propto \frac{1}{\text{atomic / ionic radius}}$ and $IE_1 < IE_2 < IE_3$. Hence the order is correct.

(3) N has stable half filled configuration thus has higher ionization energy than O. Hence the order is correct.

(4) The correct order is Mg > Al < Si. Mg (ns^2) has higher ionization energy than Al ($ns^2 np^1$) because s-sub shell electrons have higher penetration power than that of p-sub shell electrons. Further across the period size decreases and nuclear charge increases and therefore, ionization energy increases.

12. Removal of Ist electron is easier because of bigger size but 2nd electron is to be removed from $ns^2\ np^6$ configuration i.e. stable noble gas configuration. So $IE_2 >> IE_1$.

13. Second ionisation energy of potassium is greater than that of Ca. In case of potassium ion (i.e. K^+) the electron removal from the stable inert gas configuration ($1s^2\ 2s^2\ 2p^6 3s^2 3p^6$) requires much higher energy.

14. Since, K contains only one electron in its outermost shell, it has higher tendency to donate it, ie, it has higher tendency to get oxidised. Therefore, it is the strongest reducing agent among the given elements.

SECTION (E)

4. Electron affinity is the measure of the ease with which an atom receives the additional electron in its valence shell in gaseous phase.
Generally down the group, the electron affinity decreases due to increase in atomic size.

6. In chlorine, the addition of additional electron to larger 3p-subshell experiences less electron-electron repulsion than smaller 2p-subshell of fluorine. Phosphorus has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.

9. In chlorine, the addition of additional electron to larger 3p-subshell experiences less electron-electron repulsion than smaller 2p-subshell of fluorine. Phosphorus has very low electron affinity because there is high electron repulsion when the incoming electron enters an orbital that is already half filled.

10. O has exceptionally smaller value of electron affinity (minimum in family) due to smaller atomic size than sulphur (weaker electron-electron repulsion in larger 3p-subshell).

SECTION (F)

1. According to Mulliken's, the electronegativity = $\frac{\text{Ionisation energy} + \text{Electron affinity}}{2}$

2. As size of atom decreases across the period, the attraction between the nucleus and shared pair of electrons increases. So electronegativity increases across the period.

3. Eelctronegativity values are as given below N = 3.0; C = 2.5; Si = 1.8; P = 2.1

6. Halogens have valence shell electron configuration ns^2np^5. They have highest electronegative values in their respective period.

7. The addition of extra electron is difficult to the atom having stable configuration and so electron gain enthalpy will be positive. Similarly the removal of electron is quite difficult from stable configuration and so ionisation enthalpy is higher. However EN remains unaffected because it neither involves gain nor loss of electron.

8. As size of atom decreases across the period, the attraction between the nucleus and shared pair of electrons increases. So electronegativity increases across the period.

9. Electronegativity of elements generally increases across the period (less increase) and decreases down the group (more decrease).
Si = 1.8, P = 2.1, C = 2.5, N = 3.0. So, the correct increasing order is Si < P < C < N.

10. Electronegativity of elements generally increases across the period (less increase) and decreases down the group (more decrease).

11. Non metals are more electronegative than metals.

SECTION (G)

10. BiI_5 does not exists because of I^- being very strong reducing agent. So it reduces Bi^{5+} to Bi^{3+} and forms BiI_3.

EXERCISE # 2

1. For transition elements, the 3d-orbitals are filled with electrons after 4s-orbitals and before 4p-orbitals.

2. Electron configuration is $[Ar]^{18}$ $3d^{10}$ $4s^2$ $4p^3$. As last electron enters in p-subshell it is p-block element and thus its group number is equal to 10 + 5 = 15. As principal quantum number of valence shell is 4, so it is 4th period element.

4. He^+, $1s^1$; Li^{2+}, $1s^1$; Be^{3+}, $1s^1$. All these ions have only one electron. So they do not have any inner orbital and thus do not experience any screening effect.

5. Across the period ionic size decreases as nuclear charge increases for successive addition of an electron but down the group increases due to increase in the number of atomic shells (effective nuclear charge remains nearly same).
O^{2-} = 140 pm, Se^{2-} = 198 pm, F^- = 133 pm, Br^- = 196 pm, I^- = 220 pm.
So, the correct decreasing order of ionic radii. **$I^- > Se^{2-} > Br^- > O^{2-} > F^-$.**

6. K^+ has more number of shells than Mg^{2+} and Al^{3+}. Al^{3+} and Mg^{2+} are isoelectronic but Al^{3+} has higher nuclear charge, so $Al^{3+} < Mg^{2+}$. Mg^{2+} and Li^+ have diagonal relationship. But due to +2 charge in Mg^{2+}, the Mg^{2+} is smaller than Li^+. Hence Al^{3+} is the smallest one.
K^+ = 1.38 Å, Li^{2+} = 0.76 Å, Mg^{2+} = 0.72 Å and Al^{3+} = 0.535 Å.

7. Larger the positive charge on the ion, smaller will be the size of ion for the same element due to increasing Z_{eff}. For isoelectronic species, ionic radius $\propto \frac{1}{\text{Nuclear charge}}$

Ions formed by isotopes have very similar ionic radii due to same number of electrons and same nuclear charge.

8. All are isoelectronic species having 10 electrons in each species but different nuclear charge and thus ionic radius $\propto \frac{1}{\text{Nuclear charge}}$.
So correct order is ${}_9F^- < {}_8O^{2-} < {}_7N^{3-}$.

9. (1) Be has completely filled stable valence shell configuration i.e. $2s^2$ while in Be^+ because of positive charge, the removal of electron requires much higher energy. So, ionisation energy of Be^+ is greater than Be.
(3) Across the period, atomic size decreases and nuclear charge increases and thus valence shell electron(s) is/are tightly held by nucleus. So, ionisation energy of C is greater than Be.

11. IE_2 of Na > Mg as in Na, second electron is to be removed from stable inert gas configuration i.e.,$1s^2 2s^2 2p^6$.

12. There is more interelectronic repulsion in 2p-subshell of fluorine than chlorine (3p). So extra electron will be added easily in 3p-subshell of chlorine as compared to 2p-subshell of fluorine.

13. The tendency to attract bonded pair of electron in case of hybrid orbitals increases with increase in % s-character and so the order : $sp > sp^2 > sp^3$
The electron affinity values for 2p-series elements is less than that for 3p-series elements on account of small size and high inter electronic repulsions. Statements (2) and (3) are facts. Every cation releases more energy than neutral atom upon gain of an electrons.

14. ${}_4Be^-$ - $1s^2$ $2s^2$ $2p^1$ Addition of electron to a completely filled stable configuration, so least stable.

15. Only Na & F will show one non-zero oxidation state. These are Na^+ & F^-.

18. TlI_3 exists as Tl^+ and I_3^- while PbF_4 exists because of F^- being very weak reducing agent.

EXERCISE # 3

PART - I

1. Correct order of electron gain enthalpy is O < S < F < Cl since F and O have more electron electron density with respect to Cl and S.

2. Electronic configuration of ${}_{23}V$ = [Ar] $3d^3 4s^2$
Electronic configuration of ${}_{24}Cr$ = [Ar] $3d^5 4s^1$

Electronic configuration of ${}_{25}Mn$ = [Ar] $3d^5 4s^2$
Electronic configuration of ${}_{26}Fe$ = [Ar] $3d^6 4s^2$
In Mn (Z = 25) third electron will remove from most stable d^5 configuration.
So it's third ionisation enthalpy will be highest.

3. As percentage of s character increases in hybridisation, that atom will be more electronegative.
So electronegativity order of carbon will be in order $sp > sp^2 > sp^3$.

4. As the positive charge increases in isoelectronic species, size decreases.
So correct order is $S^{2-} > Cl^- > Ar > K^+ > Ca^{2+}$

5. $1s^2$, $2s^2\ 2p^5$, $3s^1$ configuration belongs to neon atom in excited state, so It is easy to remove electron from 3s sub shell in comparison to 2p sub shell.

6. It can be seen according to electronic configuration.
electronic configuration of ${}_{22}Ti$ = [Ar] $3d^2\ 4s^2$
electronic configuration of ${}_{23}V$ = [Ar] $3d^3\ 4s^2$
electronic configuration of ${}_{24}Cr$ = [Ar] $3d^5\ 4s^1$
electronic configuration of ${}_{25}Mn$ = [Ar] $3d^5\ 4s^2$
In ${}_{24}Cr$ second electron will remove from $3d^5$ stable configuration.

7. [Ne] $3s^2\ 3p^3$ is a half filled configuration which is more stable and removal of electron from $3p^3$ is more difficult than the $4p^3$.

8. The atomic radii decreases on moving from left to right in a period, thus order of sizes for Cl, P and Mg is
Cl < P < Mg. Down the group size increases. Thus overall order is : Cl < P < Mg < Ca.

9. The order of size of given cations is

$Li^+ < Na^+ < Cs^+$

and the order of size of given anions is

$I^- > F^-$

Thus, when the cation is large and anion is smallest, the cation to anion size ratio is maximum.
Hence, cation to anion size ratio is maximum for CsF.

10. Ti^{3+} (Z = 22)
Ions which have unpaired electrons exhibit colour in solution. Ti^{3+} has an outer electronic configuration of $4s^0\ 3d^1$, i.e., 1 unpaired electron. Thus its solution will be coloured.
$Sc^{3+} \rightarrow d^0$
In case of La^{3+}, $4f^0$ configuration is present and in Lu^{3+}, $4f^{14}$ is present. So, there is no possibility of f-f transition, hence these ions do not appear coloured.

11. Write the electronic configurations of given ions and find the correct answer.
Ni^{3+} (28) = [Ar] $3d^7$
Mn^{3+} (25) = [Ar] $3d^4$
Fe^{3+} (26) = [Ar] $3d^5$
Co^{3+} (27) = [Ar] $3d^6$

12. In general, the atomic and ionic radii increases on moving down a group. But the elements of second series (eg, Zr, Nb, Mo etc.) have the almost same radii as the elements of third transition serries (eg,Hf, Ta, W etc). This is because of lanthanide contraction ie, imperfect sheilding of one 4f - electron by another.

13. Ionic radii $\propto$ charge on anion $\propto \dfrac{1}{\text{charge on cation}}$

During the formation of a cation, the electrons are lost from the outer shell and the remaining electrons experience a great force of attraction by the nucleous, ie, attracted more towards the nucleous. In other words, nucleous hold the remaining electrons more tightly and this results in decreased radii.

However, in case of anion formation, the addition of electron(s) takes place in the same outer shell, thus the hold of nucleous on the electrons of outer shell decreases and this results in increased ionic radii.

Thus, the correct order of ionic radii is $S^{2-} > Cl^{-} > K^{+} > Ca^{2+}$

14. Electron gain enthalpy, generally, increases in a period from left to right and decreases in a group on moving downwards. However, members of III period have samewhat higher electron gain enthalpy as compared to the coressponding members of second period, because of their small size.

O and S belong to VI A (16) group and Cl and F belong to VII A (17) group. Thus, the electron gain enthalpy of Cl and F is higher as compared to O and S.

Cl and F > O and S

Between Cl and F, Cl has higher electron gain enthalpy as in F, the incoming electron experiences a greater force of repulsion because of small size of F atom. Similar is true in case of O and S ie, the electron gain enthalpy of S is higher as compared to O due to its small size. Thus, the correct order of electron gain enthalpy of given elements is

O < S < F< Cl

15. IE_1 of Na = – Electron given enthalpy of Na^{+} = – 5.1 Volt.

16. As the positive charge increases on metal cation, redius decreases.

17. $F^{-} = 133\ P_m$

$O^2 = 140\ P_m$

$Na^{+} = 102\ P_m$

There is no correct option.

18. Poor screening effect of *f*-orbital.

19. In case of isoelectronic species the radius decrease with increase in nuclear charge

	Ca^{2+}	<	K^{+}	<	Ar
electrons	18		18		18
Proton	20		19		18
$\frac{z}{e}$	$\frac{20}{18}=1.11$		$\frac{19}{18}=1.05$		$\frac{18}{18}=1$

Ionic radius $\alpha\ \frac{1}{(Z/e)}$

20. Zr and Hf has nearly same radius due to lanthanide contraction.

21. Incorrect option are 3 & 4

Correct order of increasing I^{st} I.E → B < C < O < N

correct order of increasing electron gain

Enthalpy → I < Br < F < Cl (in magnitude)

Values (in KJ/mol) → 296, 325, 333, 349.

22. Bond dissociation enthalpy

Cl_2	>	Br_2	>	F_2	>	I_2	
242.6		192.8		158.8		151.1	(kJ/mole)

23. $_{63}Eu \rightarrow [Xe]\ 4f^7\ 6s^2$

$_{64}Gd \rightarrow [Xe]\ 4f^7\ 5d^1 6s^2$

$_{65}Tb \rightarrow [Xe]\ 4f^9\ 5d^0 6s^2$

24. The element Z = 114 belongs to carbon family and its general electronic configuration is ,[Rn] $5f^{14}6d^{10}7s^2p^2$

25. $x = 1s^2\ 2s^2\ 2p^3$
so it is nitrogen.
$3\ Mg + N_2 \longrightarrow Mg_3N_2$

26.

$$\underline{\text{Li}\ \boxed{\text{Be} > \text{B}}\quad \text{C}\quad \boxed{\text{N} > \text{O}}\ \text{F}\ \ \text{Ne}}$$

L IP ↑

Correct order of IP
Ne > F > N > O > C > Be > B > Li
So, Answer is (3)

27. $Na_2O \longrightarrow$ Basic
$Al_2O_3 \longrightarrow$ Amphoteric
$N_2O \longrightarrow$ Neutral
$Cl_2O_7 \longrightarrow$ Acidic

PART - II

1. Electron affinity is defined as the amount of energy released when an electron is added to the outermost shell of gaseous, neutral isolated atom. Electron affinity decrease in a group from top to bottom due to increase in size, so the expected order of EA for halogens is F > Cl > Br > I.
But it is not so. It is observed that EA of elements of III period > corresponding II period, this is because elements of second period have small size and greater electron density so the incoming electrons suffer a repulsive force.
Thus, The order of EA is C l> F > Br >I.

2. In a group from left to right atomic number or nuclear charge or effective nuclear charge increase and size decrease, as a result ionisation energy increase. But elements of IIA and VA group on the account of their half filed and completely filled orbitals respectively have stable configuration and so to remove electron from them comparatively larger energy has to be supplied.
IA < IIA > IIIA < IV A < VA > VIA < VIIA
The ionisation energy of N(VA) is more than O (VIA)

4. The species CO,NO^+, CN^- and C_2^{2-} contain 14 electrons each.

6. **R** is the correct explanation of **A.**

PART - III

1. O^{2-} and F^- have two shells while Li^+ and B^{3+} have only one shell. Also, $O^{2-} > F^-$ (for isoelectronic species, as Z increases, size decreases).

2. The addition of second electron in an atom or ion is always endothermic because of repulsion between two negative charges.

3. Nitrogen has half filled stable configuration, ns^2np^3. So, ionization enthalpy of nitrogen is greater than oxygen. On moving down the group, metallic radius increases due to increase in number of shells.

4. Lanthanide contraction is due to poor shielding of one of $4f$ electron by another in the sub-shell.

5. The atomic radii of the second and third transition series are almost the same. This phenomenon is associated with the intervention of the 4f orbitals which must be filled before the 5d series of elements begin. The filling of 4f before 5d orbital results in a regular decrease in atomic radii called **Lanthanide contraction** which essentially compensates for the expected increase in atomic size with increasing atomic number. The net result of the lanthanide contraction is that the second and the third d series exhibit similar radii (e.g., Zr 160 pm, Hf 159 pm).

6.

Element :	B	S	P	F
I.E.(kJ mol^{-1}) :	801	1000	1011	1681

In general as we move from left to right in a period, the ionization enthalpy increases with increasing atomic number. The ionization enthalpy decreases as we move down a group. P ($1s^2$, $2s^2$, $3s^2$ $3p^3$) has a stable half filled electronic configuration than S ($1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^4$). For this reason, ionization enthalpy of P is greater than S.

7. Lanthanoid contraction is due to ineffective shielding produced by larger f-subshell.

8. Due to the inert pair effect (the reluctance of ns^2 electrons of outermost shell to participate in bonding) the stability of M^{2+} ions (of group 14 elements) increases as we go down the group.

9. Down the group, ionic radii increases with increasing atomic number because of the increase in the number of shells. But across the period, the ionic radii decreases due to increase in effective nuclear charge as electrons are added in the same shell. Li^+ and Mg^{2+} are diagonally related but Mg^{2+} having higher charge is smaller than Li^+, so correct order is $Na^+ > Li^+ > Mg^{2+} > Be^{2+}$.

Be^{2+} = 0.31 Å
Mg^{2+} = 0.72 Å
Li^+ = 0.76 Å
Na^+ = 1.02 Å

10. For isoelectronic species, ionic radii $\propto \frac{1}{\text{nuclear charge}}$.

So, correct order of ionic radii is ${}_8O^{2-} > {}_9F^- > {}_{11}Na^+ > {}_{12}Mg^{2+} > {}_{13}Al^{3+}$.

11. Gadolinium (${}_{64}Gd$) = $[Xe]^{54}$ $4f^7 5d^1 6s^2$

12. As we move in a group from top to bottom, electron gain enthalpy becomes less negative because the size of the atom increases and the added electron would be at larger distance from the nucleus.

Negative electron gain enthalpy of F is less than Cl. This is due to the fact that when an electron is added to F, the added electron goes to the smaller n = 2 energy level and experiences significant repulsion from the other electrons present in this level. In Cl, the electron goes to the larger n = 3 energy level and consequently occupies a larger region of space leading to much less electron-electron repulsion. So the correct order is Cl > F > Br > I.

13. Order of ionic radii $Ca^{2+} < K^+ < Cl^- < S^{2-}$

In isoelectronic species, as Z increases, size decreases.

14. Order of increasing ΔH_{IE_1} : Ba < Ca < Se < S < Ar

Ba < Ca ; Se < S : On moving top to bottom in a group, size increases. So ionisation enthalpy decreases.

Ar : Maximum value of ionisation enthalpy, since it is an inert gas.

15. $Na \longrightarrow Na^+ + e^-$ I^{st} I.E. = 5.1 eV

$Na^+ + e^- \longrightarrow Na$ Electron gain enthalpy of Na^+

Because reaction is reverse, so $\Delta_{eg}H = -5.1$ eV.

16. These are isoelectronic species.

As negative charge increases, ionic radius increases.

17. $I.P_1$ = Sc > Na > K > Rb

18. Isoelectronic species :

O^{2-}, F^-, Na^+, Mg^{2+} (All contain 10 electrons)

19. Second electron gain enthalpy of oxygen is a positive value.

20. Due to inert pair effect $T\ell$ is stable in +1 state

21. Down the group electronegativity decreases and atomic radius increases

22. Decrease in both atomic and ionic radii

23. Be (diagonal relationship).

24.

Element	Al	Si	P	S	Ga	Ge	Se	Te
E.N.	1.5	1.8	2.1	2.5	1.6	1.8	2.4	2.1

25. On going down the group size increases. From left to right in a period size decreases.