



CLASS 11 & 12 • NEET (UG) • JEE MAIN • JEE ADVANCED

SHORT NOTES

on

INORGANIC CHEMISTRY

Periodic Table • Chemical Bonding • Coordination Compounds •
Salt Analysis • d- and f-Block Elements • Metallurgy • Group 13 •
Group 14 • Group 15 • Group 16 • Group 17 • Group 18

Key results, shortcuts, traps and solved examples



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Short Notes on Inorganic Chemistry

for Class 11 and 12, NEET (UG), JEE Main and JEE Advanced

MTA Team • Master Testing Agency

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Numerical answers to the solved examples have been checked a second time by computer; the book is not printed if any of these checks fails. If you still find a mistake, please write to us.

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How to use these notes

These are short revision notes, not a first textbook. Read the NCERT chapter first; then use these notes to fix the results in memory and to see how they turn up in NEET and JEE questions. The rule behind every page is *minimum input, maximum output*: the fewest words that carry the whole idea.

Boxed text holds key definitions and results. Learn these.

Red text marks shortcuts, traps and common mistakes. **SHORTCUT** marks a quick way to the answer. **TRAP** marks a place where students lose marks.

Mnemonics are given in the form students actually use in class; say them aloud once and they stay.

Solved examples. Each chapter carries NEET- and JEE-type examples with full solutions. Cover the solution, try the question, then check. Every answer has been worked again independently.

Last week before the exam. Reading only the boxes and the red lines of every chapter is a complete revision of inorganic chemistry.

Which exam needs what.

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Periodic Table, Chemical Bonding	•	•	•	11
Coordination Compounds	•	•	•	12
Salt Analysis	•	•	◦	Practical
d- and f-Block Elements	•	•	•	12
Metallurgy	•	◦	◦	12
Groups 13 and 14	•	•	•	11
Groups 15 to 18	•	•	•	12

• core ◦ lighter weight in recent papers

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Periodic Table

Arranging elements by atomic number makes their properties repeat. Learn the classification, then the four trends (Z_{eff} , size, ionisation enthalpy, electron gain enthalpy) and electronegativity; almost every question is one of these trends with one exception.

1.1 Early classification

1.1.1 Döbereiner's law of triads

A group of three similar elements in which the atomic mass of the middle one is the average of the other two.

Triad	Atomic masses	Check
Li, Na, K	7, 23, 39	$(7 + 39)/2 = 23$
Ca, Sr, Ba	40, 88, 137	$(40 + 137)/2 = 88.5$
Cl, Br, I	35.5, 80, 127	$(35.5 + 127)/2 = 81.2$
P, As, Sb	31, 75, 122	$(31 + 122)/2 = 76.5$
S, Se, Te	32, 79, 128	$(32 + 128)/2 = 80$

SHORTCUT No atomic masses are needed: if the atomic numbers of A, B, C are x, y, z , they form a triad when

$$(y - x) = (z - y)$$

1.1.2 Lothar Meyer curve

Lothar Meyer plotted physical properties (atomic volume, m.p., b.p.) against atomic mass.

- Maxima: group 1 alkali metals. Descending part: group 2 alkaline earth metals.
- Ascending part: halogens. Minima: heavy (transition) metals.

1.1.3 Modern periodic table

- Proposed by Moseley; it is based on the **atomic number** of elements.
- Moseley bombarded different metals with high-speed electrons and studied the X-rays given out. He found $\sqrt{\nu} = a(Z - b)$, i.e. $\sqrt{\nu} \propto Z$.

Modern periodic law

The physical and chemical properties of elements are periodic functions of their atomic numbers.

1.2 Electronic configuration and the table

Elements of a group generally have the same outer-shell configuration. **Exceptions:** groups 5, 6, 8, 9, 10 and 18. For example, in group 10: Ni $3d^8 4s^2$, Pd $4d^{10} 5s^0$, Pt $5d^9 6s^1$.

1.2.1 General outer configurations

Group	General configuration
IA (alkali metals)	ns^1 (here $n \neq 1$, as hydrogen is not an alkali metal)
IIA (alkaline earth metals)	ns^2
IIIA (boron family)	ns^2np^1
IVA (carbon family)	ns^2np^2
VA (pnictogens)	ns^2np^3
VIA (chalcogens)	ns^2np^4
VIIA (halogens)	ns^2np^5
Inert gases	ns^2np^6 (except He, $1s^2$)

1.2.2 d-block configurations

While filling electrons in d-block elements we never write the following configurations; they are the exceptions:

3d series: $3d^4$, $3d^9$ 4d series: $4d^3$, $4d^6$, $4d^9$ 5d series: $5d^8$ (these are replaced by the configurations below)

Element	Configuration (not the expected one)
$_{24}\text{Cr}$	$3d^54s^1$ (not $3d^44s^2$)
$_{29}\text{Cu}$	$3d^{10}4s^1$ (not $3d^94s^2$)
$_{41}\text{Nb}$	$4d^45s^1$ (not $4d^35s^2$)
$_{42}\text{Mo}$	$4d^55s^1$ ($4d^5$ more stable)
$_{43}\text{Tc}$	$4d^55s^2$ (Class XI NCERT) / $4d^65s^1$ (Class XII NCERT)
$_{44}\text{Ru}$	$4d^75s^1$ (not $4d^65s^2$)
$_{45}\text{Rh}$	$4d^85s^1$ (t_{2g} , e_g argument)
$_{46}\text{Pd}$	$4d^{10}5s^0$ (not $4d^85s^2$)
$_{47}\text{Ag}$	$4d^{10}5s^1$ (not $4d^95s^2$)
$_{78}\text{Pt}$	$5d^96s^1$ (not $5d^86s^2$)
$_{79}\text{Au}$	$5d^{10}6s^1$ (d^{10} more stable)
$_{111}\text{Rg}$	$6d^{10}7s^1$ by analogy with Au (d^{10} more stable)

1.2.3 f-block configurations

MNEMONIC Lanthanoid names (Ce to Lu): “CP Nandu Prasad Singh, English Garde; Tution dete ho, Engineer Tumhare Yaha Lallu” = Ce Pr Nd Pm Sm Eu Gd Tb Dy Ho Er Tm Yb Lu.

MNEMONIC Actinoid names (Th to Lr): “Thapa U Nepali, Pure aam (taste) chakh chukh; Kyo fisse fir maan na Ladu?” = Th Pa U Np Pu Am Cm Bk Cf Es Fm Md No Lr.

Elements with $Z > 92$ are called **transuranic elements**.

Lanthanoids (4f series). Remember $[\text{Xe}] + 6s^2 = 56$ electrons. Only $_{58}\text{Ce}$, $_{64}\text{Gd}$ and $_{71}\text{Lu}$ have $5d^1$ in their configuration; in the remaining 4f elements keep $5d^0$ and keep filling electrons in 4f.

$_{58}\text{Ce} [\text{Xe}] 4f^1 5d^1 6s^2$	$_{65}\text{Tb} [\text{Xe}] 4f^9 5d^0 6s^2$
$_{59}\text{Pr} [\text{Xe}] 4f^3 5d^0 6s^2$	$_{66}\text{Dy} [\text{Xe}] 4f^{10} 5d^0 6s^2$
$_{60}\text{Nd} [\text{Xe}] 4f^4 5d^0 6s^2$	$_{67}\text{Ho} [\text{Xe}] 4f^{11} 5d^0 6s^2$
$_{61}\text{Pm} [\text{Xe}] 4f^5 5d^0 6s^2$	$_{68}\text{Er} [\text{Xe}] 4f^{12} 5d^0 6s^2$
$_{62}\text{Sm} [\text{Xe}] 4f^6 5d^0 6s^2$	$_{69}\text{Tm} [\text{Xe}] 4f^{13} 5d^0 6s^2$
$_{63}\text{Eu} [\text{Xe}] 4f^7 5d^0 6s^2$	$_{70}\text{Yb} [\text{Xe}] 4f^{14} 5d^0 6s^2$
$_{64}\text{Gd} [\text{Xe}] 4f^7 5d^1 6s^2$	$_{71}\text{Lu} [\text{Xe}] 4f^{14} 5d^1 6s^2$

Actinoids (5f series). Remember $[\text{Rn}] + 7s^2 = 88$ electrons. The first four elements of the series (Th, Pa, U, Np) and $_{96}\text{Cm}$ have exceptional configurations (a 6d electron).

$_{90}\text{Th} [\text{Rn}] 5f^0 6d^2 7s^2$	$_{97}\text{Bk} [\text{Rn}] 5f^9 6d^0 7s^2$
$_{91}\text{Pa} [\text{Rn}] 5f^2 6d^1 7s^2$	$_{98}\text{Cf} [\text{Rn}] 5f^{10} 6d^0 7s^2$
$_{92}\text{U} [\text{Rn}] 5f^3 6d^1 7s^2$	$_{99}\text{Es} [\text{Rn}] 5f^{11} 6d^0 7s^2$
$_{93}\text{Np} [\text{Rn}] 5f^4 6d^1 7s^2$	$_{100}\text{Fm} [\text{Rn}] 5f^{12} 6d^0 7s^2$
$_{94}\text{Pu} [\text{Rn}] 5f^6 6d^0 7s^2$	$_{101}\text{Md} [\text{Rn}] 5f^{13} 6d^0 7s^2$
$_{95}\text{Am} [\text{Rn}] 5f^7 6d^0 7s^2$	$_{102}\text{No} [\text{Rn}] 5f^{14} 6d^0 7s^2$
$_{96}\text{Cm} [\text{Rn}] 5f^7 6d^1 7s^2$	$_{103}\text{Lr} [\text{Rn}] 5f^{14} 6d^1 7s^2$

1.3 Periods of the modern periodic table

There are seven periods.

Period	Outer shell n	Subshells filled	Elements	Name
1	1	1s	2	shortest
2	2	2s 2p	8	short
3	3	3s 3p	8	short
4	4	4s 3d 4p	18	long
5	5	5s 4d 5p	18	long
6	6	6s 4f 5d 6p	32	longest
7	7	7s 5f 6d 7p	32	longest

Every period starts with ns and ends with np : $ns (n-2)f (n-1)d np$. Period number = outermost shell number.

Number of elements in a period = maximum number of electrons in that period's subshells (e.g. period 6: $6s^2 4f^{14} 5d^{10} 6p^6 = 32$).

$$\text{No. of elements in period } n = \frac{(n+2)^2}{2} \text{ (} n \text{ even), } \frac{(n+1)^2}{2} \text{ (} n \text{ odd)}$$

$$\text{No. of subshells in a period} = \sqrt{\frac{\text{no. of elements}}{2}}$$

1.4 IUPAC names of elements with $Z > 100$

The names end in the suffix *-ium* and are built from these roots:

0 nil	1 un	2 bi	3 tri	4 quad
5 pent	6 hex	7 sept	8 oct	9 enn

Example: $Z = 143 \rightarrow$ un-quad-tri-ium = **unquadtrium**, symbol Uqt.

1.5 Locating an element (group and period)

1.5.1 When the atomic number is given

1. First check whether Z belongs to the f-block: $58 \leq Z \leq 71$ lanthanoids (group 3, period 6); $90 \leq Z \leq 103$ actinoids (group 3, period 7).
2. Then check whether Z is an inert gas: He 2, Ne 10, Ar 18, Kr 36, Xe 54, Rn 86, Og 118 (periods 1 to 7; gaps 8, 8, 18, 18, 32, 32).
3. For any other Z : take the inert gas just to the right of Z . The element is in the same period as that inert gas, and

$$\text{Group of } Z = 18 - [Z_{\text{inert}} - Z_{\text{given}}]$$

If the group comes out negative (f-block region), use 32 in place of 18. **SHORTCUT** The numbers 8, 8, 18, 18, 32, 32 are the “magic numbers” used to find atomic numbers down a group.

1.5.2 When the electronic configuration is given

Period number = outermost shell number (n of ns or np).

1. If np electrons are present: p-block, Group = $12 + (\text{number of } np \text{ electrons})$.
2. If only ns^{1-2} is present [$(n-2)f$ and $(n-1)d$ absent]: s-block, Group = number of ns electrons.
3. $(n-2)f^{1-14}(n-1)d^{0-1}ns^2$ (f present): f-block, Group = IIIB (3).
4. Any other configuration, $(n-1)d^{1-10}ns^{0-2}$: d-block, Group = $(n-1)d$ electrons + ns electrons.

1.6 Bohr's classification of elements

- **Inert gases:** completely filled n th shell (group 18).
- **Normal / representative elements:** partially filled n th shell (all s- and p-block except group 18).
- **Transition elements:** partially filled n th and $(n-1)$ th shells, in the atom or in a common ion (whole d-block except Zn, Cd, Hg, Cn).
- **Inner transition elements:** partially filled n th, $(n-1)$ th and $(n-2)$ th shells (whole f-block, 28 elements).

1.7 Some important definitions

1. **Typical elements:** elements of the 3rd period (except the inert gas): Na, Mg, Al, Si, P, S, Cl.
2. **Diagonal relationship:** similarity in properties with the diagonal element of the adjacent group (Li–Mg, Be–Al, B–Si). Reason: almost the same ionic potential (charge/size). Observed only up to group 14.
3. **Bridge elements:** in Mendeleev's table, 3rd-period elements (Na, Mg, Al, ...); in the modern table, Li, Be, B. **If the table is not named, answer according to Mendeleev's table.**
4. **Transuranium elements:** elements with $Z > 92$; all are radioactive and man-made.
5. **Periodicity:** repetition of properties after regular intervals, due to similar outer-shell configuration. Atomic radius, ionisation enthalpy, electron affinity and electronegativity are all periodic properties.

1.8 Effective nuclear charge (Z_{eff})

In a many-electron species the net nuclear charge experienced by the outermost electron is the effective nuclear charge.

$$Z_{\text{eff}} = Z - \sigma \quad (Z = \text{nuclear charge, } \sigma = \text{screening (shielding) constant})$$

Screening is caused only by electrons of the same shell or inner shells. Electrons closer to the nucleus shield more; this is the **penetration effect**. Shielding power: $s > p > d > f$.

1.8.1 Slater's rules (theoretical σ)

Arrange the configuration as $(1s)(2s, 2p)(3s, 3p)(3d)(4s, 4p)(4d)(4f) \dots$

1. For an ns or np electron: other ns, np electrons 0.35 each; $(n-1)$ shell electrons 0.85 each; $(n-2)$ and inner electrons 1.00 each.
2. For an nd or nf electron: other electrons of the same nd/nf group 0.35 each; all electrons to the left 1.00 each.
3. **Electrons to the right of (outer to) the test electron do not shield: $\sigma = 0$.**
4. In two-electron species (He, Li^+ , Be^{2+}) σ of one $1s$ electron on the other is 0.30: $Z_{\text{eff}}(\text{He}) = 2 - 0.30 = 1.70$, $Z_{\text{eff}}(\text{Be}^{2+}) = 4 - 0.30 = 3.70$.
5. Z_{eff} on ns and np electrons of the same shell is the same; e.g. Z^* on a $3s$ electron of Fe = Z^* on a $3p$ electron of Fe = $26 - 11.25 = 14.75$.

Example 1.1 Find Z_{eff} for (a) a $3d$ electron of Fe, (b) the last electron of Na.

Solution. (a) Fe = $(1s^2)(2s^2 2p^6)(3s^2 3p^6)(3d^6)(4s^2)$. For a $3d$ electron: 18 inner electrons $\times 1.00$ + 5 other $3d$ electrons $\times 0.35$; $4s$ electrons are to the right, $\sigma = 0$. $\sigma = 18 + 1.75 = 19.75$, $Z_{\text{eff}} = 26 - 19.75 = 6.25$.

(b) Na = $(1s^2)(2s^2 2p^6)(3s^1)$: $\sigma = 2(1.00) + 8(0.85) = 8.8$, $Z_{\text{eff}} = 11 - 8.8 = 2.2$.

Answer: (a) 6.25 (b) 2.2

1.8.2 Important observations

1. Z^* increases by 0.65 from one element to the next across the 2nd and 3rd periods (Li $\rightarrow$ Ne, Na $\rightarrow$ Ar).
2. Z^* increases by about 0.9 on moving down a group from the 2nd to the 3rd period (Li 1.3 $\rightarrow$ Na 2.2).
3. Z_{eff} increases on moving down a group in the d-block.
4. Z_{eff} increases from Ce to Lu because of the poor shielding of $4f$ electrons.

Isoelectronic species

$$Z_{\text{eff}} \propto +\text{ve charge} \propto \frac{1}{-\text{ve charge}} \quad \text{e.g. } \text{O}^{2-} < \text{F}^- < \text{Ne} < \text{Na}^+ < \text{Mg}^{2+} < \text{Al}^{3+}$$

1.9 Atomic radius

The average distance of the valence-shell electrons from the nucleus. It is of four kinds: covalent, metallic, van der Waals and ionic radius.

1. **Covalent radius:** half the internuclear distance between two identical atoms joined by a single covalent bond (non-metal–non-metal).
2. **Metallic radius:** half the internuclear distance between two adjacent atoms in a metallic crystal lattice.
3. **van der Waals radius:** half the distance between two non-bonded atoms (e.g. noble gas atoms in the solid), held by van der Waals forces.

Non-metals have both a covalent and a van der Waals radius; van der Waals radius > covalent radius.

1.9.1 Along a period

Moving left to right Z_{eff} increases, so size decreases.

$$Z_{\text{eff}} \propto \frac{1}{\text{atomic radius}}$$

TRAP In every period the inert gas has the largest radius, because its radius is a van der Waals radius.

Exception: covalent radii of the 3d series are irregular: $\text{Sc} > \text{Ti} > \text{V} > \text{Cr} \approx \text{Mn} \approx \text{Fe} \approx \text{Co} \approx \text{Ni} < \text{Cu} < \text{Zn}$ (start: Z_{eff} effect $> \sigma$; middle: $Z_{\text{eff}} \approx \sigma$; end: $Z_{\text{eff}} < \sigma$).

1.9.2 Down a group

Moving top to bottom the number of shells increases, so the radius increases.

atomic radius $\propto$ number of shells

- If $\Delta Z = 32$ between two elements of a group, their size remains almost the same.
- The last two elements of a group have almost the same size (up to group 14), except group 12.
- In the d-block, radius contractions are caused by poor shielding of 4f and 5f electrons: size $3d < 4d \approx 5d \approx 6d$ (lanthanoid and actinoid contraction).
- **Two exceptions:** group 3 (IIIB): $\text{Sc} < \text{Y} < \text{La} \approx \text{Ac}$; group 12 (IIB): $\text{Zn} < \text{Cd} < \text{Hg}$ (metallic bonding weakens).

Transition contraction is seen only in group 13: $\text{B} < \text{Ga} \lesssim \text{Al} < \text{In} \approx \text{Tl}$ (Ga after the 3d series; Tl after the 4f series: lanthanoid contraction).

Isoelectronic species

$$\text{Ionic radius} \propto \frac{1}{Z_{\text{eff}}} \propto \frac{-\text{ve charge}}{+\text{ve charge}}$$

s-block: size increases down the group ($\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs}$; $\text{Be} < \text{Mg} < \text{Ca} < \text{Sr} < \text{Ba}$) and the group 1 atom is larger than its group 2 neighbour; ionisation enthalpy runs the opposite way.

1.10 Ionisation enthalpy (IE)

The minimum energy needed to remove the most loosely bound valence electron of an isolated gaseous atom. Energy supplied as heat: ionisation enthalpy; as an electric field: ionisation potential.

- IE is endothermic for a neutral atom.
- As the net attractive force increases, IE increases. Each successive IE is greater than the previous one:
 $\text{Al(g)} \xrightarrow{\text{IE}_1} \text{Al}^+(\text{g}) \xrightarrow{\text{IE}_2} \text{Al}^{2+}(\text{g}) \xrightarrow{\text{IE}_3} \text{Al}^{3+}(\text{g}), \text{IE}_1 < \text{IE}_2 < \text{IE}_3.$

$$\text{Isoelectronic species: IE} \propto Z_{\text{eff}} \propto \frac{+\text{ve charge}}{-\text{ve charge}} \quad \text{In a group: IE} \propto \frac{1}{\text{size}}$$

- Down a group, if size stays nearly the same (lanthanoid, actinoid or transition contraction), IE increases because Z and hence Z_{eff} increase.
 Group 13: size $\text{B} < \text{Ga} < \text{Al} < \text{In} < \text{Tl}$; IE_1 : $\text{B} (801) > \text{Al} (577) < \text{Ga} (579) > \text{In} (558) < \text{Tl} (589) \text{ kJ mol}^{-1}.$
- d-block size: $3d < 4d \approx 5d$. d-block IE: generally $5d > 3d > 4d$; **in groups 5, 6 and 10 the order is $3d < 4d < 5d$.**

1.10.1 Across a period

1. Check the electronic configuration and the outermost shell.
 2. IE: $np^3 > np^4$ (half-filled stability; applicable up to the 4th period), e.g. $\text{N} > \text{O}, \text{P} > \text{S}$.
 3. IE: $ns^2 > np^1$ (penetration effect; applicable only when $\Delta Z = 1$), e.g. $\text{Be} > \text{B}, \text{Mg} > \text{Al}$.
 4. If the charge and the outermost shell are the same, the inert-gas configuration has the maximum IE.
 5. $\text{IE} \propto Z_{\text{eff}}$.
- IE of lanthanoids generally increases from Ce to Lu because Z_{eff} increases.

- In the 3d series IE increases left to right except at intervals 2, 6 and 7:
 $\text{Sc} < \text{Ti} > \text{V} < \text{Cr} < \text{Mn} < \text{Fe} > \text{Co} > \text{Ni} < \text{Cu} < \text{Zn}$.

1.10.2 Applications of ionisation enthalpy

- Number of valence electrons:** a huge jump in a particular IE shows that the electron is now being removed from the $(n - 1)$ th shell. Example: Mg, $\text{IE}_1 = 738$, $\text{IE}_2 = 1451$, $\text{IE}_3 = 7733 \text{ kJ mol}^{-1}$: the jump after IE_2 shows 2 valence electrons.
- Stable oxidation state:** if the difference between two successive IE values is $< 11 \text{ eV}$, the higher oxidation state is more stable; if it is $> 16 \text{ eV}$, the lower oxidation state is more stable.
- Metallic character:** electropositive character $\propto \frac{1}{\text{IE}}$.

1.11 Electron affinity (EA) and electron gain enthalpy ($\Delta_{eg}H$)

The energy released when an electron is added to the outermost shell of an isolated gaseous atom is its electron affinity.

$$\Delta_{eg}H = -(\text{electron affinity}) \quad |\Delta_{eg}H| = |\text{EA}| \text{ (same magnitude)}$$

- Energy has to be supplied for adding an electron only to four kinds of species: **Be, Mg, N and the inert gases**.
- So the first $\Delta_{eg}H$ is negative for most species except Be, Mg, N and inert gases. From the second $\Delta_{eg}H$ onwards it is generally positive.
- Examples: $\text{Cl(g)} + \text{e}^- \longrightarrow \text{Cl}^-(\text{g}) + 349 \text{ kJ mol}^{-1}$ ($\Delta_{eg}H = -349$); $\text{Ne(g)} + \text{e}^- \longrightarrow \text{Ne}^-(\text{g})$ ($\Delta_{eg}H = +116$); $\text{He(g)} + \text{e}^- \longrightarrow \text{He}^-(\text{g})$ ($\Delta_{eg}H = +48 \text{ kJ mol}^{-1}$).
- Formation of a polynegative ion is endothermic (EA negative, $\Delta_{eg}H$ positive).

$$\text{Isoelectronic: EA} \propto Z_{\text{eff}} \propto \frac{+\text{ve charge}}{-\text{ve charge}} \quad \text{Down a group: EA} \propto \frac{1}{\text{size}}$$

Exception (very important). EA of a 2nd-period element is less than EA of the 3rd-period element of the same group, because the small 2p shell gives the incoming electron more repulsion:



Chlorine has the highest electron affinity of all elements (349 kJ mol^{-1}).

Across a period: (1) check the outer configuration; (2) adding an electron to Be, Mg, N and inert gases is endothermic; (3) $\text{EA } ns^1 > np^1$ and $\text{EA } np^2 > np^3$ (e.g. $\text{C } 122 > \text{N} \approx -7 \text{ kJ mol}^{-1}$; so EA of group 14 $>$ group 15); (4) $\text{EA} \propto Z_{\text{eff}}$.

Relation between IE and EA



1.12 Electronegativity (EN)

The tendency of an atom to attract the shared pair of electrons towards itself in a covalent bond. EN is a unitless quantity; it is not a property of an isolated atom (it is measured for a bonded atom).

$$\begin{array}{lll} \text{In a period: EN} \propto Z_{\text{eff}} & \text{In a group: EN} \propto \frac{1}{\text{size}} & \text{Isoelectronic species: EN} \propto \frac{+\text{ve O.S.}}{-\text{ve O.S.}} \\ \text{EN} \propto \% \text{ s-character: } sp > sp^2 > sp^3 & & \end{array}$$

1.12.1 Applications of EN

1. Non-metallic character $\propto$ EN.
2. Bond polarity $\propto \Delta$ EN.
3. Bond parameters: as Δ EN increases, ionic character increases, bond length decreases, bond strength increases.
4. Acidic nature of oxides and oxyacids $\propto$ EN of the central atom.

Solved examples

Example 1.2 Find the group and period of the element with $Z = 42$.

Solution. The next inert gas is Xe ($Z = 54$, period 5). Group = $18 - (54 - 42) = 6$.

Answer: Period 5, group 6 (it is Mo).

Example 1.3 Name the element with $Z = 119$ and give its symbol.

Solution. 1-1-9 $\rightarrow$ un-un-enn-ium.

Answer: Ununennium, Uue.

Example 1.4 An element has the configuration $[\text{Ar}]3d^{10}4s^24p^3$. Give its block, period and group.

Solution. np electrons present: p-block. $n = 4$: period 4. Group = $12 + 3 = 15$.

Answer: p-block, period 4, group 15 (As).

Example 1.5 Arrange in increasing ionic radius: Na^+ , F^- , O^{2-} , Mg^{2+} , N^{3-} .

Solution. All have 10 electrons; radius $\propto$ -ve charge/+ve charge.

Answer: $\text{Mg}^{2+} < \text{Na}^+ < \text{F}^- < \text{O}^{2-} < \text{N}^{3-}$

Example 1.6 Which has the higher first ionisation enthalpy: (a) N or O, (b) Be or B, (c) Mg or Al?

Solution. (a) N ($2p^3$, half-filled). (b) Be ($2s^2$; penetration, $\Delta Z = 1$). (c) Mg ($3s^2$).

Answer: (a) N (b) Be (c) Mg

Example 1.7 Successive IEs of an element are 520, 7298, 11815 kJ mol^{-1} . How many valence electrons does it have?

Solution. The huge jump comes after IE_1 .

Answer: One (it is Li).

Example 1.8 Arrange F, Cl, Br, I in increasing negative electron gain enthalpy.

Solution. F is out of order because of its small 2p shell.

Answer: $\text{I} < \text{Br} < \text{F} < \text{Cl}$

Example 1.9 How many elements are there in the 6th period? Use the formula.

Solution. $n = 6$ is even: $(6 + 2)^2/2 = 32$.

Answer: 32