

PERIODIC CLASSIFICATION OF ELEMENT

DOBERNIER TRIADS:-

- Arranged elements in the group of 3.
- Atomic mass of middle element is average of atomic mass of other two elements.



Drawbacks:-

- as known elements were not grouped

Lothar Meyer's Curves:-



- Alkali metals → top position
- Alkaline earth metals → descending
- Halogen → ascending position.

MODERN PERIODIC TABLE:-

- long form of periodic table
- based on Moseley law: $\sqrt{\lambda} = a - b(Z - c)$
- Periodic properties are periodic function of their atomic number.

Periodicity - Repetition of similar properties of elements after regular interval of time.

Cause of periodicity - due to similar e⁻ configuration.

Transition elements; they have incompletely filled d-orbitals in their ground state.

Zn, Cd, Hg → are not transition elements because of completely filled d-orbitals. (pseudo transition elements).

Representative elements - Combined elements of s and p block are known as representative elements.

NEWLAND'S OCTAVE RULE:-

- When elements are arranged in increasing order of atomic mass, then every 8th element standing from first will show same property.



Drawbacks:-

1. Valid upto calcium only
2. No position of inert gases
3. Not worked on heavier elements.

MENDELÉEV'S PERIODIC

- Periodic property of elements are periodic function of their atomic masses.

Merits:-

- Simplified study
- Some vacant space was left for the elements which were not known at that time.

Eka aluminium → Ga
Eka silicon → Ge

- Corrected atomic mass of element

Demerits:-

- (a) Position of hydrogen was not fixed
- (b) Could not explain periodicity.

MODERN PERIODIC TABLE:-

18 groups and 7 periods.

4-Blocks

s-Block = G₁ and G₂ = ns¹⁻²

p-Block = G₁₃ - G₁₈ = ns² np¹⁻⁶

d-Block = G₃ - G₁₀ = (n-1)d¹⁻¹⁰ ns²

f-Block = lanthanoid = 2 = 58 - 71

G₃ and P₆

actinoids = 2 = 90 - 103

G₃ and P₆

Diagonal Relation;

Li → Be → B → similar
Na → Mg → Al → change ratio size

Nomenclature;

0 → nil	5 → pent	Suffix = turn
1 → un	6 → hex	100 unnilpentium
2 → bi	7 → sept	
3 → tri	8 → oct	
4 → quad	9 → non	

EFFECTIVE NUCLEAR CHARGE

$Z_{\text{eff}} = Z - \sigma$

Trends - Period → $Z_{\text{eff}} \uparrow$

Group → Z_{eff} almost remains same. or $Z_{\text{eff}} \uparrow \downarrow$

Shielding → s > p > d > f

→ more +ve charge, more is Z_{eff}

→ more -ve charge, less is Z_{eff}

IONISATION ENERGY

- Energy required to remove an electron from an isolated gaseous atom.



IE = IE₁ < IE₂ < IE₃ ...

Trends - $Z_{\text{eff}} \propto \text{IE}$

Period - $Z_{\text{eff}} \uparrow = \text{IE} \uparrow$

Group - $Z_{\text{eff}} \downarrow = \text{IE} \downarrow$

G₁₃ = G > Al > Ga > In

G₁₄ = C > Si > Ge > Pb > Sn

Period 2:-

Li < Be < B < C < N < O < F < Ne

ATOMIC RADIUS:-

- Distance b/w nucleus and electron of outermost shell.
- But it's impossible to find exact atomic radii.

Reason-1 = atoms too small

2 = atom doesn't have any boundary

Covalent Radius - Half of the internuclear distance b/w 2 atoms bonded by covalent bond.



Metallic Radius - Half of the internuclear distance b/w 2 atoms bonded by a metallic bond.



Van der Waals Radius - Half of the internuclear distance b/w 2 nearest molecules in non-bonded state.



Trends; Atomic Radius $\propto \frac{1}{Z_{\text{eff}}}$

Period → $Z_{\text{eff}} \uparrow = \text{Atomic Radius} \downarrow$

Group → $Z_{\text{eff}} \downarrow = \text{Atomic Radius} \uparrow$

Ionic Radius + Radius < R_{neutral} < R_{anion}

ELECTRON GAIN ENTHALPY:-

- Energy required or released when an e⁻ is added to outermost orbit of an isolated gaseous atom is known as Electron gain enthalpy.



Note - In majority cases Δ_{egH}2 is -ve (exothermic)

Exception = B, N } Δ_{egH}2 = +ve due to stable configuration

Note - Successive electron gain enthalpy is always +ve.

Trends; Δ_{egH} $\propto Z_{\text{eff}}$

Period - $Z_{\text{eff}} \uparrow \Rightarrow \Delta_{\text{egH}} \uparrow$

Group - $Z_{\text{eff}} \downarrow \Rightarrow \Delta_{\text{egH}} \downarrow$

G₁₃ = S > Se > Te > Po > O

→ due to small size Reason!

→ e⁻ e⁻ repulsion.

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DJ Losen



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1. Understanding the eBook Periodic Classification 3rd Chapter 11th Class
 - The Rise of Digital Reading Periodic Classification 3rd Chapter 11th Class
 - Advantages of eBooks Over Traditional Books
2. Identifying Periodic Classification 3rd Chapter 11th Class
 - Exploring Different Genres
 - Considering Fiction vs. Non-Fiction
 - Determining Your Reading Goals
3. Choosing the Right eBook Platform
 - Popular eBook Platforms
 - Features to Look for in an Periodic Classification 3rd Chapter 11th Class
 - User-Friendly Interface
4. Exploring eBook Recommendations from Periodic Classification 3rd Chapter 11th Class
 - Personalized Recommendations
 - Periodic Classification 3rd Chapter 11th Class User Reviews and Ratings
 - Periodic Classification 3rd Chapter 11th Class and Bestseller Lists
5. Accessing Periodic Classification 3rd Chapter 11th Class Free and Paid eBooks

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 - Periodic Classification 3rd Chapter 11th Class eBook Subscription Services
 - Periodic Classification 3rd Chapter 11th Class Budget-Friendly Options
6. Navigating Periodic Classification 3rd Chapter 11th Class eBook Formats
- ePub, PDF, MOBI, and More
 - Periodic Classification 3rd Chapter 11th Class Compatibility with Devices
 - Periodic Classification 3rd Chapter 11th Class Enhanced eBook Features
7. Enhancing Your Reading Experience
- Adjustable Fonts and Text Sizes of Periodic Classification 3rd Chapter 11th Class
 - Highlighting and Note-Taking Periodic Classification 3rd Chapter 11th Class
 - Interactive Elements Periodic Classification 3rd Chapter 11th Class
8. Staying Engaged with Periodic Classification 3rd Chapter 11th Class
- Joining Online Reading Communities
 - Participating in Virtual Book Clubs
 - Following Authors and Publishers Periodic Classification 3rd Chapter 11th Class
9. Balancing eBooks and Physical Books Periodic Classification 3rd Chapter 11th Class
- Benefits of a Digital Library
 - Creating a Diverse Reading Collection Periodic Classification 3rd Chapter 11th Class
10. Overcoming Reading Challenges
- Dealing with Digital Eye Strain
 - Minimizing Distractions
 - Managing Screen Time
11. Cultivating a Reading Routine Periodic Classification 3rd Chapter 11th Class
- Setting Reading Goals Periodic Classification 3rd Chapter 11th Class
 - Carving Out Dedicated Reading Time
12. Sourcing Reliable Information of Periodic Classification 3rd Chapter 11th Class
- Fact-Checking eBook Content of Periodic Classification 3rd Chapter 11th Class
 - Distinguishing Credible Sources
13. Promoting Lifelong Learning
- Utilizing eBooks for Skill Development

- Exploring Educational eBooks

14. Embracing eBook Trends

- Integration of Multimedia Elements
- Interactive and Gamified eBooks

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