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Periodic Table Importance

Significance of Periodic Table

- The periodic table is the foundational concept in chemistry, essential for understanding element behavior.
- It organizes vast chemical information into logical, simple patterns.
- Elements are not random; they show trends and group into families with similar properties.

Solved Examples

- **Ex 1:** Identify the importance of periodic table in predicting element properties.
Step 1: Recognize element groups.
Step 2: Use group trends to predict chemical behavior.

Practice Set

- Q1: What does the periodic table primarily organize?
A. Atomic masses
B. Elements by properties
C. Compounds
D. Isotopes
- Q2: Which statement about elements in the periodic table is true?
A. Elements are randomly placed
B. Elements in a group have similar properties
C. Elements in a period have identical properties
D. Elements are arranged by color
- Q3: Why is the periodic table considered a great scientific achievement?
A. It lists all compounds
B. It groups elements by similar chemical behavior
C. It shows atomic numbers only
D. It arranges elements alphabetically

Answer Key

- Q1: B - The periodic table organizes elements based on their chemical properties.
- Q2: B - Elements in the same group share similar chemical properties.
- Q3: B - It groups elements to reveal chemical behavior patterns.

Memory Aids

- Remember: "Groups = Families" with similar properties.
- Periodic table = "Chemical behavior map".

Need for Classification

Why Classify Elements?

- Elements are basic units of matter; their number increased from 31 (1800) to 118 (present).
- Studying each element and its compounds individually is impractical.
- Classification organizes elements systematically by similar properties.
- This organization leads to periodicity and the creation of the periodic table.
- Periodic table groups elements with similar properties together for easy study.

Solved Examples

- **Ex 1:** Explain why classification of elements is necessary.
Step 1: Note the increasing number of elements.
Step 2: Understand difficulty in individual study.
Step 3: Use classification to group similar elements.

Practice Set

- Q1: What was the number of known elements around 1800?
A. 31

- B. 63
- C. 118
- D. 50
- Q2: What is the main purpose of classifying elements?
 - A. To memorize elements
 - B. To organize elements by properties
 - C. To increase element count
 - D. To separate metals and non-metals only
- Q3: What does classification of elements lead to?
 - A. Random grouping
 - B. Chemical periodicity
 - C. Element isolation
 - D. Atomic mass increase

Answer Key

- Q1: A - 31 elements were known in 1800.
- Q2: B - Classification organizes elements by similar properties.
- Q3: B - Classification results in chemical periodicity.

Memory Aids

- Remember: "Classify to simplify" study of elements.
- Periodic table = "Organized element family tree".

Genesis of Periodic Classification

Dobereiner's Triads (1829)

- Elements grouped in triads with similar properties.
- Atomic weight of middle element $\approx$ arithmetic mean of other two.
- Example: Cl (35.5), Br (80), I (127); $Br \approx (35.5 + 127)/2 = 81.25$

Triads	Mean of atomic masses of 1st and 3rd elements
Li (7), Na (23), K (39)	$(7 + 39) / 2 = 23$
Cl (35.5), Br (80), I (127)	$(127 + 35.5) / 2 = 81.25$
Ca (40), Sr (88), Ba (137)	$(137 + 40) / 2 = 88.5$

Newland's Octaves Law (1865)

- Elements arranged by increasing atomic weight.
- Every 8th element has similar properties, like musical notes.
- Example: Li, Na, K show similar properties.

Lothar Meyer's Atomic Volume Curve (1869)

- Graph plotted: Atomic volume vs atomic weight.
- Elements with similar properties occupy similar graph positions.
- Alkali metals at peaks; alkaline earth metals mid-points; halogens on ascending curve.

Mendeleev's Periodic Law (1869)

- Properties of elements are periodic functions of atomic weights.
- Elements arranged in increasing atomic weight; similar properties recur periodically.
- Table: 9 vertical columns (groups), 7 horizontal rows (periods).
- Elements with similar properties placed in same group.
- Atomic weight order adjusted to maintain chemical properties.
- Vacant spaces left for undiscovered elements.

Demerits of Mendeleev's Periodic Table

- Hydrogen's position anomalous; resembles group 1 and group 17 but placed in group 1.
- Cobalt (58.9) placed before nickel (58.6) despite higher atomic mass.
- Isotopes not given separate places.
- Could not explain cause of periodicity.

- Lanthanoids and actinoids excluded.

Solved Examples

Ex 1

- Find the atomic weight of middle element in triad: Li (7), K (39).
- Calculation: $(7 + 39)/2 = 23$.
- Middle element: Na (23), confirming Dobereiner's triad.

Practice Set

- Q1: In Dobereiner's triads, the atomic weight of the middle element is:
 - A) Geometric mean of other two
 - B) Arithmetic mean of other two
 - C) Harmonic mean of other two
 - D) Median of other two
- Q2: According to Newland's octaves, elements show similar properties after every:
 - A) 6 elements
 - B) 7 elements
 - C) 8 elements
 - D) 9 elements
- Q3: Which is a limitation of Mendeleev's periodic table?
 - A) Inclusion of isotopes
 - B) Position of hydrogen
 - C) Use of atomic number
 - D) Explanation of periodicity

Answer Key

- Q1: B - Middle element's atomic weight is arithmetic mean of other two.
- Q2: C - Every 8th element has similar properties.
- Q3: B - Hydrogen's position was anomalous in Mendeleev's table.

Memory Aids

- Dobereiner's triads: Middle atomic weight = average of outer two.
- Newland's octaves: Every 8th element repeats properties like musical notes.
- Mendeleev's table: Left gaps for undiscovered elements; adjusted atomic weight order.

Modern Periodic Law

Explanation

- Henry Moseley (1913) showed atomic number (Z) is fundamental, not atomic mass.
- Plot of $\sqrt{\nu}$ (X-ray frequency) vs atomic number is linear.
- Modern periodic law: Properties of elements are periodic functions of atomic numbers.
- Elements arranged by increasing atomic number show repeating properties at regular intervals.

Solved Examples

- **Ex 1:** Moseley's experiment showed atomic number is more fundamental than atomic mass by plotting $\sqrt{\nu}$ vs Z, resulting in a straight line.

Practice Set

- Q1: What did Moseley's X-ray spectra study prove?
 - A. Atomic mass is fundamental
 - B. Atomic number is fundamental
 - C. Atomic volume is fundamental
 - D. Atomic radius is fundamental
- Q2: Modern periodic law states properties are periodic functions of?
 - A. Atomic mass
 - B. Atomic number
 - C. Atomic volume
 - D. Atomic radius
- Q3: Which plot gave a straight line in Moseley's experiment?

- A. $\sqrt{\nu}$ vs atomic mass
- B. $\sqrt{\nu}$ vs atomic number
- C. Frequency vs atomic mass
- D. Frequency vs atomic number

Answer Key

- Q1: B - Moseley proved atomic number is fundamental, not atomic mass.
- Q2: B - Modern periodic law is based on atomic number.
- Q3: B - Plot of $\sqrt{\nu}$ vs atomic number was linear.

Memory Aids

- "Moseley's Law": Atomic number defines element properties.
- Modern periodic law = Properties repeat with atomic number, not mass.

Periods and Groups

Explanation

- Periodic table has 7 horizontal rows called **periods**.
- There are 18 vertical columns called **groups**.
- Period number = highest principal quantum number (n) of elements in that period.
- Elements in a group have similar chemical properties.

Solved Examples

- **Ex 1:** Period 3 elements have highest principal quantum number $n = 3$.

Practice Set

- Q1: How many periods are in the long form periodic table?

- A. 5
- B. 6
- C. 7
- D. 8
- Q2: What does the period number represent?
 - A. Number of valence electrons
 - B. Highest principal quantum number (n)
 - C. Number of groups
 - D. Atomic number
- Q3: How many groups are there in the long form periodic table?
 - A. 16
 - B. 17
 - C. 18
 - D. 19

Answer Key

- Q1: C - There are 7 periods in the long form table.
- Q2: B - Period number equals highest principal quantum number.
- Q3: C - There are 18 groups in the long form table.

Memory Aids

- Periods = Rows, Groups = Columns.
- Period number = highest n (shell) filled.
- Groups have similar properties.

Periodic Table Periods

Explanation

- Periods correspond to electron filling of orbitals.
- Period 1: 2 elements, fills 1s orbital.
- Period 2: 8 elements, fills 2s and 2p orbitals.
- Period 3: 8 elements, fills 3s and 3p orbitals.

- Period 4 & 5: 18 elements, d orbitals start filling.
- Period 6 & 7: 32 elements, include f orbitals (lanthanoids and actinoids).
- Length of periods increases due to more orbitals involved.
- Lanthanides and actinides placed separately at bottom.

Period	(n)	Orbital filled up	Number of elements	Atomic number From	Atomic number To	Remark
First	(1)	1s	2	H(1)	He(2)	Very short period
Second	(2)	2s 2p	8	Li(3)	Ne(10)	Short period
Third	(3)	3s 3p	8	Na(11)	Ar(18)	Short period
Fourth	(4)	4s 3d 4p	18	K(19)	Kr(36)	Long period
Fifth	(5)	5s 4d 5p	18	Rb(37)	Xe(54)	Long period
Sixth	(6)	6s 4f 5d 6p	32	Cs(55)	Rn(86)	Very long period
Seventh	(7)	7s 5f 6d 7p	32	Fr(87)	Og(118)	Very long period

Solved Examples

- **Ex 1:** Period 4 has 18 elements because 4s, 3d, and 4p orbitals fill.

Practice Set

- Q1: How many elements are in period 2?
 - A. 2
 - B. 8
 - C. 18
 - D. 32
- Q2: Which orbitals fill in period 3?
 - A. 1s
 - B. 2s, 2p

- C. 3s, 3p
- D. 4s, 3d
- Q3: Why do periods increase in length from period 4 onwards?
 - A. More s orbitals fill
 - B. d and f orbitals start filling
 - C. Less electrons
 - D. Atomic number decreases

Answer Key

- Q1: B - Period 2 has 8 elements filling 2s and 2p orbitals.
- Q2: C - Period 3 fills 3s and 3p orbitals.
- Q3: B - d and f orbitals start filling, increasing period length.

Memory Aids

- Period length = number of orbitals filled.
- Period 1 = 2 elements (1s), Period 2 & 3 = 8 elements (s and p orbitals).
- Period 4 & 5 = 18 elements (d orbitals), Period 6 & 7 = 32 elements (f orbitals).

Classification of Elements

Explanation

- Elements classified as metals, non-metals, and metalloids.
- Zig-zag line separates metals (left/center) and non-metals (right).
- Metals: pink color, include alkali metals, alkaline earth metals, transition metals.
- Non-metals: pink color on right, include noble gases.
- Metalloids: grey color along zig-zag line, have mixed properties.
- Lanthanides and actinides are inner transition elements placed separately at bottom.
- Classification helps predict element properties by position.

The zig-zag line separates the metals from the non-metals.

Metals																		Metalloids										Non-metals						Noble gases	
Periods (No. of elements)	s-block elements (Representative elements)										d-block elements (Transition elements)										p-block elements (Representative elements)										Noble gases				
1 (2) (₁ H – ₂ He)	1 H																	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar												
2 (8) (₃ Li – ₁₀ Ne)	3 Li	4 Be																	5 B	6 C	7 N	8 O	9 F	10 Ne											
3 (8) (₁₁ Na – ₁₈ Ar)	11 Na	12 Mg	3	4	5	6	7	8	9	10	11	12	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar																	
4 (18) (₁₉ K – ₃₆ Kr)	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																	
5 (18) (₃₇ Rb – ₅₄ Xe)	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																	
6 (32) (₅₅ Cs – ₈₆ Rn)	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																	
7 (32) (₈₇ Fr – ₁₁₈ Og)	87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og																	
f-block elements (Inner-transition elements)																																			
Lanthanoid series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu																					
Actinoid series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr																					

Solved Examples

- **Ex 1:** Alkali metals are in group 1, leftmost column, and are metals.

Practice Set

- Q1: Which side of the periodic table contains metals?
 - A. Right
 - B. Left and center
 - C. Only center
 - D. Bottom
- Q2: Metalloids are located along?
 - A. Left edge
 - B. Right edge
 - C. Zig-zag line
 - D. Bottom rows
- Q3: Which group contains noble gases?
 - A. Group 1

- B. Group 2
- C. Group 17
- D. Group 18

Answer Key

- Q1: B - Metals are on left and center of the table.
- Q2: C - Metalloids lie along the zig-zag line.
- Q3: D - Noble gases are in group 18.

Memory Aids

- Metals = left/center, Non-metals = right, Metalloids = zig-zag line.
- Group 1 = Alkali metals, Group 18 = Noble gases.

Nomenclature of Elements with Atomic Numbers > 100

System Overview

- Elements with atomic number (Z) > 100 are named using the IUPAC systematic method.
- Names are derived from the atomic number using Latin and Greek roots for each digit.
- The suffix **-ium** is added to the derived name.
- Element symbols are formed from the first letter of each numerical root word.

Solved Examples

- **Ex 1:** Name element with $Z = 102$.
 - Digits: 1 (un), 0 (nil), 2 (bi)
 - Name: Unnilbium
 - Symbol: Unb (U from un, n from nil, b from bi)
- **Ex 2:** Name element with $Z = 115$.
 - Digits: 1 (un), 1 (un), 5 (pent)
 - Name: Ununpentium
 - Symbol: Uup

Practice Set

- Q1: What is the IUPAC name for element with atomic number 104?
 - A. Unquadium
 - B. Unnilquadium
 - C. Unquadnilium
 - D. Unquadium
- Q2: The symbol for element 113 is:
 - A. Uut
 - B. Uup
 - C. Uus
 - D. Uuo
- Q3: Which suffix is added to the systematic name of elements with $Z > 100$?
 - A. -ium
 - B. -ine
 - C. -on
 - D. -ide

Answer Key

- Q1: A. Unquadium - Derived from digits 1 (un), 0 (nil), 4 (quad) + ium suffix.
- Q2: A. Uut - Symbol formed from un (U), un (u), tri (t) for 113.
- Q3: A. -ium - Standard suffix for elements with atomic number > 100 .

Memory Aids

- Remember digit roots: 0 = nil, 1 = un, 2 = bi, 3 = tri, 4 = quad, 5 = pent, 6 = hex, 7 = sept, 8 = oct, 9 = enn.
- Symbol letters come from first letter of each digit root.
- Always add suffix **-ium** to complete the element name.

Electronic Configurations and Periodic Table

Electronic Configuration Basics

- Electronic configuration: distribution of electrons in atomic orbitals.
- Element's position in periodic table reflects quantum numbers of last filled orbital.
- Period number = principal quantum number (n) of valence shell.
- Successive periods fill higher principal energy levels (n=1, 2, 3...).

Period	Orbitals being filled (with no. of electrons)	Elements
First	$1s^{1-2}$	${}_1\text{H}$ and ${}_2\text{He}$
Second	$2s^{1-2} 2p^{0-6}$	${}_3\text{Li}$ to ${}_{10}\text{Ne}$
Third	$3s^{1-2} 3p^{0-6}$	${}_{11}\text{Na}$ to ${}_{18}\text{Ar}$
Fourth	$4s^{1-2} 3d^{0-10} 4p^{0-6}$	${}_{19}\text{K}$ – ${}_{36}\text{Kr}$
Fifth	$5s^{1-2} 4d^{0-10} 5p^{0-6}$	${}_{37}\text{Rb}$ – ${}_{54}\text{Xe}$
Sixth	$6s^{1-2} 4f^{0-14} 5d^{0-10} 6p^{0-6}$	${}_{55}\text{Cs}$ – ${}_{86}\text{Rn}$
Seventh	$7s^{1-2} 5f^{0-14} 6d^{0-10} 7p^{0-6}$	${}_{87}\text{Fr}$ – ${}_{118}\text{Og}$

Solved Examples

- **Ex 1:** Identify valence shell for period 3 element.
Solution: Period 3 $\rightarrow n=3$, valence shell is 3s and 3p orbitals.
- **Ex 2:** Write electronic configuration for element in period 2.
Solution: Period 2 $\rightarrow n=2$, fill 2s and 2p orbitals, e.g., Carbon: $1s^2 2s^2 2p^2$.

Practice Set

- Q1: Which principal quantum number corresponds to period 4?
A. 2
B. 3
C. 4
D. 5
- Q2: Which orbitals are filled in period 3?
A. 3s only

- B. 3s and 3p
 - C. 3p and 3d
 - D. 3d only
- Q3: Element with configuration ending in $4s^2 3d^5$ belongs to which period?
 - A. 3
 - B. 4
 - C. 5
 - D. 6

Answer Key

- Q1: C. Period number equals principal quantum number n .
- Q2: B. Period 3 fills 3s and 3p orbitals.
- Q3: B. 4s and 3d orbitals filled $\rightarrow$ period 4.

Memory Aids

- Period number = valence shell principal quantum number (n).
- Orbitals fill in order: $s \rightarrow p \rightarrow d \rightarrow f$ as period increases.

Groupwise Electronic Configuration

Group Similarities

- Elements in same group have similar valence shell configurations.
- Same number of electrons in outermost orbitals $\rightarrow$ similar chemical properties.
- Example: Group 1 (alkali metals) have ns^1 valence configuration.

Atomic number	Symbol	Electronic configuration
3	Li	$1s^2 2s^1$ (or) $[\text{He}]2s^1$
11	Na	$1s^2 2s^2 2p^6 3s^1$ (or) $[\text{Ne}]3s^1$
19	K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$ (or) $[\text{Ar}]4s^1$
37	Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^1$ (or) $[\text{Kr}]5s^1$
55	Cs	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6 6s^1$ (or) $[\text{Xe}]6s^1$
87	Fr	$[\text{Rn}]7s^1$

Solved Examples

- Ex 1:** Write valence configuration of sodium (Na).

Solution: Na atomic number 11 → valence shell $3s^1$.

- Ex 2:** Identify group of element with valence configuration $4s^1$.

Solution: Group 1 (alkali metals).

Practice Set

- Q1: Which valence configuration corresponds to Group 1?

A. ns^2
 B. ns^1
 C. np^5
 D. nd^{10}

- Q2: Lithium (Li) electronic configuration is?

A. $1s^2 2s^1$
 B. $1s^2 2p^1$
 C. $1s^2 2s^2$
 D. $1s^2 2p^6$

- Q3: Which element has valence configuration $3s^1$?

A. Sodium (Na)
 B. Magnesium (Mg)
 C. Aluminium (Al)
 D. Neon (Ne)

Answer Key

- Q1: B. Group 1 elements have ns^1 valence configuration.
- Q2: A. Lithium: $1s^2 2s^1$.
- Q3: A. Sodium has valence $3s^1$.

Memory Aids

- Group 1 = ns^1 valence electrons.
- Valence electrons determine group properties.

Periodic Table Groups and Names

Group Names and Properties

- Periodic table columns = groups with similar properties.
- Group 1: Alkali metals (ns^1 valence).
- Group 2: Alkaline earth metals (ns^2 valence).
- Group 11: Coinage metals.
- Group 15: Pnictogens.
- Group 16: Chalcogens.
- Group 17: Halogens.
- Group 18: Noble gases (aerogens).

Group 1 elements	Alkali metals
Group 2 elements	Alkaline earth metals
Group 11 elements	Coinage metals
Group 15 elements	Pnictogens
Group 16 elements	Chalcogens
Group 17 elements	Halogens
Group 18 elements	Noble gases or Aerogens

Solved Examples

- **Ex 1:** Name group 17 elements.
Solution: Halogens.
- **Ex 2:** Identify group of noble gases.
Solution: Group 18.

Practice Set

- Q1: Which group is called alkaline earth metals?
A. Group 1
B. Group 2
C. Group 15
D. Group 18
- Q2: Group 16 elements are known as?
A. Halogens
B. Pnictogens
C. Chalcogens
D. Noble gases
- Q3: Coinage metals belong to which group?
A. Group 11
B. Group 12
C. Group 13
D. Group 14

Answer Key

- Q1: B. Group 2 are alkaline earth metals.
- Q2: C. Group 16 are chalcogens.
- Q3: A. Group 11 are coinage metals.

Memory Aids

- Group 1: Alkali metals (ns^1).
- Group 2: Alkaline earth metals (ns^2).
- Group 17: Halogens ($ns^2 np^5$).
- Group 18: Noble gases (full valence shell).

Blocks of Elements

Classification by Orbital Type

- Elements classified into s-, p-, d-, and f-blocks based on orbitals being filled.
- s-block: Groups 1 and 2, filling s orbitals.
- p-block: Groups 13 to 18, filling p orbitals.
- d-block: Transition metals, filling d orbitals.
- f-block: Lanthanides and actinides, filling f orbitals.

Solved Examples

- **Ex 1:** Identify block of element with valence $4s^2 3d^5$.
Solution: d-block (transition metal).
- **Ex 2:** Which block contains noble gases?
Solution: p-block.

Practice Set

- Q1: Which block contains alkali metals?
A. s-block
B. p-block
C. d-block
D. f-block
- Q2: Lanthanides belong to which block?
A. s-block
B. p-block
C. d-block
D. f-block
- Q3: Elements filling p orbitals are in which groups?
A. 1-2
B. 3-12
C. 13-18
D. Lanthanides

Answer Key

- Q1: A. Alkali metals are s-block elements.
- Q2: D. Lanthanides are f-block elements.
- Q3: C. p-block elements are in groups 13-18.

Memory Aids

- s-block: Groups 1-2, ns electrons.
- p-block: Groups 13-18, np electrons.
- d-block: Transition metals, (n-1)d electrons.
- f-block: Lanthanides/actinides, (n-2)f electrons.

s-Block Elements

Characteristics

- General outer electronic configuration: ns^{1-2} .
- Soft metals with low melting and boiling points.
- Highly reactive with low ionization enthalpies.
- Form +1 ions (alkali metals) or +2 ions (alkaline earth metals) by losing outermost electrons.
- Highly electropositive and strong reducing agents.
- Metallic character and reactivity increase down the group.
- Never found pure in nature due to high reactivity.
- Compounds are mostly ionic except lithium and beryllium.

Solved Examples

Ex 1: Write the electronic configuration of Calcium (Ca).

- Atomic number = 20
- Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Practice Set

- **Q1:** What is the general electronic configuration of s-block elements?
 - A. ns^{1-2}
 - B. $ns^2 np^{1-6}$
 - C. $(n-1)d^{1-10} ns^{0-2}$
 - D. $(n-2)f^{0-14} (n-1)d^{0-1} ns^2$
- **Q2:** Which ion is formed by alkali metals?
 - A. +2
 - B. +1
 - C. -1
 - D. 0
- **Q3:** Which of the following is NOT a characteristic of s-block elements?

- A. Soft metals
- B. High ionization enthalpy
- C. Highly electropositive
- D. Form predominantly ionic compounds

Answer Key

- **Q1:** A. ns^{1-2} - This is the defining electronic configuration of s-block elements.
- **Q2:** B. +1 - Alkali metals lose one electron to form +1 ions.
- **Q3:** B. High ionization enthalpy - s-block elements have low ionization enthalpies.

Memory Aids

- s-block = Groups 1 & 2, valence electrons in s-orbital.
- Alkali metals form +1 ions; alkaline earth metals form +2 ions.
- Soft metals, highly reactive, never found free.

p-Block Elements

Characteristics

- General outer electronic configuration: $ns^2 np^{1-6}$.
- Includes non-metals, poor metals, and metalloids; non-metals predominate.
- Exhibit covalency, electrovalency, and coordinate valency.
- Show allotropy (e.g., carbon, silicon, phosphorus, sulfur).
- Noble gases ($ns^2 np^6$) are inert due to full valence shell.
- Halogens (Group 17) and chalcogens (Group 16) have high negative electron gain enthalpies and gain electrons to achieve noble gas configuration.
- Non-metallic character increases left to right; metallic character increases down the group.
- s- and p-block elements are called representative or main group elements.

Solved Examples

Ex 1: Write the electronic configuration of Chlorine (Cl).

- Atomic number = 17
- Configuration: $1s^2 2s^2 2p^6 3s^2 3p^5$

Practice Set

- **Q1:** What is the general electronic configuration of p-block elements?
 - A. ns^{1-2}
 - B. $ns^2 np^{1-6}$
 - C. $(n-1)d^{1-10} ns^{0-2}$
 - D. $(n-2)f^{0-14} (n-1)d^{0-1} ns^2$
- **Q2:** Which group contains the halogens?
 - A. Group 16
 - B. Group 17
 - C. Group 18
 - D. Group 15
- **Q3:** Why are noble gases inert?
 - A. They have empty valence shells
 - B. They have completely filled valence shells ($ns^2 np^6$)
 - C. They have low ionization energy
 - D. They are metals

Answer Key

- **Q1:** B. $ns^2 np^{1-6}$ - Defines p-block elements.
- **Q2:** B. Group 17 - Halogens belong to group 17.
- **Q3:** B. Completely filled valence shells - Noble gases are inert due to stable electronic configuration.

Memory Aids

- p-block = Groups 13-18, valence electrons in p-orbitals.
- Halogens (Group 17) gain 1 electron; chalcogens (Group 16) gain 2 electrons.
- Noble gases are inert due to $ns^2 np^6$ configuration.

d-Block Elements

Characteristics

- General outer electronic configuration: $(n-1)d^{1-10} ns^{0-2}$.
- Elements of groups 3 to 12; called transition elements.
- Classified into 3d, 4d, 5d, and 6d series.
- High melting and boiling points; metals.
- Form coloured salts.
- Paramagnetic due to unpaired electrons.
- Form both ionic and covalent compounds.
- Good catalysts.
- Show variable oxidation states (e.g., Fe^{2+} , Fe^{3+} , Cu^+ , Cu^{2+}).
- Zn, Cd, Hg ($d^{10} s^2$) do not show typical transition properties.
- Properties intermediate between s- and p-block elements.

Solved Examples

Ex 1: Write the electronic configuration of Iron (Fe).

- Atomic number = 26
- Configuration: $[Ar] 3d^6 4s^2$

Practice Set

- **Q1:** What is the general electronic configuration of d-block elements?
 - A. ns^{1-2}

- B. $ns^2 np^{1-6}$
- C. $(n-1)d^{1-10} ns^{0-2}$
- D. $(n-2)f^{0-14} (n-1)d^{0-1} ns^2$
- **Q2:** Which of the following is NOT a property of d-block elements?
 - A. Form coloured salts
 - B. Show variable oxidation states
 - C. Are non-metals
 - D. Are good catalysts
- **Q3:** Which elements do NOT show typical transition element properties?
 - A. Fe, Cu
 - B. Zn, Cd, Hg
 - C. Sc, Ti
 - D. Mn, Co

Answer Key

- **Q1:** C. $(n-1)d^{1-10} ns^{0-2}$ - Defines d-block elements.
- **Q2:** C. Are non-metals - d-block elements are metals.
- **Q3:** B. Zn, Cd, Hg - These have $d^{10} s^2$ configuration and lack typical transition properties.

Memory Aids

- d-block = Groups 3-12, valence electrons in d-orbitals.
- Variable oxidation states and coloured compounds are key features.
- Zn, Cd, Hg are exceptions with filled d-orbitals.

f-Block Elements

Characteristics

- General outer electronic configuration: $(n-2)f^{0-14} (n-1)d^{0-1} ns^2$.
- Called inner-transition elements; divided into Lanthanoids and Actinoids (14 elements each).
- Heavy metals with high melting and boiling points.

- Exhibit variable oxidation states; commonly +3.
- Form coloured salts due to light absorption.
- All actinoids are radioactive.

Solved Examples

Ex 1: Identify the series to which Uranium (U) belongs.

- Atomic number = 92
- Belongs to Actinoids series of f-block elements.

Practice Set

- **Q1:** What is the general electronic configuration of f-block elements?
 - A. ns^{1-2}
 - B. $ns^2 np^{1-6}$
 - C. $(n-1)d^{1-10} ns^{0-2}$
 - D. $(n-2)f^{0-14} (n-1)d^{0-1} ns^2$
- **Q2:** Which series are f-block elements divided into?
 - A. Lanthanoids and Actinoids
 - B. Alkali and Alkaline earth metals
 - C. Halogens and Noble gases
 - D. Transition and Representative elements
- **Q3:** What is a common oxidation state of f-block elements?
 - A. +1
 - B. +2
 - C. +3
 - D. +4

Answer Key

- **Q1:** D. $(n-2)f^{0-14} (n-1)d^{0-1} ns^2$ - Defines f-block elements.
- **Q2:** A. Lanthanoids and Actinoids - The two series of f-block elements.
- **Q3:** C. +3 - Most common oxidation state in f-block.

Memory Aids

- f-block = inner transition metals, valence electrons in f-orbitals.
- Lanthanoids and Actinoids each have 14 elements.
- Actinoids are radioactive; common oxidation state +3.

Metals, Non-metals and Metalloids

Metals

- More than 78% of elements are metals, located on the left side of the periodic table.
- Usually solids at room temperature (exceptions: mercury liquid; gallium and caesium low melting points).
- High melting and boiling points.
- Good conductors of heat and electricity.
- Malleable and ductile.

Non-metals

- Located on the top right side of the periodic table.
- Usually solids or gases at room temperature with low melting and boiling points (exceptions: boron, carbon).
- Poor conductors of heat and electricity.
- Most non-metallic solids are brittle, not malleable or ductile.

Metalloids or Semi-metals

- Elements like silicon, germanium, arsenic, antimony, and tellurium show properties of both metals and non-metals.
- Called semi-metals or metalloids.

Periodicity

Concept

- Periodicity: Recurrence of similar properties after definite intervals in increasing atomic number order.
- Caused by repetition of similar valence shell electronic configurations at intervals: 2, 8, 8, 18, 18, 32.
- Observed in properties linked to electronic configuration.

Solved Examples

- Ex 1: Identify periodicity cause in elements: Valence shell electronic configuration repeats every 2, 8, 8, 18, 18, 32 elements.

Practice Set

- Q1: What causes periodicity in elements?
 - A. Atomic mass variation
 - B. Repetition of valence electronic configuration
 - C. Nuclear charge variation
 - D. Isotopic abundance
- Q2: Periodicity intervals in electronic configuration are?
 - A. 2, 8, 8, 18, 18, 32
 - B. 2, 8, 18, 18, 32, 32
 - C. 2, 8, 8, 16, 18, 32
 - D. 2, 8, 8, 18, 16, 32
- Q3: Periodicity is observed in properties linked to?
 - A. Atomic mass
 - B. Electronic configuration
 - C. Isotopes
 - D. Atomic volume

Answer Key

- Q1: B - Periodicity arises due to repetition of valence electronic configuration.
- Q2: A - Correct intervals are 2, 8, 8, 18, 18, 32.
- Q3: B - Properties linked to electronic configuration show periodicity.

Memory Aids

- Periodic intervals: 2, 8, 8, 18, 18, 32 (electron shell capacities).
- Periodicity = repetition of valence shell configuration.

Atomic Radius

Types and Trends

- Atomic radius decreases across a period (left to right).
- Atomic radius increases down a group (top to bottom).
- Types of atomic radius:
 - **Metallic radius:** Half distance between nearest atoms in metallic lattice (metals).
 - **Covalent radius:** Half distance between nuclei of two atoms joined by single covalent bond (non-metals).
 - **van der Waals radius:** Half distance between adjacent atoms of neighboring molecules.
- van der Waals radius > metallic radius > covalent radius.
- Noble gases have van der Waals radii as they do not form molecules.

Solved Examples

- Ex 1: Compare radii types for chlorine: Covalent radius < Metallic radius < van der Waals radius.

Practice Set

- Q1: Atomic radius trend across a period is?
 - A. Increases
 - B. Decreases
 - C. Remains constant
 - D. Random
- Q2: Which radius is largest for an atom?
 - A. Covalent radius
 - B. Metallic radius
 - C. van der Waals radius
 - D. Ionic radius
- Q3: Atomic radius increases when moving?
 - A. Left to right in a period
 - B. Top to bottom in a group
 - C. Right to left in a period
 - D. Bottom to top in a group

Answer Key

- Q1: B - Atomic radius decreases across a period due to increasing nuclear charge.
- Q2: C - van der Waals radius is largest.
- Q3: B - Atomic radius increases down a group due to added shells.

Memory Aids

- Radius order: van der Waals > metallic > covalent.
- Across period: radius decreases; down group: radius increases.

Ionic Radius

Definition and Trends

- Ionic radius: Distance from nucleus to outermost electron shell in an ion.

- Follows same trend as atomic radius.
- Cations are smaller than parent atoms (fewer electrons, same nuclear charge).
- More electrons removed $\rightarrow$ smaller cation (e.g., $\text{Mg} > \text{Mg}^+ > \text{Mg}^{2+}$).
- Anions are larger than parent atoms (more electrons, same nuclear charge).
- Size explained by Z/e ratio (Nuclear charge / number of electrons): higher ratio $\rightarrow$ smaller size.
- Isoelectronic species: Same electron number, different nuclear charge; size decreases with increasing nuclear charge.

Solved Examples

- Ex 1: Arrange isoelectronic ions by size: $\text{N}^{3-} > \text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+} > \text{Al}^{3+}$.

Practice Set

- Q1: Which ion is smallest?
 - A. Mg^{2+}
 - B. Na^+
 - C. F^-
 - D. O^{2-}
- Q2: Cation size compared to parent atom is?
 - A. Larger
 - B. Smaller
 - C. Same
 - D. Variable
- Q3: Isoelectronic species have?
 - A. Same nuclear charge
 - B. Same number of electrons
 - C. Same size
 - D. Same atomic number

Answer Key

- Q1: A - Mg^{2+} has highest nuclear charge, smallest size.

- Q2: B - Cations are smaller due to electron loss.
- Q3: B - Isoelectronic species have same electron count.

Memory Aids

- Cations smaller than parent atoms; anions larger.
- Size $\propto 1 / (Z/e)$ ratio.
- Isoelectronic size decreases with increasing nuclear charge.

Ionisation Enthalpy

Definition and Factors

- Energy required to remove an electron from a gaseous atom in ground state.
- Reaction: $M_g + IE \rightarrow M_g^+ + e^-$
- Successive ionisation energies increase: $IE_3 > IE_2 > IE_1$.
- Screening effect: Inner electrons repel valence electrons, reducing nuclear attraction.
- Factors affecting IE:
 - Size: $IE \propto 1 / \text{size}$
 - Nuclear charge: $IE \propto \text{nuclear charge}$
 - Screening: $IE \propto 1 / \text{screening}$
 - Stable orbitals (half/full filled) increase IE.
 - Orbital type: IE decreases $s > p > d > f$.
- IE decreases down group; increases across period.
- Inert gases have exceptionally high IE due to stable configuration.
- Slater's Rule: $Z_{\text{eff}} = Z - \sigma^*$; σ^* calculated from electron distribution.
- Atomic radius $\propto 1 / Z_{\text{eff}}$, $IE \propto Z_{\text{eff}}$.

Solved Examples

- Ex 1: Calculate Z_{eff} for an electron using Slater's rule.

Practice Set

- Q1: Ionisation enthalpy is the energy to?
 - A. Add an electron
 - B. Remove an electron
 - C. Break a bond
 - D. Form an ion
- Q2: Which factor increases ionisation enthalpy?
 - A. Larger atomic size
 - B. Higher nuclear charge
 - C. More screening
 - D. Filled orbitals
- Q3: Ionisation enthalpy trend down a group is?
 - A. Increases
 - B. Decreases
 - C. Constant
 - D. Random

Answer Key

- Q1: B - IE is energy to remove an electron.
- Q2: B - Higher nuclear charge increases IE.
- Q3: B - IE decreases down group due to size and screening.

Memory Aids

- $IE \propto Z_{\text{eff}} / \text{size}$.
- Successive IE increases: $IE_{n+1} > IE_n$.
- Stable half/full orbitals raise IE.

Electron Gain Enthalpy

Definition and Factors

- Change in enthalpy when an electron is added to a gaseous atom: $A_g + e^- \rightarrow A_g^-$.
- Factors affecting electron gain enthalpy ($\Delta_{eg}H$):
 - Size: $\Delta_{eg}H \propto 1 / \text{size}$
 - Nuclear charge: $\Delta_{eg}H \propto \text{effective nuclear charge}$
 - Screening: $\Delta_{eg}H \propto 1 / \text{screening}$
- $\Delta_{eg}H$ becomes less negative down group; more negative across period (except noble gases).
- F and O have less negative $\Delta_{eg}H$ than Cl and S due to electron-electron repulsion in smaller orbitals.
- Elements with half/full orbitals (Be, Mg, N, inert gases) have positive $\Delta_{eg}H$.

Solved Examples

- Ex 1: Explain why $\Delta_{eg}H$ of F is less negative than Cl.

Practice Set

- Q1: Electron gain enthalpy is energy change when?
 - A. Electron is removed
 - B. Electron is added
 - C. Atom ionizes
 - D. Atom forms bond
- Q2: Electron gain enthalpy becomes more negative?
 - A. Down group
 - B. Across period
 - C. Randomly
 - D. For noble gases
- Q3: Why is $\Delta_{eg}H$ of O less negative than S?
 - A. Larger size of O
 - B. Electron repulsion in smaller orbitals
 - C. Higher nuclear charge of O

- D. Screening effect

Answer Key

- Q1: B - Electron gain enthalpy is energy change on adding electron.
- Q2: B - Becomes more negative across period.
- Q3: B - Electron-electron repulsion in smaller $n=2$ orbitals reduces negativity.

Memory Aids

- $\Delta_{\text{eg}}H \propto 1/\text{size}$, $\propto$ nuclear charge, $\propto 1/\text{screening}$.
- Less negative down group, more negative across period.
- Repulsion lowers negativity in small orbitals (O, F).

Electronegativity

Definition and Trends

- Tendency of atom to attract shared electrons in covalent bond.
- Increases across period (left to right).
- Decreases down group (top to bottom).
- Measured by Pauling, Mulliken, and Allred-Rochow scales.
- Bond ionic/covalent character depends on electronegativity difference.

Solved Examples

- Ex 1: Calculate electronegativity difference using Pauling scale from bond energies.

Practice Set

- Q1: Electronegativity trend across period is?
 - A. Decreases

- B. Increases
 - C. Constant
 - D. Random
- Q2: Which scale uses average of IE and EA?
 - A. Pauling
 - B. Mulliken
 - C. Allred-Rochow
 - D. Born-Haber
- Q3: Bond is mostly ionic if electronegativity difference is?
 - A. 0.5
 - B. 1.0
 - C. 1.9
 - D. >1.9

Answer Key

- Q1: B - Electronegativity increases across period.
- Q2: B - Mulliken scale averages IE and EA.
- Q3: D - Difference >1.9 indicates mostly ionic bond.

Memory Aids

- Electronegativity: Increases across period, decreases down group.
- Pauling scale uses bond energies.
- Mulliken = $(IE + EA)/2$.

Chemical Properties

Valence and Oxidation States

- Valence usually equals number of outer electrons or 8 minus outer electrons.
- Oxidation state often used as valence.
- Transition elements and actinoids show variable valencies.
- Valency increases 1 to 4 then decreases 4 to 0 across period.
- Valence electrons same down group; valency same in group.

Anomalous Properties of Second Period

- First elements of groups 1, 2, 13–17 differ from rest (Li, Be, B–F).
- Li and Be show diagonal relationship with Mg and Al.
- Due to small size, high charge/radius ratio, high electronegativity.

Chemical Reactivity Trends

- Reactivity highest at extremes of periodic table, lowest in center.
- Alkali metals reactive due to electron loss; halogens due to electron gain.
- Metallic character decreases left to right; non-metallic increases.
- Metallic character increases down group; non-metallic decreases.
- Oxides: Left extreme form basic oxides (Na_2O), right extreme acidic (Cl_2O_7), center amphoteric/neutral (Al_2O_3 , CO).

Solved Examples

- Ex 1: Identify oxide nature of Al_2O_3 : Amphoteric.

Practice Set

- Q1: Valency of elements in a group?
 - A. Varies
 - B. Same
 - C. Increases down group
 - D. Decreases down group
- Q2: Which element shows diagonal relationship?
 - A. Na
 - B. Be
 - C. Mg
 - D. Al
- Q3: Oxide of element on extreme right of period is?
 - A. Basic
 - B. Amphoteric

- C. Acidic
- D. Neutral

Answer Key

- Q1: B - Valency same down group due to same valence electrons.
- Q2: B - Be shows diagonal relationship with Mg.
- Q3: C - Extreme right oxides are acidic.

Memory Aids

- Valency same in group; varies across period.
- Diagonal relationship: Li-Mg, Be-Al.
- Oxides: Left basic, right acidic, center amphoteric/neutral.

General Trends

Summary of Periodic Properties

- Effective nuclear charge (Z_{eff}) increases across period.
- Atomic size decreases across period; increases down group.
- Ionisation and electron gain enthalpy increase across period.
- Electronegativity increases across period; decreases down group.
- Screening effect slightly increases down group.
- Metallic character decreases across period; increases down group.
- Oxidising nature increases across period; reducing nature decreases.
- Basic character of oxides decreases across period; acidity increases.
- Thermal stability of carbonates decreases across period.

Solved Examples

- Ex 1: Predict trend of ionisation enthalpy from Na to Ar: Increases.

Practice Set

- Q1: Atomic size trend across period?
 - A. Increases
 - B. Decreases
 - C. Constant
 - D. Random
- Q2: Metallic character trend down group?
 - A. Increases
 - B. Decreases
 - C. Constant
 - D. Random
- Q3: Oxide nature across period?
 - A. Basic to acidic
 - B. Acidic to basic
 - C. Neutral only
 - D. Amphoteric only

Answer Key

- Q1: B - Atomic size decreases across period.
- Q2: A - Metallic character increases down group.
- Q3: A - Oxides change from basic to acidic across period.

Memory Aids

- Across period: size ↓, IE ↑, EN ↑, metallic character ↓.
- Down group: size ↑, metallic character ↑, IE ↓.
- Oxides: basic → amphoteric → acidic across period.