CHEMISTRY

Unit 1: Periodic Table and Periodic Properties

Complete MCQs • Short Questions • Long Questions

Compiled & Presented by

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SECTION 1

Multiple Choice Questions

47 MCQs with correct answers, reasons & solutions

Exercise MCQs

1. Which scientist first time observed the periodicity in the elements?

(a) J. Newlands (b) L. Meyer (c) Dobereiner (d) D. Mendeleev

✓ **Correct Answer: (c) Dobereiner**

2. Recognize the element if it has 3 electron shells, belongs to "s" block and has 2 electrons in its outermost shell.

(a) Calcium (b) Sodium (c) Magnesium (d) Potassium

✓ **Correct Answer: (c) Magnesium**

3. Which one do you think is correct about metallic character?

(a) It decreases from top to bottom in a group (b) It increases from top to bottom in a group (c) It remains constant from left to right in a period (d) It increases from left to right in a period

✓ **Correct Answer: (b) It increases from top to bottom in a group**

4. Which property increases as you go down a group in the periodic table?

(a) Atomic radius (b) Electron Affinity (c) Electronegativity (d) Ionization Energy

✓ **Correct Answer: (a) Atomic radius**

5. Which set of the following conditions results in higher ionization energy?

(a) Smaller atom and greater nuclear charge (b) Smaller atom and smaller nuclear charge (c) Larger atom and greater nuclear charge (d) Larger atom and the smaller nuclear charge

✓ **Correct Answer: (a) Smaller atom and greater nuclear charge**

6. Which of the following atoms show more than one (variable) oxidation states?

(a) Sodium (b) Magnesium (c) Aluminium (d) Phosphorous

✓ **Correct Answer: (d) Phosphorous**

7. Which is the correct general trend in the variation of electron affinity in a group?

(a) It becomes less negative from top to bottom (b) It becomes more negative from top to bottom (c) It remains the same (d) It has no definite trend and changes irregularly

✓ **Correct Answer: (a) It becomes less negative from top to bottom**

8. What is the oxidation state of sulfur in the sulfate ion SO4(2-)?

(a) +4 (b) +2 (c) +6 (d) 0

✓ **Correct Answer: (c) +6**

Solution: Let O.N of S = x | x + 4(-2) = -2 | x - 8 = -2 | x = +6

9. Which is the correct trend in variation of electronegativity along a period of the periodic table?

(a) It decreases from left to right across a period (b) It increases from left to right across a period (c) It remains constant (d) It has no definite trend

✓ **Correct Answer: (b) It increases from left to right across a period**

10. The atomic radius generally _______ across a period in the periodic table.

(a) Increases (b) Decreases (c) Remains constant (d) First increases then decreases

✓ **Correct Answer: (b) Decreases**

11. Which one of the following elements has the highest ionization energy?

(a) Sodium (Na) (b) Magnesium (Mg) (c) Aluminium (Al) (d) Argon (Ar)

✓ **Correct Answer: (d) Argon (Ar)**

SLO Based Multiple Choice Questions

Modern Periodic Table and its History

12. Which one is the correct statement among the following?

(a) Anionic radius is generally smaller than atomic radius (b) Cationic radius is generally bigger than atomic radius (c) Cationic radius is generally smaller than atomic radius (d) Both anionic and cationic radii are smaller than atomic radius

✓ **Correct Answer: (c) Cationic radius is generally smaller than atomic radius**

13. Modern periodic table provides information about _______ elements.

(a) 100 (b) 110 (c) 118 (d) 130

✓ **Correct Answer: (c) 118**

14. By 1700 A.D. only _______ elements were recognized.

(a) 6 (b) 12 (c) 10 (d) 3

✓ **Correct Answer: (b) 12**

15. Concept of Triads was given by:

(a) Mendeleev (b) L. Meyer (c) Newlands (d) Dobereiner

✓ **Correct Answer: (d) Dobereiner**

16. John Newlands gave the idea of:

(a) Law of triads (b) Law of Octaves (c) Modern periodic law (d) Curves between weight and volume

✓ **Correct Answer: (b) Law of Octaves**

17. Who is considered as the father of the periodic table?

(a) Moseley (b) Newland (c) Mendeleev (d) Dobereiner

✓ **Correct Answer: (c) Mendeleev**

18. Select metalloid from the following:

(a) Po (b) P (c) Cl (d) Fe

✓ **Correct Answer: (a) Po (Polonium)**

19. Group 16 elements are also called:

(a) Pnictogens (b) Chalcogens (c) Halogens (d) Alkali metals

✓ **Correct Answer: (b) Chalcogens**

20. Select which is not alkaline earth metal:

(a) Be (b) Co (c) Sr (d) Ba

✓ **Correct Answer: (b) Co (Cobalt)**

21. The term "Halogen" means:

(a) Ash former (b) Salt former (c) Copper giver (d) Sulphur giver

✓ **Correct Answer: (b) Salt former**

22. Noble gases are present in 18th group, also called:

(a) IA (b) IIB (c) VIIIA (d) VIIIB

✓ **Correct Answer: (c) VIIIA**

23. Which ion will have greater size?

(a) Na+ (b) Mg2+ (c) Al3+ (d) Si4+

✓ **Correct Answer: (a) Na+**

Reason: All are iso-electronic species with 10 electrons. Size decreases with increasing nuclear charge.

Ionization Energy, Electron Affinity and Electronegativity

24. Which of the following has highest ionization energy?

(a) P (b) S (c) Cl (d) Ar

✓ **Correct Answer: (d) Ar**

25. Select the element which has highest 1st ionization energy:

(a) Boron (b) Carbon (c) Nitrogen (d) Oxygen

✓ **Correct Answer: (c) Nitrogen**

Reason: Nitrogen has stable half-filled 2p3 configuration.

26. Cl has more electron affinity value than F and its value is:

(a) +328 kJ/mol (b) +349 kJ/mol (c) -328 kJ/mol (d) -349 kJ/mol

✓ **Correct Answer: (d) -349 kJ/mol**

27. 2nd value of electron affinity:

(a) will be negative always (b) will always be positive (c) equals to 1st value (d) double than 1st value

✓ **Correct Answer: (b) will always be positive**

Reason: Adding electron to negatively charged ion requires energy.

28. L. Pauling developed a scale for:

(a) Ionization energy (b) Electron affinity (c) Electronegativity (d) None

✓ **Correct Answer: (c) Electronegativity**

29. Highest electronegative element is:

(a) F (b) Cl (c) Br (d) I

✓ **Correct Answer: (a) F**

30. Least electronegative element is:

(a) H (b) He (c) Cs (d) Li

✓ **Correct Answer: (c) Cs**

31. Electronegativity value of "Cs" is:

(a) 4 (b) 2.1 (c) 0.1 (d) 0.79

✓ **Correct Answer: (d) 0.79**

Metallic and Non-Metallic Character

32. Tendency to lose electrons is called:

(a) Electronegativity (b) Ionization energy (c) Electron affinity (d) Metallic character

✓ **Correct Answer: (d) Metallic character**

33. Metallic character increases:

(a) Downward and forward from left to right (b) Upward and forward from left to right (c) Downward and backward from right to left (d) Upward and backward from right to left

✓ **Correct Answer: (c) Downward and backward from right to left**

34. The most reactive metals of the periodic table are:

(a) Coinage metals (b) Rare earth metals (c) Alkali metals (d) Alkaline metals

✓ Correct Answer: (c) Alkali metals

Oxides and Hydroxides

35. Which of the following will not give hydroxide upon treating with water?

(a) Na(s) and $H_2O(l)$ (b) Mg(s) and $H_2O(g)$ (c) Mg(s) and $H_2O(l)$ (d) Ca(s) and $H_2O(l)$

✓ Correct Answer: (b) Mg(s) and $H_2O(g)$

Explanation: Mg reacts with steam $H_2O(g)$ to give MgO, not hydroxide. Mg reacts slowly with cold water to give $Mg(OH)_2$.

36. Peroxides are those in which oxidation state of oxygen is:

(a) -2 (b) -1 (c) -1/2 (d) +2

✓ Correct Answer: (b) -1

Example: Na_2O_2 (Sodium peroxide)

37. Which element is used in S.O.S flares and in fireworks?

(a) Na powder (b) $CaCO_3$ (c) Mg powder (d) KO_2

✓ Correct Answer: (c) Mg powder

38. Na burns with oxygen to give:

(a) Golden yellow flame (b) Golden brown flame (c) Crimson red flame (d) Intense white flame

✓ Correct Answer: (a) Golden yellow flame

39. Select the one which is acidic oxide:

(a) Na_2O (b) CaO (c) Al_2O_3 (d) CO_2

✓ Correct Answer: (d) CO_2

40. Metallic oxides are ________ in nature.

(a) Acidic (b) Basic (c) Amphoteric (d) Depend upon oxide

✓ Correct Answer: (b) Basic

41. Select basic oxide:

(a) CO_2 (b) SiO_2 (c) Al_2O_3 (d) MgO

✓ Correct Answer: (d) MgO

42. Which one is amphoteric oxide?

(a) CaO (b) MgO (c) Al_2O_3 (d) Na_2O

✓ Correct Answer: (c) Al_2O_3

43. Acidic oxides are:

(a) Metallic oxides (b) Non-metallic oxides (c) Metalloid oxides (d) All

✓ Correct Answer: (b) Non-metallic oxides

44. Acidic chlorides have pH:

(a) more than 7 (b) less than 7 (c) equals to 7 (d) may be any

✓ Correct Answer: (b) less than 7

45. Oxidation state of P in P4O10 is:

(a) +2 (b) +3 (c) +4 (d) +5

✓ Correct Answer: (d) +5

Solution: Let O.N of P = x | 4x + 10(-2) = 0 | 4x - 20 = 0 | x = +5

46. Element that can show variable oxidation states:

(a) Na (b) Mg (c) S (d) Ca

✓ Correct Answer: (c) S

47. Which oxidation state will not be shown by S?

(a) -2 (b) +4 (c) +2 (d) +8

✓ Correct Answer: (d) +8

Reason: Sulfur has maximum oxidation state of +6 (as in H2SO4 or SO3).

SECTION 2

Short Questions

Concise answers, equations & comparison tables

Exercise Short Question Answers

Q1. What is 1st ionization energy? Give an example.

The amount of energy which is required to remove one valence electron from one mole of isolated gaseous atoms to form mono-positive gaseous ions is called 1st ionization energy.

```
Na(g) → Na+(g) + e- dHi1 = +496 kJ/mol
```

Q2. Explain why sulfur has a lower first ionization energy than phosphorus?

Sulfur has four electrons in its 3p sub-shell while phosphorus has three electrons in its 3p sub-shell which makes phosphorus more stable than sulfur due to the half-filled stability rule. It is easy to remove one electron from sulfur but difficult to ionize phosphorus due to its stable half-filled configuration.

Element	Configuration
15P	[Ne] 3s2 3px1 3py1 3pz1
16S	[Ne] 3s2 3px2 3py1 3pz1

Q3. Why the elements in group 13 to 17 are called p-block elements?

The elements in group 13 to 17 are called p-block elements because their valence electrons are ended up in the p-subshell. According to valence electrons, the elements are classified as s, p, d and f-block elements.

Q4. What are the factors that affect electronegativity?

(i) Atomic Size - Greater the atomic size, lesser will be the electronegativity (E.N is proportional to 1/Atomic Size). (ii) Effective Nuclear Charge (Z-effect) - Greater the value of effective nuclear charge, larger will be the electronegativity (E.N is proportional to Effective Nuclear Charge).

Q5. What factors are responsible for the increasing reactivity of alkali metals as you move down the group?

The main factors responsible are: 1) Increase in atomic size down the group. 2) Low ionization energy down the group. As alkali metals have only one valence electron, a low ionization energy is required to remove it, so their reactivity increases. Increase in atomic size down the group also makes them more reactive.

Q6. Why some of the elements show variable oxidation numbers while others do not?

Some elements can show variable oxidation numbers because they can expand their octet by exciting their electrons to empty orbitals, if available. If elements do not have empty orbitals, they cannot show variable oxidation numbers. Na, Mg, Ca cannot show variable oxidation states because they lack empty d-orbitals.

Element	Compound	Oxidation State
P	PCl_3	+3
P	PCl_5	+5
Na	All compounds	+1 (fixed)
Mg	All compounds	+2 (fixed)
Ca	All compounds	+2 (fixed)

Q7. Identify the element which is in period 5 and group 15?

The element is in the 5th period, so its valence shell configuration ends with principal quantum number 5. Its group number is 15.

$X = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^3$

n = 5 ($5s^2\ 5p^3$); 5th period; 15th group ($5s^2\ 4d^{10}\ 5p^3$: 2+10+3=15). X = Sb (Antimony), atomic number 51.

Q8. Why the oxides of sodium and magnesium are more ionic than the oxides of nitrogen and phosphorus?

Na and Mg are present on the left side of the periodic table and are metals, so their oxides are more ionic in nature. Moving to the right side of the periodic table, ionic character decreases and the elements become non-metals, so ionic character decreases as in nitrogen and phosphorus oxides.

Q9. Give reasons for the different chemical reactivities of Na and Mg toward oxygen and chlorine.

Reaction with Oxygen - Sodium gives a golden yellow flame, forms peroxide in excess oxygen and normal oxide in limited oxygen. Magnesium gives an intense white flame and forms only normal oxide. Reaction with Chlorine - both react to give ionic chlorides.

$2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$ (excess oxygen - peroxide)

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$ (limited oxygen - normal oxide)

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ (white solid - normal oxide)

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$ (white solid)

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$ (white to colourless solid)

Q10. Why the ionization energy of lithium is much lower than that of helium despite the fact that the nuclear charge of lithium is +3 and that of helium is +2?

Lithium is a metal and has the tendency to lose an electron easily. Helium is a noble gas with a filled valence shell ($1s^2$), making it very stable and difficult to remove an electron from, so its ionization energy is much higher than lithium's despite its lower nuclear charge.

Q11. The ionization energy of Be (atomic no. 4) is higher than that of B (atomic no. 5) despite the fact that the nuclear charge of Be is +4 and that of B is +5.

The reason lies in the stability of the electronic configuration of Be, which has a filled 2s sub-shell, compared to the unstable configuration of B, which has one electron in its 2p sub-shell.

Element	Configuration
4Be	1s2 2s2 (stable - filled subshell)
5B	1s2 2s2 2px1 (unstable)

Q12. What is common in Na+, Mg2+, Al3+, Ne and F-? Arrange them in increasing order of sizes.

All these species are iso-electronic and have 10 electrons each. As nuclear charge increases, the electron cloud is pulled more tightly towards the nucleus, decreasing the ionic radius.

```
Al3+ < Mg2+ < Na+ < Ne < F- (increasing size)
```

Q13. Consider the chlorides of sodium, magnesium and phosphorus: NaCl, MgCl2 and PCl5. Classify each as acidic, basic or neutral and explain.

Neutral chlorides simply ionize in water without reacting and have pH near 7. Acidic chlorides react with water (hydrolysis) to produce H+ ions, giving acidic solutions; chlorides of Al to S (Group IIIA-VIA) are acidic.

```
NaCl(s) → Na+(aq) + Cl-(aq) (pH = 7, Neutral)

MgCl2(s) → Mg2+(aq) + 2Cl-(aq) (pH = 6.5, Neutral)

PCl5(s) + 4H2O(l) → H3PO4(aq) + 5HCl(aq) (pH = 0, Acidic - hydrolysis)
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SLO Based Short Question Answers - History of Periodic Table

Q14. Which elements were discovered by the year 1700 AD?

By 1700 A.D., only 12 elements were recognized: Gold (Au), Silver (Ag), Copper (Cu), Iron (Fe), Lead (Pb), Tin (Sn), Mercury (Hg), Phosphorus (P), Sulfur (S), Carbon (C), Zinc (Zn) and Arsenic (As).

Q15. What is Dobereiner's Triads?

In 1829, Dobereiner grouped three elements such that the atomic mass of the middle element was the average of the atomic masses of the first and third elements, noticing periodicity among elements.

```
7Li, 23Na, 39K → Atomic mass of Na = (7+39)/2 = 23
```

Q16. Explain Newland's Octaves.

After every eight elements, the next element shows periodicity with the first element when arranged by atomic mass; this is called Newland's Octaves. Every 8th element resembles the 1st element in properties.

Q17. State Mendeleev's periodic law.

In 1869, Dimitri Mendeleev stated: "Properties of the elements are the periodic functions of their atomic masses."

Q18. What do you know about Lother Meyer's curves law?

In 1869, Lother Meyer developed curves between atomic weight and atomic volume of elements, which also showed periodicity. Elements with similar properties occupied similar positions on the curve.

Q19. State Moseley's Periodic Law / Modern Periodic Law.

"Properties of the elements are the periodic functions of their atomic numbers." This is also called the Modern Periodic Law.

Q20. How many groups and periods are present in the modern periodic table?

There are 7 periods (horizontal rows) and 18 groups (vertical columns). The 18 groups are further divided into 8 groups of A-type and 10 groups of B-type.

Metals, Non-Metals and Metalloids

Q21. What are metals and non-metals? Give examples.

Metals are elements that have the tendency to lose electrons easily, e.g. Li, K, Cs, Na, Mg, Al. Non-metals are elements that have the tendency to accept electrons easily, e.g. N, O, F, Cl, S.

Q22. What are metalloids? Give their positions in the periodic table.

Metalloids are elements that have some properties of metals and some of non-metals; also called semi-metals. They occupy the p-block in a "stair-step line" from B (Group IIIA) to Po (Group VIA): B, Si, Ge, As, Sb, Te, Po.

Blocks in Periodic Table

Q23. Why are some elements called s-block elements and some p-block?

Elements are classified as s, p, d and f-block based on which sub-shell their valence electrons occupy. If valence electrons end up in the s-subshell, they are s-block elements; if in the p-subshell, they are p-block elements.

Q24. What are Lanthanides and Actinides?

Lanthanides are the series of 14 elements after Lanthanum (Z=57), from Ce to Lu. Actinides are the 14 elements after Actinium (Z=89), from Th to Lr. Both series are also called f-block elements.

Families in Periodic Table

Q25. Name some families present in the periodic table.

Some families are Alkali Metals, Alkaline Earth Metals, Transition Metals, Chalcogens, Halogens and Noble Gases.

Family	Elements
Alkali Metals	Li, Na, K, Rb, Cs, Fr
Alkaline Earth Metals	Be, Mg, Ca, Sr, Ba, Ra
Transition Metals	Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn
Chalcogens	O, S, Se, Te, Po, Lv
Halogens	F, Cl, Br, I, At, Ts
Noble Gases	He, Ne, Ar, Kr, Xe, Rn, Og

Q26. Why are group 16 elements called Chalcogens?

"Chalcogen" derives from the Greek "Chalcos" (copper) and "Gen" (form), meaning "copper giver". Elements: O, S, Se, Te, Po, Lv.

Q27. Why are group 17 elements called "Halogens"?

"Halogen" means "salt former" because these elements readily react with alkali and alkaline earth metals to form stable halide salts. Elements: F, Cl, Br, I, At, Ts.

Q29. Differentiate between Cations and Anions.

A cation carries positive charge and forms by losing an electron; its size is smaller than the parent atom (e.g. Li+, Na+). An anion carries negative charge and forms by gaining an electron; its size is greater than the parent atom (e.g. F-, Cl-).

Q30. Explain "Diagonal relationship" of the elements with a suitable example.

Elements diagonally positioned in the periodic table show similarities in physical and chemical properties despite being in different groups, e.g. Li and Mg, Be and Al, B and Si - similar atomic/ionic sizes, charge densities and polarizing powers.

Ionization Energy, Electron Affinity

Q31. Define ionization energy. Give an example.

The amount of energy required to remove an electron from the outermost shell of an isolated gaseous atom to form a positive ion.

```
Na(g) → Na+(g) + e- dHi1 = +496 kJ/mol
```

Q32. What is "Spin-pair repulsion"?

Electrons present in the same orbital experience repulsion; it is easier to remove a paired electron than an unpaired one. This is called spin-pair repulsion. Oxygen has lower ionization energy than nitrogen due to spin-pair repulsion in its 2p4 configuration.

```
N(g) → N+(g) + e- dHi1 = +1403 kJ/mol
```

```
O(g) → O+(g) + e- dHi1 = +1365 kJ/mol
```

Q33. Why is the ionization energy of Be (899 kJ/mol) higher than that of B (801 kJ/mol) although B is smaller than Be?

Be (1s2 2s2) has a stable filled s-subshell, while B (1s2 2s2 2px1) is relatively unstable due to one electron in its p-subshell. Hence Be has a higher ionization energy than B, even though B has a smaller atomic size.

Q34. Give the trend of electron affinity in a period.

Electron affinity increases across a period (left to right) due to: 1) smaller size 2) increase in effective nuclear charge.

```
Li < Be < B < C < N < O < F (increasing electron affinity, 2nd period)
```

Electronegativity

Q35. Define electronegativity. Give trend in the 2nd period.

Electronegativity is the power of an atom to attract the shared pair of electrons towards itself in a molecule. It increases from left to right across the 2nd period.

Element	Li	Be	B	C	N	O	F
E.N Value	1.0	1.5	2.0	2.5	3.0	3.5	4.0

Metallic Character

Q36. Define metallic character. Give its trend.

Metallic character is the tendency of an element to lose electrons. It increases down a group and decreases across a period.

Oxides

Q37. Na can form normal and peroxides. Give reactions.

Sodium reacts with O2 vigorously in air to give a peroxide, while in a limited amount of oxygen it forms a normal oxide.

```
2Na + O2 → Na2O2 (peroxide - excess oxygen)
4Na + O2 → 2Na2O (normal oxide - limited oxygen)
```

Q38. What are acidic oxides?

Non-metallic oxides are acidic because they form acids when they react with water.

```
CO2 + H2O → H2CO3 (carbonic acid)
SO2 + H2O → H2SO3 (sulphurous acid)
```

Q39. What are basic oxides?

Metallic oxides are basic because they form bases when they react with water.

$Na_2O + H_2O \rightarrow 2NaOH$ (sodium hydroxide)

$MgO + H_2O \rightarrow Mg(OH)_2$ (magnesium hydroxide)

Q40. What are amphoteric oxides?

Oxides that act as both an acid and a base are called amphoteric oxides, e.g. ZnO, Al_2O_3, BeO, PbO.

$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$ (acting as base)

$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$ (acting as acid)

Oxidation State

Q41. Calculate the oxidation state of P in P_4O_6 and in P_4O_{10}.

For P_4O_6: $4x + 6(-2) = 0$, $4x - 12 = 0$, $x = +3$. For P_4O_{10}: $4x + 10(-2) = 0$, $4x - 20 = 0$, $x = +5$.

Compound	Oxidation State of P
P_4O_6	+3
P_4O_{10}	+5

Descriptive Questions

Q3D. Write equations for the reactions of Na and Mg with oxygen, chlorine and water. Compare their reactivity in terms of metallic character.

$2Na + O_2 \rightarrow Na_2O_2$ (excess oxygen - peroxide)

$4Na + O_2 \rightarrow 2Na_2O$ (limited oxygen - normal oxide)

$2Mg + O_2 \rightarrow 2MgO$ (normal oxide)

$2Na + Cl_2 \rightarrow 2NaCl$

$Mg + Cl_2 \rightarrow MgCl_2$

$2Na + 2H_2O \rightarrow 2NaOH + H_2$ (up arrow)

$Mg + 2H_2O \rightarrow Mg(OH)_2 + H_2$ (up, slow with cold water)

$Mg + H_2O \rightarrow MgO + H_2$ (up, with steam)

Property	Sodium (Na)	Magnesium (Mg)
Metallic Character	More metallic	Less metallic
Reactivity	More reactive	Less reactive
With O_2	Vigorous, golden flame	Intense white flame
With H_2O	Violent reaction	Slow with cold, fast with steam
With Cl_2	Vigorous reaction	Vigorous reaction

Q4D. Explain with equations the acidic and basic behaviour of oxides and chlorides.

Non-metallic oxides are acidic; metallic oxides are basic; chlorides of non-metals hydrolyze to give acids while chlorides of metals simply ionize (neutral).

$CO_2 + H_2O \rightarrow H_2CO_3$; $SO_2 + H_2O \rightarrow H_2SO_3$; $N_2O_5 + H_2O \rightarrow 2HNO_3$ (acidic oxides)

$Na_2O + H_2O \rightarrow 2NaOH$; $CaO + H_2O \rightarrow Ca(OH)_2$; $MgO + H_2O \rightarrow Mg(OH)_2$ (basic oxides)

$PCl_5 + 4H_2O \rightarrow H_3PO_4 + 5HCl$; $SiCl_4 + 4H_2O \rightarrow H_4SiO_4 + 4HCl$ (acidic chlorides)

$NaCl \rightarrow Na^+ + Cl^-$; $MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$ (neutral chlorides)

Q5D. Describe the factors affecting and periodic trends of electron affinity.

Electron affinity is the energy released when an electron is added to a neutral gaseous atom to form a negative ion. Factors: (i) Atomic size - E.A is proportional to 1/Atomic Size (ii) Effective nuclear charge - E.A is proportional to nuclear charge (iii) Stable half-filled/fully-filled configurations lower electron affinity. It increases across a period (smaller size, greater nuclear charge) and decreases down a group (larger atomic size).

Element	Li	Be	B	C	N	O	F
E.A (kJ/mol)	59.6	~0	27	122	~0	141	328

Special cases: N and Be have very low electron affinity due to stable half-filled ($2p^3$) and filled ($2s^2$) configurations respectively.

Q6D. Define ionization energy. Discuss the factors affecting and periodic trends of ionization energy.

Ionization energy is the energy required to remove the most loosely bound (valence) electron from one mole of isolated gaseous atoms to form gaseous positive ions.

$X(g) \rightarrow X^+(g) + e^-$ dH_i = positive

Factors: (i) Atomic size - I.E is proportional to 1/Atomic Size (ii) Effective nuclear charge - I.E is proportional to nuclear charge (iii) Shielding effect - I.E is proportional to 1/Shielding Effect (iv) Half/fully-filled subshells increase stability and ionization energy. It increases across a period and decreases down a group.

Element	Li	Be	B	C	N	O	F	Ne
I.E (kJ/mol)	520	899	801	1086	1402	1314	1681	2081

Special cases: Be > B (Be has stable 2s2); N > O (N has stable half-filled 2p3).

Element	Li	Na	K	Rb	Cs	Fr
I.E (kJ/mol)	520	496	419	403	376	380

Summary of Periodic Trends

Property	Across a Period (L to R)	Down a Group (T to B)
Atomic Size	Decreases	Increases
Ionization Energy	Increases	Decreases
Electron Affinity	Increases	Decreases
Electronegativity	Increases	Decreases
Metallic Character	Decreases	Increases
Non-Metallic Character	Increases	Decreases

SECTION 3

Long Questions

Detailed explanations, Quick Checks & summary tables

Q1. Why was the periodic table created? Explain its historical background.

Reasons for the creation of the periodic table: 1) It is considered the symbol of chemistry. 2) It provides a systematic categorization and arrangement of elements, making the wide subject of chemistry easier to study. 3) Periodicity - properties of elements repeat after regular intervals. 4) It was a major turning point in the history of science. 5) All 118 elements are arranged in tabular form based on atomic number, electronic configuration and recurring chemical characteristics. 6) It offers a framework for researching the periodic behaviour of elements.

Historical Background of the Periodic Table

(i) Dobereiner's Triads (1829) - Dobereiner grouped elements into triads of three with similar properties, noting the atomic weight of the middle element was roughly the average of the other two, e.g. Li, Na, K.

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Atomic mass of Na = (7 + 39) / 2 = 23
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(ii) Newland's Law of Octaves (1864) - English chemist John Newland first observed periodicity in the 62 known elements. Properties of every eighth element were similar when arranged by increasing atomic mass.

(iii) Lother Meyer's Curves (1869) - Lother Meyer plotted atomic weight against atomic volume; these curves also showed periodicity.

(iv) Mendeleev's Periodic Table (1869) - Russian chemist Dmitri Mendeleev, considered the father of the periodic table, arranged 63 elements into 8 groups and 7 periods. His periodic law: properties of elements are periodic functions of their atomic masses. He famously left gaps for undiscovered elements, predicting their properties accurately.

(v) Modern Periodic Table (1913) - Moseley determined exact atomic numbers using X-ray emission, resolving flaws in Mendeleev's table by arranging elements by atomic number. Modern Periodic Law: properties of elements are periodic functions of their atomic numbers.

Q2. Explain the main features of the Modern Periodic Table.

Atomic Number - all 118 elements are arranged in ascending order of atomic number. Groups and Periods - 7 horizontal periods and 18 vertical groups (older versions: 8 A-groups and 10 B-groups). Properties - elements in the same group share similar chemical properties (same valence electrons) but show gradual change in physical properties top to bottom.

Location of Metalloids - Metalloids separate metals and non-metals via a "stair-step line" beginning at Boron (B) and extending to Polonium (Po), including Si, Ge, As, Sb and Te.

Q3. Explain the blocks in the periodic table.

Elements are classified based on the subshell containing their valence electrons, into four blocks: s-block (Groups 1-2, valence electrons in s-subshell, extreme left), p-block (Groups 13-18, valence electrons in p-subshell, right side), d-block (Groups 3-12, filling d-subshell, called transition elements, located between s and p blocks), f-block (Lanthanides and Actinides at the bottom, filling f-subshell).

Block	Groups	Subshell	Position
s	1-2	s	Extreme Left
p	13-18	p	Right Side
d	3-12	d	Middle (between s and p)
f	Lanthanides and Actinides	f	Bottom

Q4. What types of families are present in the periodic table?

An element family is a set of elements sharing common properties, placed in the same group.

Family	Elements
Alkali Metals	Li, Na, K, Rb, Cs, Fr
Alkaline Earth Metals	Be, Mg, Ca, Sr, Ba, Ra
Transition Metals	Sc, Ti, V, Cr, Mn, Fe, Co, Ni, Cu, Zn etc.
Chalcogens	O, S, Se, Te, Po, Lv
Halogens	F, Cl, Br, I, At, Ts
Noble Gases	He, Ne, Ar, Kr, Xe, Rn, Og

(i) Alkali Metals - Group 1 elements, called alkali metals because they produce alkalis when reacting with water, e.g. Na, K. Characteristics: one valence electron, low density, low melting point, low ionization energy, most reactive metals, soft.

(ii) Alkaline Earth Metals - Group 2 elements found in the earth and forming alkalis, e.g. Ca, Mg. Characteristics: two valence electrons (divalent), metallic solids, harder and denser than alkali metals, easily oxidized, high thermal/electrical conductivity, soft but harder than alkali metals.

(iii) Transition Elements - the largest family, middle of the table, comprising d-block (Groups 3-12) plus f-block Lanthanides and Actinides. Characteristics: high thermal/electrical conductivity, high melting points, high density, variable oxidation states, mostly coloured compounds, form complex compounds.

(iv) Chalcogens - Group 16, called chalcogens because most copper ores are oxides or sulphides (Greek: Chalkos=ore, gen=forming). O and S are non-metals, Se/Te/Po are metalloids, Lv is a metal. Characteristics: six valence electrons, mostly -2 oxidation state.

(v) Halogens - Group 17, non-metallic, "salt-former" as they react with alkali/alkaline earth metals to form stable halide salts, e.g. F, Cl, Br, I. Characteristics: highly reactive non-metals, high electron affinities, easily accept one electron to complete their outer shell.

(vi) Noble Gases - Group 18, unreactive elements at the extreme right, e.g. He, Ar, Kr. Characteristics: stable complete outer shell, almost entirely unreactive, rarely form compounds, mono-atomic. Note: some compounds exist, e.g. $XePtF_6$; "inert gases" was renamed "noble gases".

Quick Check 1.1

(a) Why are the elements in Groups 1 and 2 known as s-block elements?

Elements of Groups 1 and 2 have valence electrons in the s-sub-shell, so they are called s-block elements, e.g. Na, Mg.

(b) Name the elements in the chalcogen family. Give their two characteristics.

O, S, Se, Te, Po and Lv. Characteristics: 1) all have 6 valence electrons 2) all can accept 2 electrons to form -2 ions.

Q5. Explain how elements are arranged in the periodic table according to their electronic configuration.

Electronic configuration is the arrangement of electrons in sub-shells or orbitals around the nucleus. The periodic arrangement of elements provides information about physical properties (physical state and atomic radii), electronic structure, and chemical reactivity.

Q6. Define atomic radius and ionic radius. Give their factors and trends in the periodic table.

Atomic Radius - a measure of the size of an atom; half the distance between two identical bonded atoms. It varies with bond type (covalent, metallic, van der Waals) or state of the atom. Example: atomic radius of sodium is 186 pm.

Factor	Effect
Atomic Number	Increases effective nuclear charge
Effective Nuclear Charge	Greater charge = smaller radius
Shielding Effect of Inner Electrons	More shielding = larger radius

Periodic Trends - Across a period, atomic radius decreases (left to right) because increasing nuclear charge pulls the electron cloud closer. Down a group, atomic radius increases because additional electron shells add more shielding, outweighing the increase in nuclear charge.

Ionic Radius - the distance from the nucleus of an ion to the outermost electron shell. Example: cationic radius of K^+ is 133 pm. Cationic radius: when an atom loses electrons, it becomes smaller (less electron repulsion pulls remaining electrons closer). Anionic radius: when an atom gains electrons, it becomes larger (increased electron repulsion expands the electron cloud).

Periodic Trends in Ionic Radius - Across a period, both cationic and anionic radii decrease due to increasing nuclear charge. Down a group, both increase because principal quantum number (n) increases, adding electron shells and increasing distance from the nucleus.

Quick Check 1.3

(a) Which factors affect atomic and ionic radii?

1) Atomic number 2) Effective nuclear charge 3) Shielding effect. Increasing atomic number increases effective nuclear charge, so radii decrease across a period and increase down a group; shielding effect stays constant across a period but increases down a group.

(b) Using knowledge of Period 2 elements, predict and explain the relative sizes of:

(i) The atomic radii of lithium and fluorine

Both are in the same period; size decreases left to right, so Li (152 pm) is larger than F (64 pm).

(ii) A lithium atom and its ion, Li+

The cation is always smaller than the parent atom: Li = 152 pm, Li+ = 60 pm.

(iii) An oxygen atom and its ion, O2-

The anion is always larger than the parent atom: O = 66 pm, O2- = 140 pm.

(iv) A Nitride ion, N3-, and a fluoride ion, F-

N3- has a greater negative charge than F-, so N3- (171 pm) is larger than F- (136 pm).

Q7. Define and explain ionization energy. Write its factors and trends in the periodic table.

1st Ionization Energy - the energy needed to remove one electron from each atom in one mole of gaseous atoms to form one mole of gaseous 1+ ions.

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Na(g) → Na+(g) + e- dHi1 = 496 kJ/mol
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Ca(g) → Ca+(g) + e- dHi1 = 590 kJ/mol
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2nd Ionization Energy - removing a second electron from each ion in a mole of gaseous 1+ ions.

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Ca+(g) → Ca2+(g) + e- dHi2 = 1150 kJ/mol
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3rd Ionization Energy - removing a third electron from each ion in a mole of gaseous 2+ ions.

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Ca2+(g) → Ca3+(g) + e- dHi3 = 4940 kJ/mol
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Note: an element can have several ionization energies, the number corresponding to its atomic number.

Factor	Effect
Nuclear Charge	Greater effective nuclear charge = higher ionization energy
Size of Atom/Ion	Larger atom = lower ionization energy
Electronic Arrangement	Half-filled and fully-filled orbitals = higher ionization energy (more stable)

Periodic Trends - Across a period, ionization energy increases (nuclear charge increases, shells unchanged). Down a group, ionization energy decreases (atomic size, shells and shielding increase; nuclear charge effect decreases). Groups 13 and 16 show abnormalities.

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Li < B < Be < C < O < N < F < Ne (actual trend, 2nd period)
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Noble gases have the highest ionization energies (complete outer shell); alkali metals have the lowest.

Quick Check 1.4

(a) Explain with reasoning:

(i) 1st ionization energy of Boron is lesser than Beryllium.

Be has a stable filled 2s2 subshell, while B has one electron in an unstable 2p subshell, so Boron's ionization energy is lower.

(ii) 1st ionization energy of Aluminium is lower than Magnesium.

Al has one electron in its 3p subshell (unstable), while Mg has a filled 3s subshell (stable), so Aluminium's ionization energy is lower.

(b) What trend is observed in ionization energy down group 3? Give reason.

Ionization energy decreases down group 3 because: 1) atomic size increases 2) number of shells increases 3) effective nuclear charge decreases 4) shielding effect increases.

Q8. Define and explain electron affinity. Give its factors and trends in the periodic table.

1st Electron Affinity - the enthalpy change when 1 mole of electrons is added to 1 mole of gaseous atoms to form 1 mole of gaseous uni-negative ions.

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Cl(g) + e- → Cl-(g) dHea1 = -348.8 kJ/mol
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Since energy is released, the 1st electron affinity carries a negative sign. 2nd Electron Affinity - energy required to add an electron to 1 mole of uni-negative gaseous ions to form 1 mole of gaseous 2- ions.

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O(g) + e- → O-(g) dHea1 = -141 kJ/mol
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O-(g) + e- → O2-(g) dHea2 = +844 kJ/mol
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Net: dHea1 + dHea2 = (-141) + (844) = +701 kJ/mol
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Factor	Effect
Size of Atom	Smaller atom = greater electron affinity
Nuclear Charge	Greater nuclear charge = greater electron affinity
Electronic Configuration	Half-filled or fully-filled subshells = lower electron affinity (more stable)

Periodic Trends - Down a group, electron affinity generally decreases (larger atomic size, less attraction for incoming electron). Trend in halogens: At < I < Br < F < Cl. Across a period, electron affinity generally increases (smaller size, greater effective nuclear charge).

Group 17 Element	E.A (kJ/mol)	Group 1 Element	E.A (kJ/mol)
Fluorine	-328.0	Lithium	-60.0
Chlorine	-349.0	Sodium	-53.0
Bromine	-324.0	Potassium	-48.0
Iodine	-295.0	Rubidium	-47.0
Astatine	-270.1	Cesium	-46.0

Quick Check 1.5

(a) 1st electron affinity of Oxygen is -141 kJ/mol but 2nd is +844.0 kJ/mol. Explain.

Adding the 1st electron to form a uni-negative ion releases energy (negative enthalpy). Adding a 2nd electron to a uni-negative ion to form a di-negative ion requires energy (against repulsion), shown as positive.

(b) Which of nitrogen and phosphorus has the higher electron affinity? Justify.

Nitrogen, because electron affinity decreases down a group due to increasing atomic size, more shells, greater shielding and decreasing effective nuclear charge.

(c) F has lower electron affinity than Cl although its size is smaller. Explain why.

1) Small size of F's 2p orbital causes high electron density and repulsion for the incoming electron. 2) F has only two shells despite highest electronegativity. 3) Cl's larger 3p orbital has lower electron charge density, allowing easier addition of an electron.

(d) Why do noble gases (group 18) have positive 1st electron affinities? Explain in terms of electronic configuration.

Noble gases have a stable ns2np6 configuration (except He), a complete octet/duplet, so they never lose or gain electrons under ordinary conditions, giving them positive electron affinities.

Q9. Write a note on electronegativity in detail.

Electronegativity is the power of an atom to attract the shared pair of electrons toward itself in a molecule. Example: electronegativity of F is 4.0 on the Pauling scale (dimensionless, no unit). Linus Pauling's scale ranges from just below one for alkali metals to a maximum of four for fluorine; higher values mean stronger electron attraction.

Factors: (i) Atomic Size - larger size gives lower electronegativity, since electrons far from the nucleus feel a weaker pull, e.g. F > Cl > Br > I. (ii) Effective Nuclear Charge - higher effective nuclear charge gives greater electronegativity, e.g. Li (1.0) to F (4.0) across period 2.

Periodic Trends - Across a period (left to right), electronegativity increases (increasing nuclear charge, decreasing atomic size). Down a group, it decreases (increasing size and shielding effect), e.g. F(4.0) > Cl(3.0) > Br(2.8) > I(2.5).

Period	IA	IIA	IIIA	IVA	VA	VIA	VIIA
2	Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
3	Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
4	K 0.8	Ca 1.0	-	-	-	Br 2.8	-
5	Rb 0.8	Sr 1.0	-	-	-	I 2.5	-

Metals (left side) have lower electronegativity and are electropositive; non-metals (right side) have higher electronegativity and are electronegative.

Q10. Explain metallic character with its trend in the periodic table.

Metallic character is the tendency of elements to lose electrons. Elements on the left side of the table lose electrons easily to attain noble gas configuration and form positive ions (metals); elements on the right, especially the right corner, gain electrons and form negative ions (non-metals). It largely depends on valence shell electronic configuration.

Periodic Trends - Across a period, metallic character decreases left to right, since increasing nuclear charge pulls the electron cloud closer, making it harder to lose electrons. Down a group, metallic character increases, since increasing atomic size and shielding effect reduce nuclear attraction on valence electrons. Example: Cesium is far more reactive and electropositive than sodium or lithium.

Quick Check 1.6

(a) Illustrate how metallic character varies in group 14.

Metallic character increases down group 14 because of greater atomic size and higher shielding effect, so electrons can be removed more easily due to weaker nucleus-electron attraction.

(b) Identify semi-metals in groups 14, 15 and 16. Why are they semi-metals?

Group 14: Si, Ge. Group 15: As, Sb. Group 16: Te, Po. They are semi-metals (semi-conductors) because their properties lie between those of metals and non-metals.

Q11. Describe the reactions of Na and Mg with oxygen, chlorine and water.

(i) With Water - Sodium reacts vigorously with water to form sodium hydroxide and hydrogen gas. Magnesium reacts slowly with cold water to form magnesium hydroxide and hydrogen, but reacts vigorously with steam to form magnesium oxide and hydrogen.

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

(ii) With Oxygen - Sodium burns with a golden yellow flame to give a mixture of sodium oxide and sodium peroxide (kept under kerosene to prevent reaction with air). Magnesium burns with an intense white flame to give magnesium oxide, reacting more slowly than sodium.

$2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$ (excess oxygen - peroxide)

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$ (limited oxygen - normal oxide)

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(iii) With Chlorine - Both react exothermically to give soluble salts; sodium gives a golden yellow flame forming white NaCl, magnesium forms white $MgCl_2$. Magnesium powder burns rapidly with an intense white flame, useful in fireworks and S.O.S flares.

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

Quick Check 1.7

(a) What is the nature of oxides and hydroxides of Na and Mg?

Na reacts with oxygen more quickly than Mg, giving a golden yellow flame to form Na_2O_2 or Na_2O. The oxides and hydroxides of both Na and Mg are strongly alkaline.

$2Na(s) + O_2(g) \rightarrow Na_2O_2(s)$; $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$; $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$; $Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(aq) + H_2(g)$

(b) What could you predict about the reactivity of Ca, a group 2 element, when reacted with water and oxygen?

(i) With water: Ca reacts with water at 100C to give calcium hydroxide and hydrogen gas. (ii) With oxygen: Ca reacts with excess oxygen to give calcium oxide.

$Ca + 2H_2O \rightarrow (100C) \rightarrow Ca(OH)_2 + H_2$

$2Ca + O_2 \rightarrow 2CaO$

Q12. Explain the trends of bonding and classification of chlorides and oxides of period 3.

Oxides - binary compounds of oxygen with other elements. Groups 1, 2 and 3 oxides (Na_2O, MgO, Al_2O_3) are more ionic; Groups 4, 5, 6 and 7 oxides (SiO_2, P_2O_5, SO_2, SO_3) are more covalent, due to increasing electronegativity and decreasing ionic character.

(i) Basic Oxides - form an alkali when combined with water, e.g. Na_2O, CaO, BaO. Formed by metals reacting with oxygen; usually ionic. Group 1 and 2 form basic oxides; Group 2 hydroxide solubility (and alkalinity) increases down the group.

$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

(ii) Amphoteric Oxides - react with both acids and bases, e.g. Al_2O_3, ZnO, BeO.

$Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(s) + 3H_2O(l)$ (acting as base)

$Al_2O_3(s) + 2NaOH(aq) \rightarrow 2NaAlO_2(s) + H_2O(l)$ (acting as acid)

(iii) Acidic Oxides - form an acid when combined with water, e.g. P_2O_3, P_2O_5, SO_3, SO_2, SiO_2.

$P_2O_3(s) + H_2O(l) \rightarrow H_3PO_3(aq)$

$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

$SiO_2(s) + 2NaOH(aq) \rightarrow Na_2SiO_3(aq) + H_2O(l)$

$P_2O_3(s) + 6NaOH(aq) \rightarrow 2Na_3PO_3(aq) + 3H_2O(l)$

$SO_2(g) + 2NaOH(aq) \rightarrow Na_2SO_3(aq) + H_2O(l)$

Chlorides - binary compounds of chlorine. Groups 1, 2, 3 chlorides (NaCl, $MgCl_2$, $AlCl_3$) are predominantly ionic; Groups 4-7 chlorides ($SiCl_4$, PCl_3, PCl_5, SCl_2) are covalent, due to decreasing electronegativity difference between the halogen and the other atom.

Neutral Chlorides - dissolve to give a solution near pH 7, e.g. NaCl, $MgCl_2$.

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ (pH = 7)

$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$ (pH = 6.5)

Acidic Chlorides - react with water (hydrolysis) to give an acidic solution with pH less than 7, e.g. $AlCl_3$, $SiCl_4$, PCl_3, PCl_5.

$AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(aq)$

$SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(aq)$

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PCl3(l) + 3H2O(l) → H3PO3(aq) + 3HCl(aq)
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Oxide	Oxidation Number	Chloride	Oxidation Number
Na in Na_2O	+1	Na in NaCl	+1
Mg in MgO	+2	Mg in $MgCl_2$	+2
Al in Al_2O_3	+3	Al in $AlCl_3$	+3
Si in SiO_2	+4	Si in $SiCl_4$	+4
P in P_4O_{10}	+5	P in PCl_5	+5
P in P_4O_6	+3	P in PCl_3	+3
S in SO_3	+6	S in SCl_2	+2
S in SO_2	+4	-	-

Key Point: Phosphorus and sulfur show several oxidation numbers because they can expand their octet by exciting electrons into empty 3d orbitals.

Quick Check 1.9

(a) Calculate the oxidation number of sulphur in SO_2 and SO_3.

For SO_2: $x + 2(-2) = 0$, $x = +4$. For SO_3: $x + 3(-2) = 0$, $x = +6$.

(b) Why do some p-block elements show variable oxidation states?

Because of the presence of an empty d-subshell allowing octet expansion, e.g. P and S. S: -2 in H_2S, +4 in SO_2, +6 in SO_3. P: +3 in PCl_3, +5 in PCl_5.

Summary of Periodic Trends

Property	Across Period (L to R)	Down Group (T to B)
Atomic Radius	Decreases	Increases
Ionic Radius (Cation)	Decreases	Increases
Ionic Radius (Anion)	Decreases	Increases
Ionization Energy	Increases	Decreases
Electron Affinity	Increases	Decreases
Electronegativity	Increases	Decreases
Metallic Character	Decreases	Increases
Non-Metallic Character	Increases	Decreases