

Modern Periodic Table : (1912)

Modern Periodic Law :

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

SALIENT FEATURES OF MODERN PERIODIC TABLE

1. The classification of elements is based on increasing order of the **atomic number** of elements, which are more fundamental than the atomic mass.
2. The periodic table correlates the position of each element with its **electronic configurations**.
3. Elements belonging to the **same group** have a **same number of electrons** in their **valence shells** and show **similar chemical properties**.
4. It explains regular gradation in **chemical properties** of elements along a period in terms of their valency electrons. Hence it can **explain the cause of periodicity**.
5. **Inert gases**, in which valence shells are completely filled, have been placed at the end of each period.

DEMERITS OF MODERN PERIODIC TABLE

1. **Failed to give Proper Position for Hydrogen** as it exhibits properties similar to both alkali metals and halogens.
2. Failed to accommodate **Lanthanides and Actinides** in the main body of the periodic table and are placed separately at the bottom of the table.

CHARACTERISTICS OF LONG FORM OF PERIODIC TABLE**GROUPS :**

- The vertical columns in the periodic table are called **GROUPS**.
- There are eighteen groups in all (**1 – 18**).
- Elements having same number of **VALENCY ELECTRONS** are placed in the same group.
- Elements having same outer **ELECTRONIC CONFIGURATION** are placed in the same group.
- Elements placed in the same group show similar **CHEMICAL PROPERTIES**.
- Elements in **group 1 (I – A)** are called **ALKALI METALS**.
- Elements in **group 2 (II – A)** are called **ALKALINE EARTH METALS**.
- The elements placed in **group 17 (VII – A)** are called **HALOGENS**.
- The elements placed **group 18 (VIII – A)** are called **NOBLE GASES / INERT GASES**.
- The elements placed in groups 1, 2, 13, 14, 15, 16 and 17 are called **NORMAL ELEMENTS**.
- The elements placed groups 3, 4, 5, 6, 7, 8, 9, 10, 11 and 12 are called **TRANSITION ELEMENTS**.

Transition Elements

- Transition elements are placed in between **s - block** elements and **p - block** elements.
- Transition elements mostly form coloured compounds.
- Most of the Transition elements are used as catalysts.
- Transition elements have two incomplete outermost shells.

- Some elements belonging to **GROUP 3** with At. No. **57 to 71** and **89 to 103** are placed separately below in the periodic table and are called **INNER TRANSITION ELEMENTS**.
- Inner transition elements belonging to **GROUP 3** and **PERIOD 6** are called **LANTHANIDES**.
- Inner transition elements belonging to **GROUP 3** and **PERIOD 7** are called **ACTINIDES**.

MAJOR TYPES OF ELEMENTS

S - BLOCK ELEMENTS		D - BLOCK	P - BLOCK ELEMENTS					
Normal Elements			Normal Elements				Inert gases	
Active Metals								
Alkali Metals	Alkaline Earth Metals						Halogens	Noble Gases
Group 1	Group 2	3 - 12	13	14	15	16	Group 17	Group 18
Hydrogen							Helium	
Lithium	Beryllium	Transition Elements	B				Fluorine	Neon
Sodium	Magnesium			Si			Chlorine	Argon
Potassium	Calcium						Bromine	Krypton
Rubidium	Strontium						Iodine	Xenon
Caesium	Barium						Astatine	Radon
Francium	Radium							
F - BLOCK	Inner transition Elements	3 rd Group		6 th Period		Lanthanides		
		3 rd Group		7 th Period		Actinides		

PERIODS :

- The horizontal rows in the modern periodic table are called as **PERIODS**.
- There are **seven periods** in all in Modern Periodic Table.
- The **period number** equals the **number of electron shells** that elements in that period possess.
- All elements present in the same period have **same number of energy shells (electron shells)**.

	1	2		13	14	15	16	17	18
1	1 H								2 He
2	3 Li	4 Be		5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg		13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	Transition Elements	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr

Shortest period	1 st	2 elements	Hydrogen & Helium.
Short periods	2 nd	8 elements	Lithium – Neon
	3 rd	8 elements	Sodium – Argon
Long periods	4 th	18 elements	Potassium – Krypton
	5 th	18 elements	Rubidium – Xenon
Longest period	6 th	32 elements	Cesium – Radon + 14 Lanthanides
Longest period	7 th	32 elements	Francium – Oganesson + 14 Actinides

Electron Configuration

The **arrangement of electrons** of an atom in different energy levels is called **Electronic Configuration**.

- The electron configuration is useful in understanding the chemical properties of elements.
- The electron configuration is useful for describing the type of chemical bonds that atoms form.
- The elements with similar electronic configuration have similar chemical properties.
- The elements with similar electronic configuration occupy same group.

ELECTRONIC CONFIGURATION OF ELEMENTS BELONGING TO **GROUP – 1** AND **GROUP – 17**

1 st Group Elements	Electronic Configuration	Valency with respect to Hydrogen		Valency with respect to Oxygen	
H	1	+ 1	H ₂	+ 1	H ₂ O
Li	2, 1	+ 1	LiH	+ 1	Li ₂ O
Na	2, 8, 1	+ 1	NaH	+ 1	Na ₂ O
K	2, 8, 8, 1	+ 1	KH	+ 1	K ₂ O
Rb	2, 8, 18, 8, 1	+ 1		+ 1	Rb ₂ O
Cs	2, 8, 18, 18, 8, 1	+ 1		+ 1	Cs ₂ O

17 th Group Elements	Electronic Configuration	Valency Electrons	Valency with respect to					
			Hydrogen		Alkali metals		Oxygen	
F	2, 7	7	- 1	HF	- 1	NaF	7	F ₂ O ₇
Cl	2, 8, 7	7	- 1	HCl	- 1	NaCl	7	Cl ₂ O ₇
Br	2, 8, 18, 7	7	- 1	HBr	- 1	KBr	7	Br ₂ O ₇
I	2, 8, 18, 18, 7	7	- 1	HI	- 1	KI	7	I ₂ O ₇

ELECTRONIC CONFIGURATION OF ELEMENTS BELONGING TO PERIOD – 2 AND PERIOD – 3

	1 (I A)	2 (II A)	13 (III A)	14 (IV A)	15 (V A)	16 (VI A)	17 (VIIA)
2nd Period Elements	Li	Be	B	C	N	O	F
Electronic Configuration	2, 1	2, 2	2, 3	2, 4	2, 5	2, 6	2, 7
Valency w.r.t. Hydrogen	1	2	3	4	3	2	1
	LiH	BeH ₂	BH ₃	CH ₄	NH ₃	H ₂ O	HF
Valency w.r.t. Oxygen	1	2	3	4	5		7
	Li ₂ O	BeO	B ₂ O ₃	CO ₂	N ₂ O ₅	O ₂	F ₂ O ₇

3rd Period Elements	Na	Mg	Al	Si	P	S	Cl
Electronic Configuration	2, 8, 1	2, 8, 2	2, 8, 3	2, 8, 4	2, 8, 5	2, 8, 6	2, 8, 7
Valency w.r.t. Hydrogen	1	2	3	4	3	2	1
	NaH	MgH ₂	AlH ₃	SiH ₄	PH ₃	H ₂ S	HCl
Valency w.r.t. Oxygen	1	2	3	4	5	6	7
	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇

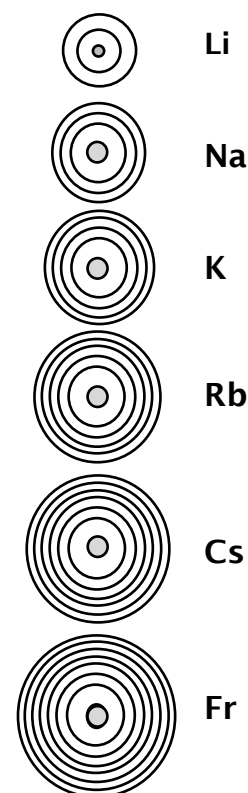
TRENDS IN PERIODIC TABLE
ATOMIC RADIUS / ATOMIC SIZE

The atomic radius of a chemical element is a measure of the size of its atoms, usually the mean distance from the centre of the nucleus to the boundary of the surrounding cloud of electrons.

→ **The size of atom increases as we move down the group.**

Reason :

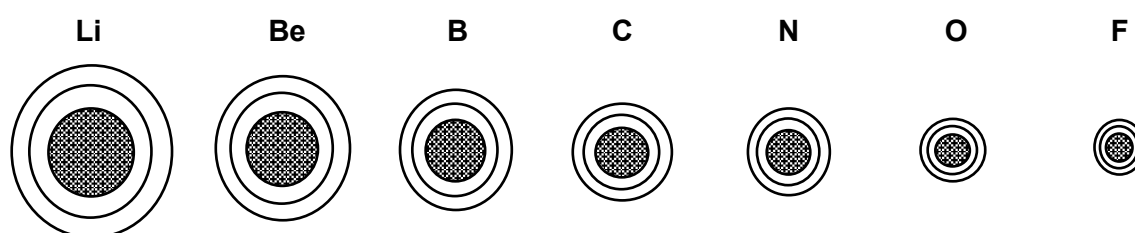
The number of electron shells increases as you move down a group. Therefore, the atomic radius increases as the energy levels increase.



→ **The size of atom decreases as we move across the period.**

Reason :

As we go across a period in a periodic table, electrons are added to the same energy level. At the same time, protons are being added to the nucleus. The concentration of more protons in the nucleus creates a "**higher effective nuclear charge.**" Therefore, a stronger force of attraction pulls the electrons closer to the nucleus, resulting in a smaller atomic radius.



→ **Noble gases have comparatively large atomic sizes. Give reason.**

Reason :

In noble gases, all the electron shells are completely filled, resulting in increased repulsions between electrons within the same shell. This leads to a weaker effective nuclear attraction on the outermost electrons compared to elements like halogens. Since the **van der Waals forces** between noble gas atoms are weak, this results in a larger distance between atoms, hence a larger atomic radius.

→ **Anions (— ve ions) are larger than their respective neutral atoms. Give reason.**

Reason :

When an atom gains electrons, the number of electrons exceeds the number of protons. As a result, the nuclear pull on each electron decreases, and increased electron–electron repulsion pushes the electrons farther apart. This causes the size of the anion to be larger than its neutral atom.

→ **Cations (+ ve ions) are smaller than their respective neutral atoms. Give reason.**

Reason :

When an atom loses electrons, the number of protons becomes greater than the number of electrons. This increases the effective nuclear pull on the remaining electrons, drawing them closer to the nucleus and reducing electron–electron repulsion. As a result, the size of the cation decreases.

METALLIC CHARACTER

The tendency of the atoms of an element to form positive ions by losing electrons from its outermost energy level is a measure of its metallic character.

→ **The metallic nature gradually decreases as we move across the period.**

Reason :

From left to right in a period, the nuclear charge increases while the atomic size decreases. The stronger attraction between the nucleus and the valence electrons makes it difficult for atoms to lose electrons. Since the tendency to lose electrons is the basis of metallic character, metallic nature decreases across a period.

2nd Period	Li	Be	B	C	N	O	F
Nature	Metal	Metal	Metalloid	Non metal	Non metal	Non metal	Non metal

3rd Period	Na	Mg	Al	Si	P	S	Cl
Nature	Metal	Metal	Metal	Metalloid	Non metal	Non metal	Non metal

→ The metallic nature gradually increases as we move down in a group.

Reason :

Down the group, atomic size increases due to the addition of new shells. As a result, the effective nuclear attraction on the valence electrons decreases. This makes it easier for atoms to lose electrons and form positive ions, which is the characteristic property of metals. Hence, metallic nature increases down the group.

Group – 14 (IV A)	
C → Carbon	Non-metal
Si → Silicon	Metalloid
Ge → Germanium	Metalloid
Sn → Tin	Metal
Pb → Lead	Metal

NON-METALLIC CHARACTER

The tendency of the atoms of an element to form negative ions by gaining electrons into its outermost energy level is a measure of its non-metallic character.

Group – 17 (VII A)	
F → Fluorine	Gas
Cl → Chlorine	Gas
Br → Bromine	Liquid
I → Iodine	Solid – with metallic lustre

→ Non-metallic character decreases as we move from top to bottom in a group.

Reason :

As we move down a group, the atomic radius increases due to the addition of new electron shells. This causes the effective nuclear attraction for incoming electrons to decrease. As a result, the tendency of atoms to gain electrons reduces, and hence, the non-metallic character decreases down the group.

→ Non-metallic character gradually increases as we move across the period.

Reason :

From left to right across a period, the nuclear charge increases while the atomic size decreases. The stronger attraction between the nucleus and the valence electrons makes it easier for atoms to gain electrons. Since the tendency to gain electrons is the basis of non-metallic character, non-metallic nature gradually increases across a period.

- **Metallic nature** → increases down a group & decreases across a period.
- **Non-metallic nature** → decreases down a group & increases across a period.

IONIZATION ENERGY

The amount of energy required to remove valence electron/s from a gaseous atom in its ground state to form a cation is called ionization energy / ionization potential.

- Metals have **low** ionization energy and
- Non-metals have **high** ionization energy.

→ **Ionization energy decreases going down in a group.**

Reason :

As we move down a group, the atomic radius increases due to the addition of new electron shells. The increased distance between the nucleus and the valence electrons reduces the effective nuclear attraction. As a result, it becomes easier to remove the outermost electron, and therefore, the ionization energy decreases.

→ **Ionization energy increases going from left to right in a period.**

Reason :

From left to right in a period, the atomic radius decreases while the nuclear charge increases. The stronger attraction between the nucleus and the valence electrons makes it more difficult to remove an electron. Therefore, the ionization energy increases across a period.

→ **Alkaline earth metals have greater ionization energy than alkali metals.**

Reason :

Alkali metals need to lose only one electron to form a stable cation, whereas alkaline earth metals need to lose two electrons. Additionally, alkaline earth metals have a smaller atomic radius and a higher nuclear charge than alkali metals in the same period, so their valence electrons are held more tightly. This makes it harder to remove electrons, giving them a higher ionization energy.

→ **The noble gases possess very high ionization energies**

Reason :

They have completely filled valence shells, which makes them extremely stable. As a result, they do not easily lose electrons, requiring a very large amount of energy to remove an electron.

ELECTRON AFFINITY

Electron affinity is the amount of energy released when an electron is added to a neutral atom in its isolated gaseous state to form a negative ion.

- Metals have **low** electron affinities and
- Non-metals have **high** electron affinities.

→ **Electron affinity decreases going down a group**

Reason :

As we move down a group, the atomic radius increases and the valence electrons are farther from the nucleus. This reduces the effective nuclear attraction on an incoming electron, making it more difficult for the atom to gain electrons. As a result, electron affinity decreases down the group.

→ **Electron affinity increases going from left to the right in a period**

Reason :

From left to right in a period, the atomic radius decreases and the nuclear charge increases. The stronger attraction between the nucleus and the valence shell makes it easier for atoms to gain electrons. As a result, electron affinity increases across the period.

→ **Electron affinity of noble gases is nearly zero because :**

Reason :

Noble gases already have completely filled valence shells, which makes them very stable. Since they do not need extra electrons to achieve stability, they show almost no tendency to gain electrons, so their electron affinity is nearly zero.

- Noble gases have **zero** electron affinities, because they are stable.
- Halogens have **higher** electron affinities value. (**chlorine** has highest electron affinity)
- Among nonmetals, the elements in the first period have **lower** electron affinities than the elements below them in their respective groups.

CHLORINE HAS HIGHER ELECTRON AFFINITY THAN FLUORINE. GIVE REASON.

Although fluorine is smaller and has a higher nuclear charge, its tiny size causes strong electron–electron repulsion when a new electron is added. Chlorine, being larger, has more space in its valence shell, so the added electron experiences less repulsion. This makes it easier for chlorine to gain an electron, giving it a higher electron affinity than fluorine.

ELECTRONEGATIVITY

Electronegativity is the relative tendency of an atom in a molecule to attract the shared pair of electrons more towards itself.

- Metals have low **electronegativity** and
- Non-metals have high **electronegativity**.

→ **Electronegativity decreases going down in a group. Give reason.**

Reason :

As we move down a group, the atomic radius increases and the valence electrons are farther from the nucleus. The weaker nuclear attraction makes it harder for an atom to pull the shared pair of electrons in a bond toward itself. Therefore, electronegativity decreases down the group.

→ **Electronegativity increases going from left to the right in a period. Give reason.**

Reason :

From left to right in a period, the atomic radius decreases while the nuclear charge increases. The stronger attraction between the nucleus and the valence electrons makes it easier for the atom to pull the shared pair of electrons in a bond toward itself. Therefore, electronegativity increases across the period.

ATOMS OF EQUAL ELECTRONEGATIVITY BOND TOGETHER AND FORM A NON-POLAR COVALENT BOND.

When two atoms have the same electronegativity, they attract the shared pair of electrons equally. As a result, the electrons are located midway between the two nuclei, forming a non-polar covalent bond.

ATOMS OF SMALL ELECTRONEGATIVITY DIFFERENCE LEADS TO A POLAR COVALENT BOND.

The more electronegative atom attracts the shared electron pair more strongly, gaining a partial negative charge (δ^-), while the less electronegative atom acquires a partial positive charge (δ^+). This separation of charges gives the bond a dipole character, making it polar.

ATOMS OF LARGE ELECTRONEGATIVITY DIFFERENCE LEADS TO AN IONIC BOND.

The more electronegative atom completely attracts the valence electron(s) from the less electronegative atom. This results in the complete transfer of electrons, forming positively charged cations and negatively charged anions, which are held together by strong electrostatic forces.

REACTIVITY :

Reactivity of an element is the measure of its ability to lose electrons from or gain electrons into the valance shell.

Metals have low ionization energy so it's fairly easy for them to lose electrons and metals react by losing electrons.

→ **The reactivity of metals increases as we move down the group.**

Reason :

As we move down a group, the atomic size increases and the valence electrons are farther from the nucleus. The weaker nuclear attraction makes it easier for atoms to lose electrons, which increases the reactivity of metals.

→ **The reactivity of metals decreases as we move left to right in a period.**

Reason :

From left to right across a period, the nuclear charge increases while the number of electron shells remains the same. This stronger attraction between the nucleus and valence electrons makes it harder for atoms to lose electrons. As a result, the reactivity of metals decreases across the period.

Most reactive metals: Caesium (Cs) and Francium (Fr)

→ **The reactivity of non-metals decreases as we move down the group.**

Reason :

As we move down a group, atoms have more electron shells, which increases the atomic size. The nucleus is farther from the valence shell, so its ability to attract electrons weakens. This makes it harder for non-metals to gain electrons, decreasing their reactivity.

→ **The reactivity of non-metals increases as we move left to right in a period.**

Reason :

From left to right across a period, the nuclear charge increases while the number of electron shells remains the same. This stronger attraction makes it easier for atoms to gain electrons. As a result, the reactivity of non-metals increases across the period.

Most reactive non-metal : Fluorine (F)

BASIC AND ACIDIC PROPERTIES
BASIC AND ACIDIC PROPERTIES OF OXIDES OF ELEMENTS IN GROUP - 1

<p>→ Gradual increase in metallic character down the group reflects gradual increase in basic nature of oxides of elements in that group.</p> <p>→ Gradual decrease in non-metallic character down the group reflects in gradual decrease in acidic nature of the oxides of the elements in that group.</p>	Oxides		Nature
	Li ₂ O	BeO	Least basic
	Na ₂ O	MgO	↓
	K ₂ O	CaO	
	Rb ₂ O	SrO	
	Cs ₂ O	BaO	Most basic

BASIC AND ACIDIC PROPERTIES OF OXIDES OF ELEMENTS IN PERIOD - 3

- Gradual increase in non-metallic character across the period reflects gradual increase in acidic nature of oxides of elements in that period.
- Gradual decrease in metallic character across the period reflects in gradual decrease in basic nature of the oxides of the elements in that period.

Oxides of 3 rd Period	Na ₂ O	MgO	Al ₂ O ₃	SiO ₂	P ₂ O ₅	SO ₃	Cl ₂ O ₇
Nature of oxides	Strongly basic	basic	Amphoteric	Mild acidic	Acidic	Highly acidic	Most acidic

HEAT AND ELECTRICITY CONDUCTIVITY

- Heat and electricity conductivity vary regularly down the group.
- Heat and electrical conductivity increase going across period 3 from sodium to aluminium, then decreases to silicon and then remaining elements have negligible conductivity.

Example :

3 rd Period	Na	Mg	Al	Si	P	S	Cl
Conductivity	Good conductor	Good conductor	Good conductor	Fair conductor	Poor conductor	Poor conductor	Poor conductor

MELTING AND BOILING POINTS
MELTING POINTS

Melting points are the amount of energy required to break a bond/s to change the solid phase of a substance to a liquid.

- Melting and boiling points decrease gradually down the group for metals
- Melting and boiling points increase gradually down the group for non-metals
- Melting and boiling points may increase gradually across a period then gradually decreases.

3rd Period	Na	Mg	Al	Si	P	S	Cl
M.P / B.P	Low	Medium	High	High	Medium	Medium	Low

- Metals generally possess a high melting point and boiling point.
- Transition metals have high melting / boiling point due to strong metallic bonds.
- Non-metals possess low melting points and boiling point.
- The metalloid boron also possesses a high melting point and boiling point.
- The non-metal carbon possesses the highest boiling point of all the elements.

→ Most reactive non-metal	:	Fluorine
→ Most electronegative element	:	Fluorine
→ Element with highest Ionization Energy in group 7	:	Fluorine
→ Radioactive element with the lowest electronegativity	:	Francium
→ Metal with the lowest Electronegativity	:	Caesium
→ Most reactive metal	:	Caesium
→ Most electropositive metal	:	Caesium
→ Metal element with Lowest Ionization Energy	:	Caesium
→ Element with higher Electron Affinity	:	Chlorine
→ Elements with Highest Ionization Energy	:	Helium
→ Metal in liquid state at room temperature	:	Mercury
→ Metal change to liquid state when handled	:	Gallium
→ Non-metal in liquid state at room temperature	:	Bromine

PROPERTIES OF ALKALI METALS (Group 1 Elements)

1. Alkali metals, found in **group 1** of the periodic table
2. Alkali metals have only **one electron** in their outer shell, hence placed in group 1.
3. Alkali metals are **very reactive metals** that do not occur freely in nature.
4. Alkali metals **react explosively** with water liberating **hydrogen gas**.
5. Alkali metals are **stored in kerosene oil** / mineral oil because of their high reactivity.
6. Alkali metals are characterized by their **soft** texture and **silvery** colour.
7. Alkali metals are **malleable, ductile**, and are **good conductors** of heat and electricity.
8. Alkali metals have **low boiling** and **low melting** points.
9. Boiling points and melting points decreases from top to bottom of the group
10. Alkali metals are **less dense** than most elements.
11. **Lithium, Sodium, and Potassium float** on water because of their low density.
12. Alkali metals have the lowest **electronegativity** and ionization energy.

PROPERTIES OF HALOGENS (Group 17 Elements)

1. Halogen means "**salt-forming tendency**" and halogens compounds are called "salts".
2. Halogens are five non-metallic elements found in **group 17** of the periodic table.
3. Halogens have **7 electrons** in their valence shell, and their oxidation number of **-1**.
4. Halogens are known to have very high degrees of **electron affinity**.
5. Halogens are considered to be the **most reactive** non-metals.
6. **Fluorine** is the most reactive halogen in this group.
7. Halogens are **diatomic** molecules at room temperature. i.e. **F₂, Cl₂, Br₂ & I₂**.
8. Hydrogen halides are extremely strong acids. i.e. **HI > HBr > HCl > HF**
9. Halogens exist, at room temperature, in all three states of matter.

→ Gases	– pale yellow gas	Fluorine	– pale green gas	Chlorine
→ Liquid	– reddish brown liquid	Bromine		
→ Solid	– dark purple lustrous solid	Iodine		

MAIN REASONS FOR GRADATIONS IN PROPERTIES
DOWN THE GROUP :

Atomic radius increases and nuclear force of attraction on valence electrons decreases

ACROSS THE PERIOD :

Atomic radius decreases and nuclear force of attraction on valence electrons increases

PROPERTY	ACROSS THE PERIOD	DOWN THE GROUP
1. NUMBER OF VALANCE ELECTRONS	INCREASES	REMAIN SAME
2. ATOMIC RADIUS	DECREASES	INCREASES
3. METALLIC CHARACTER	DECREASES	INCREASES
4. BASIC NATURE OF OXIDES	DECREASES	INCREASES
5. REDUCING NATURE	DECREASES	INCREASES

1. OXIDIZING NATURE	INCREASES	DECREASES
2. ACIDIC NATURE OF OXIDES	INCREASES	DECREASES
3. NON-METALLIC CHARACTER	INCREASES	DECREASES
4. ELECTRON AFFINITY	INCREASES	DECREASES
5. IONIZATION ENERGY	INCREASES	DECREASES
6. ELECTRONEGATIVITY	INCREASES	DECREASES

1. MELTING POINT	METALS	INCREASES	DECREASES
	NON-METALS	DECREASES	INCREASES
2. BOILING POINT	METALS	INCREASES	DECREASES
	NON-METALS	DECREASES	INCREASES
3. NUCLEAR CHARGE		INCREASE	INCREASE

