THE PERIODIC TABLE

BASIC CONCEPTS

- A vertical column of elements in the Periodic Table is called a **Group**.

- A horizontal row of elements in the Periodic Table is called a **Period**.

- An element that has the properties of both a metal and a non-metal is called a **Metalloid**.

- The Group I elements in the Periodic Table is called **Alkali Metals**.

- The Group VII elements in the Periodic Table is called **Halogens**.

- The Group 0 or VIII elements in the Periodic Table is called **Noble Gases**.

- A block of metallic elements between Groups II and III in the Periodic Table is called **Transition Elements**.

- A substance that increases (enhances) the speed of a chemical reaction is called a **Catalyst**.

ANALYSE:

- ✓ **Describe the Periodic Table**
- ✓ **Relate the position of an element in the Periodic Table to its portion number and electronic structure**
- ✓ **Recognize the similarities between the elements in the same group of the Periodic Table in terms of their electronic structure**

Periodic table is defined as " The arrangement of known elements in certain groups and periods in such a way that elements with similar properties are grouped together."

The early attempt to classify elements meaningfully was done by Dmitri Mendeleev, a Russian chemist in 1869. He classified the elements based on the periodic law. Mendeleev's periodic law states that "the physical and chemical properties of the elements are a periodic function of their atomic masses". In Mendeleev's periodic table when elements were arranged in the increasing order of their atomic masses the physical and chemical properties of elements got repeated after definite intervals. The elements in vertical rows were called groups and the elements in horizontal columns were called periods. In Mendeleev's table there were nine groups and seven periods. About 60 elements known at that time were arranged in the periodic table.

In 1913, Moseley by his studies on x-ray diffraction pointed that it is the atomic number not the atomic mass is responsible for the periodicity of chemical properties. He proposed the Modern Periodic Law. It states that:

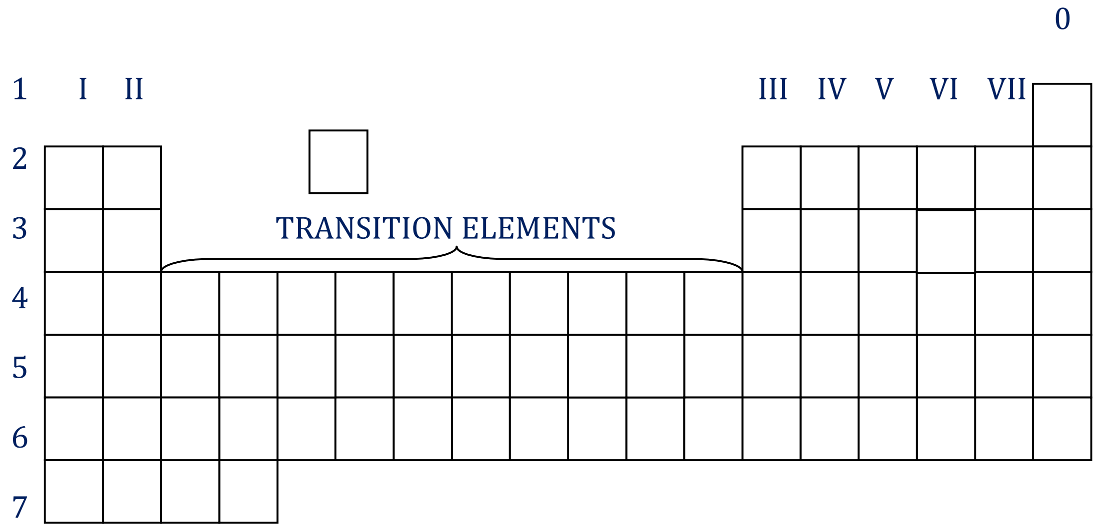
"The chemical properties of the elements are the periodic function of their atomic numbers".

1.1 FEATURES OF THE PERIODIC TABLE:

1. **The list of elements arranged in order of their increasing Proton(Atomic) Numbers in a table is called Periodic Table .**

A Period is a horizontal row of elements and a group is a vertical column of elements.

2. **The Periodic Table consists of Seven Periods and Eight Groups of elements.**



3. Periods:

a) The Periodic Table consists of seven periods of elements, numbered from 1 to 7.

b) The Period run horizontally from left to right.

4. Groups:

a) The Periodic Table has eight groups of elements, numbered from 1 to 0.

b) They run vertically from top to bottom.

c) Group 0 is also called group VIII.

TIPS FOR STUDENTS:

With the exception of beryllium, elements in groups I and II form ionic compounds. Elements of Groups V, VI and VII in Periods 2 and 3 can form both ionic and covalent compounds. For example, Na_2 is an ionic compound but H_2S is a covalent compound.

5. The block of element between Group II and Group III is called the transition

elements or transition metals.

6. The position of an element and its electronic configuration

a) Elements in the same group have the same number of valence electrons.

This number is the same as the group number.

For example: oxygen and sulfur from Group VI have six valence electrons.

b) Elements in the same group have similar chemical properties.

c) The table below shows the relationship between group number and electronic configuration.

GROUP	**I**	**II**	**III**	**IV**	**V**	**VI**	**VII**	**0**
NUMBER OF VALENCE ELECTRONS	1	2	3	4	5	6	7	8 (except He: 2 valence electron)
EXAMPLE	Li 2,1 Lithium	Be 2,2 Beryllium	B 2,3 Boron	C 2,4 Carbon	N 2,5 Nitrogen	O 2,6 Oxygen	F 2,7 Fluorine	Ne 2,8 Neon

d) The electronic configuration of an element can be worked out if we know the position of the element in the Periodic Table.

e) The group number of an element tells us the number of valence electrons.

The period in which the element is placed tells us the number of shells of electrons in the electronic configuration.

ADDITIONAL PORTION:

Periodicity is the repeatition of similar properties of elements placed in a group after of atomic numbers. The periodic table given by Moseley is called "long form of periodic table." At present 118 elements are known and they are arranged in 18 groups and 7 periods in the modern periodic table. In long form of periodic table the roman numerals for the groups are now replaced by groups 1 to 18 according to the recommendation made by International Union of Pure and Applied Chemistry (IUPAC).

EXAMPLE:

An element in Group V and Period 3 has the following electronic configuration:

2, 8, 5

Five valence electrons: Group V element

Electronic configuration: $\underbrace{2, 8, 5}$

Four electron shells: Period 3 elements

f) Conversely, from the electronic configuration of an element, we can work out the position of the Periodic Table .

EXAMPLE:

An element with the electronic configuration (2, 8, 8,1) is a Group I and Period 4 element.

One valence electrons: Group I element

Electronic configuration: $\underbrace{2, 8, 8, 1}$

Four electron shells: Period 4

7. The position of an element and its proton

If the proton number of an element is given, we can work out the period and group to which this element belongs .

EXAMPLE:

An element has the proton number 16. Its electronic configuration is (2, 8,6).

Six valence electrons: Group VI element

Electronic configuration: 2, 8, 6

Three electron shells : Period 3

TEST YOURSELF:

- ✓ Recognise the change from metallic to non-metallic character of the elements from left to right across a period of the Periodic Table
- ✓ Relate the group number to the number of valency electrons and metallic or non-metallic character of the elements
- ✓ Relate the group number to the ionic charge of an element

1.2 PERIODIC TRENDS:

1. The bold line in the Periodic Table divides the metals from the non-metals.

KEY:

☐ METALS

■ NON - METALS

Line dividing metals and Non - metals

	I	II											III	IV	V	VI	VII	0
1																		He
2	Li	Be											B	C	N	O	F	Ne
3	Na	Mg											Al	Si	P	S	Cl	Ar
4	K	Ca	Sc	Ti	V	Cr	Mn	Fe	C	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
	Rb	Sr	Y	Zr	Nb	Mo	Tc	RU	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe

5

6	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	T*l*	Pb	Bi	Po	At	R
7	Fr	Ra																

2. **Metallic and non-metallic character across a period.**

a) Elements on the left of the periodic Table are metals, while those on the right are non-metals.

b) From left to right across a period, there is a decrease in the metallic properties and an increase in non-metallic properties .

c) A metalloid is an element with properties that are intermediate between those of metals and non-metals. Examples are silicon, Si (Period 3), and germanium ,Ge (Period 4).

d) If the proton number or electronic configuration of an element is given, we can predict the metallic or non-metallic character of the element of the element and hence, predict the properties of the elements.

3. **Nature of oxides:**

a) In general,

i) metals form basic oxides,

ii)non-metals form acidic oxides and

iii)metalloids form amphoteric oxides .

b) However, there are exceptions. Some metals such as aluminium form an amphoteric oxide, Al_2O_3, and the metalloid, silicon, form an acidic oxide, SiO_2 .

4. **Group number and the ionic charge of an element**

a) In The table below shows the relationship between group number and the ions formed by the element .

Group	I	II	III	IV	V	VI	VII	0
Element	Na	Mg	Al	Si	P	S	Cl	Ar
Formula of ion	Na^+	Mg^{2+}	Al^{3+}	–	P^{3-}	S^{2-}	Cl^-	-

b) The elements in Group I, II, and III are metals. They form ionic compounds.

c) In the ionic compounds, *the metals lose electrons to form positive ions*. The charges on the ions of the elements in Group I, II, and III correspond to the group numbers .

d) The elements in Group V, VI, and VII are non-metals. They tend to form covalent compounds when they combine with non-metals .

e) However, when they combine with Groups I and II metals, they form ionic compounds.

f) In the ionic compounds, the non-metals form negative ions. The charge on the ion corresponds to group number minus eight .

EXAMPLE :

Nitrogen is a Group V non-metal. The charge on its ion corresponds to:

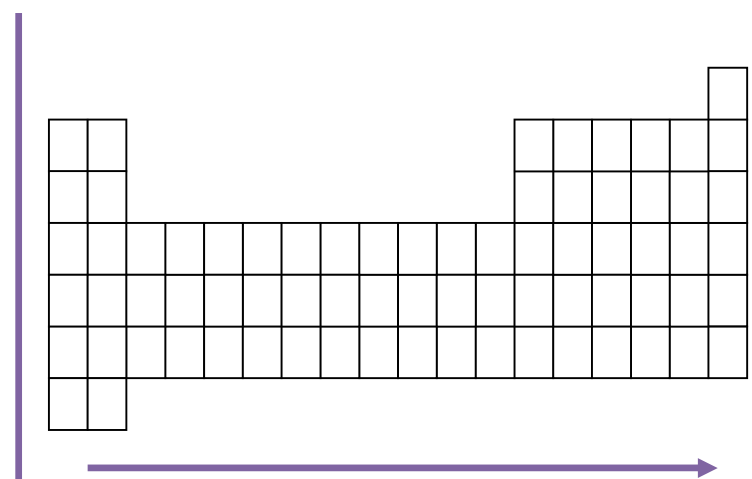
$$(\underbrace{5}_{\textbf{Group number}} - 8) = -3$$

That is, N^{3-} .

g) The element in Groups IV, V, VI, and VII form covalent bonds by the sharing of electrons.

COMMON ERROR	ACTUAL FACTS
✗ Group IV elements are all non-metals.	✓ On going down a group, the metallic property increases. For example, in Group IV, carbon is a non-metal, silicon and germanium are metelloids,but tin and lead are metals.

5. The diagram below summaries the trends of the propreties of the elements in the Periodic Table



ACROSS A PERIOD:

- ✓ Metallic properties decrease
- ✓ Non – Metallic properties increase

ON GOING DOWN A GROUP:

- ✓ **Proton number increases**
- ✓ **Size of atom increases**
- ✓ **Metallic properties increase**

PERIODIC PROPERTIES:

Periodic properties of elements are those properties which are directly or indirectly related to the electronic configuration of the elements and show gradation in moving down the group or along the period.

Some of the important periodic properties and their trends are discussed below.

ATOMIC RADIUS AND IONIC RADIUS:

ATOMIC RADIUS:

Atomic radius is defined as "the distance between the centre of its nucleus and electron in the last orbit."

The exact value of atomic radius is difficult to determine because it is not possible to determine the exact position of an electron. In the atom, some times the electron may be very close to nucleus and sometimes it may be far away from the nucleus.

There are 3 forms of atomic radius based on the nature of the bonding. They are

- COVALENT RADIUS
- VANDER WAAL'S RADIUS
- METALLIC RADIUS

✓ COVALENT RADIUS:

In homoatomic molecules having same type of atoms the covalent radius may be defined as " half the distance between the centres of the nucleus of two adjacent atoms which are similar joined by a singal convalent bond ".

COVALENT RADIUS = $\frac{\textit{Inter nuclear distance in single bond}}{2}$

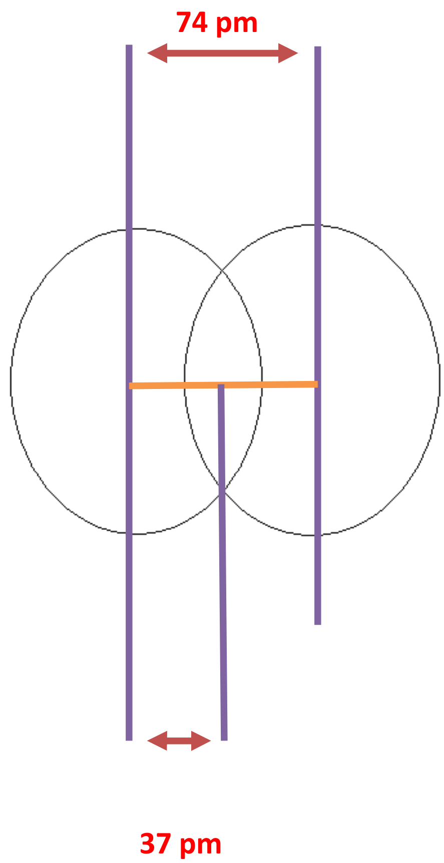


FIG. COVALENT RADIUS IN H_2

EXAMPLE:

the inter nuclear distance between two hydrogen atoms is 74 pm. Therefore its covalent radius is $\frac{74}{2}$ = 37 pm.

- ✓ **VANDER WAAL'S RADIUS:**

It is defined as the " One half the inter nuclear distance between two similar atoms belonging to two neighbouring molecules of same substance in the solid state ".

The Vander Waal's forces of attraction is very weak in the gases and liquids hence it can only be determined in the solid state. The magnitude of force is more hence the Vander Waal's

covalent radius is always more than the corresponding value of covalent radius. For example Vander Waal's radius of chlorine is 180 pm while its covalent radius is 90 pm .

- ✓ **METALLIC RADIUS:**

It is half the distance between the centres of the nuclei of two adjacent metal atoms in a crystal. Since the metallic bond is weaker it must have more radius when compared to covalent radius.

VARIATION OF ATOMIC RADIUS IN THE PERIODIC TABLE:

The atomic radii (covalent radii) of elements generally decrease from left to right in a period, while when we move down the group the atomic radii increases.

1. In the Second period elements the covalent radii decrease from lithium to Fluorine. It is given below.

ELEMENT	Li	Be	B	C	N	O	F
COVALENT RADIUS IN PM	123	90	80	77	75	74	72

2. In the group I i.e., alkali metals the variation of covalent radius is given below.

ELEMENT	Li	Na	K	Rb	Cs	Fr*
COVALENT RADIUS IN PM	123	157	202	217	235	

REASONS:

On moving left to right across the period the effective nuclear charge increases and electrons are strongly attracted towards the nucleus from one element to another. This strong attraction is mainly because of the entry of new electron to the same shell. Thus the covalent radius decreases.

As we move down a group there is increase in the number of shells and atomic size increases. The added electron enters the new shell. Even though a proton is also added, the effective nuclear charge is less, on the added electron in the new shell. Therefore the atomic radius increases down the group.

IONIC RADIUS:

Ions are the charged particles which are formed by the loss or gain of one or more electrons from a neutral atom. A cation has positive charge while an anion has negative charge. Ionic radius is defined as follows.

"It is the distance between the centre of the nucleus to the region up to which it can attract the electron cloud ".

Since the electron cloud is extended to larger distance from the nucleus it is not possible to determine the ionic radius experimentally. Pauling by his x-ray diffraction ,suggested that if the radius of one of the ions is known, then the ionic radius of the other can be calculated by subtracting the value of known ionic radius, from the inter nuclear distance.

In NaCl , the inter- nuclear distance is 276 pm. The radius of Na^+ ion is 95 pm, then the radius of Cl^- is 276 – 95 = 181 pm.

The radius of a cation is always smaller than the neutral atom from which it is formed. This is because when an electron is removed from the outermost shell the neutral atom has become unipositive ion. It means the nucleus gains a proton. Now the nucleus can attract the remaining electrons more strongly, hence the electron cloud shrinks, resulting in decrease in ionic radius. It is shown below

EXAMPLE:

ELEMENT	Li	Na	K	Mg	Ca	Al
ATOMIC RADIUS (PM)	123	157	202	136	114	125
IONIC RADIUS OF CATION (pm)	60	95	133	65	99	50

The radius of Anion is always larger than the neutral atom from which it is formed. This is because the added electron is less attracted by the nucleus and it also experiences shielding effect by the inner electrons. Thus it results in expanding of the electron cloud. This fact is given in the following table.

ELEMENT	**O**	**N**	**F**	**Cl**	**Br**	**I**
ATOMIC NUMBER	**74**	**75**	**99**	**114**	**133**	**125**
IONIC RADIUS OF CATION (pm)	**142**	**171**	**136**	**181**	**196**	**219**

ISOELECTRONIC IONS:

They are the species of different ions which have same number of electrons but different magnitude of nuclear charges. If the ion has greater value of nuclear charge it has lesser ionic radius. The size of isoelectronic ions decrease with increase in atomic number. It is shown in the following table.

ELEMENT	N^{3-}	O^{2-}	Na+	Mg^{2+}	Al^{3+}
ATOMIC NUMBER	7	8	11	12	13
NUMBER OF ELECTRONS IN THE ION	10	10	10	10	10
IONIC RADII (pm)	171	140	95	65	50

IONISATION ENERGY OR IONISATION POTENTIAL:

It is defined as " The minimum amount of energy needed to remove the most loosely bound electron from a neutral gaseous atom to form a positive ion ".

Ionisation energy is denoted as IE

$$Na_{(g)} + IE_1 \longrightarrow Na^+_{(g)} + e^-$$

The energy needed to remove the first electron form a neutral gaseous atom called first ionisation potential IE_1 .The ionisation energies required to remove second and third electrons are IE_2 and IE_3 respectively.

The ionization energy in kJ mol^{-1}. The first ionization energy is always less than the second IE because, when one electron is removed from the neutral gaseous atom forms a unipositive ion The extra proton in the nucleus attracts the remaining electrons more strongly . Therefore it is difficult to remove second electron as more energy is required to overcome the nuclear attraction. In summary $IE_3 > IE_2 > IE_1$ which is given as follows.

ELEMENT	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**
ATOMIC NUMBER	**3**	**4**	**5**	**6**	**7**	**8**	**9**
IE_1 kJ mol^{-1}	**520**	**900**	**800**	**1086**	**1402**	**1314**	**1681**

IE_2 kJ mol^{-1}	7297	1757	2427	2352	2858	3388	3375
IE_3 kJ mol^{-1}	11810	14850	3638	4619	4576	5296	6045

Variation of IE in the periodic table:

Ionisation energy is expected to increase on moving across the period and the value decrease down the group.

Example:

1. In the second period the IE_1 values increases as we move from left to right of the periodic table.

ELEMENT	Li	Be	B	C	N	O	F	Ne
IE_1 (kJ mol^{-1})	520	899	800	1086	1402	1314	1681	2080

2. In the I group i.e., Alkali metals, the first ionisation energy values decreases.

ELEMENT	H	Li	Na	K	Rb	Cs
IE_1 (kJ mol^{-1})	1312	520	496	419	403	475

In a period we notice that noble gases have highest value while the alkali metals have lowest values.

REASONS:

On moving along a period, the electron enters the same shell. It leads to increase in effective nuclear charge and decrease in atomic size. This results in increase in the forces binding the

electron with the nucleus. Hence IE increases, when we move down a group the IE decreases because the new shell is added and electron is present far away from the nucleus. It is less attracted by the nucleus. Moreover the shielding effects (Repulsion by inner electrons) causes the electron less tightly bound to nucleus, it requires very less energy to pull the electron. Therefore the ionisation energy decreases. The variations of IE1 for second period and first group are represented in the following graphical representation.

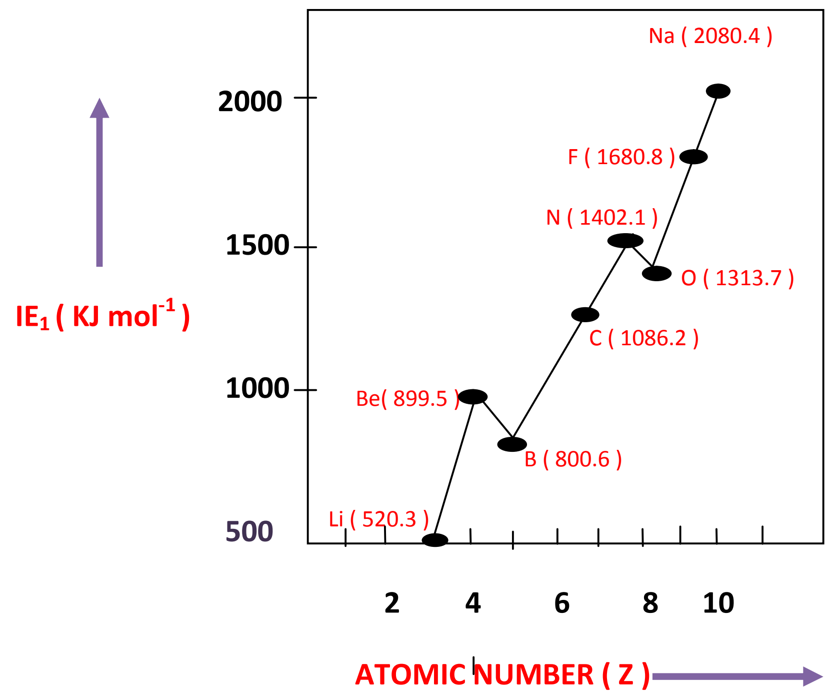
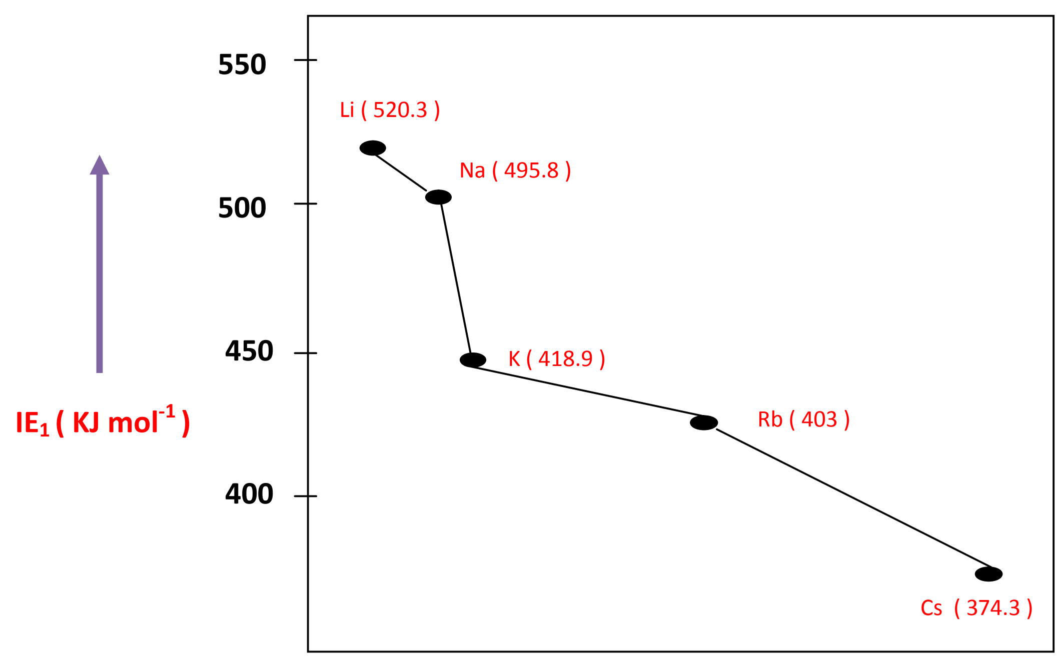


FIG. VARIATION OF IE ACROSS SECOND PERIOD



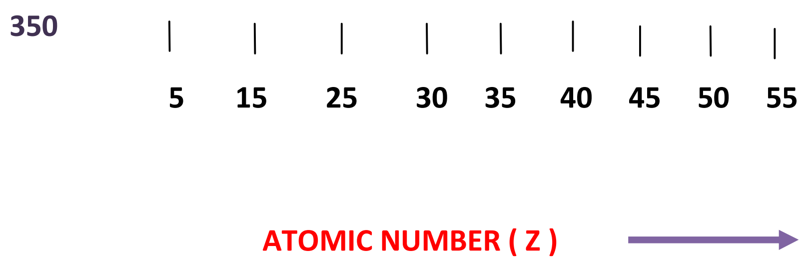


FIG. VARIATION OF IE DOWN THE FIRST GROUP

ELECTRON AFFINITY:

It is defined as " The energy released when a neutral gaseous atom accepts an electron to form an anion "

Electron affinity is denoted as EA

$$F_{(g)} + e^- \longrightarrow F^-_{(g)} + EA$$

Similar to ionisation energies, when a uninegative anion is formed it is called first electron affinity EA_1 . The EA_1 energy released while EA_2 energy is absorbed. This is because a uninegative anion has to take up electron cloud against the repulsion by the electron could. Therefore energy must be supplied.

VARIATION OF ELECTRON AFFINITY IN THE PERIODIC TABLE:

Electron affinities normally increase along a period and decrease down the group with few exceptions.

EXAMPLE:

1. In the elements of second period, it increases from lithium to Neon.

ELEMENT	Li	Be	B	C	N	O	F	Ne
EA_1 kJ mol^{-1}	60	0	83	122	0	141	333	0

2. In the group I Alkali metals, the values decrease which is given in the following table

ELEMENT	H	Li	Na	K	Rb	Cs
EA_1 kJ mol^{-1}	73	60	53	48	46.9	45.5

Electron affinity value is the measure of tightness of extra electron to the atom.

REASONS:

Along a period, there is increase in the nuclear charge and decrease in atomic size.

Therefore the electron is more tightly bound and greater is the energy released EA_1 values of Be, N and Ne are very low because of high symmetric electron configurations and high ionisation energies. Hence EA_1 value is zero.

In a group EA decrease with atomic number. This is because the added electron is far away from nucleus hence loosely bound. Fluorine has lower electron affinity than chlorine because of smaller size. (Figure No. 1.5)

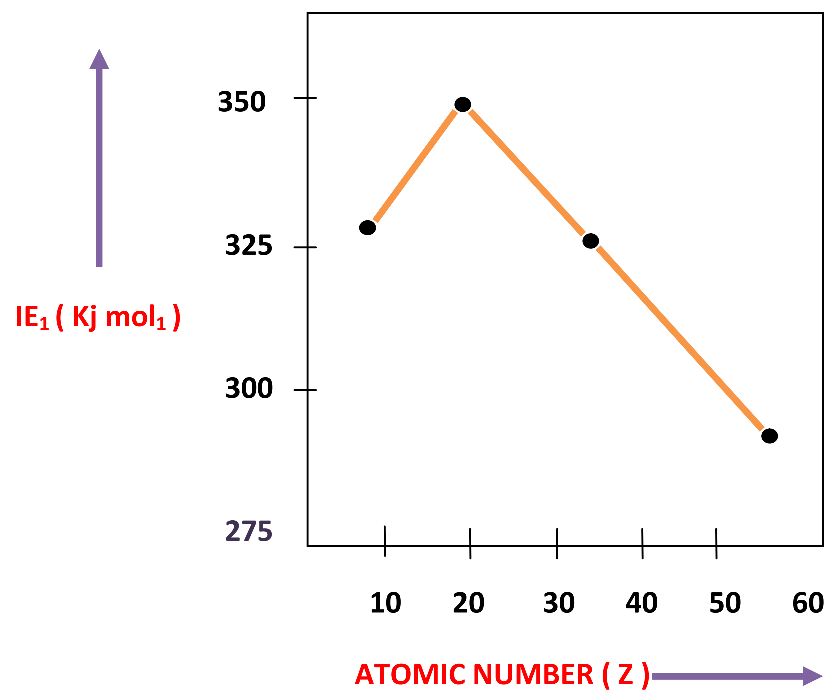


FIG. ELECTRON AFFINITY IN HALOGENS

Higher values of EA indicate greater tendency to gain electrons and hence good oxidizing capacity.

ELECTRONEGATIVITY:

It is defined as " The relative electron attracting tendency of an atom for the shared pair of electron in a covalent bond "

It has no specific units. Greater the value of electronegativity , higher is its tendency to attract the shared pair of electrons towards itself. It is a relative value.

Fluorine (4.0) is highest electronegative element.

Where as Cesium (0.7) is least electronegative element.

VARIATION OF ELECTRONEGATIVITY IN PERIODIC TABLE:

Electronegativity value increases across the period and it decreases down the group.

EXAMPLE:

1. In the second period, the value of electronegativity increases from Lithium to Fluorine.

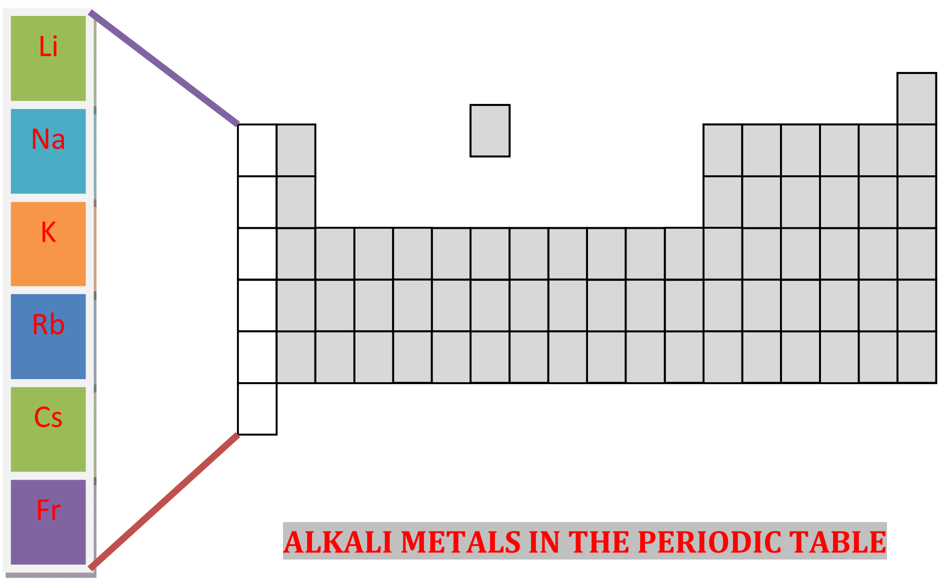
ELEMENT	Li	Be	B	C	N	O	F	Ne
ELECTRONEGATIVITY	1.0	1.5	2.0	2.5	3.0	3.5	4.0	0

ANALYSE:

- ✓ Describe the Group I elements (the alkali metals) and predict their properties using the Periodic Table.

1. 3 GROUP I ELEMENTS - ALKALI METALS:

1. Group I elements are called alkali metals because they react with water to form alkalis.
2. The diagram below shows the properties of Group I elements in the Periodic Table.



ALKALI METALS IN THE PERIODIC TABLE

3. **Physical properties of alkali metals:**

a) The physical properties of Group I is metal charge gradually down the group.

b) The density of the alkali metal generally increases down the group.

Lithium, sodium and potassium are less dense than water. Rubidium and caseium are denser than water.

c) They are relatively soft and can be cut easily with a knife.

d) They have low melting and boiling points compared with other metals.

The melting boiling points decreased down the group.

e) They are good conductors of heat and electricity.

f) **The following table shows the melting and boiling points, density and appearance of same Group I elements.**

ALKALIS METAL	**MELTING POINT (^0C)**	**BOILING POINT (^0C)**	**DENSITY (g\cm3)**	**APPREANCE**
LITHIUM, L*I*	**180**	**1330**	**0.53**	**Shiny, silver solid**
SODIUM, NA	**98**	**890**	**0.97**	**Shiny, silver solid**
POTASSIUM, K	**64**	**688**	**0.86**	**Shiny, silver solid**
RUBIDIUM, RB	**38**	**688**	**1.53**	**Shiny, silver solid**

4. **Chemical properties of alkali metal:**

a) **They have similar chemical properties because all the elements have similar electronic configuration , that is, one electron in the outermost shell.**

By losing their valence electron, an alkali metal attains a noble gas configuration.

b) **The Group I metals are very reactive. They react rapidly with air and vigorously with water.**

They are kept under oil to prevent coming into contact with air and water.

5. **Reaction of alkali metals with water:**

a) **All the alkali metals reacts with cold water to form alkali (soluble metal hydroxides) and hydrogen gas.**

ALKALI METAL	OBSERVATION	EQUATION
LITHIUM	· Reacts quickly. · Lithium floats on the water · No flame is seen.	2Li(s) + $2H_20$(l) ⟶ N2Li (aq) + H_2(g) Lithium water lithium Hydroxide hydrogen
SODIUM	· Reacts very quickly. · Sodium melts and burns with a yellow flame. The molten sodium darts around the water surface.	2Na (s) + $2H_20$(I) ⟶ 2 NaOH (aq) + H_2 (g) Sodium water sodium Hydroxide hydrogen
POTASSIUM	· Reacts violently. · Potassium melts and burns a lilac flame.	2K (s) + $2H_2O$ (I) ⟶ 2KOH(aq) + H_2 (g) Potassium water potassium Hydroxide hydrogen

c) The reaction between alkali metals and water is a redox reaction.

EXAMPLE:

In the reaction between alkali metals and water is a redox reaction.

Na (s) ⟶ Na^+(aq) + e^- … oxidation

Sodium acts as the reducing agent when it reacts with water .

$2H_2O$ (I) + $2e^-$ ⟶ H_2 (g) + $2OH^-$(aq) … reduction

Overall reaction:

2Na (s) + $2H_2O$(I) ⟶ $2Na^+$ (aq) + $2OH^-$ (aq) + H_2(g)

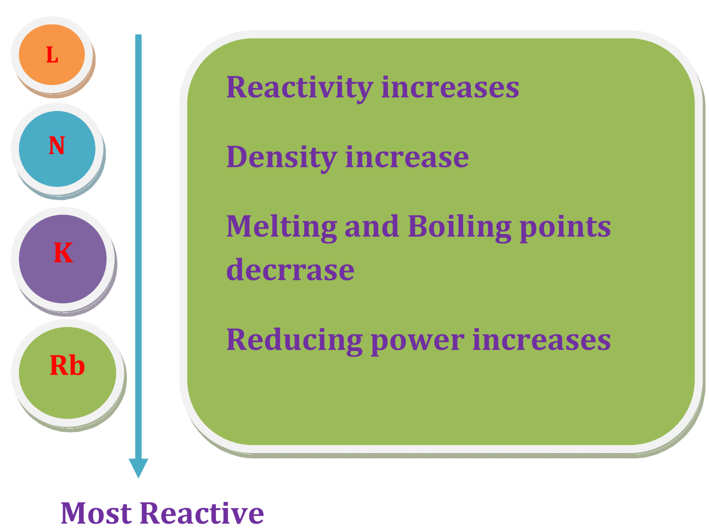
6. The alkali metals are powerful reducing agent because they give up their valence electrons easily.

EXAMPLE:

Li ⟶ Li^+ + e^-

7. The reactivity of Group I elements increases down the group because the atomic size increases.

The larger the atomic size, the further the valence electron is from the nucleus, and thus, the easier for the atom to lose the valence electron.



EXAMPLE1:

How and under what conditions do you expect rubidium to react with water?

SOLUTION

Rubidium is below potassium in Group I of the Periodic Table. It reacts violently with water at room temperature to give rubidium hydroxide and hydrogen.

$$2Rb(s) + 2H_2O(l) \longrightarrow 2RbOH(aq) + H_2(g)$$

TEST YOURSELF:

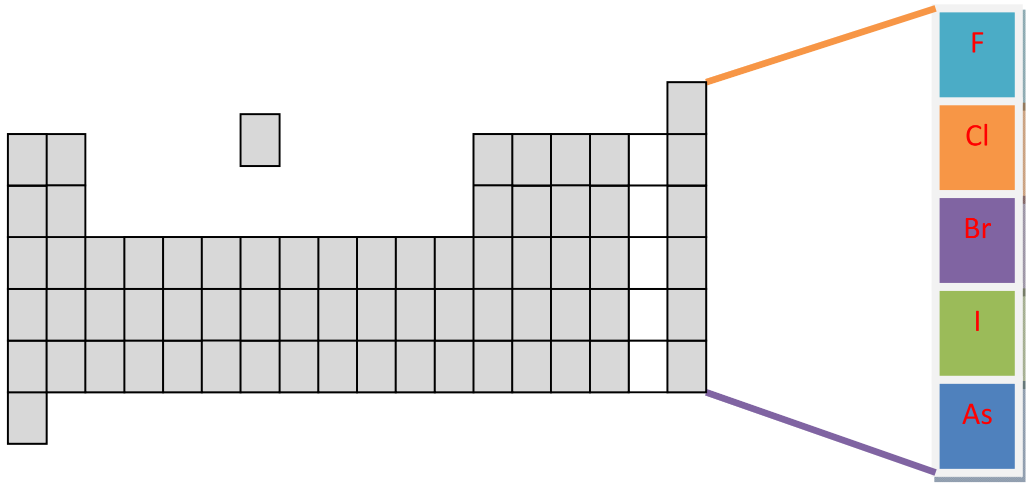
- ✓ **Describe the group VII elements (the halogens) and predict their properties using the Period Table**

1.4 GROUP VII ELEMENTS – HALOGEN:

1. The Group VII elements are also called halogen in the Periodic Table.

2. The diagram on the right shows the position of halogens in the Periodic Table.

3. The halogens are non-metals and exit as diatomic covalent molecules, F_2, CI_2, Br_2 and I_2. Their ions, such as fluoride ion (F^-), bromide ion (Br^-) and iodide ion (I^-).



HALOGENS IN THE PERIODIC TABLE

4. **Physical properties of halogens:**

a) All halogens have low melting and boiling points.

b) All the halogens are coloured and are non-conductors of electricity.

c) On going down the group,

i) The colour intensity increases ,

ii) The melting and boiling points increase.

d) The following table shows the physical properties of some halogens.

HALOGEN	COLOUR\PHYSICAL STATE	MELTING POINT(^{0}C)	BOILING POINT(^{0}C)
FLUORINE, F_2	Pale yellow gas	-219	-188

CHLORINE, Cl_2	Greenish-yellow gas	-101	-35
BROMINE, Br_2	Reddish-brown liquid	-7	59
Iodine, I_2	Purplish-black solid	114	184

5. Chemical properties of Halogens:

a) **Halogens have similar chemical properties because they similar electronic structure, that is seven electrons in the outermost shell .**

b) **Halogens react vigorously with metals to form salts called halides. The metal halides are ionic compounds.**

EXAMPLE:

$$2\,Fe\,(s) \quad + \quad 3Cl_2\,(g) \longrightarrow FeCl_3(s)$$

Iron Chlorine Iron (III) Chloride (ionic compound)

c) **They also react with non-metals to form covalent compounds.**

EXAMPLE:

$$H_2(g) \quad + \quad Cl_2(g) \longrightarrow 2HCl(g)$$

Hydrogen chlorine hydrogen chloride (covalent compound)

d) Halogens are powerful oxidising agents because they gain electrons readily to form halide ions.

EXAMPLE:

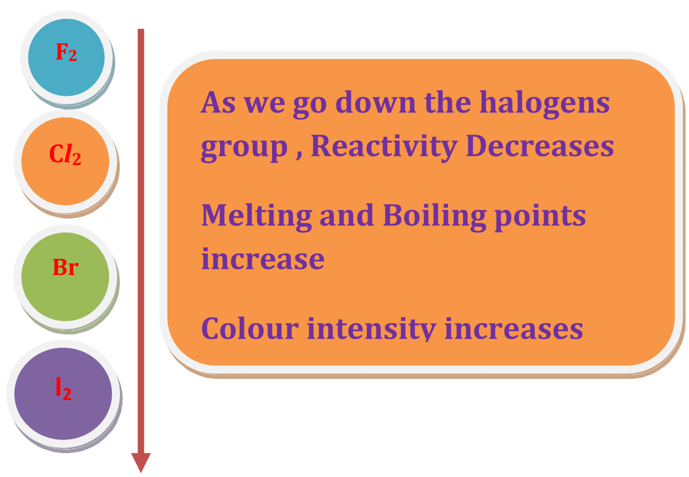
$$\underset{\text{Chlorine}}{Cl_2(aq)} + 2e^- \longrightarrow \underset{\text{Chloride}}{2Cl^-(aq)}$$

6. Unlike Group I elements, the reactivity of halogens decreases on going down the group.

7. Displacement reactions of halogens:

a) A displacement reaction is a reaction in which one element takes the place of another element.

b) A more reaction halogen will displace a less reactive halogen from its halide solution as shown in the following table.



Least Reactive

HALOGEN	KCl(aq)	KBr(aq)	KI(aq)
CHLORINE	-	Reddish-brown solution formed.	Brown solution formed.

		$Cl_2(aq) + 2Br\text{-}(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$	$Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$
BROMINE	**NO REACTION**	-	**Brown solution formed.** $Br_2(aq) + 2I^-(aq) \rightarrow 2Br^-(aq) + I_2(aq)$
IODINE	**NO REACTION**	**No Reaction**	-

Tips for Students:

1. **When writing the chemical equations for the displacement reactions of Halogens, the state symbols for the reactants are all (aq).this is because the Halogens and the halides are dissolved in water.**

2. **halogens with halides (ions of halogens) are two different identities.**

   For example: Chlorine is Cl_2 and chloride is Cl^-. Similarly,bromine is Br_2 and bromide is Br^- .

8. **A displacement reaction is a Redox Reaction.**

EXAMPLE:

In the displacement reaction between chlorine and potassium bromide, chlorine oxidises bromide ions to bromine and is itself reduced to chloride ions .

Reduced

$$Cl_2\,(aq) + 2Br^-\,(aq) \qquad 2Cl^-\,(aq) + Br_2\,(aq)$$

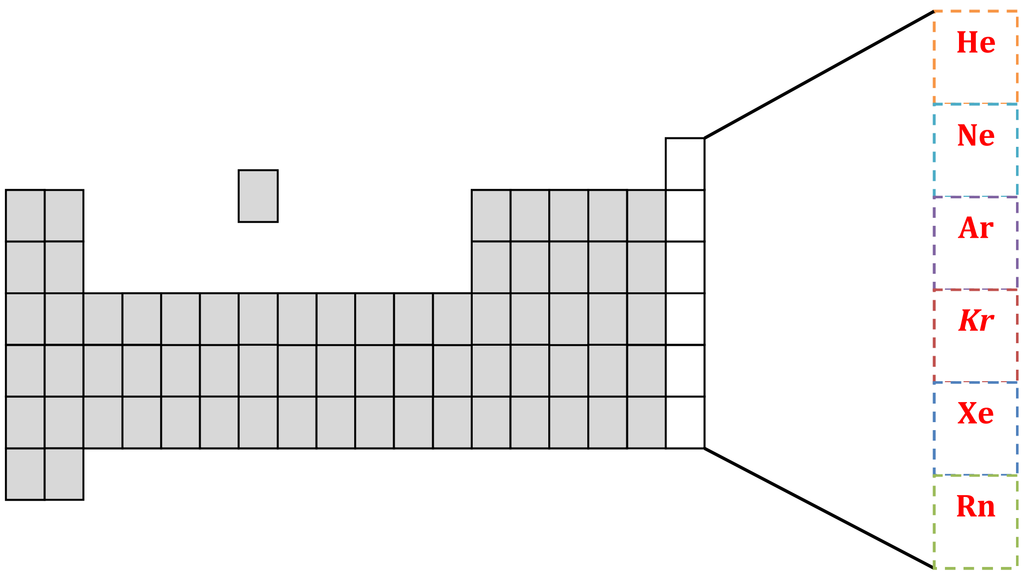
Oxidised

1.5 COMMON ERRORS	ACTUAL FACTS
✗ Potassium fluoride (KF), potassium bromide (KBr) and potassium iodide (KI) are coloured substances.	✓ The halogens, F_2, Cl_2, Br and I_2, are coloured substances, but KF, KBr and KI are white solids and their aqueous solutions are colourless.

GROUP VIII ELEMENTS –NOBLE GASES:

1. **The Group 0 or Group VIII elements are called Noble Gases.**

2. **The diagram below shows the positions of Group 0 elements in the Periodic Table.**



NOBLE GASES IN THE PERIODIC TABLE

3. PHYSICAL PROPERTIES OF NOBLE GASES:

a) **They exist as single atom and are known as Monatomic Elements.**

b) **They are Colourless gases at room temperature and pressure.**

c) They have Low Melting and Boling Points that increase on going down the group.

d) They are insoluble in Water.

4. CHEMICAL PROPERTIES OF NOBLE GASES:

a) The noble gases are Inert (Chemically Unreactive) because they have stable electronic configurations .

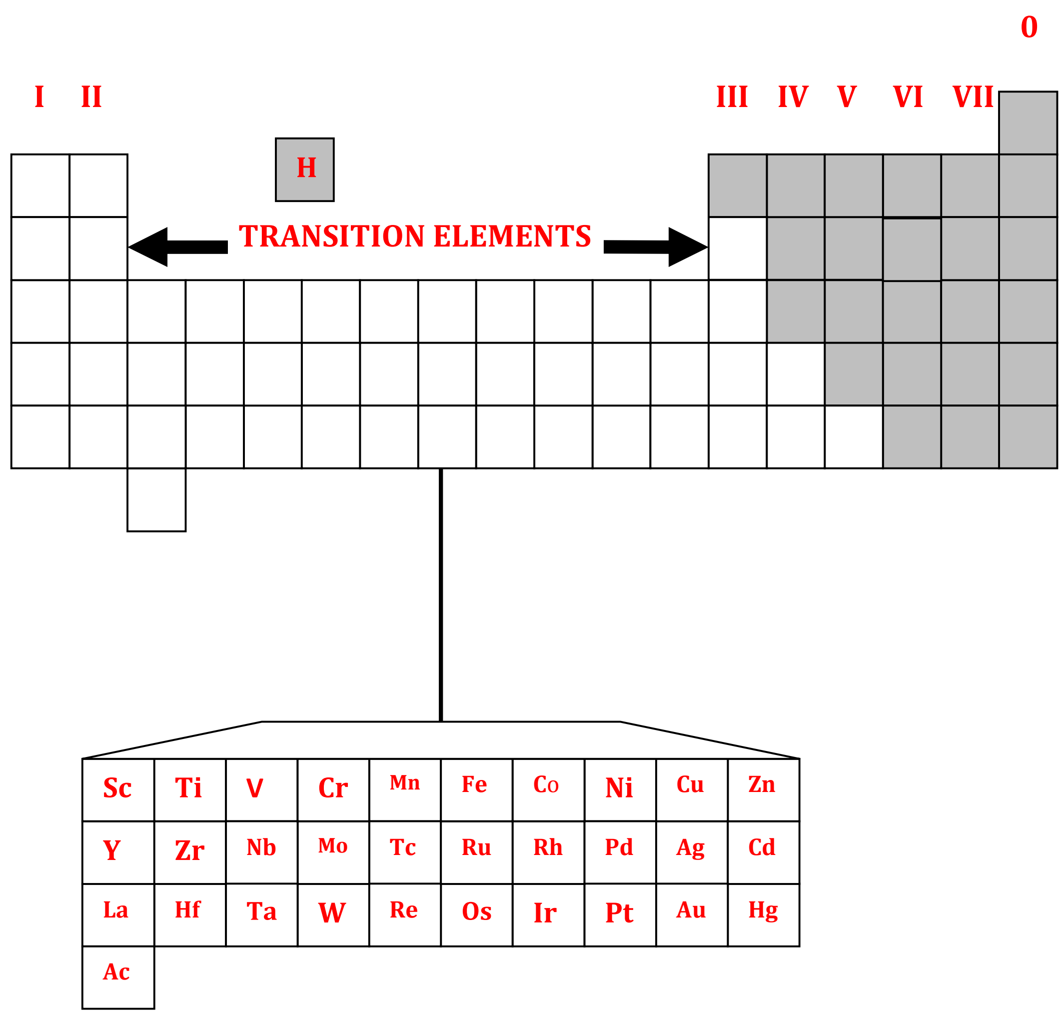
b) Their outermost shells are fully occupied, with two valence electrons for helium atom and eight valence electrons for the other noble gases.

5. USES OF NOBLE GASES:

NOBLE GAS	USES	PROPERTY
HELIUM	For filling weather balloons, advertisement balloons and airships	Low density, non - flammable
NEON	In advertising signs (neon lights)	Glows brightly when electricity is passed through
ARGON	In light bulbs, in welding and making steel	Provides and inert atmosphere to prevent the filament in light in light bulb or the steel in the furnace from reaction with oxygen

1.6 TRANSITION ELEMENTS:

1. The transition elements (or transition metals) are located between Groups II and III of the Periodic Table . Iron, Copper, Nickel, Silver and Gold are some examples of Transition Elements .



TRANSITION METALS IN THE PERIOD TABLE

2. Melting Point and Density:

All transition elements are metals. Unlike the alkali metals, transition elements have high melting points and high density . The table shows the melting points and density of three transition elements, iron, nickel and copper, as compared to that sodium.

ELEMENT	MELTING POINT (^0C)	DENSITY (g/cm^3)
IRON (Fe)	**1535**	**7.86**
NICKEL (Ni)	**1453**	**8.90**
COPPER (Cu)	**1083**	**8.96**
SODIUM (Na)	**98**	**0.97**

3. Oxidation States:

a) Group I metals have a fixed oxidation state of + 1 and group II elements have a fixed oxidation state of + 2, but a transition element has variable oxidation states (i.e., more than one oxidation state) in its compounds . The oxidation states of the transition elements, chromium and manganese are shown below .

COMPOUNDS OF TRANSITION ELEMENTS	CHEMICAL FORMULAE	OXIDATION STATE OF THE TRANSITION ELEMENT
Chromium (III) Oxide **Potassium Dichromate (VI)**	Cr_2O_3 $K_2Cr_2O_7$	**+ 3** **+ 6**
Manganese (II) Chloride **Manganese (IV) Oxide**	$MnCl_2$ MnO_2	**+ 2** **+ 4**

Potassium Manganate (VII)	**$KMnO_4$**	**+ 7**

b)Because of their variable oxidation states, transition elements can form more than one compound with another element . For EXAMPLE:

- **Iron (II) Chloride, $FeCl_2$ and Iron (III) Chloride,$FeCl_3$.**
- **Copper (I) Oxide, Cu_{20} and Copper (II) Oxide, CuO.**

4. Coloured Compounds:

- **Transition elements form coloured compounds . For example, Iron(II) sulfate is green, Copper (II) sulphate is blue and potassium manganate (VII) is purple.**

- **The colour of the anhydrous compound of a transition element may be different from its hydrated (Crystalline) compound.**

- **For example, anhydrous copper (II) sulfate, $CuSO_4$, is white, but the hydrated (Crystalline) copper (II) sulfate, $CuSo_4.5H_2O$, is blue.**

- **Similarly, anhydrous cobalt (II) chloride is blue, but hydrated cobalt (II) chloride is pink .**

- **Because of this property, both anhydrous copper (II) sulfate and anhydrous cobalt (II) chloride are used as a test for the presence of water .**

COMMON ERRORS	ACTUAL FACTS
✗ Transition elements are coloured.	✓ Transition elements are not coloured . It is their compounds that are coloured.
✗ Only transition metals form coloured compounds.	✓ Few non – transition metals also form coloured compounds besides transition metals. For example, led (II) oxide(PbO) is a yellow solid.

5. As Catalysts:

a) A Catalyst is a substance that increases the speed of a reaction.

b) Transition elements are often used as catalysts.

For Example, platinum is used as a catalyst to speed up the reaction between nitrogen and hydrogen to form ammonia .

$$N_2\,(g) + 3H_2\,(g) \underset{450^0C,\ 250\ atm}{\overset{Pt}{\rightleftharpoons}} 2NH_3$$

c) Compounds of transition elements are also used as catalysts to speed up chemical reactions. For Example, copper (II) sulphate is used to speed up the reaction between zinc and dilute sulfuric acid to form zinc sulfate and hydrogen gas.

$$\mathbf{Zn\ (s) + H_2SO_4\ (aq) \xrightarrow{CuSo_4\ (Catalyst)} ZnSO_4\ (aq) + H_2\ (g)}$$

d) Transition elements and their compounds are often used as catalysts in industrial processes.

TIP FOR STUDENTS:

Compounds of some non-transition metals can also act as catalysts.

STRUCTURAL QUESTIONS AND ANSWERS

1.) In A modern Periodic Table:

a) What is the link between an element's electronic configuration and the Group it is in?
b) Describe how the elements are arranged in the modern Periodic Table in terms of atomic number.

Answer**:**

a.) The group number of an element (for groups I to VII) corresponds to the number of electrons in its outermost shell.

b.) The elements are arranged in the order of increasing atomic number.

2) (You may refer to the Periodic Table to answer this question). Give the chemical symbol of:
   a) A metal used as the catalyst in the Haber process:
   b) A metal which is a liquid at room temperature and pressure;
   c) The most reactive non-metal;
   d) An element which consists of diatomic molecules and constitutes 78% of air
   e) An element in Period 2 which forms an ion of the type X^+
   f) An element in Period 2 which has four electrons available for bonding.

Answer:

2. a) Fe   b) Hg   c) F   d) N   e) Li   f) C

3) The properties of some elements are described below. For each of the elements, state which position of the Periodic Table, A, B, C, D, E or F, it is found in
   a) This element exists as a colourless gas at room temperature. It is a poor heat conductor. It is used as a gas to fill up glass light bulbs in incandescent lighting because it does not react with their filaments.
      Position : ______________

   b) This element has a melting point of 2334 ^{0} C. it is a versatile catalyst. It has oxidation states of +2, +3 and +4 in its compounds.
      Position : ______________

   c) This element exists as a silvery-white solid at room temperature. It is strong and light and is used in the manufacture of electronic components. It is highly flammable when shaved into thin strips and produces a brilliant white flame when burnt in air. It has oxidation state of +2 in its compounds.
      Position : ______________

   d) This element is used in the galvanization of steel to prevent corrosion. It burns in air to give an oxide which reacts with both acids and alkalis. It has an oxidation state of +2 in its compounds.
      Position : ______________

   e) This element is used in the galvanization of steel to prevent corrosion. It burns in air to give an oxide which reacts with both acids and alkalis. It has an oxidation state of +2 in its compounds.
      Position : ______________

   f) This element has a melting point of 221^{0} C. it is used widely in the manufacture of glass and ceramics. It reacts with oxygen to give a colourless, flammable gas.
      Position : ______________

g) This element exists as a very soft, silvery-white solid at room temperature. It is a component of many low-melting alloys. It has an oxidation state of +3 in many of its compounds.
Position : ____________

h) This element exists as a very soft, silvery-white solid at room temperature. It is a component of many low-melting alloys. It has an oxidation state of +3 in many of its compounds.
Position : ____________

Answer:

3. a) F b) B c) A d) C e) E f) D

4. (a)

i) Which element in Period 3 has the highest melting point? Explain in terms of structure and bonding, why this element has a high melting point.

ii) Which element in Period 3 has the lowest melting point? Explain, in terms of structure and bonding. Why this element has a low melting point.

b) Name the Period 3 elements which are good conductors of electricity.

c) An element X, in Period 3, reacts with oxygen to give an ionic compound with the formula X0. Identify the element X.

a) An element Y, in Period 3, burns in air on gentle heating to produce a colourless gas with the formula $Y0_2$. Identify the element Y.

Answer:

4. a)

i.)Silicon has the highest melting point. Silicon has a giant molecular structure. A large amount of energy is needed to overcome the strong covalent bonds holding the Si atoms together. Hence, silicon has a high melting point.

ii. Argon has the lowest melting point. Argon has a simple atomic structure. A small amount of energy is needed to overcome the weak forces discrete argon atoms being together. Hence, argon has a low melting point.

b) Sodium, Magnesium and aluminium.

c) Magnesium

d) Sulfur.

5.The table below shows the formulae and melting point of oxides formed from six elements in Period 3.

Formula of element	Group	Formula of oxide	Melting point /℃
Na	I	$Na_2 0$	1280
Mg	II	Mg0	2900
Al	III	Al_20_3	2040
Si	IV	$Si0_2$	1610
P	V	P_40_6	24
P	V	P_40_{10}	Sublimes at 300 ℃
S	Vl	$S0_2$	-75
S	VI	$S0_3$	17

a) Explain, in terms of structure and bonding, the variation in melting points of the oxides listed in the table.
b) Explain why when some magnesium oxide powder is mixed with water at room temperature, a solution of around pH 9 is obtained.

c) Write balanced chemical equations for the reactions of water with the oxides of sulphur.Suggest the pH of the solutions formed.

d) One of the oxides listed in the table does not dissolve in water but reacts with both acids and bases. Identify this oxide.

e) Write the formulae of the two oxides of selenium (Group VI).

Answer:
5.

a) Na_2O, MgO and Al_2O_3 have giant ionic structures (lattices) in which the positively charged ions and the negatively charged oxide ions (O^{2-}) are held together by strong electrostatic attraction (ionic bonds). A large amount of energy is needed to overcome the strong ionic bonds holding the positive ions and O^{2-} ions together. Hence, these ionic oxides have high melting points.

SiO_2 has a giant molecular structure. A large amount of energy is needed to overcome the strong covalent bonds holding Si and O atoms together. Hence, SiO_2 has a high melting point.

P_4O_6, P_4O_{10}, SO_2 and SO_3 have simple molecular structure. A small amount of energy is needed to overcome the weak intermolecular forces holding the discrete molecules together. Hence, these covalent oxides with simple molecular structures have low melting points.

b) Magnesium oxide reacts with water to form magnesiu hydroxide.
$MgO + H2O \longrightarrow Mg(OH)2$

Magnesium hydroxide is only sparingly soluble in water. Only a small amount of hydroxide ions gets into the solution. Hence, the resulting solution is a weak alkaline (pH 9)

c) $SO_2 + H_2O \longrightarrow H_2SO_3$
$SO_3 + H_2O \longrightarrow H_2SO_4$
In both cases, acidic solutions are produced. The pH of the solutions is around 2.

d) Aluminium oxide.

e) SeO_2, SeO_3

6.) Phosphorous (group V) reacts with oxygen to give two different oxides.

(a) Fill the relevant formulae and oxidation states in the table below.

(b) Explain why the phosphorous oxides do not conduct electricity when molten.

(c) Arsenic (As) is in the same Group of the Periodic Table as phosphorous. Write the empirical formula of the two oxides of arsenic.

Answer:

6.) a)

Molecular formula of oxide	P_4O_6	P_4O_{10}
Empirical formula of oxide	P_2O_3	P_2O_5
Oxidation state of phosphorous in oxide	+3	+5

b) The phosphorous oxides have simple molecular structures in which the molecules in which the molecules are held together by weak intermolecular forces. They do not conduct electricity because of absence of mobile charged particles (such as ions or electrons).

c) As_2O_3 and As_2O_5

7.) Some physical and chemical properties of metallic element M are listed below.

Colour	Density/ gcm^{-3}	Melting point/°C	Boiling point/°C	Reaction with water
Silvery white	0.89	63	759	Bubbles vigorously and catches fire to form aqueous MOH

(a) (i) Deduce the Group of the Periodic Table to which element M belongs.
(ii) Write a balanced chemical equation,with state symbols , for the reaction between element M and water.

(b) M reacts vigorously with fluorine to form a salt MF.When 0.78 g of M is reacted with fluorine,1.16g of MF is obtained .

(i) Write a balanced chemical equation , with state symbols , for the t reaction of element M and fluorine.

(ii) Calculate the number of moles of fluoride ions in the salf formed , and hence determine the relative atomic mass of M.

Answer:

7.)

a) i) M belongs to Group I

ii) $2M (s) + 2H_2O(I) \longrightarrow 2MOH (aq) + H_2 (g)$

b) i) $2M (s) + F_2 (g) \longrightarrow 2MF (s)$

ii) Mass iof fluoruide ions = 1.16 – 0.78 = 0.38

No. of moles of fluoride ions = 0.38 /19 = 0.02

No. of moles of M^+ ions = 0.02

Relative atomic mass of M = 0.78 / 0.02 = 39

8.)Sodium is an element in Group l of the Periodic Table.the atomic number of sodium is 11.

(a) Describe what will be observed when a small piece of sodium is put into a bowl of cold water containing a small amount of Universal Indicator.

(b) Liquid sodium reacts with hydrogen gas to give an ionic solid , sodium hydride(NaH)

(i) Explain why this is a redox reaction.

(ii) Draw a ‘ dot and cross’ diagram to show the electronic structure of sodium hydride .You Only need to show the valence electrons.

(iii) Explain in the terms of structures and bonding, why sodium hydride has a high melting point.

(c) Rubidium is another element in Group l. the atomic number of rubidium is 37.

(i) Use your knowledge of Group l elements to complete the table below.

Symbol of rubidium	Rb
Formula of rubidium chloride	
Formula of rubidium hydride	
Formula of rubidium sulfide	
Formula of rubidium sulfate	
Formula of product of the reaction between rubidium and water	

(ii) Explain why the reaction between rubidium and water is more vigorous than the reaction between sodium and water.

Answer:

a) The piece of sodium rapidly melts into a silvery-white ball, which runs on the surface of the water making a hissing sound, and soon catches fire. The Universal Indicator turns from green to blue.

b)

i) In this reaction, Na (oxidation state 0) is oxidized to Na^+ (oxidation state +1) while H_2 (oxidation state o) is reduced to H^- (oxidation state -1). Hence, this is a redox reaction.

iii) Sodium hydride has a giant ionic structure (lattice) in which the positively-charged sodium ions (Na^+) and the negatively-charged hydride ions (H^-) are held together by strong electrostatic attraction (ionic bonds). A large amount of energy is needed to overcome the strong ionic bonds holding the positive ions and negative ions together. Hence, sodium hydride has a high melting point.

c) i)

Formula of rubidium chloride	RbCl
Formula of rubidium hydride	RbH
Formula of rubidium sulfide	Rb_2S
Formula of rubidium sulfate	Rb_2SO_4
Formula of products of the reaction between rubidium and water	RbOH and H_2

ii)Rubidium is in period 5 while sodium is in period. Rubidium has a larger atomic radius than sodium. The outermiost electron of rubidium is further away from the nucleus as compared to eth outermost electron of sodium. The attraction between the nucleus and the outermost electron in rubidium is not as strong as that in sodium. Hence, this electron can be removed more easily. Thus, rubidium is more reactive than sodium and the reaction between rubidium and water is more vigorous.

8.) Cesium(Cs) is an element in Group l of the periodic Table.The atomic numbers of caesium is 55.

(a) Explain why cesium must be kept under dry argon.

(b) Explain why cesium reacts with ice at - 110°C but lithium does not.

Answer:

a) Cesium is a very reactive metal which reacts vigorously with oxygen and water in air. Thus, it should be stored under dry argon (an inert gas).

b) Cesium is in period 6 while lithium is in period 2.Cesium has a larger atomic radium than lithium. The outermost electron of cesium is further away from the nucleus as compared to the outermost electron in cesium is not as strong as that in lithium. Hence, this electron can be removed more easily.Thus,cesium is more reactive than lithium.

9.) The elements in group VII of the periodic Table are also known as halogens.

(a) Complete the table below.

Atomic numbers	Elements	Molecular formula	State at room temperature and pressure	Colour
9	F	F_2	gas	Pale Yellow
17	Cl			
35	Br		Liquid	
53	L	L_2		Black

(b) The halogen react with hydrogen to form hydrogen halides.

(i) Explain why all the hydrogen halides have low melting and boiling points.

(ii) Pure hydrogen halides are poor conductors of electricity but hydrogen halides diluted with water are good electrical conductors. Explain this phenomenon.

(iii) The table below shows the conditions required for the hydrogen and some halogens to take place.

Halogen	Conditions and rate of reaction
Fluorine	Explosive under all conditions; rapid reaction

Chlorine	Explodes in the presence of light, rapid reaction
Bromine	Reacts when heated in the presence of appropriate catalyst; slow reactoion

The reactivity of the halogen show a trend down the group. Describe the trend.

(c) Astatine is another group VII elements, found below iodine in the periodic Table.

(i) How many electrons are there in the outermost shell of astatine?

(ii) Suggest the appearance of astatine at room condition.

(iii) Write the formula of hydrogen astatide.

Answer:

9.) a)

Atomic number	Element	Molecular formula	State at room temperature and pressure	Colour
9	F	F_2	Gas	Pale yellow
17	Cl	Cl_2	Gas	pale yellow-green
35	Br	Br_2	Liquid	Red
53	I	I_2	Solid	Black

b)

i)The hydrogen halides have simple molecular structures. A small amount of energy is needed to overcome the weak intermolecular forces holding the discrete molecules together. Hence, the hydrogen halides have low melting points.

ii)The hydrogen halides have simple molecular structures in which the molecules are held together by weak intermolecular forces. They do not conduct electricity because of the absence of mobile charged particles (such as ions or electrons).

However, when diluted with water, the hydrogen halides react with water are able to conduct electricity because of the presence of mobile ions.

iii)The reactivity of halogens decreases down the group.

c) i) 7 ii) Black solid iii) HAt

10.) Chlorine, bromine and iodine are halogens. Given the following chemicals, describe how will you carry out a series of experiments to show the relative chemical reactivity of the three halogens.

Chemicals provided;

aqueous chlorine	aqueous potassium chloride
aqueous bromine	aqueous potassium bromide
aqueous iodine	aqueous potassium iodide

your answer should include;

- A table with the reagents which you will use and the observation you will see in the each experiment
- ionic equation for any reaction which takes place

Answer:

10.)Aqueous chlorine is mixed sequentially with aqueous potassium chloride, aqueous potassium bromide and aqueous potassium iodide. This experiment is repeated with aqueous bromine and aqueous iodine. The observations are shown below.

	KCl (aq)	KbR (aq)	Kl (aq)
Aqueous chlorine	No visible change; solution remains pale green	Solution turns orange due to the formation of Br_2	Solution turns dars brown; I2 formed.
Aqueous bromine	No visible change; solution remains orange	No visible change; solution remains orange	Solution turns dark brown; I_2 formed
Aqueous iodine	No visible change; solution remains dark brown	No visible change; solution remains dark brown	No visible change; solution remains dark brown

The experiments above show that:
Chlorine displaces bromine from potassium bromide.
$Cl_2 + 2\,Br^- \longrightarrow Br_2 + 2\,Cl^-$

Chlorine displaces iodine from potassium iodide.
$Cl_2 + 2I^- \longrightarrow I_2 + 2\,Cl^-$

Bromine displaces iodine from potassium iodide.
$Br_2 + 2I^- \longrightarrow I_2 + 2Br^-$

Hence, chlorine is the most reactive while iodine is the least reactive among the three halogens.

11.) For each of the followings reactions involving Group VII elements, predict whether or not a reaction will take place ,write a balanced chemicals equation for a the reaction.

(a) Lithium is heated with chlorine gas.

(b) Sodium is heated with astatine.

(c) Aqueous bromine is added to aqueous potassium bromide.

(d) Aqueous iodine is added to aqueous potassium bromide.

(e) Chlorine is bubbled through aqueous potassium iodide.

Answer:

a) Yes, a reaction takes place. $2Li + Cl_2 \longrightarrow 2\,LiCl$

b) Yes, a reaction takes place. $2Na + At_2 \longrightarrow 2NaAt$

c) Yes, a reaction takes place. $Br_2 + 2KAt \longrightarrow At_2 + 2KBr$

d) No reaction takes place.

e) Yes, a reaction takes place. $Cl_2 + 2KI \longrightarrow I_2 + 2KCl$

12.) Argon is the most abundant noble gas on Earth.

(a) element chlorine exist as diatomic molecules while argon is a monoatomic gas. give reason

(b) Argon is often used in light bulbs.
Give one property of argon that makes it suitable for this application.

(c) Describe how pure argon is obtained commercially.

Answer:

a) Argon has eight electrons in its outermost shell. The electronic configuration of argon (2, 8, 8) is very stable. Thus, argon is unreactive and exists as single atoms. On the other hand, the electron configuration of chlorine is 2, 8, 7. To achieve the stable electronic configuration of argon, each chlorine atom bonds covalently with another chlorine atom, forming diatomic molecules.

b) Argon is inert and will not react with the hot tungsten filament.

c) Argon is obtained from the fractional distillation liquified air.

13.) The table below contains about the first three Group 0 elements.

Name	Atomic number	Relative atomic mass	Density at r.t,p,/g dm^{-3}
Helium	2	4	0.167
Neon	10	20	0.833
Argon	18	40	1.67

a) Suggest why group elements are also known as noble gases.

b) Describe how the boiling points of group 0 elements change down the group.

c) Helium is used to fill small balloons. Explain what properties of helium make it suitable for this application.

d) Krypton is another group 0 element, found below argon in the periodic table.

i. State the relative atomic mass of krypton
ii.Calculate the density of krypton at r.t.p.

Answer:

a) Helium has two electrons in its outermost shell while the other Group 0 elements have eight electrons in their outermost shells. All the Group 0 elements have stable electronic configurations. Thus these elements are uncreative and seldom react with other substances. In addition, these elements have vey low melting and boiling point and exist as gases under room conditions.

b) The boiling points of the elements increase down the group.

c) helium has a very low density and is unreactive (not flammable)

d) i) relative mass of crypton is 84
ii) Density of krypton at r.t.p = 84 / 24 = 3.5 g/dm^3.

14.)The table below shows the physical properties of some transition metals.

Metal	Density gm^{-3}	Melting point / ^0C	Boiling point / ^0C	Relative electrical conductivity
Titanium	4.50	1660	3287	1
Vanadium	5.96	1890	3380	1.5
Chromium	7.20	1244	1962	5
Manganese	7.20	1244	1962	1
Iron	7.86	1535	2750	8.5
Cobalt	8.90	1495	2870	12
Nickel	8.90	1455	2730	12
Copper	8.92	1083	2567	49

a) List two differences between the physical properties of Group I metals and the transition metals.

b) Identify the metal with the highest electrical conductivity. Explain why it is often used to make electrical wiring but is NOT usually used to make overhead power lines.

c) Titanium is used to make many aircraft parts. Relate this application of titanium to its physical properties.

d) Calculate the number of moles of nickel atoms that are found in a 10 cm^3 block of nickel.

Answer:

a) The densities of Group I metals are low while the densities of transition metals are high. The melting points of Group I metals are also lower than the melting points of the transition metals.

b) Copper. Density of Copper is too high so it is usually not used to make overhead power lines.

c) Titanium has a low density (compared to the other transition metals) and is strong.

d) Mass of a 10 cm^3 block of nickel = 10 x 8.90 = 89 g
No. of mol of nickel = 89 / 59 = 1.51 (3 s.f.)

15.) Rubidium (Rb) and ruthenium (Ru) are metals.

Some properties of rubidium and ruthenium and their compounds are given in the table below.

Metal	Density /kg m^{-3}	Melting point / ^0C	Compounds formed (colour of compound)
Rb	1532	39	RdF (white) RbCl (white)
Ru	12370	2334	RuF_3 (brown, RuF_4 (yellow), RuF_6 (brown), $RuCl_2$ (brown), $RuCl_3$ (brown)

Based on the data above, suggest where the metals could be found in the periodic table. Support your answer by making reference to the physical and chemical properties of the metals.

Answer:

Rb can be found in Group I. it has low melting point and low density. The oxidation state of Rb is its compounds is +1) (each atom of Rb gives away one electron).
Ru is a transition metal. It has high melting point and high density.It forms compounds of variable oxidation states and these compounds are coloured.

16.) The Contact process is an individual process for producing sulfuric acid. During this process, sulfur dioxide is converted to sulfur trioxide. This is done by reacting sulfur dioxide gas and oxygen gas in the presence of vanadium (V) oxide.

In the absence of vanadium (V) oxide, the conversion of sulfur dioxide and sulfur trioxide is very slow. The chemical equation for this reaction is

$$2SO_2 + O_2 \longrightarrow 2\,SO_3$$

In the presence o vanadium (V) oxide, this conversion takes place faster and in two steps.

Step 1: $SO_2 + V_2O_5 \longrightarrow SO_3 + 2CO_2$

Step 2: $4\,VO_2 + O_2 \longrightarrow 2V_2O_5$

a) What is the role of vanadium (v) oxide in the contact process?

b) Based on the information given above, explain why only a small amount of vanadium (V) oxide is needed in the process.

c) Identify the oxidizing and reducing agents in steps 1 and 2.

Answer:

a) It acts as a catalyst.
b) V_2O_5 used up in step 1 is regenerated in step 2.
c) Step 1: Oxidising agent is V_2O_5 and reducing agent is SO_2
Step 2: Oxidising agent is O_2 and reducing agent is VO_2

17.) Antonie Lavoisier, a French scientist is often cited as the father of modern chemistry. In 1789, he published a book – Elementary Treatise of Chemistry – and in this text, he explained that an element is a substance that could not be broken down by any known method of chemical analysis. The table of elements proposed by Lavoisier is show below. Based on the experiments he had conducted and the above observations made, he classified the substances into four element groups.

Simple gas-like substances	Non-metallic elements	Metallic elements	Simple earthy elements
Light Heat Oxygen Nitrogen Hydrogen	Sulfur Phosphorous Carbon	Antimony Silver Arsnic Bismuth Cobalt Copper Tin Iron Manganese Mercurys Molybdenum Nickel Gold Platinum Lead Tungsten Zinc	Lime (calcium oxide) Magnesia (magnesium oxide) Barite (barium sulfate) Alumine (Aluminium oxide) Silice (silicon dioxide)

a) Your knowledge of the modern periodic table is that all the metallic elements listed by Lavoiusier metals?

b) i) From your knowledge of the modern periodic table, which one of the four element groups proposed by Lavousier are definitely not elements but compounds?

ii) Suggest why the scientists of the 18th century thought that these substances were elements.

iii) Identify two terms in the list which do not refer to an element or a compound.

Answer:

17.)(a) Yes

b)
i) The simple earthy elements.
ii) The scientists did not have the instruments to separate the elements in these compounds. For instance, electrolysis could not be done to extract the reactive meals from the salts.

iii) Light and heat.

18.) 0.78 g of a Group I element M reacts with oxygen to give 0.94 g of an oxide M_2O

a) Calculate the relative atomic mass of M and identify M.

b) 0.94 g of M_2O reacts further with more oxygen gas on heating to give a peroxide with the formula M_2O_2. Write a balanced chemical equation with state symbols, for the reaction between M_2O and oxygen.

c) i) Write the formula of the anion present in M_2O_2 and state the oxidation state of oxygen in this anion.
ii) Explain why this anion is expected to behave as an oxidizing agent.

Answer:

18.)
a) $4M + O_2 \longrightarrow 2M_2O$
Mass of O_2 reacted = 0.94 - 0.78 = 0.16 g
No. of moles of O_2 reacted = 0.16 / 32 = 0.005
No. of moles of M reacted = 4 x 0.005 = 0.02
Ar of M = 0.78 / 0.02 = 39
Thus, M is Potassium.

b) $2M_2O\ (s) + O_2\ (g) \longrightarrow 2M_2O_2\ (s)$

c)

i) O_2^{2-}, the oxidation state of O is -1

ii) O_2^{2-} is expected to behave as an oxidizing agent because it can be easily reduced to the stable oxide ion O_2^{-}. From O_2^{2-} to O_2^{-} the oxidation state of O is changed from -1 to -2.

19.) When chlorine is bubbled into a solution of iodine in hot aqueous sodium hydroxide, the two halogens react, forming a white precipitate W, sodium chloride and water. W has the following composition by mass:

Na	16.9 %
I	46.7 %
H	1.1%
O	35.3%

a) Calculate the empirical formula of W.
b) What s the oxidation state of iodine in W?
c) i) Given that 7 moles of chlorine reacts with 1 mole of iodine in the given reaction, construct a balanced chemical equation for this reaction.
   ii) Explain why this is a redox reaction.

Answer:

19.)(a)

	Na	I	H	O
% by mass	16.9	46.7	1.1	35.3
Ar	23	127	1	16
No. of moles	0.735	0.368	1.1	2.206
Simplest mole ratio	2	1	3	6

Hence, the empirical formula is $Na_2IH_3O_6$.

b) Let the oxidation state of I in $Na_2IH_3O_6$ be x.
$2(+1) + x + 3(+1) + 6(-2) = 0$

Hence, the oxidation state of I = x = +7

c)

i) $7Cl_2 + I_2 + 18NaOH \qquad Na_2IH_3O_6 + 14\ NaCl + 6H_2O$

ii) In this reaction, iodine is oxidized. The oxidation state of iodine changes from 0 in I_2 to +7 in $Na_2IH_3O_6$. On the other hand, chlorine is reduced. The oxidation state of chlorine changes from 0 in Cl_2 to -1 in NaCl.

Hence,this is a redox reaction.

SUMMARY AND KEY POINTS

1.) A vertical column of elements in the Periodic Table is called a Group.

2.) A horizontal row of elements in the Periodic Table is called a Period.

3.) An element that has the properties of both a metal and a non-metal is called a Metalloid.

4.) The Group I elements in the Periodic Table is called Alkali Metals.

5.) The Group VII elements in the Periodic Table is called Halogens.

6.) The Group 0 or VIII elements in the Periodic Table is called Noble Gases.

7.) A substance that increases (enhances) the speed of a chemical reaction is called a Catalyst.

8.) Periodic table is defined as " The arrangement of elements in certain groups and periods in such a way that elements with similar properties are grouped together."

9.) The chemical properties of the elements are the periodic function of their atomic numbers.

10.)The Periodic Table consists of Seven Periods and Eight Groups of elements.

11.) The block of element between Group II and Group III is called the transition elements or transition metals.

13.) Elements in the same group have the same number of valence electrons. This number is the same as the group number. Elements in the same group have similar chemical properties.

14.) The electronic configuration of an element can be worked out if we know the position of the element in the Periodic Table.

15.) The group number of an element tells us the number of valence electrons. The period in which the element is placed tells us the number of shells of electrons in the electronic configuration.

16.) Periodicity of elements is the repeatition of similar properties of elements placed in a group according to atomic numbers. The periodic table given by Moseley is called "long form of periodic table."

17.) Elements on the left of the periodic Table are metals, while those on the right side are non- metals .

18.) From left to right across a period, there is a decrease in the metallic properties and an increase in non-metallic properties .

19.) A metalloid is an element with properties that are intermediate between those of metals and non-metals. Examples are silicon, Si(Period 3), and germanium,Ge (Period 4).

20.) If the proton number or electronic configuration of an element is given, we can predict the metallic or non-metallic character of the element of the element and hence, predict the properties of the elements.

21.) In general,

i) metals form basic oxides,

ii) non-metals form acidic oxides and

iii) metalloids form amphoteric oxides .

22.) In the ionic compounds, *the metals lose electrons to form positive ions*.

The charges on the ions of the elements in Group I, II, and III correspond to the group numbers .

23.) The elements in Group V, VI, and VII are non-metals. They tend to form covalent compounds when they combine with non-metals .

24.) Periodic properties of elements are those properties which are directly or indirectly related to the electronic configuration of the elements and show gradation in moving down the group or along the period.

25.) Atomic radius is defined as "the distance between the centre of its nucleus and electron in the last orbit."

There are 3 forms of atomic radius based on the nature of the bonding. They are

i) COVALENT RADIUS

ii) VANDER WAAL'S RADIUS

iii) METALLIC RADIUS

26.) The atomic radii (covalent radii) of elements generally decrease from left to right in a period, while when we move down the group the atomic radii increases.

27.) Ionic radius is defined as " the distance between the centre of the nucleus to

the region up to which it can attract the electron cloud ".

28.) The radius of a cation is always smaller than the neutral atom from which it is formed. This is because when an electron is removed from the outermost shell, the neutral atom has become unipositive ion. It means the nucleus gains a proton. Now the nucleus can attract the remaining electrons more strongly, hence the electron cloud shrinks, resulting in decrease in ionic radius.

29.) Isoelectronic Ions are the species of different ions which have same number of electrons but different magnitude of nuclear charges.

The size of isoelectronic ions decrease with increase in atomic number.

30.) Inert gases are elements present in the zero group. They are called noble gases because they hardly react with other elements.

31.) The elements with atomic numbers 58 (Ce) to 71(Lu) are called lanthanides.

they are also called Rare earth elements.

32.) The elements with atomic numbers 90(Th) to 103(Lr) are called actinides.

they are radioactive elements.

33.) Lanthanides and Actinides series of elements have similar properties and have been placed in two rows at the bottom of the periodic table. They are collectively known as Inner- transition elements.

34.) Metals have a tendency to form a cation, thus reacting easily with an anion.

35.) Non-metals have a tendency to form an anion, thus reacting easily with a cation.

36.) Shortest period consisting of 2 elements H and He is period 1.

37.) Elements of the same period have the same number of electron shells.

38.) Ionisation Potential is the amount of energy required to remove an electron from the outermost shell of an atom and to make it a positive charge ion.

Ionisation potential increases from left to right across a period. It decreases from top to bottom going down a group.

39.) Electron affinity is the amount of energy released when an electron is gained by a neutral atom.

Electron affinity increases from left to right across a period. It decreases from top to bottom going down a group.

40.) Electronegativity is the tendency of an atom of an element in a molecule to attract the shared pairs of electrons towards itself.

Electronegativity increases from left to right across a period. It decreases from top to bottom going down a group.