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CHEMISTRY

Class 9th (KPK)

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CHEMICAL REACTIVITY

(Topic Wise Questions Answers)

Q1. Write the characteristic of Metal and Non-Metals.

Ans:

Electropositive Character:

All elements have the ability to lose electrons easily from their valence shells and get (+ive) charged to form cation. Electron losing ability is called electro-positivity.

Example:

Na + energy ➔ $Na^+ + 1e^-$ ΔE = 496 kj/mol

On other hand, non-metals have the ability to accept electrons in their valence shell to get (-ive) charged particle called Anion.

Example: Cl + e ➔ Cl^-

2. Electrical Conductance:

Metals are good conductor of heat and electricity. While non-metals are insulator.

The conductance in metal is due to mobile sea of electrons which are loosely held are responsible for the conduction of electric current.

3. Nature of Oxides:

Metal are basic in nature while non-metals oxide are acidic nature.

Example:

$Na + O_2$ ➔ Na_2O

$Na_2O + H_2O$ ➔ 2NaOH

Similarly

$2S + 3O_2$ ➔ 2SO3

$SO3 + H_2O$ ➔ H_2SO_4

Q2. What are Alkali Metals? Also explain occurrence of alkali metals.

Ans:

Alkali Metals:

The elements of the group IA except Hydrogen are called alkali metals.

The name Alkali came from Arabic language. It means Ashes. These metal were first found in the ashes of plants.

Some chemist had the opinion that the word alkali is given due to the fact that these elements react with water and forming the strong Alkalies. Alkali metal include the elements Lithium (Li), Sodium (Na), Potassium (K), Rubidium (Rb), Cesium (Cs) and Francium (Fr).

These metals have only one electron in their valence shell. Their valence sub-shell is ‘s’. They are highly electropositive elements. The alkali metals lose their one electron and form mono-positive ions. The ionization energy of alkali metals is low. The electron thus removed is provided to an electronegative element to form ionic compounds. Elements of group IA form ionic compounds with elements of group VIA and group VIIA.

Occurrence of Alkali Metals:

Alkali metals have low ionization energies. They are very reactive metals in nature that is why they do not occur in free state. Lithium found in the form of complex minerals. It mostly occurs in the form of spodumene, $LiAL(SiO_3)_2$. Sodium and potassium are abundantly (2.4%) found on the earth crust. Rubidium and Cesium occurs in small amounts in the potassium salts deposits. Francium is not found in nature. It is prepared in laboratory.

Q3. What are Alkaline earth Metals? Also explain occurrence of alkaline earth metals.

Ans: **Alkaline Earth Metals:**

The elements of group IIA are called Alkaline earth metals.

The name of this group is due to they produce the alkalies and are widely distributed in the earth crust. The Alkaline earth metals have two electrons in their valence shells. Their valence sub-shell is "s". They are electropositive metals. They lose the two valence electrons and form M^{+2} ions. Their ionization energies are low.

There are six alkaline earth metals, including Beryllium (Br), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). They become stable by gaining the electronic configuration of noble gases by losing their outermost electrons. These metals are often found in the form of sulphate in nature, Examples include the minerals such as gypsum (calcium sulphate), epsomite (magnesium sulphate) and barite (barium sulphate).

Occurrence of Alkaline earth metals:

Aklaline earth metals have low ionization energies, so they are very reactive metals. That is why they do not occur free in nature. Beryllium occurs in nature in small amount in the form of beryl. Magnesium and calcium are very abundant in the earth crust. Magnesium and calcium are present with sodium and potassium in rocks as cations. Magnesium halides are found in the sea waters. Magnesium is an important constituent of chlorophyll. Calcium is found in nature in the form of calcium phosphate and calcium fluoride. Calcium is the important constituent of living organism. It occurs as skeletal materials in bones, teeth, egg shells, etc. Radium is a rare element. It is radioactive in nature.

Q4. How ionization potential values vary for Group I and group II elements on descending the group?

Ans: **i. Energies of Group I and II elements:**

Ionization Energy:

The amount of energy required to remove an electron from an isolated gaseous atom of an element is called ionization energy.

Example: (Alkali Metals)

The alkali metals have one electron in their outer most shell. E.g.

$Na \rightarrow Na^{+} + 1e$ $\Delta E = 496$ kj/mole

(2.8.1) (2.8)

In alkali metals, sodium (Na) has the highest I. Energy in its own group due to smaller size and the small distance b/w the nuclear charge and valence electron. Down the group I.E is decrease due to increasing number of shells by increasing in atomic number.

Now the distance b/w the nuclear charge and the valence shell electrons are also increase. So, it is easier to remove an electron due to less bonded.

I.E of Group IA

Elements	Atomic No	Atomic Radius	I.E Jk/mol
Li	3	$1.52A^0$	520
Na	11	$1.86A^0$	496
K	19	$2.27A^0$	419
Rb	37	$2.48A^0$	403
Cs	55	$2.68A^0$	375

Alkaline earth metals:

Alkaline earth metals have two electrons in their valence shell. Since atomic radii decrease due to increase of nuclear charge therefore high amount of energy will be required to remove an electron from the valence shell.

$Mg \longrightarrow Mg^{+1} + 1e \quad \Delta E = 738$ kj/mol

Removal of second e after first one then

$Mg^{+1} \longrightarrow Mg^{+2} + 1e \quad \Delta E = 1450$ kj/mol

Ionization energy of element increase from left to right in a period due to smaller size and increasing nuclear charges.

Similarly:

It decreases down the group due to increasing of shell and the distance b/w the nucleus and valence shell electron also increase, so down the group ionization energy decreases. There is general decreasing order to melting and boiling points, hardness, conductivity and ionization energy down the group.

Group I	Group II	Properties			
Li	Be	Decreasing ↓	Melting point Boiling Point Hardness Conductivity Ionization Potential	Increasing	Electro positivity Atomic radii Atomic volume Reactivity Reducing power density
Na	Mg				
K	Cs				
Rb	Sr				
Cs	Ba				

Periods

		Decreasing ➔	Increasing ➔
Group I	Group II	Electro positivity	Melting point
Li	Be	Atomic radii	Boiling point
Na	Mg	Conductivity	Density
K	Ca	Atomic volume	Hardness
Rb	Sr	Reactivity	
Cs	Ba	Reducing agent	

Q5. What is the difference in the reactivities of Group I and Group II elements? Describe with Respect to the variation in atomic number and ionization potential.

Ans: **Differences in the reactivity of group IA and IIA.**

The reactivity of element shows that how much the element is reactive when it is reacted with other substances especially air, acids and water. The differences in the reactivities of group IA and group IIA with respective to atomic number and I.P is as follow.

Differences in the reactivities with respect variation in atomic number and I.P

As we go from group IA to group IIA, along a period, the atomic number increases, but the number of shells remain the same. Thus, the nuclear charge increases and the atomic size decreases.

Therefore, the valence electrons of group IIA are more tightly bound to the nucleus as compared to alkali metals and hence group IA elements are more reactive as compared to group IIA. Down the group both in group IA and IIA, the atomic size increases due to the addition of new shells although the atomic number increases which increase the nuclear charge. Thus, the valence electrons become farther from the nucleus and they can be easily removed. Therefore, the IE values decrease down the group due to which the reactivity is increased down the group.

Q6. Describe the position, properties and uses of Sodium.

Ans: **Sodium (Na):**

Sodium does not occur as a free metal in nature because it is too reactive metal and readily combines with other elements and compounds. It is found in sea as sodium chloride, sodium bromide and sodium iodide. It is also found in deposits as rock salt.

Its Latin name is “Natrium”

Position of Sodium in Periodic table:

Sodium belongs to alkali metals. Sodium atomic number is "11" and mass number is "23" and its symbol is "Na". It occupies first position in 3rd period and 3rd position in Alkali metal (Group IA) it has three electronic shell have only one electron in their valence shell.

Physical Properties:

i. It is silvery white solid.
ii. Na is soft metal and can be cut with a knife
iii. Its density ($\frac{m}{y}$) is 0.971g/cm^3
iv. Its melting point in 97.6^{0}C and boiling point is 880^{0}C.
v. It has relatively low tensile strength
vi. It is lighter than water and therefore floats on the surface of water.
vii. It ductile (which can be drawn to form wires) and malleable.
viii. It is good conductor of electricity due to the free movement of valence electrons.

Chemical Properties:

Sodium (Na) is highly reactive and can react with water (H_2O), hydrogen ((H_2), Oxygen (O_2) and halogens (Group VIIA)

1. Reaction with water:

Sodium react vigorously with cold water forming metal hydroxide with the liberation of hydrogen gas.

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

The reaction is exothermic (heat released) as a result hydrogen produced catch fire on the surface of water.

2. Reaction with hydrogen:

Sodium react with hydrogen to form hydrides.

$2Na + H_2 \rightarrow 2NaH($ ➔ metal hydrides)

3. Reaction with oxygen:

Sodium (which is metal) react with oxygen from basic oxide and react with water form Alkali (Base).

Examples:

$4Na\ (s) + O2\ (g) \rightarrow 2Na_2O$ (sodium oxide)

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

4. Reaction with halogens:

Sodium react with halogens (Group VIIA) to form sodium halide

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

$2Na\ (s) + Br_2(l) \rightarrow 2Na\ Br\ (s)$

5. Reaction with Sulphur:

Sodium is powerful reducing agent. It reduces the other substance but itself oxide.

$2Na\ (s) + S\ (s) \rightarrow Na2S(s)$

6. As a reducing agent:

Sodium is powerful reducing agent. It reduces the other but itself oxide.

$2Na^0(s) + Mg^{+2}O^{-2}\ (s) \rightarrow Na_2{}^{+1}\ O\text{-}2\ (s) + Mg^0\ (s)$

$4Na^0\ (s) + Ti^{+4}\ Cl_4\text{-}4\ (s) \rightarrow 4Na^{+1}\ Cl^{-1}\ (s)$ of Ti^0

In above case, the oxidation state of Na^0 is zero and in Na_2O, the oxidation state change to (+2) increasing oxidation state occur is reducing agent.

Uses of Sodium:

i. It is used in the preparation of important compounds such as sodium carbonate (Na_2CO_3), sodium bicarbonate ($NaHCO_3$), sodium hydroxide (NaOH), sodamide (NaNH2).

ii. It is used in sodium vapour lamps (which gave a bright orange-yellow light) for street lighting.
iii. It is used as coolant in nuclear reactors.
iv. It is used in purification of petroleum, in order to remove Sulphur from it. This process is called desulphurization.
v. It is used as reducing agent to prepare metals such as Titanium (Ti), Zirconium (Zr) from chlorides or oxides.
vi. It forms alloys with other metal. Its most useful alloy is with mercury (Hg) called sodium amalgam and with metal silver.

Q7. Write the position, properties and uses of Magnesium and Calcium.

Ans:

Magnesium:

Magnesium is the member of alkaline earth metals. It occurs in nature only in combined state, as Dolomite ($CaCO_3$, $MgCO_3$), kieserite ($MgSO_4$), Epsom salt ($MgSO_4.7H_2O$), in many silicates including talc and asbestos. Magnesium is present in sea water as chlorides and bromides. It is responsible for permanent hardness of water. It is also essential constituent of chlorophyll in green plants.

Position of Magnesium in Periodic Table:

Magnesium atomic number is 12 and its symbol is "Mg". It occupies second position in 3rd period and second in group IIA as it has three electronic shells and two electrons in their valence shell.

Calcium:

Calcium is too reactive to occur as free metal in nature. It occurs abundantly in the combined state in minerals such as calcium carbonate ($CaCO_3$), in lime stone, marble, chalk and as calcium sulphate ($CaSO_4$) in gypsum etc.

Position of Calcium in Periodic table:

Calcium atomic number is 20 and its symbols is "Ca". It occupies second position in 4th period and third position in group IIA as it has four electronic shells and two electrons in their valence shell.

Physical Properties of Magnesium

i. Magnesium is silvery grey solid.
ii. Its density is $1.74 g/cm^3$.
iii. Its melting point 651^0C and boiling point is 1106^0C.
iv. It is malleable and ductie.
v. It is good conductor of heat and electricity.

Physical properties of Calcium:

i. Calcium is silvery white solid.
ii. Its density is $1.55 g/cm^3$.
iii. Its melting point is $851C^0$ and boiling point is $1106C^0$.
iv. It is malleable and ductile.
v. It is good conductor of heat and electricity.

Chemical properties of Magnesium and calcium:

1. Reaction with H2:

Both "Mg" and "Ca" combined directly with hydrogen formed hydrides.

$Mg (s) + H_2 \rightarrow MgH_2$ (Magnesium hydride)

$Ca (s) + H_2 (g) \rightarrow CaH_2$ (Calcium hydride)

2. Reaction with Oxygen:

Both (Mg, Ca) burn in air. Magnesium burns with a dazzling flame forming MgO called magnesia.
$2Mg (S) + O_2(g) \rightarrow 2MgO (s)$
While calcium form CaO produce brick red coloured flame.
$2Ca (s) + O_2(g)$ heating $2CaO (s)$
Both (Mg, Ca) form base, when dissolved in water.
$2MgO (s) + H_2O (I) \rightarrow 2Mg (OH)_2$
$2CaO (s) + H_2O (I) \rightarrow 2Ca (OH)_2 (aq)$

3. Reaction with Nitrogen:

Both (Mg, Ca) react with "N_2" form nitrides
$3Mg (s) + N_2 (g) \rightarrow Mg_3N_2 (s)$
$3Ca (s) + N_2 (g) \rightarrow$ heat $Ca_3N_2 (s)$

4. Reaction with Acids:

Both (Mg, Ca) react with strong & dill acids give us hydrogen (H2) gas.
$Mg (s) + 2HCl (aq) \rightarrow MgCl_2 (s) + H_2 (g)$
$Mg (s) + H_2SO_4 (dil) \rightarrow MgSO_4 (s) + H_2 (g)$
$Ca (s) + 2HCl (dil) \rightarrow Cacl_2 (s) + H_2 (g)$
$Ca (s) + H_2SO_4 (aq) \rightarrow CaSO_4 (s) + H_2 (g)$

Reaction with Halogens:

They react with halogens form halides
$Mg (s) + Cl_2 (g) \rightarrow MgCl_2 (s)$
$Mg (s) + Br_2 (l) \rightarrow MgBr_2 (s)$
$Ca (s) + Cl_2(g) \rightarrow CaCl_2 (s)$
$Ca (s) + Br_2(l) \rightarrow CaBr_2 (s)$

Uses of Magnesium:

vi. Mg is low density metal, so it is used in the formation of light but tough alloys, such as Duralumin (a mixture of Al, Cu, Mg and Mn) Magnesium (a mixture of Al, Mg). These alloys are used for construction of aircrafts, cars and moving parts of machines.
vii. It is also used in phogrophic flashlight powder, flames and fireworks.
viii. It is used deoxidant in metallurgy and the extraction of Titanium and Uranium.
ix. Its compounds such as magnesium oxide (MgO) are mixed with clay, to make refractory bricks for furnace lining.
x. Magnesium sulphate (MgSO4) is used in textile, paper industry, soap formation and pharmaceutical industries etc.

Uses of Calcium:

i. Calcium is used as dioxide in steel coatings and copper alloys.
ii. It is used in the making of calcium and calcium hydride and in extraction of uranium.
iii. Their compounds such as lime CaO is added to soil in the form of fertilizers to decrease its acidity. It is also used for softening, pollution control and in pulp, paper, sugar and glass manufacturing industries.
iv. It is used in steel making.

Q8. What are Soft and Hard metals?

Ans: **Soft and Hard metals:**

Soft Metals:

Metals of group IA elements are quite soft, they react quickly with H2 and O2 and violently with H2o, and such metals are called soft metals. They are soft and have low melting and boiling point.

Example: Na, Li, k etc.

Hard metals:

The metal of "d" and "f" block elements are hard metals. They are hard in their physical appearance. Iron (fe), Copper (Cu) Silver (Ag), Cobalt (Co), Nickel (Ni), Tungsten (W) are hard, their melting point, boiling points and density show much higher values.

They do not react readily under ordinary conditions of temperature and pressure.

Both soft metals and hard have their own importance. Such as iron is used to prepare steel which is harder form of iron also used in heavy machinery locomotives, railway tracks in the construction of bridges.

Q9. Write down the comparison properties of Sodium (Na) and Iron (Fe).

Ans: Comparison properties of Sodium (Na) and Iron (Fe)

Sodium (Na)	Iron (Fe)
Sodium is an alkali metal, with atomic number 11	Iron is a transition metal, with atomic number 26.
One electron in its outermost shell and is very soft and can be cut with knife.	It is hard and requires great energy to break.
It has weak attractive force between the atoms of sodium.	It has strong attractive force between the atoms of iron.
The melting point of sodium is 97.6^0C	The melting point of iron is 1538^0C.
The boiling point of sodium is 880^0C.	The boiling point of iron is 2862^0C.
It has low density $0.927g/cm^3$	It has higher density $7.874g/cm^3$
It is lighter and floats on the surface of water.	It is heavy and settles at the bottom of water.
It has low tensile strength, cannot be used where stress is required.	It has high tensile strength which can be used in construction of builing and bridges. It is also used to prepare steel.
It is very reactive, stored in kerosene oil.	It is less reactive than sodium.

Q10. Write a note on commercial value of silver (Ag), Platinum (Pt) and Gold (Au)?

Ans: Those metals which are expensive and have great commercial and economic value are precious metals. In noble metals particularly silver (Ag), Platinum (Pt) and Gold (Au) are considered as precious metals.

i. Silver (Ag):

Silver is soft, white metal that usually occurs in nature in one of four forms, a) A native element, b) as a primary constituent in silver mineral, c) as a natural alloy with other metals, d) as a minor constituent in the ores of other metals.

Silver is known as a precious metal because it is rare and high economic value. It is valuable because it has a number of physical properties that make it the best possible metal for many different uses.

Pure silver is very soft. It is usually mixed with copper to form an alloy for making commercial articles. This alloy is used to make coins, jewellery and tableware. Silver chloride combine with silver bromide is used in photography. Silver is drawn into sheets and wires. It has higher electrical and thermal conductance and reflectivity than any other metal.

ii. Platinum (Pt):

The name platinum comes from the Spanish word patina meaning little silver. Platinum is the 72nd most common element in earth crust. That is why, platinum is an expensive metal. Platinum is heavy, soft, malleable, and ductile and has a fairly high melting point (1770^0C). It is noble metal because it is un-reactive. It does not even react with oxygen in air and resistant to react with acids.

Platinum is used in the catalytic converters to remove pollutants from the car engine exhaust gases. But as an expensive metal, so metals such as palladium etc. are used in its place.

The ease with which platinum can be shaped, its strength, colour, hardness and inertness make it suitable for jewellery and gem setting. Un-reactivity also makes it useful in dental fillings, making surgical tools and apparatus for scientific laboratories. Apart from that, platinum is also used in the electrical industry, in lasers and in making photographic materials.

iii. Gold:

Gold has been used to make ornamental objects and jewellery for thousands of years. Special properties of gold like high luster, attractive colour, inertness, tarnish resistivity, ability to be drawn into wires, hammered into sheets or cast into shapes etc. Make it perfect for manufacturing of jewellery.

Pure gold is too soft to resist the stress applied to many jewellery items. Alloying gold with other metals such as copper, silver and platinum increases its durability. Older than the coins were made of gold. Gold coins were commonly used in transactions up to paper currency because a more common form of exchange.

Gold is using a standard desktop or laptop computers. The rapid and accurate transmission of digital information from one component to another requires an-efficient and reliable conductor. Gold meets these requirements better than any other metal. The importance of high-quality reliable performance justifies the high cost. Gold alloys are used for dental filling, tooth crowns and orthodontic appliance. Gold is used in dentistry because it is chemically inert, non-allergic and easy for the dentist to work.

Q11. Write the electronegative characters of non-metals.

Ans: **Electronegative character of non-metals:**

The tendency of an atom of an element to gain electron from other element in order to become stable electronic configuration is called electronegative character.

Every element tries to complete its valence shell to become stable. The electronegative character increase across the period so all elements of group VIIA are most electronegative in their respective period. They accepted electron from less electronegative elements to complete its valence shell by octet rule. The general electronic configuration of group VIIA is ns^2np^5.

$X + e^- \rightarrow x$

Where x is called halide ion

$F + e^- \rightarrow F^-$ (Fluorine ion)

Group VIIA are collectively called halogen which mean salt formers. They contain fluorine (f), chlorine (cl), bromine (Br), Iodine (I) and Astatine (At)

Electro configuration and physical state of halogens

Name	AT No.	Symbol	Physical State	Electron Affinity	Electro negativity	Atomic Radium
Fluorine	9	F	Pale yellow gas	-322	4.0	72 x 10^{-9} m
Chlorine	17	Cl	Green Gas	-349	3.0	99 x 10^{-9} m
Bromine	35	Br	Dark red liquid	-325	2.80	114pm
Iodine	53	I	Dark crumble solid	-295	2.50	133 pm
Astatine	85	At	Black solid	-270	2.20	150pm

Q12. **Write the physical properties of Halogens.**

Ans: **Physical Properties:**

1. Fluorine and chlorine are gases.
2. Bromine is a fuming gas.
3. Fluorine is yellowish gas.
4. Cl_2 is greenish yellow colour gas.
5. Bromine is radish brown liquid.
6. Iodine is deep violet solid.
7. Astatine is Black solid.

They all are electronegative in all over the periodic table. They form (-ive) ion when react with metals. Fluorine is the most electronegative atom and form hydrogen bonding. Their melting and boiling point increase down the group.

CHEMICAL REACTIVITY
(Long Questions Answers)

Q1. **Compare and contrast the properties of alkali and alkaline earth metals, with reactions.**

Ans: **Comparison between physical properties of alkali and alkaline earth metals:**

ALKALI METALS	ALKALINE EARTH METALS
They are all silvery white metals.	They are all silvery white metals but be is grayish white.
They are soft metals.	They are soft metals but harder than alkali metals.
They have large atomic sizes	They have small atomic sizes
They have large atomic radii and ionic radii	They have small atomic radii and ionic radii
They have lower melting points and boiling points.	They have higher melting and boiling points.
They have lower densities	They have higher densities
They have low ionization energies and electronegativity values	They have higher ionization energies and electronegativity values.
They have lower electron affinity but higher than alkaline earth metals.	They have exceptionally lower electron affinity than alkali metals.
They are less conductor of heat and electricity	They are more conductor of heat and electricity

Comparison between chemical properties of alkali and alkaline earth metals:

Alkali metals are more reactive than the alkaline than the alkaline earth metals because they have one electron in valence shell while alkaline earth metals have two electrons. Therefore, alkali metals can easily loss their electron and are more reactive.

i. Reactive with hydrogen:

Elements of group 1A and group IIA react with hydrogen and forms their respective hydrides.

$2Na + H_2 \rightarrow 2NaH$

$Mg + H_2 \rightarrow MgH_2$

ii. Reactive with Oxygen:

Elements of group IA and group IIA with oxygen and forms their respective oxides.

$2Na + O_2 \rightarrow 2NaO$

$Mg + O_2 \rightarrow MgO_2$

iii. Reaction with Halogens:

Both reacts with halogens forming halides.

$2Na + Cl_2 \rightarrow 2NaCl$

$Mg + Cl_2 \rightarrow MgCl_2$

iv. Reaction with water:

Most of alkali metals reacts with water liberating hydrogen gas while alkaline earth metals react slowly except beryllium which do not react with water.

$2Na + H_2O \rightarrow 2NaOH + H_2$

$Mg + H_2O \rightarrow MgO + H_2$

v. Reaction with Nitrogen:

Among the alkali metals only lithium reacts with nitrogen, while all alkaline earth metals react with nitrogen forming nitrides.

$6Li + N_2 \rightarrow 2Li_2N$

$3Mg + N_2 \rightarrow Mg_3N_2$

Q2. a **Differentiate between soft and hard metals.**

Ans: **a. Differences between soft and hard metals.**

SOFT METALS	HARD METALS
The metals present in group IA and IIA are very soft, very reactive, low ionization energies and very electropositive are known as soft metals.	Metals like copper, silver, iron etc. are less reactive, having high ionization energies, less electropositive are known as hard metals.
They are very soft in nature even some of them can be cut with knife.	They are very hard in nature and requires greater energy to break.
They are very reactive in nature.	They are less reactive in nature
They have low ionization energies	They have high ionization energies
They are very electropositive in nature	They are less electropositive in nature.
They have low densities and are lighter which can float on the surface of water.	They have higher densities and are heavy which settles as the bottom of water.
They have low melting points and boiling points.	They have higher melting and boiling points.
They readily react with H_2 and O_2	They do not readily react with H_2 and O_2 at normal condition.
They violently react with H_2O	They may or may not react with H_2O, some react with H_2O but very slowly
They have weak attractive force between the atoms.	They have strong attractive force between the atoms.

Q2. b **Give the reaction of magnesium with: i. H_2 ii. HCl iii. O_2 iv. H_2O v. Cl_2**

Ans: **b. Chemical properties of Magnesium:**

1. Reaction with H2:

Magnesium combined directly with hydrogen formed hydrides.

$Mg (s) + H_2 (g) \rightarrow MgH_2 (s)$ (Magnesium hydride)

2. Reaction with Oxygen:

Magnesium burns in air with a dazzling flame forming MgO called magnesia.

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

3. Reaction with H2O:

When magnesium oxide (MgO) is dissolved in water, it forms basic solution.

$2MgO (s) + H_2O (I) \rightarrow 2Mg (OH)_2 (aq)$

4. Reaction with HCl:

Magnesium (Mg) reacts with HCl give us hydrogen (H_2) gas.

$Mg (s) + 2HCl (aq) \rightarrow MgCl_2 (s) + H_2 (g)$

5. Reaction with Cl2:

Magnesium reacts with chlorine and form halides

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

Q3. Discuss the reasons why some elements exist as free elements in nature while other occurs In combined states as compounds. Give two examples of each.

Ans: **Elements exist in Free state:**

Some elements exist in free state because they have completed their outermost shells and are stable. These elements have high ionization energies and having low reactivity. So that's why they cannot easily take part in a chemical reaction and do not form chemical bond with other elements, and exist freely.

Examples:
Group VIIIA (noble gases)
Noble material like Ag, Cu, Hg and Au.

Elements in combined state:
Some elements exist in combined state because they have incomplete their outermost shells and are unstable. These elements have low ionization energies and large atomic size. They having high reactivity. So, to complete their outermost shell they easily take part in a chemical reaction and forms chemical bond with other elements, and cannot exist freely i.e. occur in combined state.

Examples:
Alkali metals and alkaline earth metals.
Halogens, Carbon family and Oxygen family etc.

Q4. Define metal and non-metal and compare the properties (both physical and chemical) of metals and non-metals.

Ans:

Metal:
A metal is an element which loses an electron and forms a cation.

Explanation:
Metals are those substances which are good conductor of heat and electricity. Their oxides and hydroxides are basic in nature. When a metal reacts with in oxygen it produces a basic oxide. When it is dissolved in water it forms an alkaline solution which turns red litmus paper into blue.

Examples:
Elements of group IA except hydrogen, Group IIA, transition elements, lanthanides and actinides.

Non-metal:
A non-metal is an element which gains an electron and forms an anion.

Explanation:
Non-metals are those substances which are non-conductor of heat and electricity. Their oxides and hydroxides are acidic in nature. When a non-metal reacts with in oxygen it produces an acidic oxide. When it is dissolved in water it forms an acidic solution which turns blue litmus paper into red.

Examples:
Hydrogen, boron of group IIIA, C and Si of group IVA, N and P of group VA, Group VIA, Group VIIA and Group VIIIA.

Comparison between the properties (both physical and chemical) of metals and non-metals.

METALS	NON-METALS
They are good conductor of heat and electricity.	They are non-conductor of heat and electricity.
Their oxides and hydroxides are basic in nature.	Their oxides and hydroxides are acidic in nature.
They are ductile, malleable and sonorous.	They are not ductile, malleable and sonorous.
They are usually solids at room temperature except mercury.	They are present in all three states of matter i.e. solid, liquids and gases.
They electron donor in chemical reactions. They are reducing agents.	They are electron acceptor in chemical reactions. They are oxidizing agents.
They become positively charged ion in solution.	They become negatively charged ion in solution.
They are electropositive in nature.	They are electronegative in nature.
They form electrovalent (ionic) chlorides.	They form covalent chlorides.

Some metals can replace hydrogen from acids to form salts	Non-metals cannot replace hydrogen from acids.
They do not combine easily with Hydrogen. Few hydrides are formed are electrovalent. $2Na + H_2 \rightarrow 2NaH$ $2K + H_2 \rightarrow 2KH$	They combine easily with hydrogen to form many stable hydrides. $H_2 + Cl_2 \rightarrow 2HCl$ $H_2 + F_2 \rightarrow 2HF$ $H_2 + L_2 \rightarrow 2Hl$

Q5. Halogens are very reactive elements, write down halogen's reactions with hydrogen, oxygen,

metals, non-metals and other compounds along with displacement reaction.

Ans: **Chemical Properties of Halogens:**

All halogens are very reactive elements and exist in diatomic state with single covalent bond.

i. Reaction with H2:

H_2 (g) + F_2 (g) quick 2HF (aq)

Fluorine react with hydrogen vigorously

H_2 (g) + Cl_2 (g) quick 2HCl (aq)

H_2 (g) + Br_2 (I) light 2HBr (aq)

H_2 (g) + l_2 (s) <===> 2Hl (aq)

ii. Reaction O2:

Fluorine react with oxygen to form monoxide and dioxide (di-oxygen – di fluoride)

O_2 (g) + $2F_2$ (g) $\rightarrow$ $2OF_2$ = (g) (monoxide)

O_2 (g) + F_2 (g) $\rightarrow$ $O2F_2$ (g) (dioxide)

iii. Reaction with metals:

Halogens react with metals and form corresponding halides.

Cu (s) + Br_2 (I) $\rightarrow$ $CuBr_2$ (Copper bromide)

2K (s) + l_2 (s) $\rightarrow$ 2Kl (s) (Potassium iodide)

iv. Reaction with non-metals:

Halogens react with non-metals such as phosphorous to form PCl3 (tri-chloride) and PCl5 (Penta chloride)

2P(s) + $3Cl_2$ (g) $\rightarrow$ $2PCl_3$ (aq)

2P(s) + $3Br_2$(g) $\rightarrow$ $2PBr_3$(aq)

2P(s) + $5Cl_2$ (g)$\rightarrow$ $2PCl_5$ (aq)

2P(s) + $5Br_2$ (g)$\rightarrow$ 2PBrs (aq)

v. Reaction with other compounds:

Halogens oxidized other compounds but itself reduce, during reaction

H_2S (aq) + Cl_2 (g) $\longrightarrow$ 2HCl (Aq) + S (s)

$2NH_3$ (aq) + $3Cl_2$ (g) $\longrightarrow$ 6HCl (aq) + N_2 (g)

vi. Displacement Reactions:

During displacement reaction a more reactive halogen will displace a less reactive halogen from its halide solution. They reactivity of halogen decreases down the group.

Order of reactivity F > Cl > Br > I > As.

Examples:

i. $2NaBr + Cl_2 \rightarrow 2Nacl + Br_2$

ii. $2Kl + Br_2 \rightarrow 2Kbr + l_2$

CHEMICAL REACTIVITY
(Short Questions Answers)

Q1. Identify at least wo groups which contain only metallic elements.

Ans: In periodic table most of the metal elements are present at the left side of the periodic table. All The elements of group IA (Except hydrogen) and group IIA are metals. Group IA contain Li, Na, K, Rb, Cs and Fr while group IIA contains Be, Mg, Ca, Sr, Ba and Ra.

Q2. Write the reaction of group IA metals with oxygen, with balance equations.

Ans: **Alkali metals react with oxygen and forms various types of oxides:**

i. In presence of oxygen lithium burns with red flame and give lithium oxide, which is white solid.

1. $4Li\ (s) + O_2\ (g) \rightarrow 2Li_2O\ (s)$

ii. In presence of oxygen sodium burns with bright yellow flame and give white sodium oxide.

2. $4Na\ (s) + O_2\ (g) \rightarrow Na_2O_2\ (s)$

iii. In presence of oxygen potassium burns violently with a little coloured flame and give white potassium oxide.

3. $K\ (s) + O_2\ (g) \rightarrow KO_2\ (s)$

iv. Similarly rubidium and cesium catch fire in air and produce superoxide.

4. $4Rb\ (s) + O_2\ (g) \rightarrow 2RbO_2\ (s)$
5. $4Cs\ (s) + O_2\ (g) \rightarrow 2CsO_2\ (s)$

Q3. State the physical properties of metals.

Ans: **Physical properties of metals:**

i. All metals are solid at room temperature and on atmospheric pressure except mercury.
ii. Metals are malleable i.e. they can be beaten into sheets and foils.
iii. Metals are ductile i.e. they can be drawn into wires.
iv. All the metals are good conductors of heat and electricity.
v. Metals are lustrous i.e. they have shiny surfaces.
vi. Metals are sonorous i.e. they produce ringing sound when struck.
vii. They have high melting points and boiling points.
viii. They have low I.E, E.A and E.N.
ix. They have large atomic masses as compare to nonmetals.

Q4. How does sodium act as reducing agent and write down its reaction also?

Ans: Sodium is powerful reducing agent. It reduces the metal oxides into metals and itself oxidize.

$2Na^0\ (s) + Mg^{+2}\ O\text{-}2\ (s) \rightarrow Na_2^{+1}\ O^{-2}\ (s) + Mg^0\ (s)$

$4Na^0\ (s) + Ti^{+4}\ Cl_4^{-4}\ (s) \rightarrow 4Na^{+1}\ Cl^{-1}\ (s) + Ti^0\ (s)$

In above case, the oxidation state of Na0 is zero and in Na_2O, the oxidation state change to (+2) increasing oxidation state occur is reducing agent.

Q5. Ionization energy of Alkaline earth metals is higher than alkali metals, why?

Ans: The amount of energy required to remove an electron from isolated gaseous atom of an element is called ionization energy.

I.E of alkaline earth metals (group IIA) is higher than alkali metals (group IA). Because the atomic size of alkaline earth metal is smaller than the atomic size of alkali metals.

Akali metals have one electron in their outer most shell while alkaline earth metals have two electrons. Therefore, higher energy is needed to remove two electrons from alkaline earth metals as compare to alkali metals.

Example: Group 1A

After removing one electron 1e
$Na \rightarrow Na^+ + 1e$
I.E = 496 kj/mol

after removing one electron
$Mg \rightarrow Mg^+ 1e$
I.E = 738 kj/mol

Q6. Pure gold is not used for ornaments, why?

Ans: Gold has been used to make ornamental objects for thousands of years. Special properties of gold Like very high luster, attractive colour, inertness, resistivity etc. makes it perfect for manufacturing jewelry. Pure gold is not used for ornaments because pure gold is too soft and malleable to resist the stresses applied due to which it can easily be, deshaped, by applying a little force. Therefore, alloying gold with other metals such as copper, silver and platinum increase its strength, hardness and durability.

Q7. What are the uses of Magnesium?

Ans: **Uses of Magnesium (Mg):**

i. Mg is low density metal, so it is used in the formation of light but tough alloys, such as Duralumin (a mixture of Al, Cu, Mg and Mn) Magnesium (a mixture of Al, Mg).
These alloys are used for construction of aircrafts, cars and moving parts of machines.
ii. It is also used in photographic flashlight powder, flames and fireworks.
iii. It is used as deoxidant in metallurgy and in the extraction of titanium and Uranium.
iv. Its compounds such as magnesium oxide (MgO) are mixed with clay, to make refractory bricks for furnace lining.
v. Magnesium Sulphate (MgSO4) is used in textile, paper industry, soap formation and pharmaceutical industries etc.

Q8. Write down the reaction of chlorine with sodium hydroxide with balance equation.

Ans: **Reaction of chlorine with sodium hydroxide:**

Chlorine reacts with sodium hydroxide into two ways:
i. When chlorine is passed through the cold solution of sodium hydroxide, then the sodium hypochlorite is formed.
$Cl_2 + 2NaOH \rightarrow NaCl + NaClO + H_2O$
ii. When chlorine is passed through the hot solution of sodium hydroxide, then the sodium chlorate is formed.
$3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$

Q9. How does ionization energies values vary in a group?

Ans: As we know that the atomic size increases down the group due to increasing number of shell and the distance between the nucleus and valence shell electrons. As a result of force of attraction between the nucleus and valence electrons decreases so, less amount of energy will be required to remove and electron. So down the group ionization energy decreases.

Examples:

LE of Li = 520 kj/mol
I.E of Na = 496 kj/mol

Q10. What happens during displacement reaction in halogens?

Ans: During displacement reaction a more reactive halogen will displace a less reactive halogen from its halide solution. The reactivity of halogen decreases down the group.
Order of reactivity: F > CL > Br > I > As.

Examples:

$2NaBr + Cl_2 \rightarrow 2Nacl + Br_2$
$2Kl + Br_2 \rightarrow 2Kbr + I_2$