**Pakistan School , Kingdom of Bahrain**

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**Subject: CHEMISTRY Grade : 12**

**Book: CHEMISTRY Grade 12 (NBF) FIRST TERM**

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**Chapter 13: s – and p – block elements.**

**PERIOD 3(Na to Ar)**

**Atomic And Physical Properties Of The Period 3 Elements:** This period contains Sodium (Na), Magnesium (Mg), Aluminum (Al), Silicon (Si), Phosphorus (P) , Sulphur (S), Chlorine (Cl) and Argon (Ar)

**Atomic Properties: Electronic Structures:** In period 3 of the Periodic Table, the valence orbitals are 3s and 3p.They are filled with electrons.

**TRENDS IN ATOMIC RADIUS:**

* The distance of outermost electron from the nucleus of an atom while considering it spherical is called atomic radius.
* The size of atom decreases from left to right along the period. In a given period, the number of shells in all the elements remains the same. However, the value of effective nuclear charge increases from left to right due to increase in the number of protons. The increased effective nuclear charge pulls the electron cloud of the atom closer to the nucleus. Thus, the size of the atoms and the ions goes on decreasing from left to right. Along the period the atomic and ionic radii of s and p block elements decreases with the increase of atomic number.
* In any period, the alkali metals are present at the extreme left of the periodic table. Hence, they have largest size in any period.
* In any period, the halogens are present at the extreme right of the periodic table (excluding the noble gases). Hence they have smallest size in any period.
* In any period, the size of noble gases is larger than the previous halogen.

**Trends in the first ionization energy**: the first ionization is the energy required to remove the most loosely held electron from 1 mole of gases atoms to produce 1 mole of gaseous ion.(1+)

A -----🡪 A+ + 1e-

**Factors affecting ionization energy**. The ionization energy depends upon the following factors.

**(i)Atomic size:** Increase in atomic size decreases the hold on outer electrons, therefore Ionization energy decreases.

**(ii)Nuclear charge:** Increase in nuclear charge increases the hold on electrons, therefore Ionization energy increases.

**(iii)Shielding effect:** Increase in shielding effect decreases the hold on outer electrons, therefore Ionization energy decreases.

**(iv)Nature of orbital:** The order of ionization energy for different orbitals is: s>p>d>f.

**TRENDS:**

* Generally the ionization energy of the elements increases the left to right in a periodic table due to successive increase in the nuclear charge and decrease in atomic size.
* However, some irregular trends in period 3 are present e.g. ionization energy of Mg and P are higher than those of Al and S respectively.
* Mg has higher value of ionization energy as compared to Al. It is due to the following reasons.

In case of Mg(1s2,2s2,2p6,3s2) it is more difficult to remove an electron from the completely filled 3s-orbital. In case of Al (1s2,2s2,2p6,3s2,3p1) it is easier to remove from partially filled 3p-orbital. Since to remove an electron from a 3s-orbital of Mg atom requires more energy than to remove the same from 3p-orbital of Al atom, thus ionization energy of Mg is higher than that of Al.

**P has higher value of ionization energy as compared to S. It is due to following reasons.**

In case of P (1s2,2s2,2p6,3s2,3px1,3py1,3pz1) it is more difficult to remove an electron from the 3p-orbital,since all electrons are unpaired.

In case of S (1s2,2s2,2p6,3s2,3px2,3py1,3pz1)it is easier to remove one electron from the electron pair in 3px orbital due to mutual repulsions of the electron pair. Hence, ionization energy of P is higher than that of S.

**Trends in Electronegativity:** Electronegativity is the measure of the tendency of an atom to attract a bonding pair of electrons.

**Scale of Electronegativity:**

* The Pauling scale is the most commonly used.
* Fluorine is the most electronegative element. It is assigned a value of 4.
* The values ranges down to cesium and francium which are least electronegative at 0.7.

**Trend or variation in Electronegativity along the period:** THE ELECTRONEGATIVITY VALUES **INCREASES FROM LEFT TO RIGHT** IN A PERIOD OF S- AND P-BLOCK ELEMENTS.

**REASONS:**

1. In a period ,the atomic size decreases from left .Since smaller atoms have greater tendency to attracts the electrons towards themselves. Thus smaller atoms have higher electronegativity values.
2. In a period, values of ionization energy and electron affinity of elements increases from left to right .Thus the atoms of the elements with higher values of ionization energy and electron affinities also have higher electronegativities.
3. In a period, the number of protons in the nucleus increases from left to right .Thus, due to increase in nuclear charge, it attracts the bonding pair more closely. Hence electronegativity increases from left to right.

**Electronegativity cannot be assigned to Argon . Why?**

It is because electronegativity is the tendency of an atom to attract a bonding pair of electrons. Since Argon does form a covalent bonds, electronegativity cannot be assigned to it.

**Physical Properties:** The properties like electrical conductivities, melting point, boiling point, etc. depends upon the structure of elements.

**Structures of the elements: T**he structure of elements changes along the period. In period 3 the first three elements (i.e. Na, Mg ,Al ) are metallic, silicon is a giant covalent, and the remaining (i.e .P, S, Cl, Ar ) are simple molecules.

(1) **Three metallic structures :**

* Sodium, magnesium and aluminium all have metallic structure.
* In sodium, only 1 electron per atom is involved in the metallic bond i.e the single 3s electron. In magnesium both of its outermost electrons are involved. In Aluminium all 3 valence electrons are involved.
* The other difference is the way of packing of atoms in the metal crystal
  + - Sodium is 8-co-ordinated. Thus each sodium atom is touched by only 8 other atoms.
    - Both magnesium and aluminium are 12-co-ordinated. However both are packed in slightly different ways. This is a more efficient way to pack atoms. Hence less space is wasted .In the metallic structure stronger bonding is present.

**(2) A giant covalent structure:**

* Silicon has a giant covalent structure just like diamond.
* The structure is held together by strong covalent bonds in all three dimensions.

**(3) Four simple molecular structure :**

* The structure of phosphorous (i.e. white etc.) and Sulphur (i.e. rhombic or monoclinic etc.) depends upon the allotropes forms of phosphorus and sulphur.
* The atoms in each of these molecules are held together by covalent bonds.
* In argon there is no bond.
* In the liquids or solid state, the molecules are held close to each other by Vander Waals dispersion.

**Electrical Conductivity:** It is the capability of a substance to conduct the electric current due to free electrons with changing the composition of the substance.

**Reason For Electrical Conductivity :**

* Presence of relatively large number of electrons in the valence shell of the elements.
* The easy movement of valence electrons in their respective lattice.
* Sodium, magnesium and aluminium are all good conductors of electricity. Among these conductivity increase from left to right (i.e. from sodium to magnesium to aluminium ).It is because the number of valence (free) electrons increases from sodium to magnesium to aluminium. (1,2 &3 )
* Silicon is a semiconductor.
* The rest of elements of period 3 (phosphorus, Sulphur, chlorine and argon) do not conduct electricity.

**Explanation:**

* The three metals (Na,Mg,Al) conduct electricity. It is because the delocalized electrons (electrons sea) are free to move throughout the solid or the liquid metal.
* In the silicon case, the semiconductor conducts electricity depending upon conditions of temperature etc. Generally conductivity increases with increase in temperature.
* The remaining elements do not conduct electricity because they are molecular substances. They do not have free electrons to carry current.

**The melting and boiling point of the elements increase from left to right upto the middle in the period 3elements and decreases onwards. Why?**

**Trends in Melting and Boiling Points:**

* The figure 13.5 (page 11) shows the variations of melting and boiling points of elements across the period 3.
* The melting and boiling point are given in Kelvin rather than 0C to avoid negative values.
* The melting and boiling point tell about the strength of binding forces present in atoms, ions and molecules. Generally stronger the forces, higher the melting and boiling point and vice versa.

**Trends:** The melting and boiling point increases upto the middle (i.e. iv-a)then decreases from group v-a to noble gases (zero group elements).

**Reasons:**

* The melting and boiling points mainly depends upon two factors: (i) size of the molecules. (ii) strength of intermolecular forces i.e. number of binding electrons present in the valence shell.
* The number of binding electrons increases upto the middle, therefore melting and boiling point increases.
* The melting and boiling point are low from P to Ar. It is because these elements exist in molecular form and thus have weak intermolecular forces.

**CHEMICAL REACTIONS OF THE PERIOD 3 ELEMENTS:**

1. **Reactions with Water:**

**Sodium:**

* Sodium has an exothermic reaction with cold water producing hydrogen gas and a colorless solution of sodium hydroxide. The reaction is so much exothermic, that the liberated hydrogen catches fire.

**2Na + 2H2O🡪 2NaOH +H2**

**Magnesium:**

* Magnesium has a very slight reaction with cold water, but burns in steam.
* A very clean coil of magnesium dropped into cold water finally get covered in small bubbles of hydrogen which floats on the surface.
* Magnesium hydroxide is formed as a very thin layer on the magnesium. It acts as protective layer and thus tends to stop the reaction .

**Mg + 2H2O (cold) 🡪 Mg(OH)2 +H2**

* Magnesium burns in steam with its typical white magnesium oxide and hydrogen.  **Mg + H2O (steam) 🡪 MgO**

**Aluminium:**

* When Aluminium powder is heated with steam, it produces hydrogen and aluminium oxide.
* The reaction is relatively slow because of the strong aluminium oxide layer on the metal and formation of more oxide during the reaction. The oxide layer acts as a protective layer and thus slows down the reaction.

**2Al + 3H2O (steam)🡪Al2O3 +3H2**

**Phosphorous and Sulphur:** These have no reaction with water.

**Chlorine:**

* Chlorine dissolves in water to some extent to give a green solution. A reversible reaction takes place to produce a mixture of hydrochloric acid and chloric (I) acid (hypochlorous acid).

**Cl2 + H2O 🡪 HCl +HOCl**

* In the presence of sunlight, the chloric (I) acid slowly decomposes to produce more hydrochloric acid releasing oxygen gas.

**2HOCl 🡪 2HCl + O2**

**Argon:** There is no reaction between argon and water.

1. **Reactions with oxygen:**

**Sodium:**

* It burns in oxygen with orange flame to produce a white solid mixture of sodium oxide and sodium peroxide. For simple oxide : **4Na +O2 🡪2Na2O.** For peroxide  **2Na + O2 🡪Na2O2**

**Magnesium:**

* Magnesium burns in oxygen with an intense white flame to give white solid magnesium oxide. **2Mg + O2🡪 2MgO**

**Silicon:** It burns in oxygen if heated strongly, silicon dioxide is produced.

**Si + O2 🡪 SiO2**

**Phosphorus :**

* White phosphorus catches fire spontaneously in air. It burns with a white flame and producing clouds of white smoke which is a mixture of phosphorus (III) oxide and phosphorus (V) oxide. The production of two oxides depends on the amount of oxygen available.
* In an excess of oxygen, the product will be almost completely phosphorus (V) oxide.

For the phosphorus (III) oxide: **P4 +3O2 🡪 P4O6**

For the phosphorus (V) oxide: **P4 + 5O2 🡪 P4O10**

**Sulphur:** It burns in air or oxygen on gentle heating with a pale blue flame. It produces colourless sulphur dioxide gas. **S+ O2 🡪 SO2**

**Chlorine and Argon :** Despite having several oxides, chlorine will not react directly with oxygen. Argon does not react either.

1. **Reactions with Chlorine :**

**Sodium:** It burns in chlorine with a bright orange flame. White solid sodium chloride is formed. **2Na +Cl2** 🡪 **2NaCl**

**Magnesium:** It burns with an intense white flame to give white magnesium chloride.  **Mg +Cl2 🡪 MgCl2**

**Aluminium:** It often reacts with chlorine by passing dry chlorine over aluminium foil heated in a long tube. The aluminium burns in stream of chlorine to produce a very pale yellow aluminium chloride. This sublimes from solid to vapour and collects in cooler part of the tube.

**2Al +3Cl2 🡪 2AlCl3**

**Silicon:** When chlorine is passed over silicon powder heated in a tube, it reacts to produce silicon tetrachloride. Silicon tetrachloride is a colorless liquid which vaporizes and can be condensed further along the apparatus.

**Si + Cl2 🡪 SiCl4**

**Phosphorus:** White phosphorus burns in chlorine to produce a mixture of two chloride, phosphorus (III) chloride and phosphorus (V) chloride. PCl3 is a colourless fuming liquid. **P4 +6 Cl2🡪4PCl3**.

Phosphorus (V) chloride is an off-white (going towards yellow) solid.

**P4 + 10Cl2🡪 4PCl5**

**Sulphur:** When a stream of chlorine is passed over heated sulphur it reacts to form an orange, evil-smelling liquid, disulphur dichloride, S2Cl2

**2S +Cl2 🡪 S2Cl2**

**Chlorine and Argon:** Chlorine gas cannot reacts with itself. Argon does not reacts with chlorine.

**PHYSICAL PROPERTIES OF THE OXIDES :**

**The oxides:** The important oxides of period 3 elements are :

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Na2O** | **MgO** | **Al2O3** | **SiO3** | **P4O10** | **SO3** | **Cl2O7** |
|  |  |  |  | **P4O6** | **SO2** | **Cl2O** |

Those oxides in the first row are known as the highest oxides of the various elements i.e. in these oxides, the period 3 elements are in their highest oxidation states. In highest oxides, all the valence electrons in the period 3elements are involved in the bonding e.g. one electron for sodium, two for magnesium, three for aluminium and all seven for chlorine.

1. **Structures of oxides:** The trend in structure changes from the metallic oxides containing giant structure of ions on the left of the period via a giant covalent oxide (silicon dioxide) in the middle to molecular oxides on the right
2. **Melting and boiling points :**

* The giant structures (the metal oxides and silicon dioxide) will have high melting and boiling points. It is because a lot of energy is needed to break the strong bonds (ionic or covalent) operating in three dimensions.
* The oxides of phosphorus, sulphur and chlorine consist of individual molecules. Some are small and simple, others are polymeric.
* The attractive forces between these molecules will be Vander Waals dispersion and dipole-dipole interactions. These vary depending on the size, shape and polarity of various molecules. However, these forces are much weaker than ionic or covalent bonds of a giant structure. Thus, these oxides tends to be gases, liquids or low melting point solids.

**Electrical conductivity:**

* None of these oxides has any free or mobile electrons. Thus, none of them will conduct electricity when they are solid.
* The ionic oxides can conduct electricity in molten state. It is because the ions become free in molten state. So the movement of ions towards the electrodes causes electrical conductivity. Thus, they undergo electrolysis in molten state and ions are discharged at electrodes.

**Short questions and answers:**

1. **Although Na and P are present in the same period yet their oxides are different in nature. Na2O is basic while P2O5 is acidic. Why?**

Metal forms basic oxides while non-metal forms acidic oxides. Since sodium is a metal, it will form basic oxide i.e.Na2O. When this oxide is dissolved in water, it forms NaOH, that is a base and solution becomes basic.

**Na2O +H2O🡪2NaOH**

Phosphorus is a non-metal, so it will form an acidic oxide i.e.P2O5. When this oxide is dissolved in water it produces H3PO4, which is an acid and solution becomes acidic due of formation of H3PO4.

**P2O5 + 3H2O 🡪 2H3PO4**

1. **How acidic basic and amphoteric behavior oxides is explained?**

An amphoteric oxide shows both acidic and basic properties e.g. BeO. It reacts with acids to form salts. **BeO + H2SO4 🡪BeSO4 +H2O.**

It also reacts with bases to forms salts. **BeO +2NaOH 🡪 Na2BeO2 +H2O**

1. **Why the elements of group 1 are called alkali metals?**

The word “alkali” is an Arabic word means “ashes”. This word was used by Arabs because they found that ashes of plant mainly consist of Na and K. So, they called these metals as alkali metals. The alkali metals include Li, Na, K, Rb, Cs and Fr. They produces strong alkaline solution with water**.**

1. **Why all group 1 metals have low ionization energies?**

In any period, Group 1 metals have larger size and least electronegativity. So it is easy to remove outermost electron. Hence, they have low ionization. Thus they readily form positive ions, e.g. Na+1 ,K+1 etc.

1. **Why do the group 1 metals show strong electropositive character?**

Electropositivity is the ability to give electrons. Since in any period, group 1 metals have larger size and least electronegativity. So, it is easy to remove their valence electron. Hence, they have very low ionization energies and show electropositive character. In group 1 metals, the electropositivity increases down the group**.**

1. **Why do group 1 metals shows strong reducing properties ?**

In any period, group 1 metals have larger size and least electronegativity. Thus alkali metals can easily loose electrons i.e. they are easily oxidized. Thus they can reduce substances and acts as excellent reducing agents.

1. **Why different colors are imparted by the atoms of group 1 metals to the flame?**

In the alkali metal atoms the outer ns1electron is loosely held with the nucleus. So, it can be easily excited to the higher energy levels by absorbing a small amount of heat energy. When this excited electron comes back to its original position, it gives out absorbed energy in the form of light in the visible region of the electromagnetic radiations. Different amount of energy is absorbed in different atoms for excitation of electron. Thus different colors are imparted by the atoms to the flame.

1. **Why the elements of group 2 are called alkaline earth metals?**

The alkaline earth metals are given this name because they form alkalis in water and are widely distributed in earth crust. These include Be,Mg,Ca,Sr,Ba and Ra.

1. **Why do the group 2 earth metals have high melting and boiling point than alkali metals?**

It is because the group 2 elements have 2 valence electrons. So, they form greater number of bonds than group 1 elements.

1. **How do group 1 metals resemble with group 2 metals?**
2. Both alkali and alkaline earth metals are s-block elements. Both have their outermost electron in s-orbital.
3. Elements of both groups are highly electropositive.
4. Elements of both groups do not occur free in nature.
5. Hydroxides of both alkali and alkaline earth metals are strongly basic.
6. On heating in Bunsen flame, elements of both groups imparts characteristics color to flame.
7. **How do group 1 metals differ from group 2 metals?**

|  |  |
| --- | --- |
| **ALKALI METALS** | **ALKALINE EARTH METALS** |
| (i) They have one electron in their outermost s-orbital. | They have two electrons in their outermost s-orbital. |
| ii) They have low melting points than alkaline earth metals, | They have relatively high melting points. |
| iii) They have relatively larger atomic size. | They have relatively smaller atomic size. |
| iv) The alkali metals are relatively softer than alkaline earth metals. | They are relatively harder. |
| v) They have relatively low values of ionization energies. | They have relatively high values of ionization energies. |
| vi) Their oxides and hydroxides are more ionic in nature. | Their oxides and hydroxides are less ionic in nature. |
| vii)They decompose water vigorously at room temperature.  2Na +2H2O 🡪2NaOH +H2 | They decompose water less vigorously. Mg + H2O🡪MgO +H2 |
| viii) Their carbonates, sulphates and phosphates are soluble in water except Li. | Their carbonates, sulphates and phosphates are mostly insoluble in water. |
| ix) They are highly electropositive. | They are relatively less electropositive |
| x) Their hydroxides are strongly basic | Their hydroxides are less basic |
| xi) On ,heating, their nitrates gives nitrites and oxygen except Li. 2NaNO3 🡪2NaNO2 +O2 | On, heating their nitrates gives oxides, nitrogen peroxides and oxygen. 2Ca(NO)3🡪2CaO+4NO2+O2. |

1. **Discuss the metallic and non-metallic character of group 4 elements?**

The metallic character increases down the group in Group IV elements

* Carbon , Silicon and Germanium has covalent structures.
* Grey tin or alpha tin also shows covalent bonding.
* The white tin or beta-tin is metallic .In this form atoms are held together by metallic bonds in a close-packed arrangement. In close-packing, each atom is surrounded by 12 neighbors.
* Lead also show metallic bonding.

Thus a clear trend is found down the group from the typical covalency in non-metals to the metallic bonding in metals.

1. **Discuss the general group trends of group 7 elements?**

* Atomic radius, melting points and boiling points increases down the group due to increase number of shells, greater shielding effect and less nuclear charge.
* Electronegativity, electron affinity values decreases down the group due to increase in atomic size.
* The bond enthalpies of the Cl-Cl, Br-Br and I-I decreases down the group due to increase in atomic size except F-F bond.
* The oxidizing power decreases down the group.

1. **Why the term halogen is used for group 7 elements?**

These form salts on reacting with metals, so they are called halogens. (Halo=salt, Gen=former). 2Na +Cl2 🡪2NaCl.

Other examples: KBr, NaF, LiCl, etc.

1. **Why does fluorine differ from other members of its group?**

Fluorine differ from its family member due to :

* Small size of F atom and F-1ion.
* High first ionization energy and electronegativity.
* Low dissociation energy of F2 molecule as compared to Cl2 and Br2.
* Valence shell restriction to an octet.

1. **What is the structure of CO2 and SiO2 and why they differ?**

* CO2 exist in molecular form with weak intermolecular forces. Its dipole moment is zero, therefore it is a linear molecule. Due to weak forces it exist as a gas at room temperature.
* The SiO2 is a macromolecular compound. In this Silicon and oxygen atoms are held together covalently in continuous chains. Thus, silicon dioxide is non-volatile and hard solid unlike carbon dioxide.

The difference in structure of SiO2 and CO2  is due to the following two reasons:

1. Silicon atoms are much larger in size than carbon atoms and thus tends to be surrounded by more oxygen atoms.
2. Silicon forms only single bond with oxygen atoms while carbon forms double bonds.
3. **CO2 is a gas while SiO2 is a solid although C and Si belongs to the same group?**

The CO2 exist in molecular form with weak intermolecular forces. Its dipole moment is zero, therefore it is a linear molecule. Due to weak forces it exits as gas at room temperature.

The SiO2 is a macromolecular compound. In this silicon and oxygen atoms are linked together covalently in a continuous chains. Thus, silicon dioxide is a non-volatile and hard solid.

1. **SnCl2 is a solid while SnCl4 is a liquid. Why?**

According to Fajan’s Rule smaller the cation, greater is the amount of covalent character in the compound. Since Sn4+ ion is smaller than Sn2+ ion, therefore SnCl2 is an ionic compound and SnCl4 is covalent. Hence due to strong forces in SnCl2 these are tightly packed and forms a network solid. Whereas SnCl4 molecule, there are weak intermolecular forces called London dispersion forces. Hence it is a liquid.

1. **C and Si are always tetravalent but Ge, Sn and Pb shows divalency. Why?**

These are group 4 elements with general electronic configuration ns2 np2.

If all the four ns2np2 electrons are used in bonding then +4 oxidation state is observed Both C and Si forms stable compounds in +4 oxidation state. In the case of Ge, Sn and Pb only two np2 electrons may be lost. Thus these elements may show +2 oxidation states. The ns2 pair of electrons is not lost so it is called inert pair of electrons. The stability of +2 oxidation state increases from Ge+2 to Pb+2.The order of stability of M+2 and M+4 cations of Ge, Sn and Pb is: Ge+2<Ge+4 ,Sn+2 <Sn +4, Pb+2>Pb+4.

1. **CCl4 is resistant to hydrolysis but SiCl4 is easily hydrolysed. Why?**

**OR**

1. **Si-Cl bond is stronger than C-Cl bond, still SiCl4hydrolyzed easily but CCl4 is not. Why?**

The C-atom being a member of 2nd period of the periodic table, has no d-orbital in its valence shell. Thus, it is unable to accommodate the lone pair of donated by the donor oxygen atom of water molecule to form an unstable intermediate compounds. Hence the tetrahalides of C are not hydrolyzed. On the other hand Si, Ge and Sn have vacant d-orbitals. These can accept lone pair from water and thus this tetrahalides get readily hydrolyzed.

1. **Explain why nitrates and carbonates of Li are not stable?**

The nitrates and carbonates of Li are decomposed on heating.

**Li2CO3🡪 Li2O +CO2 4LiNO3🡪2Li2O +4NO2 +O2**

Li has small size. So, the lattice energy of the product oxide is considerable due to close packing. Hence the oxide is stable and carbonates/nitrates of Li are converted into oxide readily. So, they are not stable.

1. **Differentiate the behavior of Li and Na with atmospheric oxygen?**

Lithium burns in atmospheric oxygen to form the normal oxide, Li2O

**4Li +O2 🡪 2Li2O (Lithium oxide)**

Sodium burns in atmospheric oxygen to form the peroxide

**2Na + O2🡪Na2O2 (sodium peroxide)**

1. **Alkali metal carbonates are more soluble than alkaline earth metal carbonates. Why?**

The alkali metal cations have relatively larger ionic size than alkaline earth metals cations .Moreover, they have single charged M +ions. Therefore in carbonates they have loose packing of ions and have relatively weaker forces. So their lattice energy are very low. Hence they dissociate in water easily and become soluble. Hence alkali metal carbonates are soluble than alkaline earth metal carbonates.

1. **Explain why solubilities of alkaline earth metal carbonates decreases down the group?**

In alkaline earth metals, the ionic size increases down the group .Moreover, carbonate anion is itself bulky (large) anion. Generally larger the size of ion, lesser is the hydration and solubility. So, due to difficulty in hydration of ions, the solubility of alkaline earth metal carbonates decreases down the group.

1. **Oxidizing power of F2 is greater than I2. Why?**

It is due to three reasons.

* The bond dissociation energy of fluorine is quite low, so it will dissociates rapidly to take up electron.
* The electron affinity of F is greater than I, so it will take up electrons more rapidly.
* The F- ion has high hydration energy due to small size, so it will readily in solution. Actually, this is the main factor.

It means that Fluorine will take up electrons to form F- ions than iodine. Hence, fluorine is a much stronger oxidizing agent than iodine.

1. **HF is a weak acid than HI. Why?**

In HF, molecules are hydrogen-bonded in a zig-zig manner. Thus H-atom is entrapped between two F atoms .Moreover, the bond energy of H-F bond is considerably greater than H-I bond. Thus, HF cannot donate its H+ ions easily hence it is a weaker acid.

**(xxviii) On what factors does the oxidizing power of halogens depends?**

Following factors affects oxidizing power: (a) Energy of dissociation (b) Electron affinities of atoms (c) Hydration energy of ions. If a halogen has low energy of dissociation, high electron affinity and high hydration energy of its ions, then it will be a better oxidizing agent. On the basis of these factors, the order of oxidizing power is: F2>Cl2>Br2>I2

**CHOOSE THE CORRECT OPTION A/B/C/D.**

(1) Aluminium oxide is :

(A)Acidic (B) Basic (C) Amphoteric (D)Neutral

(2)The elements of group 3 and 7 are known as \_\_\_\_\_ block elements.

(A)s (B)p (C)d (D) d

(3)The radioactive element in halogen group is:

(A)Radon (B)Radium (C)Astatine (D)Bromine

(4)Green is the characteristics flame color of :

(A)Calcium (B)Barium (C)Strontium (D)Sodium

(5)The first ionization energy is higher for the :

(A)Alkaline earth metals (B) Alkali metals (C) Halogens (D)Noble gases

(6)Ge2 +compounds are:

(A) Reducing agent (B)Oxidizing agent (C)Both A&B (D)None of these

(7)The flame color of sodium metal or its compound is:

(A)Bright crimson (B)Violet (C)Golden yellow (D)Blue

(8) Lead(II)chloride PbCl2 :

(A)White solid (B)Liquid (C)volatile compound (D)None of these

(9)Which is the least reactive of all the alkali metals :

(A)Li (B)Na (C)K (D)Cs

(10)Which is not an alkali metal:

(A)Francium (B) Cesium (C)Rubidium (D)Radium