EXERCISE

Q1: Select the right answer from the choices given with each question.

1) **Oxides and hydroxides of Group1 elements are:**
   a) Acidic b) Alhaline
   c) Neutral d) Amphoteric

2) **The flame colour of sodium metal or its compounds is:**
   a) A, Bright crimson b) Voilet
   c) Golden yellow d) Bright blue

3) **When sodium burn in air , it forms sodium:**
   a) Monixide b) Peroxide
   c) Oxide d) Superoxid

4) **The carbonates of alkali metal are not affected by heat except:**
   a) $Li_2 CO_3$ b) Na_2CO_3

c) K_2CO_3 d) Rb_2CO_3

5) **Green is characteristic flame colour of:**
a) Calcium b) Barium
c) Strontium d) Sodium

6) **All the carbonates, sulphates and phosphates of alkaline earth metals are ________ in water.**
a) Sprangly b) Soluble
c) Insoluble d) Less soluble

7) **The first ioniziatio energy is higher for the:**
a) Alkaline earth metals b) Alkali metals
c) Halogens d) Nobel gases

8) **Which one of the element has the maximum electron affinity?**
a) F b) Cl c) Br d) I

9) **Which pair has both members from same period of periodic table?**
a) Na – Ca b) Na – Cl c) Ca – Cl d) Cl – Br

10) **Melting points and boiling points of alkali metals:**
a) Decreases from top to bottom b) Increases from top to bottom
c) First increases then decreases d) Remains uncharged

11) **Which of the following oxides is amphoteric in nature?**
a) Rubidium oxide b) Barium oxide
c) Antimony d) Sulphur oxide

12) **Oxidizing power of halogen depends upon:**
a) Energy of dissociation b) Electron affinity
c) Heat of vaporization d) All of above

13) **Which of following oxides amphoteric is nature?**
a) Mgo b) BeO c) CO_2 d) SnO_2

14) **Select the correct increasing order of atomic radius?**
a) Ne > O > S > Al b) Ne < O > S > Al
c) Ne < O > S > Al d) Ne > O < S > Al

15) **Due to inert pair effect ________ oxidation state is more stable than ________ For Sn and Pb.**
a) 2+, 4+ b) 1+, 4+ c) 4+, 2+ d) 2+, 3+

16) **Highest electron affinity is shown by?**
a) F_2 b) I_2 c) Br_2 d) Cl_2

17) **Which is the strongest reducing agent?**
a) HF b) HCl c) Br_2 d) HBr

18) **Substance boiling at higher temperature among following is?**
a) HI b) HF c) HCl d) HBr

19) **Group VII-A elements are generally called:**

a) Halogens b) Nobel gases
c) Inert gases d) Metallods

20) **The radioactive element in halogen group is:**

a) Radon b) Radium c) Astatine d) Bromine

Answers

1)	b)	2)	c)	3)	b)	4)	a)	5)	b)
6)	c)	7)	d)	8)	b)	9)	b)	10)	a)
11)	c)	12)	b)	13)	b)	14)	c)	15)	a)
16)	d)	17)	d)	18)	b)	19)	a)	20)	c)

Q2: Give brief answers for the following questions.

Q1. Although Na and P are present in the same period their oxides are different in nature Na_2O is basic while P_2O_5 is acidic why?

Answer

Na and P are present in same period But Na_2O is basic and P_2O_5 is acidic:

Na is an alkali metal present in IA group and has one electron in its outermost shell, It gives a strong base (NaOH) on dissolving in water.

$$Na_2O + H_2O \longrightarrow 2NaOH$$

Phosphorus is member of group – VA. It is non-metal and has five electrons on its outermost shell. P_2O_5 on dissolving in cold water produce meta phosphoric acid and on dissolving in hot water give orthophosphoric acid.

$$\underset{\text{(Cold)}}{P_2O_5 + H_2O} \longrightarrow 2HPO_3 \quad \text{(metaphosphoric acid)}$$

$$\underset{\text{(Hot)}}{P_2O_5 + 3H_2O} \longrightarrow 2H_3PO_4 \quad \text{(orthophosphoric acid)}$$

Q2. How acidic basic and amphoteric behaviour of oxides is explained?

Answer

Basic, Acidic and amphoteric behaviour of oxides:

Odides of alkali metals are basic in nature.

$$\underset{\text{(Base)}}{Na_2O} + \underset{\text{(Acid)}}{H_2SO_4} + H_2O \longrightarrow Na_2SO_4$$

Oxides of Be Saluminium are amphoteric in nature e.g $SO_3 + P_2O_5$

$$\underset{\text{(acid)}}{SO_3} + \underset{\text{(Base)}}{Na_2O} \longrightarrow \underset{\text{(Salt)}}{Na_2SO_4}$$

Oxides of Be saluminium are amphoteric

$$\underset{\text{(Acid)}}{BeO + 2NaOH} \longrightarrow \underset{\text{(Sodium beryllate)}}{Na_2BeO_2} + H_2O$$

$$\underset{\text{(Base)}}{BeO + H_2SO_4} \longrightarrow BeSO_4 + H_2O$$

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Q3. Why the elements of group 1 are called alkali metals?

Answer

Elements of group IA are metals and their oxides and hydroxides are alkaline in nature.

e.g $Na_2O + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

$2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

Q4. Why all group of metals have low ionization energies?

Answer

All IA – group elements have one electron on their outermost shell. Except lithium all alkali metal have strong shielding or screening effect. As a result pull of nucleus on outermost electron decreases thats why outermost electron of alkali methods may be removed easily using low energy.

Q5. Why do the group 1 metals show strong electropositive character?

Answer

The metals which lose outermost electrons easily and give a stable cation are called (electropositive. And group IA metals lose their outermost single electron easily to complete their outermost stable octet Therefore they are electropositive metals.

Q6. Why do group 1 metals show strong reducing properties?

Answer

All alkali metals have one electron in their outermost shell and on losing one outermost electron their octet is completed which is a stable state. Therefore they try to lose their valence electron and all as good reducing agent.

Q7. Why different colours are imported by the atoms of the group 1 metals to the flame?

Answer

The outermost electron of group IA is ns^1. This electron is loosely hold with the nucleus.

When an element is brought to a flame of a burner the electron absorbs some energy and goes to higher energy orbital and is called excited. This excited electron absorbs energy during process of excitation Atoms of different elements absorb different energies therefore on dexitation they impart different colours to flame.

Q8. Why the elements of group 2 are called alkaline earth metals?

Answer

II – A group members are metals, their salts are found in earth and their oxides & hydroxides alkaline in nature. Therefore they are called alkaline earth metals.

e.g $CaO + H_2SO_4 \longrightarrow CaSO_4 + H_2O$

$Ca(OH)_2 + H_2SO_4 \longrightarrow CaSO_4 + 2H_2O$

Q9. Why do the group 2 earth metals have high melting and boiling points than alkali metals?

Answer

The alkaline earth metals have two electrons in their outermost shell whereas group II – A have one electron in their outermost shell. Therefore there are strong metal–metal

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bonds among group IIA. Therefore II – A members are harder and stranger than group I–A members. Therefore the melting and boiling point of group II–A members are higher than group I – A members.

Q10. How do group 1 metals resemble with group 2 metals?

Answer

Both group I–A and group II–A members

1) are metals.
2) are iso-electronic up to subshell.

e.g Li = 3 $1s^2, 2S^1$

Be = 4 $1s^2, 2S^2$

3) have alkaline oxides and hydroxides

e.g Na2O, MgO

NaOH, $CaCOH_2$

4) are good conductors for heat and electricity can react with halogens to form salts. e.g NaCl, $CaCl_2$

Q11. How do group 1 metals differs from group 2 metals?

Answer

Group IA Metals	Group IIA Metals
1) Their have one electron in their outermost shell. 2) They are monovalent i.e M^{+1}.	Their atoms have two electons in their outermost shell. They are divalent i.e M^{+2}
3) They are soft metals.	They are hard metals
4) Their melting and boiling points are lone.	Their melting and boiling points are high.
5) Their carbonates are solube in water.	Their carbonates are insoluble in water.
6) Their carbonates are stable towards heat $Na_2CO_3 \xrightarrow{\Delta} X$	Their carbonates decompose on heating. $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$
7) Their hydroxides are stable towards heat $NaOH \xrightarrow{\Delta} X$	Their hydroxides decompose on heating. $Ca(OH)_2 \xrightarrow{\Delta} CaO + H_2O$

Q12. Discuss the metallic and non-metallic character of group 4 elements.

Answer

Metallic and Non Metallic charcter of Group IV A Elements.

Metallic character of Group iv – A elements increase from top to bottom.

iv – A.

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$\left.\begin{matrix} C \\ Si \end{matrix}\right]$ Non – metals

Ge – semi metals

$\left.\begin{matrix} Sn \\ Pb \end{matrix}\right]$ – Metals

Q13. Discuss the general trends of group 7 elements.

Answer

General Group Trends of Group VII – A Elements

i) **Trends in physical states:**

As attractive forces among their molecules increase from F_2 to I_2, therefore.

$\left.\begin{matrix} F_2 \\ Cl_2 \end{matrix}\right]$ gases

Br_2 – Liq

I_2 – solid

Trends in Electronegativity

Electronegativity decrease from top to bottom in halogen group. i.e. Decreases from F to I.

iii) **Trends in Electron Affinity**

Due to high electron density, the E.A value of E is low. And E.A value of remaining halogens decrease from Cl to I.

iv) **Bond Enthalpy in halogens**

Due to very small size and high electron density F atoms repell each other in F_2. As a result the bond dissociation. Energy of F_2 is lower than the remaining halogens B D E of remaining halogens decrease from Cl_2 to I_2

v) **Trends as oxidising agents**

Oxidising ability of halogens decrease from F2 to I_2.

$$F_2 > Cl_2 > Br_2 > I_2$$

vi) **Trends of Activity of Halogen halides**

Due to strong H – bonds, the HF ionizes slightly, therefore is a weak acid. Other halo acids ionize completely and are strong acids.

Their acidic strength increases from HF to HI

$$\underset{\text{(weakest acid)}}{HF} < HCl < HBr < \underset{\text{(Strong acid)}}{HI}$$

Q14. Why the term halogen is used for group 7 elements.

Answer

Halogen means salt forming. The elements of Group VII–A on reaction with alkali and

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alkaline earth metals form salts. e.g.

$$2Na + Cl_2 \longrightarrow 2NaCl \text{ (Salt)}$$
$$2K + Br_2 \longrightarrow 2K\,Br \text{ (Salt)}$$
$$Ca + Cl_2 \longrightarrow CaCl_2 \text{ (Salt)}$$

Q15. Why does fluorine differ from other members of its group?

Answer

Flourine differs from its Group members in following respects.

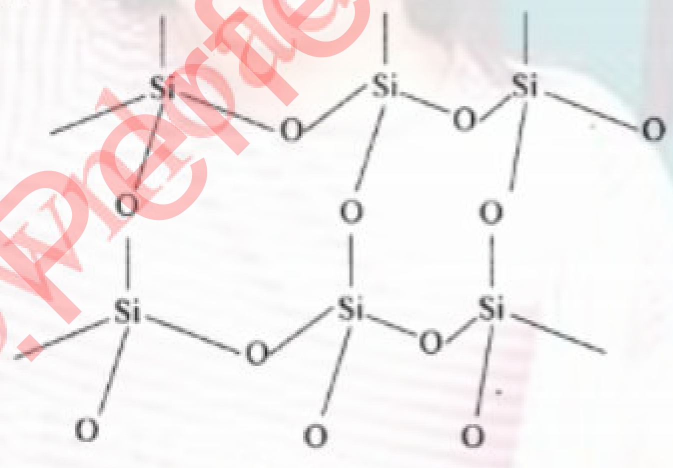
1) Very small size of F atom and $F^{\ominus}$ ion.
2) Very high 1st Ionization Energy and Elecronegativity.
3) Low BDE of F_2 as compared to Cl_2 to Br_2.
4) Restriction of valence shell to an octet.
5) CaF2 & Mg F_2 are insoluble in water whereas $CaCl_2$ f $CABr_2$ are soluble.

Q16. What is the structure of CO_2 and SiO_2 and why they differ?

Answer

Structure of CO_2 & SiO_2 & their Difference

CO_2 is a Linear shaped molecule (O = C = O) having zero dipole moment. Therefore exists as gas. Si isa big sized atom as compared to C. That is why SiO_2 molecule is tetra hederal in shape. Therfore it has a strong network. And that is why it exists i solid crystalline form.



(Network of SiO_2 molecules)

Q17. CO_2 is a gas while SiO_2 is a solid although C and Si belong to the same group.

Answer

Please see answer of Q16.

Q18. Explain why nitrates and carbonates of Li are not stable?

Answer

Li^+ is very small in size and has high charge density. It takes O^2 from nitrate as well in carbonate and forms Li_2O on heating i.e. $4Li\,NO_3 \xrightarrow{\Delta} 2Li_2O + 4NO_2 + O_2$

$$Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$$

Q19. Differentiate the behaviour of Li and Na with atmospheric oxygen.

Answer

Behaviour fo Li & Na with atmospheric O_2

On burning in air, Li forms normal oxides whereas Na forms peroxide.

$$Li + O_2 \xrightarrow{\Delta} Li_2O \quad \text{(Lithium oxide)}$$

$$2Na + O_2 \xrightarrow{\Delta} Na_2O_2 \quad \text{(Sodium peroxide)}$$

Q20. Alkali metals carbonates are more solube that alkaline earth metal carbonates. Why?

Answer

Alkaline earth metals have 2+ charge on their cations. And have high charge density. On heating they take O^{2-} from their carbonates and from Metal oxide releasing CO_2 gas.

e.g $\quad CaCO_3 + \xrightarrow{\Delta} CaO + CO_2\uparrow$

$$MgCO_3 \xrightarrow{\Delta} MgO + CO_2\uparrow$$

On the other hand alkali metal ions have 1+ charge o them and have low charge density. These ions with low charge density cannot take O^{2-} ion to form metal oxide. Therefore they are stable towards heat and don't decompose.

Q21. Explain why stability solubility of earth metal carbonates decreases down the group.

Answer

Carbonates decreases down the Group

Upper alkaline earth metals have high charge density which may by hydrated some white. But lower alkaline earth nucleus of group II – A have low charge density and are not hydrated easily.

That is why the solubility of their carbonates decreases down the group.

Q22. Oxidizing power of F_2 is greater than I_2. Why?

Answer

Fluorine atom is small sized atom which has very high Electroneagtivity. It may attract the electrons with more force than large sized, low Electronegative Iodine.

Q23. HF is weak acid than HI Why?

Answer

Please see answer of short Q. No. 56.

Q24. On what factors does the oxidizing power of halogens depends?

Answer

Factors on which oxidising power of Halogens depend

i) High Electron Affinity

Halogens have high electron affinity therefore they accept the electrons and oxidise other elements.

ii) **Tendency to accept electrons**

Outermost shell of halogens need one electron to complete then octet therefore they easily take electron and oxidise other elements.

Q3: Give detailed answers for the following questions.

Q1 (a) The pattern of first ionization energy and Melting & Boiling points is not smooth. Justify it.

Answer

Melting point of Be is very high due to its very small atomic size and very strong inter-atomic forces of attraction.

But in case of Mg, the charge to size ratio make it neutral and nonpolirizable. As a result the M.P of Mg decreases.

In case of Ca the charge density increases and inter-atomic forces also increase. As a result the M.P and boiling points increases.

In case of Sr and Ba, atomic sizes increase too much and charge density decreases. As a result the melting point and boiling point decreases down the group. That is why the melting to boiling point decrease from Ca to Ba.

Q1 (b) Why Atomic radius in Group and decreases along the period?

Answer

In periodic Table, with in a Group .There is increase in no of shells from top to bottom. As a result atomic size increase with in a period. The nuclear charge increase from left to right due to increase in no. of protons in nucleus the pull of nucleus on outer most electrons increase. As a result atomic size decrease from left to right.

Q1 (c) Describe the trends in reaction of 3rd period elements with water.

Answer

3rd period elements are:

Na, Mg, Al, Si, P, S, Cl

Rate of I

Reaction of water with 3rd period elements decrease from left to right

Na reacts with water vigorously and produced hydrogen catch fire due to evolved heat that is why Na burns when placed in water. And produce NaOH

$$Na + H_2O \longrightarrow NaOH + \frac{1}{2}H_2$$

Mg reacts with cold water slowly while it reacts with steam and produce white MgO and $H_{2(g)}$

$$Mg + H_2O \longrightarrow MgO + H_2O$$

Al and Si both are semi-metals which have no reaction water in cold or hot state. P reads without water and produce a mixture of phosphorus acid (H_3PO_3) and phosphoric acid (H_3PO_4).

$$\underset{(Hot)}{P} + \underset{(Steam)}{H_2O} \longrightarrow H_3PO_4 + H_3PO_4$$

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Cl_2 reacts with wate and produce HCl and HOCl (HYpochlorous acid)

$$Cl_2 + H_2O \longrightarrow HCl + HOCl$$

Q1 (d) The melting and boiling points of the elements increase from left to right upto the middle in 3rd period elements and decrease onward. Why?

Answer

3rd period elements are:

Na, Mg, Al, Si, P, S, Cl, Ar

Metallic character decreases upt mid of the periodic table. And inter-atomic forces of attraction increase from Na to Si due to increase in number of electrons in valance shell

Na	Mg	Al	Si
Very good metal	good metal	metal	Semi-metal

From phosphorus to argon are non-metals. P contains 5 electrons in its outermost shell and exists as white, red and black phosphorus which are soft and hard solids. Sulphur has 6 electrons in its outer-most shell which requires 2 electrons to complete its octate therefore exists as amorphous and crystalline solids.

Cl has 7- electrons in its outer-most shell and reqirs 1-electrons to complete its ocate, therefore exists as $Cl_{2(gas)}$

Ar has ompletely filled outermost shell and exist as inter gas.

Q2. Discuss the metallic oxides and silicon dioxide under the following headings:

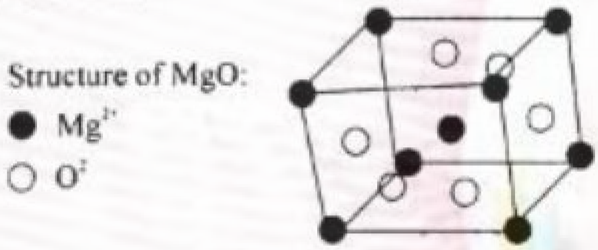
i) Structure **ii) Melting and Boiling pints**

iii) Electrical conductivity

Answer

i) Structures of Metallic Oxides:

Metallic oxides are giant structures of ions having layers of ions. Oxides are Na, Mg, Al are like structure of NaCl.



Structure of SiO_2

Silicon atom has a tetrahedral linkage with oxygen atoms SiO_2 forms a strong network of SiO_2 molecules.

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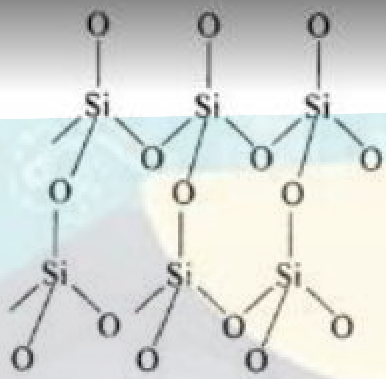


Structure of SiO_2

ii) **Melting and Boiling Points of metallic oxides**

Points of metallic oxides increase from left to right upto mid of periodic table.

If we consider the melting point of 3rd period oxides Na2O has very low melting point Mgo has higher melting point whereas Al_2O_3 has very higher melting point.

Melting and Boiling points of SiO_2

There is a strong network of Si and O_2 atoms bonded through covalent bonds. That is why the melting and boiling point of SiO_2 is very high which is about 3000°C.

iii) **Electrical conductivity of metallic oxides**

In solids state metallic coxides are non-conductor as metallic oxides are offently ionic oxides on melting they produce ions therefore, are good conductors for electricity.

Electrical conductivity of SiO_2

SiO_2 is purely covalent compound and is non-conductor in solids as well as in fused slate.

Q3. Discuss acid-base behaviour of

i) Aluminium oxide **ii) Sodium oxide**

Answer

i) **Acid base behaviour of Aluminium oxide (Al_2O_3).**

Al_2O_3 is an amphoteric oxide and behaves acid are well as base.

$$Al_2O_2 + NaOH \longrightarrow NaAlO_2\ .2H_2O$$

$$Al_2O_3 + 3H_2SO_4 \longrightarrow Al_2\ (SO_4)_3 + 3H_2O$$

ii) **Acid Base Behaviour of Na_2**

Na2O is metallic oxide and is strongly basic in nature.

$$Ns_2O + HCl \longrightarrow 2NaCl + H_2O$$

Q4. a) Why are different types of oxides formed as you go down the group?

b) How Beryllium differs from other members of its group?

c) Why is Beryllium chloride covalent and not ionic?

Answer

a) Why are different types of oxides formed as you go down the group?

Due to addition of new shells atomic size increases from top to bottom. That is why the ability of oxide to loose O^{-2} ion increase from top to bottom within a group.

If we consider the example of group I–A, LiO_2 is amphoteric i nature Na_2O is basic in nature, K2O is more strong base and the basic character increase. From top to bottom.

b) How Beryllium differs from other members of its group?

Answer

Beryllium differs from other members of its group due to its very small size. The main points of differences are as follows:

i) II–A group members are soft metals but beryllium is as hard as steel and hard enough to scratch

ii) The melting and boiling points of beryllium are much higher then remaining members of its group.

	Be	Mg	Ca	Sr	Ba
Melting points °C	1277	650	838	703	714
Boiling points °C	2770	1107	1440	1380	1640

iii) All II–A group members reduce water but beryllium due to forming of an insoluble oxide layer does not reach with water vigorously.

iv) All II–A group members reads with acids but beryllium is resistant towards complete oxidation due to formation of BeO coating on its surface.

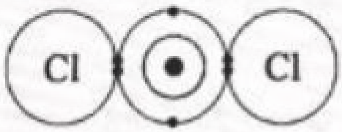
v) Beryllium is the only members of group II–A which react with NaOH.

$$2Be + 2NaOH \longrightarrow \underset{\text{(Sodium Beryllode)}}{Na_2Be_2O} + M_2$$

c) Why is Beryllium chloride covalent and not ionic?

Answer

Beryllium is a very small atomic size and does not loose an electrons to forms ions therefore $BeCl_2$ is covalent in nature.



Cl – Be – Cl

Q5. a) Why do some metals form peroxides on heating in oxygen?
b) Why do group 2 elements form nitrides on heating in air?
c) Discuss the trend in solubility of hydroxide of group 2 elements.

Answer

a) Why do some metals form peroxides on heating in oxygen?

The attachment of oxygen with a metal depends upon its charge to size ratio (charge density) small sized metal atoms high charge density forms normal oxides large sized low charge density forms peroxide and very large sized metals formed super-oxide. For example Li forms normal oxide (Li_2O). Na forms peroxide (Na_2O_2) and K forms super-oxide (K_2O) due to very low charge density and large size of metal atom.

b) Why do group 2 elements form nitrides on heating in air?

Answer

Attachment of N^{-3} with a metal depends upon charge density of metal ion high charge density cations as II–A group metal ions have high charge density so they forms nitrates. In alkali metals only Li^{+1} forms nitrate.

c) Solubility of hydroxides of group II–A elements

Answer

The solubility of hronxides depends upon the charge density of cations, of high chare density are hydrated easily. For example hydroxides of calcium (Ca) and Magnesium (Mg) are soluble in water. On the other hand the solubility of hydroxides decreases down the group.

Q6. Discuss the trends in thermal stability of carbonate and nitrates.

Answer

Thermal stability of carbonate and nitrates

The thermal stability of carbonates and nitrates depends upon the charge density of cations. Carbonates and nitrates having a metal of high charge density when are heated the metal cation takes oxide ion (O^{-2}) and forms metal oxide and release $CO_{2(gas)}$. For example:

$$LiCO_3 \xrightarrow{\Delta} Li_2O + CO_2$$

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

Trends in group I–A

In group I–A only $LiCO_3$ and $LiNO_3$ decompose on heating due to high charge density of Li^{+1} ion whereas carbonates and nitrates of other elements of group I–A don't decompose on heating due to low charge density of cations.

Trends in group II–A

As all alkaline earth metal cations have high charge density and on heating take O^{-2} iion so decompose on heating and form metal oxide. For example

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

$$2Ca(NO_3)_2 \xrightarrow{\Delta} 2CaO + 4NO_2 + O_2$$

Q7. Explain with examples that Beryllium hydroxide is amphoteric?

Answer

Beryllium hydroxide as amphoteric

Beryllium hydroxide is acid as well las a base and behaves amphoteric due to high charge density of Be^{+2} ion

$$\underset{\text{(As Acid)}}{Be(OH)_2} + 2NaOH \longrightarrow Na_2BeO_2 + 2H_2$$

$$Be(OH)_2 + H_2SO_4 \longrightarrow BeSO_4 + 2H_2O$$

(As Base)

Q8. Explain the trends in oxidation states with suitable examples.

Answer

Trends In Oxidation State

The ability of an tom to loose or gain the number of electrons is called its oxidation state.

Trends within a Period

When we move from left to right oxidation state increases from upto mid of the periodic table

Q9. Discuss the inert pair effect in the

i) Formation of ionic bonds **ii) Formation of covalent bonds**

Answer

Inert pair Effect and Ionic Bond

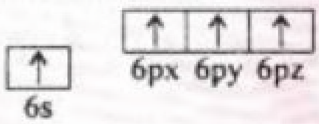
In ground state the electronic configuration of group IV-A elements is n^2p^2. When we move down the group atomic size increases due to addition of new shells. They loos p-electrons leaving the s^2 pair unused this is called inert pair. For example to form Pb^{+2} ion lead will loos two 6p– electrons but the 6s – electrons will be left unused as an inert pair. Also the ionization energy decreases to form 2 + ions therefore the chances of formation of ionic bonds increases specially in case of Sn to form Sn^{+2} and Pb to form Pb^{+2} that is whey Sn and Pb like to form ions bonds.

ii) Inner pair Effect and Covalent Bond

According to theory of relativity the behaviour elements like Sn, Pb, Pt, Si difficult to remove the electros due to relativistic contact on. This effects "S" electrons more than "P" electrons. And one electron moves from S-orbital to empty p-orbital

Electrons configuration of Pb in ground state.

Electronic configuration of Pb in exited state =



The energy is released when atoms of halogen like chlorine come to attach with tin or lead. Therefore no extra energy is needed to promote an electron from 6s orbital to 6p-orbital. Now these four unpaired electrons like to form covalent bonds.

Q10. Discuss in detail acid base trend of group IV–A oxides?

Answer

Acid-base trends in grup IV–A oxides.

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Group IV–A oxides are as follows:

Co – Acidic

SiO_2, GeO_2 – Neutral

SnO, PbO_2 – Basic

Top most oxide of group IV–A are acidic in nature. CO_2 on dissolving in water produce carbonic acid

$$CO_2 + H_2O \longrightarrow H_2CO_3 \text{ (carbonic acid)}$$

Lowermost oxide of group IV–A may gave O^- ion which on dissolving in water produce OH ion and behave as Lu- flood bases.

$$PbO_2 \longrightarrow Pb^{4+} + 2O^{2-}$$

$$O^{-2} + H_2O \longrightarrow 2OH^{-1}$$

Q11. Explain in detail the trends in group VII– of following physical properties.

i) Electronegativity **ii) Electroneffinity**

Answer

Trends of ElectronegativityIn group VIII–A

Due to increase in atomic size and increase in shielding effect the electronegativityin group VII–A decreases from top to bottom

Element	E.N
F	4.0
Cl	3.2
Br	3.0
I	2.5
At	2.2

Trends of Electron effinity in group VII–A

Electron effinity of F has thick value due to high charge density (thick electronic cloud) whereas electron effinity in group VII–A decreases from Cl to At.

Element	Electron effinity
F	–322
Cl	–349
Br	–325
I	–295
At	–270

Q12.a) Why is the bond enthalpy of F–F less as compared to Cl – Cl and Br– Br?

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b) Explain the order F>Cl> Br>I with respect to oxidizing agent or power.

Answer

a) Why is the bond enthalpy of F–F less as compared to Cl – Cl and Br– Br?

Due to very high electronegeitivity of F attract the electrons towards their nuclei with greater force which hardly allows the electrons to be shared to form a covalent bond. Small amount of supply of energy increases the Kinetic energy of atom and break the bond between atoms.

In case of Cl_2 and Br_2 the electronegativityof Cl and Br atoms are moderate and easily allow atoms to share the electrons to form stable covalent bond therefore Cl_2 and Br_2 exist stable gaseous molecules.

The order F>Cl> Br>I with respect to oxidizing agent

The electronegativityorder of halogen is:

i) Element F> Cl> Br> I

Electronegetivity 4.0>, 3.2>, 3.0>, 2.5

Due to highest value of electronegativityF may attract and lake the electron from a substance and oxidize with more force as the order of electronegativitydecreases towards I therefore oxidizing power towards I also decreases.

Q13. a) Why is flourine much stronger oxidizing agent than chlorine?
b) HCl is strong acid as compared to HF. Why?

Answer

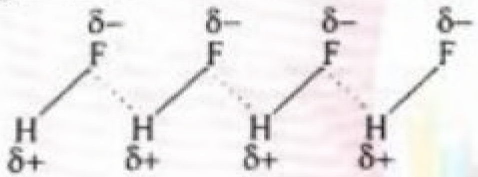
a) Why is flourine much stronger oxidizing agent than chlorine?

The electronegativityof F is 4.0 which is maximum in periodic table due to its high electonegitivity it may take an electrons from a sub stance to complete its octate whereas Cl has somewhat lower electronegativityand has low oxidizing power.

b) HCl is strong acid as compared to HF. Why?

Due to very high electronegativity F, the molecule HF is highly polar.

There are strong hydrogen bonds between HF molecules.



As a result the power of HF to release proton is decreases. As a result HF behaves as weaker acid than HCl which is completely ionized on dissolving in water.