Introduction to p-Block Elements

With so many elements in the periodic table, don't you get a little nervous? Well, it's not that tricky! The scientists possibly knew how difficult it would be remember all of the elements and their properties. That is why they grouped the elements into different groups. In this chapter, we will cover the p block elements. Nonetheless, let us just tell you, it is not THAT simple! You do have to focus on the subject honestly! Let us start with what p-block elements are.

What are p – Block Elements?

We have shown the p block elements in the diagram below. These elements are usually good conductors of electricity. They have a shiny lustre. They are good conductors, owing to the fact that they have a tendency to lose their electrons.

In this block, you will find some of the most amazing and diverse properties of various elements, like Gallium. It is a p block metal that can actually melt in your palms. On the other hand p block elements also have silicon that is a metalloid. It is a very important component in the making of glass.

More on P Block Elements

A prominent characteristic of these p block elements is that the last electron of all these elements enters the outermost p-subshell. P block elements comprise of the various families that include:

- Boron family
- Nitrogen family
- Oxygen family
- Fluorine family and
- Neon family, or the family of the inert gases.

Thus, we see that the P block starts from the 13th group and goes to 18th group in the periodic table.

Uses of the P-block Elements

P-block elements show diverse properties. You must have seen that people in villages use coal to cook food. It is nothing but a P-block element i.e. carbon. We use diamonds for the making of beautiful ornaments. This is also a product of carbon. Aluminium foil made up of aluminium is also made up of p block element.

Characteristics of this Block

The general electronic configuration of p-block elements is ns^2np^{1-6}. Except for Helium, all the other elements in this block follow this configuration. In these elements, the inner core electronic configuration may differ. Just because of this difference the inner core, there are changes in both physical and chemical properties of the elements.

The oxidation state of elements in the p-block is maximum when it is equal to a total number of valence electrons i.e. the sum of S and P electrons. One of the most interesting facts about the p-block elements is that it has both non-metals and metalloids.

The first member of the p block elements differs with other elements in two major respects:

- First is the size and every property that depends upon size.
- The second difference applies only to the p-block element which arises from the effects of d-orbitals in the valence shell of the heavier elements.

Learn more about s Block Element here.

Solved Example for You

Q: Why do noble gases not participate in chemical reactions?

Ans: The noble gases are inert in nature. They do not participate in the reactions easily because of their stable electronic configuration, high ionization energies and low electron affinity.

Some Important Compounds of Carbon and Silicon

You have made sand castles on the beach, haven't you? But, did it ever occur to you what sand is and how is it so abundant? Well, let us tell you all about carbon compounds and silicon compounds. Sand belongs to the latter category. In this chapter, we will look at these compounds in brief and learn about some important silicon uses.

Silicon Compounds and Silicon Uses

You know about the element silicon. Let us now look at its compounds. There are two main compounds, Silica and Silicates. These are the most plentifully available compounds on the earth's crust(around 95%). Let's take a look at some of the silicon compounds.

1) Silica

Silica is nothing but Silicon dioxide(SiO_2). We usually find it in crystalline forms like quartz, cristobalite, tridymite. These forms are interconvertible when we apply the correct temperature or pressure.

The structure of SiO_2 is very large. It has a very high bond enthalpy. Therefore, it is unreactive to many acids, halogens as well as alkalies. However, you must remember that it easily dissolves in HF and NaOH. Silica gel is a drying agent. We use it to support chromatographic materials and other catalysts.

2) Silicones

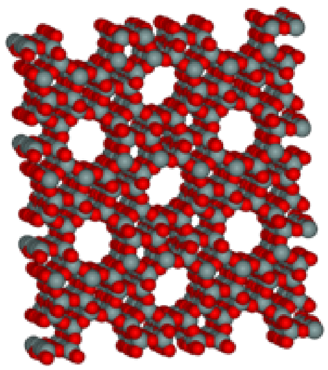
This refers to a group of organosilicon polymers with repeating units. Alkyl or Aryl substituted silicones are the basic products in the manufacture of silicones. Silicones have high thermal stability and are resistant to oxidation and other chemicals. We use this compound as an electrical insulator. It finds wide use in surgeries as well.

3) Silicates

The structure of silicate is SiO_4^{4-} in which we have four oxygen atoms attached to one silicon atom. The two most known man-made silicates include glass and cement.

4) Zeolites

Zeolites find their use in the petrochemical industries as a catalyst. We can convert alcohols directly into gasoline by using a zeolite called ZSM-5. We can also use these zeolites to soften hard water.



That is all about the compounds of silicon. Now, let us move to the other part of the story, the carbon compounds.

Some Important Carbon Compounds

Well, when it comes to carbon, we know quite a bit about its compounds already. Yes, carbon dioxide and carbon monoxide are some important carbon compounds We know their formation and structure and also, their use. Therefore, in this section, let us look at some complex carbon compounds. We can term them as organic compounds as well.

1) Ethanol (C_2H_5OH)

Ethanol goes by the general name ethyl alcohol. We commonly know it as alcohol or spirit. It mostly constitutes the alcoholic drinks. This organic compound finds its use in the manufacturing of medicines, such as tincture iodine, cough syrup, etc. However, one must never consume pure ethanol as it could prove to be fatal.

2) Ethanoic Acid (CH_3COOH)

Acetic acid is the general name of ethanoic acid. It has a melting point of 290 K. This compound freezes in winter and hence, we term it as glacial acetic acid. The vinegar that we consume is a 5% to 8% solution of acetic acid in water. It is also used to preserve pickles.

Now, we will look at a very interesting concept. Do you know what your soaps are made of? Well, organic chemistry has an important role to play in this!

Saponification

Ester of higher fatty acids gives sodium salt of a higher fatty acid. This reaction happens when we heat it with glycerol and sodium hydroxide. Sodium salts of higher fatty acid are known as soaps. This reaction is called saponification or the process of soap making.

What are Soaps and Detergents?

Soap

Soaps are esters of higher fatty acids. We manufacture these by the reaction of ester of a higher fatty acid with sodium hydroxide. The sodium salt, thus produced, has cleansing property.

Detergent

Soap has the inability to form a lather in hard water. To overcome this problem, chemists introduced detergents. These are soapless soaps. Detergent is the sodium salt of benzene sulphonic acid or sodium salt of long chain alkyl hydrogen sulphate.

Cleansing Action of Soap

A soap molecule has two ends. One end is hydrophilic and another end is hydrophobic. In other words, one end is lipophobic (hydrophilic) and another end is lipophilic (hydrophobic).

When we dissolve a soap in water and put the clothes in the soapy solution, soap molecules converge in a typical fashion to make a structure. This structure is a micelle. The hydrophobic ends of different molecules surround a particle of grease and make the micelle. It has a spherical shape.

In this, the hydrophilic end is outside the sphere and hydrophobic end is towards the centre of the sphere. That is how soap molecules wash away dirt and grease by making micelles around them.

Solved Example for You

Q: Why does soap lose its cleansing property in hard water?

Ans: Hard water contains salts of calcium and magnesium. Soap molecules react with the salts of calcium and magnesium and form a precipitate. This precipitate begins floating as an off-white layer over water. This layer is what we call scum. Soaps lose their cleansing property in hard water because of formation of scum.

Trend and Anomalous Properties of Carbon

By now, you know how important carbon is for our existence. Even our bodies are largely made up of carbon compounds. It is true that it has many harmful effects on our environment, but can't deny the fact that carbon is indispensable for life on earth. However, the properties of carbon are very interesting to note.

Carbon shows a lot of anomaly in its behaviour. Why does that happen? In this chapter, we will learn all about the same. Let us begin with a brief introduction to carbon.

An Introduction to Carbon

We are aware that the amount of carbon present in the earth's atmosphere and its crust is very less. There is only 0.02 % of carbon in the earth's crust. This carbon exists as minerals like coal, carbonates and hydrogen carbonates etc. 0.03 % of the carbon exists in the atmosphere of the earth as carbon dioxide.

Well, do not underestimate the importance of carbon based only on its percentage of availability. Carbon is utmost important for our existence and it finds an extensive usage in chemistry. Because of its indisputable importance, chemistry has been divided into two branches:

- Organic Chemistry: This deals with the various compounds containing carbon.
- Inorganic Chemistry: This branch deals with the compounds that do not have any carbon content.

Thus, you must know that carbon is an essential part of every living organism as we are all organic beings!

The Anomalous Behaviour of Carbon

We must remember that most of the first members of a group have peculiar characteristics and properties. On similar grounds, even carbon behaves differently than the other members of the group. The properties of carbon are very unique. We can attribute this behaviour to carbon mainly due to :

- The small size of the atom
- High electronegativity
- High ionization enthalpy
- Unavailability of d-orbital's

Let us now look at some of these anomalies and their causes in more details.

Unique Properties of Carbon

1) The Small Size of Carbon

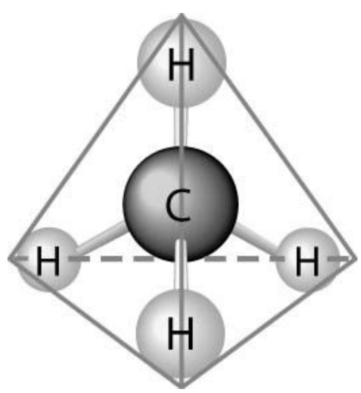
Carbon derives a lot of its properties from its small size. The compounds that carbon forms are highly stable and this is also because of its small size. Due to its small size, the nucleus effectively holds on to the bonded and nonbonded electrons.

Hence, in short tetravalency, small size and property of catenation make carbon different from other elements and so we have the whole branch of chemistry dedicated to the study of this kind of compound.

2) Tetravalency of Carbon

Carbon exhibits tetravalency. It means it can share four electrons to complete its octet. Thus, we know it bonds to four different monovalent atoms. Carbon forms a large variety of compounds with oxygen, nitrogen, hydrogen, halogens. This results in a different set of compounds which have distinctive characteristics and properties.

Carbon has the availability of only s and p orbitals. Therefore, it can hold only four pairs of electrons in its valence shell. Thus, we can restrict the covalence to four. However, the other elements in the group can easily grow their covalence due to the availability of d-orbitals.



3) Catenation

One of the unique properties of Carbon is its ability to form long carbon chains. It implies that carbon attaches with other carbon atoms to form long carbon chains. This property is what we call as catenation. Sometimes, this chain could be as big as to have a total of 70-80 carbons. This gives rise to a variety of complex compounds. Some of the compounds have a straight carbon chain while some others have branched carbon chain or even rings.

The carbon compounds having only single bond are the saturated hydrocarbons. On the other hand, the compounds with double or triple bond are the unsaturated hydrocarbons.

As we move down the group, the size of the elements increases. This results in a decreasing electronegativity. Thus, the property to show catenation also decreases. This can be clearly observed from bond enthalpy values. The catenation order is C >> Si > Ge >> Sn.

4) Electronegativity

Additionally, carbon has an extraordinary capacity to shape pp – pp multiple bonds with itself and with different molecules. This can also be related to its smaller size and high electronegativity. Some of the examples would include C = C, C° C, C = O, C = S and C° N.

As a matter of fact, the heavier elements don't shape pp – pp bonds. This is mainly because of the reason that their nuclear orbitals are too vast and diffused to have viable overlapping. For example, lead does not indicate catenation.

Solved Example for You

Q: Give some practical uses of carbon.

Ans: There are many important uses of carbon. Some of them are:

- We use impure carbon in the form of charcoal (from wood) and coke (from coal) in metal smelting.
- Graphite is a common use in pencils. We also use graphite to make brushes in electric motors and in furnace linings.
- Activated charcoal finds its usage in purification and filtration in respirators and kitchen extractor hoods.
- Industrial diamonds are a common tool for cutting rocks and drilling.

Trends and Properties of Boron and Aluminium

What do you know about the properties of Boron Family? Have you ever wondered why some properties of Boron are different from those of aluminium? Why are they then clubbed under a single group? In this chapter, we will study the properties of boron and its family. We will look at a few anomalies and also the uses of boron and aluminium.

The Boron Family

We can witness some critical patterns in the chemical and physical behaviour of the group 13 elements. The trichlorides, iodides as well as bromides of each one of these elements hydrolyse in water. This is because these are all covalent in nature. Species like tetrahedral $[M(OH)_4]^-$ and octahedral $[M(OH)_6]^{3+}$ exist in a fluid medium. An exception to this is boron.

The monomeric trihalides are strong Lewis acids. This is because they are highly electron lacking. Boron trifluoride effortlessly reacts with Lewis bases, for example, NH_3 to finish octet around boron.

$$F_3B + :NH_3 \rightarrow F_3B - NH_3$$

The greatest covalence of Boron is 4. We can attribute this to the non-attendance of d-orbitals. On the other hand, the d-orbitals are accessible with Aluminium and other elements. This is the reason their most extreme covalency can be normal past 4.

A large portion of other metal halides (e.g. $AlCl_3$) are dimerized through halide bridging (e.g. Al_2Cl_6). The metal species finishes its octet with the help of electrons accepted from halogen in these halogen bridged particles. Let us now look at the anomalous properties of boron and aluminium.

Anomalous Properties of Boron and Aluminium

- The most basic anomaly that we notice in these elements is that Boron is a non-metal. In contrast to this, aluminium is a metal.

- While boron is a non-conductor of electricity, aluminium is a very good conductor of electricity.
- We can find Boron in two structures. These are the amorphous and crystalline structures. On the other hand, aluminium is a delicate metal. It does not exist in various structures.
- The boiling point and melting point of boron are much greater as compared to those of aluminium.
- Boron forms just covalent compounds while aluminium forms even some ionic compounds.
- The oxides and hydroxides of boron are acidic in nature. As compared to this, the oxides and hydroxides of aluminium are amphoteric in nature.
- The trihalides of boron (BX_3) are monomers. On the other hand, aluminium halides exist as dimers (Al_2X_6).
- The hydrides of boron are quite inert while those of aluminium are flimsy.

Diagonal Resemblance of Boron With Silicon

It is really interesting to note that boron and silicon show a lot of similarity in their properties. Both of these elements exhibit the distinctive properties of non-metals. These don't shape cations. Both of them exist in crystalline and amorphous structures.

Boron Oxide (B_2O_3) and silica (SiO_2) are both acidic in nature. These break up in solutions that are alkaline in nature to form borates and silicates separately.

$$B_2O_3 + 6NaOH \rightarrow 2Na\ BO_3 + 3H_2O$$

$$SiO_2 + 2NaOH \rightarrow Na_2SiO_3 + H_2O$$

We can find Silicates and borates having a tetrahedral SiO_4 and BO_4 auxiliary units individually. The chlorides of both Si and B hydrolyse by water. This reaction gives rise to the formation of respective silicic acid and boric acid.

$$BCl_3 + 3H_2O \rightarrow H_3BO_3 + 3HCl$$

$$SiCl_4 + 3H_2O \rightarrow H_2SiO_3 + 4HCl$$

The Hydrides of Boron and Silicon are very stable. We know a variety of unstable hydrides that burst into flames on introduction to air and are effortlessly hydrolyzed. Both boron and silicon are semiconductor by nature. Now, we move on to study some of the uses of boron.

Uses of Boron

- Metal borides are a common part of the nuclear reactors. Here, they are used as defensive shields and control rods. This is because of the high capacity of $_{10}B$ isotope to ingest neutrons.
- Boron is an important component in the steel industry. It finds its usage in the expansion of the hardness of steel.
- Many times, we utilise boron as a semiconductor for making electronic gadgets.

- Boron compounds are turning out to be progressively vital as rocket fills. This is primarily because of their high energy/weight proportion.
- The fibres from boron are a common usage in the process of making light composite materials for airships.
- Boron is a basic element in the plant digestion system.
- Boron carbide filaments are hard, even though they are extremely light. That is why they are primarily used for making bulletproof vests.

Solved Example for You

Q: Write down some common uses of aluminium.

Ans:

- Aluminium is widely used in various industries and also regular day to day existence.
- It is a major component of the steel and iron industry. Aluminium and its amalgams are widely applied in the construction of pipe, poles, tubes, wires, plates or foils etc.

- They are of utmost importance in the pressing, utensil making, plane, construction, and transportation industry.

Ammonia

Did you ever enter a washroom that smelled so bad? What is that smell of? That smell is of nothing but ammonia. It is a compound of nitrogen, which has a pungent smell but it is of great importance to humans. From fertilizers to dyes, it is useful in day-to-day life. Let's learn more about this important compound.

What is Ammonia?

Ammonia (NH_3) is an essential compound of nitrogen and hydrogen. It is created by the regular decay of vegetable and animal bodies. The demise and rot of animals and plants cause the nitrogen compounds exhibit in them to get deteriorated, producing ammonia. Ammonia, likewise, is present in the soil as ammonium salts.

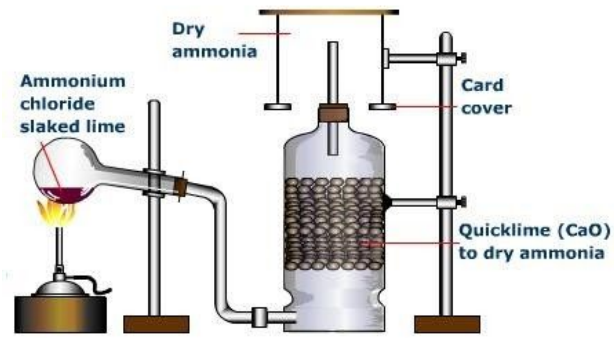
Preparation

We can manufacture the compound by the following methods:

1) From Ammonium Chloride

We can generate ammonia gas in the research centre by slowly heating ammonium chloride (NH_4Cl) and slaked lime [$Ca(OH)_2$].

$$2NH_4Cl + Ca(OH)_2 \xrightarrow{\text{Heat}} CaCl_2 + 2NH_4OH$$
$$2NH_4OH \rightarrow 2NH_{3(g)} + 2H_2O$$



Ammonia gas is lighter than air, requiring its accumulation by the descending displacement of air. Since it is quite solvent in water it can't be gathered over it.

Advancing ammonia gas through quicklime (CaO) dries it. Being an essential gas, we can't dry it by advancing it through concentrated sulphuric acid or phosphorus pentoxide (P_2O_5). This is because it reacts with them to frame ammonium sulfate or ammonium phosphate separately.

$$2NH_3 + H_2SO_4 \longrightarrow (NH_4)_2SO_4$$

$$P_2O_5 + 4NH_3 + 3H_2O \longrightarrow 2(NH_4)_3PO_4$$

We can't utilise calcium chloride for drying ammonia gas as it forms ammoniates with $CaCl_2$.

$$CaCl_2 + 8NH_3 \longrightarrow CaCl_2.8NH_3$$

2) By the Hydrolysis of Metal Nitrides

Hydrolyzing metal nitrides like magnesium and aluminium nitrides, with water or alkalis, can likewise deliver ammonia gas.

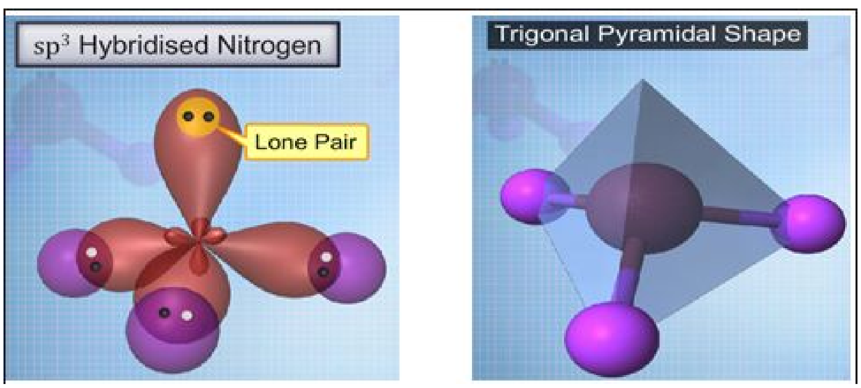
$$Mg_3N_2 + 6H_2O \xrightarrow{boil} 3Mg(OH_2) + 2NH_3$$

$$AlN + NaOH + H_2O \xrightarrow{boil} \underset{\text{sodium meta -aluminate}}{NaAlO_2} + NH_3$$

Structure

Ammonia is a covalent atom. It is seen as a dot structure. The particle is shaped because of the overlap of orbitals of three hydrogen atoms and three sp3 hybrid orbitals of nitrogen in the structure as the central atom. The fourth sp^3 hybrid orbital is involved by a lone pair.

This provides a trigonal pyramidal shape to the compound. The H-N-H bond edge is 107.3°, which is somewhat not exactly the tetrahedral edge of 109°28. This is on the grounds that the bond pair-lone pair repulsions push the N-H bonds somewhat inwards. In solid and liquid states, ammonia is related through hydrogen bonds.



Physical Properties

Ammonia is a gas. It has no colour. It has a sharp pungent odour having a soapy taste. At the point when inhaled all of a sudden, it attacks the eyes bringing tears. It is lighter than air. It is very soluble in water. Ammonia effortlessly melts at room temperature at a pressure of around 8-10 atmospheres.

Liquid ammonia bubbles at 239.6 K (- 33.5°C) under one-atmosphere pressure. It has a high value of the latent heat of vaporization (1370 J for each gram). It solidifies at 195.3 K (- 77.8°C) to give a solid that is white crystalline in appearance.

Chemical Properties

1) Thermal Stability

Ammonia is exceptionally inert. In any case, we can disintegrate it into hydrogen and nitrogen by advancing over metallic impetuses that have been heated.

$$2NH_3 \longrightarrow N_2 + 3H_2$$

2) Combustibility

It is flammable in air.

$$4NH_3 + 3O_2 \longrightarrow 2N_2 + 6H_2O$$

3) Basic Character

The compound has a natural propensity to give its lone pair of electrons of nitrogen to different atoms. Consequently, it acts like a strong Lewis base.

Uses

- We use it in the production of urea and rayon.
- We also use it in the production of composts, for example, ammonium nitrate, urea diammonium phosphate, ammonium sulfate and so on.
- More frequently, we also use it as a refrigerant, in ice plants.
- It finds its use in the furniture industry, as a purging operator for furniture and glass surfaces.
- We use it in the production of nitric acid by Ostwald's procedure.
- We also use it in the production of sodium carbonate by Solvay's procedure.

Tests

The undermentioned tests of any specimen affirm the presence of the compound:

- The ammoniacal odour of the compound is effectively perceivable having a trademark pungent smell.
- It turns wet red litmus blue and moist turmeric paper brown in colour.
- A glass bar dunked in concentrated HCl when conveyed near ammonia causes thick white exhaust.
- When added to a solution of copper sulphate, ammonia turns the solution deep blue.

$$NH_3 + CuSO_4 + nH_2O \rightarrow [Cu(NH_3)_4(H_2O)_n]SO_4$$

- When added to Nessler's reagent (basic arrangement of $K_2[HgI_4]$, ammonia gives a precipitate brown in colour.

$$NH_4^+ + 2[HgI_4]^{2-} + 4OH^- \rightarrow HgO \cdot Hg(NH_2)I \downarrow + 7I^- + 3H_2O$$

Solved Example for You

Q: How can we make the ammonia converter?

Ans: We produce the converter using chrome-vanadium steel. It is normally 1.3 meters high and 1 meter in measurement. The converter is furnished with a warmth exchanger in the upper part and the impetus is packed in the focal segment of the converter.

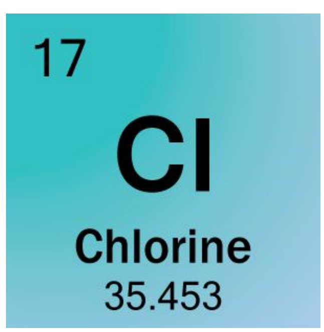
There is a process to warm the combination of gases. After the gas blend enters through the bay at the base, the gases course around the impetus kept up at 450-500°C and afterwards goes through to the heat exchanger. The gases at the end enter the chamber of catalyst to give ammonia.

Chlorine

You have heard your mother scream at you for being in the swimming pool for long as the ‘chlorine’ water could darken your skin. Right? Is it that bad? Life wouldn’t exist if it weren’t for the chlorine compounds. In this chapter, we will look at this element, its compounds and its uses.

Chlorine

Scheele manufactured a gas in 1774 by the activity of hydrochloric acid on manganese dioxide. In 1810, Davy built up its elementary nature and recommended the name Chlorine by virtue of its colour. It is greenish-yellow and it has a pungent smell.



Preparation Methods

1) We can prepare the gas by heating manganese dioxide along with concentrated hydrochloric acid. We can also prepare the gas by the activity of hydrochloric acid on bleaching powder (or) lead dioxide potassium (or) permanganate.

$$\underset{\text{Hydrochloric acid}}{16HCl} + \underset{\text{Potassium permanganate}}{2KMnO_4} \rightarrow \underset{\text{Potassium chloride}}{2KCl} + \underset{\text{Manganese dichloride}}{2MnCl_2} + \underset{\text{Water}}{8H_2O} + \underset{\text{Chlorine}}{5Cl_2}$$

$$\underset{\text{Hydrochloric acid}}{4HCl} + \underset{\text{Lead dioxide}}{PbO_2} \rightarrow \underset{\text{Lead dichloride}}{PbCl_2} + \underset{\text{water}}{2H_2O} + \underset{\text{Chloride}}{Cl_2}$$

2) Electrolytic Process

We can acquire the gas by the electrolysis of salt water in a Nelson cell. This is the least expensive technique and gives the purest form of the gas.

3) Deacon's Process

In this procedure, we can manufacture the gas by the oxidation of hydrochloric acid in the presence of cuprous chloride at 723K and a pressure of 1 atmosphere.

Physical Properties

- Chlorine is a greenish-yellow gas. It has a pungent odour.

- It has a boiling point of 239.11K and a melting point of 171.6K.
- The gas is harmful to nature.
- It is 2-5 times denser than air.
- It can be effectively condensed.
- The gas is marginally dissolvable in water.
- Its valency is 7.

Chemical Properties

- It breaks up in water to give a firmly smelling, yellow arrangement- chlorine water. It loses its yellow colour when it remains in the daylight. This is because of the arrangement of a blend of hydrochloric acid and hypochlorous acid.
- Hypochlorous acid, being unsteady, breaks down and releases nascent oxygen. The oxygen so shaped is in charge of the bleaching and oxidizing properties of chlorine.

$$\underset{\text{Chlorine}}{Cl_2} + \underset{\text{Water}}{H_2O} \xrightarrow{\text{Sun Light}} \underset{\text{Hydrochloric acid}}{HCl} + \underset{\text{Hypochlorous acid}}{HOCl}$$

$$\underset{\text{Hydrochlorous acid}}{2HOCl} + \underset{\text{hydrochloric acid}}{2HCl} \xrightarrow{\text{Decompose}} \underset{\text{Nascent Oxygen}}{2[O]}$$

- It joins particularly with all non-metals except nitrogen, carbon, and oxygen. It reacts very fast with most of the metals. This reaction results in the formation of chlorides.
- It has an extraordinary liking for hydrogen. It reacts with hydrogen in the presence of light with a blast to form hydrochloric acid.

$$\underset{\text{Hydrogen}}{H2} + \underset{\text{Chlorine}}{Cl2} \xrightarrow{\text{Light}} \underset{\text{Hydrochloric acid}}{2HCl}$$

- Chorine breaks down a few hydrogen compounds to shape hydrochloric acid.
- It is a decent oxidizing agent; it oxidizes ferrous to ferric, sulfur dioxide to sulphuric acid, sulfites to sulfates, and iodine to iodic acid.

- Moist chlorine, because of the release of nascent oxygen, goes about as a very powerful bleaching agent. It fades organic matter or vegetables.
- With slaked lime, it frames bleaching powder.

$2Ca(OH)_2$	+	$2Cl_2$	→	$Ca(OCl)_2$	+	$CaCl_2$	+	$2H_2O$
Calcium hydroxide		Cholrine		Calcium oxychloride		Calcium Chloride		Water
(Slaked Lime)		(Bleaching Powder)						

- It reacts with unsaturated hydrocarbons to give addition products and substitution products with saturated hydrocarbons.

Chlorine Poisoning

Chlorine is a chemical that counteracts bacterial development in stationary water. It’s utilized to purify sewage and commercial waste. It’s additionally an active ingredient in a few cleaning items.

Chlorine poisoning usually happens when you take in or breathe in the chemical. It reacts with water — incorporating the water in your digestive tract. it, thus, forms hydrochloric acid and hydrochlorous acid. Both of these substances are very harmful to our body.

You might be most acquainted with chlorine that is utilized as a part of pools. Be that as it may, most occurrences of chlorine poisoning come about because of ingesting household cleaners, not pool water. A couple of regular household items and substances containing chlorine include:

- Tablets utilized as a part of swimming pools
- Swimming pool water
- Mild household cleaners
- Bleaching products

Solved Example for You

Q: Mention some uses of chlorine.

Ans:

- We can use it as a bleaching agent. It is primarily used in the wood pulp, cotton and textile businesses.

- We use it regularly to clean drinking water. It is an antiseptic and disinfectant in swimming pools.
- It can be used in the extraction of gold and platinum. In the arrangement of harmful gasses, for example, phosgene, mustard gas and tear gas, we use chlorine.

Dinitrogen

Do you know what those chips packet have in them? Yes! A lot of air and some chips! Isn't it? However, that is not air! That is nitrogen! Yes! But, why do we use dinitrogen there? In this chapter, we will look at dinitrogen, its compounds and uses.

What is Dinitrogen?

Dinitrogen constitutes almost 78% of the earth's atmosphere. It is the most available element present in the air. It is the seventh most abundant uncombined element found in the universe. Scottish physician Daniel Rutherford discovered dinitrogen back in the year 1772. The symbol of this chemical element is N and its atomic number is 7.

$$|N \equiv N|$$

Preparation of Dinitrogen

We can obtain nitrogen commercially by liquefaction and fractional distillation of air. This process mainly involves two steps:

Step 1

We have to reduce air to liquid air by applying high pressure between 100 to 200 atmospheres. After this, we pass this compressed air through a fine jet where it undergoes expansion. We repeat this method several times which results in the formation of liquid air.

Step 2

The liquid formed undergoes fractional distillation.The boiling point of dinitrogen is lower than that of the liquid oxygen and hence it distils out, leaving behind liquid oxygen. We obtain nitrogen from the impure liquid. In a laboratory, we obtain dinitrogen by reacting an aqueous solution of ammonium chloride with sodium nitrite.

$$NH_4Cl(aq) + NaNO_2(aq) \rightarrow N_2(g) + 2H_2O(l) + NaCl(aq)$$

The products obtained consists of impurities such as NO and HNO_3which can be removed by thermal decomposition of ammonium dichromate. Another method in which we can remove the impurities is by passing the gaseous mixture through sulphuric acid that has potassium dichromate in it.

$$(NH_4)_2Cr_2O_7 \rightarrow N_2 + 4H_2O + Cr_2O_3$$

Decomposition of sodium or barium azide in the presence of high temperature also results in the formation of pure nitrogen.

Physical Properties of Dinitrogen

Now, we look at the physical properties of dinitrogen.

- Nitrogen is a colourless, odourless and diamagnetic in nature.
- It is a non-toxic gas.
- It is sparingly soluble in water.
- Nitrogen undergoes condensation to form a colourless liquid. This, on solidification results in the formation of snow like mass.

Chemical Properties of Dinitrogen

1) Dinitrogen has a high bond enthalpy due to the N = N bond. Due to this, it is inert at room temperature. However, the reactivity increases

as the temperature increases. At high temperatures, nitrogen molecules react with metals. This reaction results in the formation of respective ionic nitrides. The molecules react with non-metals to form covalent nitrides.

$$6Li + N_2 \xrightarrow{heat} 2Li_3N$$

2) At about 773 K, it reacts with hydrogen to form ammonia in Haber's Process.

$$N_2(g) + 3H_2(g) \overset{773k}{\leftrightarrow} 2NH_3(g)$$

3) Nitric oxide is formed by the reaction of nitrogen molecule with the oxygen molecule at a temperature of 2000 K.

$$N_2(g) + O_2(g) \leftrightarrow 2NO(g)$$

Solved Example for You

Q: Give some uses of dinitrogen.

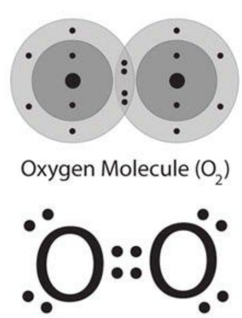
Ans: It is mainly used in the industrial manufacturing of compounds such as ammonia, calcium cyanamide etc. Dinitrogen is used in the manufacturing industries such as iron and steel to obtain an inert atmosphere. We use liquid nitrogen in food industries as a preservative and as a refrigerant.

Dioxygen

You obviously know what oxygen is. Don't you? It is what makes life possible on earth. But, do you know anything more than that? In this chapter, we will learn about Dioxygen. We will look at its properties, preparation methods and uses. Let's begin.

What is Dioxygen?

Similar to Group 14 and 15, the lightest member of group 16 has the best inclination to shape numerous bonds. In this way, elemental Oxygen is found in nature as a diatomic gas that contains a net twofold bond: O=O. Dioxygen means the normal allotrope of oxygen having two atoms of oxygen in the molecule.



Likewise, with nitrogen, electrostatic repulsion between lone pair of electrons on adjoining atoms keeps oxygen from framing stable catenated compounds. Ozone (O_3), a standout amongst the most intense oxidants known, is utilized to cleanse drinking water since it doesn't deliver the characteristic taste connected with chlorinated water.

Equation

$$2H_2O_2\ (l) \rightarrow 2H_2O(l) + O_2\ (g)$$

$$\Delta G^\circ = -119\ kJ/mol$$

Regardless of the quality of the O = O bond (D_{O_2} = 494 kJ/mol), O_2 is amazingly reactive, that reacts straightforwardly with almost all different elements aside from the noble gasses. A few properties of O_2 and related species, for example, the peroxide and superoxide particles, are mentioned in the accompanying table:

Species	Bond Order	Number of unpaired e–	O- O distance (pm)
O2	2	2	121

O2+	2.5	1	112
O2–	1.5	1	133
O22-	1	0	149

Preparation

The most advantageous technique for making oxygen in the laboratory includes the catalytic decay of potassium chlorate in the solid form where manganese dioxide works as a catalyst.

$$2KClO_3 \xrightarrow[MnO_2]{\Delta} 2KCl + 3O_2$$

Another laboratory technique includes the thermal disintegration of metal oxides from the lower part of the electrochemical arrangement. For example, the thermal decomposition of mercuric oxide or silver oxide gives oxygen.

$2HgO \rightarrow 2Hg + O_2$

Mercuric oxide ▵ Mercury Dioxygen

$2Ag_2O \rightarrow 4Ag + O_2$

Silver oxide ▵ Silver Dioxygen

We can also prepare oxygen in the laboratory by thermal treatment of the higher oxides of some metals like lead, manganese and barium.

$2PbO_2 \rightarrow 2PbO + O_2$

Lead (IV) oxide ▵ Lead (II) oxide Dioxygen

$2MnO_2 + 2H_2SO_4 \rightarrow 2MnSO_4 + 2H_2O + O_2$

Manganese Sulphuric ▵ Manganese Water Dioxygen

(IV) oxide (II) sulphate

$2BaO_2 \rightarrow 2BaO + O_2$

Barium peroxide △ Barium oxide Dioxygen

Salts that are rich in oxygen, such as permanganates and nitrates, when decomposed thermally also yields oxygen.

$2KNO_3 \rightarrow 2KNO_2 + O_2$

Potassium nitrate Potassium nitrite Dioxygen

$2KMnO_4 \rightarrow K_2MnO_4 + MnO_2 + O_2$

Potassium permanganate △ Potassium manganate manganese (IV) oxide Dioxygen

$2NaNO_3 \rightarrow 2NaNO_2 + O_2$

Sodium nitrate △ Sodium nitrate Dioxygen

Physical Properties

- Oxygen is a tasteless, colourless and scentless gas.
- It is marginally heavier than air.

- It is marginally soluble in water. This little amount of oxygen dissolved is quite adequate to support the aquatic and marine life.
- Under pressure, we can condense it to a light blue fluid by compacting the gas at 90K. It can likewise be solidified into a bluish white solid at 55K.

Chemical Properties

Oxygen is an exceptionally reactive element and reacts straightforwardly with almost all metals and non-metals. It doesn't react straightforwardly with a few metals like gold and platinum and some noble gases like helium, argon, and neon.

1) The Reaction With Metals

Most metals blaze in oxygen and frame oxides that are for the most part basic in nature.

Metal		Dioxygen	Metal-oxide
$2M$	$+$	O_2	$\rightarrow 2MO$

$$4M + O_2 \rightarrow 2M_2O$$

$$4M + 3O_2 \rightarrow 2M_2O_3$$

A large portion of non-metals burns within the presence of oxygen structures acidic oxides. Example: Sulfur blazes within the presence of oxygen gives sulfur dioxide.

$$S + O_2 \rightarrow SO_2$$

2) Reactions With Some Compounds

Sulfur dioxide experiences catalytic oxidation within the presence of vanadium pentoxide(V_2O_5) to frame sulfur trioxide. This is a critical stride in the production of sulphuric acid by the contact procedure.

$$2SO_2 + O_2 \rightarrow 2SO_3$$

Oxygen reacts with many organic compounds such as carbohydrates and hydrocarbons, at hoisted temperatures or on start, resulting in the formation of carbon dioxide and water.

High Temperature

CH_4 + $2O_2$ → CO_2 + $2H_2O$

Methane Dioxygen Carbon dioxide Water

High Temperature

$C_6H_{12}O_6$ + $6O_2$ → $6CO_2$ + $6H_2O$

Glucose Dioxygen Carbon dioxide Water

Dioxygen Difluoride

Dioxygen Difluoride with the formula O_2F_2 is a compound of fluorine and oxygen. It can exist as an orange-coloured solid which liquefies into a red fluid at −163 °C (110 K). It is to a great degree solid oxidant and breaks down into oxygen and fluorine even at −160 °C (113 K) at a rate of 4% every day: its lifetime at room temperature is in this manner amazingly short. Dioxygen difluoride reacts energetically with about each concoction it experiences – even common ice – prompting to its onomatopoeic label "FOOF."

Solved Example For You

Q: Write down the uses of oxygen.

Ans: The fundamental significance of dioxygen lies in its support to key procedures. For example, respiration and combustion. Its other uses are mentioned below.

- Dioxygen blended with carbon dioxide or helium is utilized for artificial respiration.
- It is utilized as a part of manufacturing many metals.
- It is utilized as a part of oxy-acetylene welding and metal cutting.
- Oxygen is utilized to oxidize ammonia in the nitric acid preparation.
- Oxygen barrels are broadly utilized as a part of healing facilities, high-height flying and in mountaineering.
- Liquid oxygen is an essential constituent of the fuel utilized as a part of rockets.

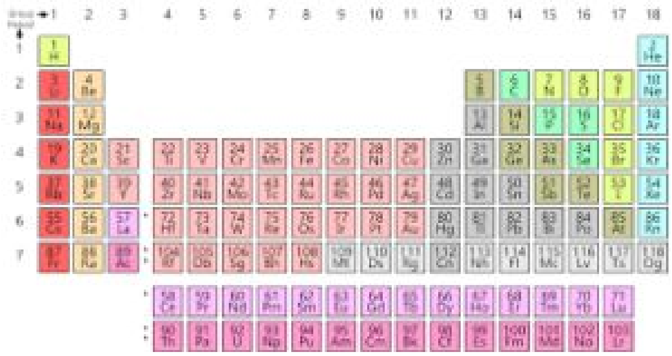
Group 13 Elements: Boron Family

Now we are all a little accustomed to the periodic table. Aren't we? Isn't it a lot difficult to remember each of the properties of the elements? That is why we group them and study their properties. In this chapter, we will study about the Boron family that forms the 13th group in the periodic table. You know boron right? The popular fibreglass is a product of boron. Let us now study about these elements in greater details.

Group 13 Elements- Boron Family

The group 13 elements consist of six elements. They are boron (B), aluminium (Al), gallium (Ga), Indium (In), thallium (Tl), and element 113. This element gets the name of ununtrium [Uut]. The commonest property of the group is that each one of the elements has three electrons in the outer shell of their nuclear structure.

Boron is the lightest of the elements in this group. It is a non-metal. Surprisingly, the others in the group are brilliant white metals. These elements have likewise been alluded to as icosagens and triads.



The Occurrence of the Boron Family

We find boron in limited quantities. It is mostly a product of the barrage of subatomic particles created from characteristic radioactivity. Aluminium is readily available on our planet. It is also the third most copious element in the Earth's outside (8.3%).

We can find Gallium in the earth with a wealth of 13 parts per molecule. Indium is the 61st richest element in the world's covering. Thallium is scattered in small amounts all over the planet. Ununtrium is not available naturally and thus, has been named a synthetic element.

Physical Properties of Group 13 Elements

In this section, we will look at the physical properties of the boron family.

- Indium has a lesser nuclear radius than Thallium. This is because of the lanthanide compression.
- As we move down the group, +1 oxidation state turns out to be steadier than +3 states. This is mainly because of the inert pair impact.
- Boron has a high melting point. This is because of the icosahedral structure. In the boron family, gallium has the lowest melting point.
- All the elements of this family blaze in oxygen at high temperatures framing M_2O_3.
- Aluminium is amphoteric. It means that the metal disintegrates in weakened mineral acids and in sodium hydroxide (aqueous).
- As we move down the group, the acidic nature of hydroxides reduces.
- Boric acid is an extremely delicate monobasic acid.

Chemical Properties of Group 13 Elements

Dissociation of the group 13 elements requires a lot of energy. This is because the compounds formed by the Group 13 elements with oxygen are inert thermodynamically.

Boron acts as a non-metal chemically. However, the rest of the elements show metallic properties. Why does this happen? A large portion of the irregularities seen in the properties of the group 13 elements is attributed to the expansion in Z_{eff}(Effective Nuclear Charge). This emerges from the weakened protection of the atomic charge by the filled (n − 1) d^{10} and (n − 2) f^{14} subshells.

Rather than shaping a metallic grid with delocalized valence electrons, boron frames special aggregates that consist of multicenter bonds. This includes metal borides, in which boron attaches to other boron iotas. This arrangement creates three-dimensional systems or bunches with consistent geometric structures.

All the neutral compounds of the group 13 elements are electron lacking and act like Lewis acids. The trivalent halides of the heavier elements shape halogen-connected dimers that consist of electron-match bonds, as opposed to the delocalized electron-lacking bonds typical for diborane.

Their oxides break down in weakened acids, in spite of the fact that the oxides of aluminium and gallium are amphoteric. The group 13 elements never react with hydrogen because the valency of hydrogen is one and that of the boron family is three. The trihalides of group 13 elements are strong Lewis acids because they have the tendency to form compounds with electron-pair donors, the Lewis bases.

Solved Example for You

Q: Write down the applications of the boron family.

Ans:

- We use Boron commonly in fibreglass. It also finds use in the ceramics market. It is used in the making of pots, plates, vases etc due to its excellent insulating properties.
- Aluminium is frequently utilized as a part of construction materials, electrical gadgets, particularly as a transmitter in links. We also use it in apparatuses and vessels for cooking and

safeguarding food. Aluminium's absence of reactivity with food items makes it especially helpful for canning.

- Aluminium is a part of alloys that we primarily use for making lightweight bodies for flying machine.
- Gallium arsenide is a common part of semiconductors, enhancers, solar cells etc.
- We also use gallium amalgams for a lot of dental purposes. Gallium ammonium chloride finds a common use in the leads in transistors. A noteworthy use of gallium is in LED lighting.
- We can usually find Indium on platings, phosphors, bearing, display gadgets, warm reflectors etc. Indium tin oxide finds an extensive variety of uses. These include glass coatings, solar panels, road lights, electrophoretic displays (EPD) etc.
- They are also common in plasma display boards (PDPs), electroluminescent light displays (ELDs), electrochemical displays (ECs), sodium lights etc.

Group 14 Elements: Carbon Family

Well, you are aware of carbon. Isn't it? Be it in your chapter on respiration or environmental protection, you have heard enough about carbon. However, that is actually not enough! There is so much more to the story. In this chapter, we will look at the carbon family or element 14. We will look at their physical and chemical properties, with examples. Let's begin.

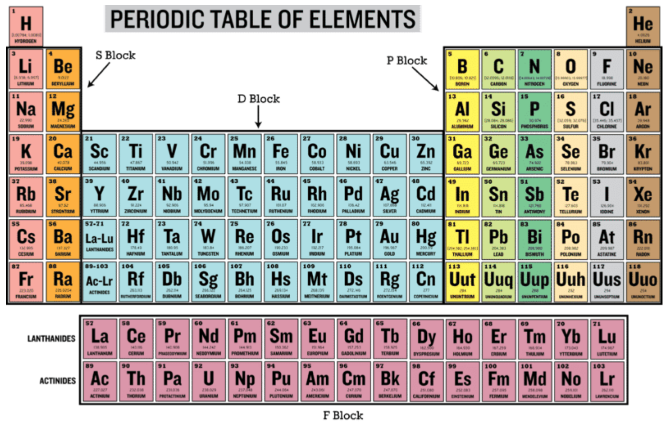
The Carbon Family

We can find the Carbon family towards the right side of the periodic table. We refer to them as the group 14 elements. The members of this family include carbon (C), silicon (Si), germanium (Ge), tin (Sn), lead (Pb), and flerovium (Fl).

These elements belong to the p-block of elements in the periodic table. We can, thus know, their electronic configuration is ns2np2. Let us first look at all the members of this group in greater detail.

Elements of Carbon Family

- Carbon is the first element in this 14th group of elements. It is one of the most plentily available elements present on our earth. We can find it in combined as well as free states. We usually find it in air, polymers, organic compounds, carbonates etc. It has three isotopes, namely, ^{12}C, ^{13}C, and ^{14}C where ^{14}C is radioactive.
- Silicon is a common element in dust, sand, clay, stone, silica and silicate minerals. We can hardly find it as a pure element. It is neither a nonmetal or a metal. In fact, it is a metalloid.
- Germanium is a rare element which we use in the manufacturing of semiconductor devices. Pure germanium is an excellent semiconductor. However, it only occurs in traces as it is too reactive to be found in the elemental state.
- Tin is a soft, malleable metal with a low melting point. It is mainly obtained from the mineral cassiterite. It has two main allotropes at regular pressure and temperature.
- Lead, also plumbate, is obtained from Galena. We find its common use in the making of lead-acid batteries, oxidizing agents, and alloys. Lead is toxic for us, the humans.



Learn more about Group 16 Elements here.

Electronic Configuration of the Carbon Family

Electronic configuration of an atom is nothing but an illustration of the layout of electrons distributed among the sub-shells and orbitals. By this configuration of electrons, we can understand the various physical and chemical properties of the elements. The chemistry behind the elements can be determined by studying the number of valence electrons in the outermost shells.

Before understanding the electronic configuration of elements, we must understand the rules for assigning the electrons into the orbitals. There are many principles that help us in doing so. These include

Pauli’s exclusion principle, Hund’srule of maximum multiplicity and Aufbau principle.

Electrons fill the orbitals in such a way that the energy of the atom is at the minimum. Hence, the electrons of an element fill the energy levels in an increasing order as per the Aufbau principle. Pauli defined a set of unique quantum numbers for each electron. Pauli exclusion principle states that all the four quantum numbers for any two electrons in an atom can never be same.

As per Hund’s rule, the pairing of electrons in an orbital takes place only when all the sub-shells have one electron each. The general electronic configuration of these group 14 elements is ns^2np^2. These elements have 2 electrons in the outermost p orbitals. The electronic configuration of group 14 elements is as follows:

Period	**Element**	**Symbol**	**Atomic Number**	**Electronic Configuration**
2	Carbon	C	6	$[He]2s^2 2p^2$

3	Silicon	Si	14	[Ne]3s2 3p2
4	Germanium	Ge	32	[Ar]3d10 4s2 4p2
5	Tin	Sn	50	[Kr]4d10 5s2 5p2
6	Lead	Pb	82	[Xe]4f14 5d10 6s2 6p2

As all the elements in group 14 have 4 electrons in the outermost shell, the valency of Group-14 elements is 4. They use these electrons in the bond formation in order to obtain octet configuration.

Learn more about Group 17 Elements here.

Properties and Trends in Element 14

1) Covalent Radius

As we move down the group, the covalent radius increases. Therefore, there is a substantial increase in radius from carbon to silicon. Post that, the difference is less considerable. The reason can be credited to the d and f orbitals which are completely filled with the heavier members.

2) Ionisation Enthalpy

Moving down the group, we notice that the ionization enthalpies decrease. This is because of the increase in the distance from the nucleus. There is a substantial decrease of ionization enthalpy from carbon to silicon. Post that, the difference is less considerable. There is a slight increase in ionization enthalpy from tin to lead due to the poor shielding effect of the d and f orbitals.

Learn more about s-Block Elements here.

Solved Example for You

Q: How does the electronegativity vary along the Group 14 elements?

Ans: As we move down the group, the electronegativity decreases in general. The reason behind this irregularity is because of the filling of intervening d and f atomic orbitals. However, the electronegativity is almost the same from silicon to lead.

Group 15 Elements

The periodic table is very vast. You certainly find it difficult to memorise the various properties of each element. That is why we have divided it into various groups or blocks. In this chapter, we will read more about the Group 15 elements of the periodic table. We will look at the various properties of the elements that belong to this group. So let's begin.

What are Group 15 Elements?

The p-block elements are put to the right-hand side of the periodic table in groups from 13 to 18. In the iotas of p-block elements, the separating electron enters the valence p subshell. In this manner, in these elements, the np subshell is filled step by step.

The general valence shell electronic setup of group fifteen elements is ns^2, np^{1-6}. The electronic design of helium is $1s^2$. It has no orbitals. However, it is a p-block element since it takes the physical and chemical properties after that of other p-block elements of the eighteenth group. P-block elements are generally non-metals, while the remaining are metalloids and metals.

P-block elements

Occurrence

Group 15 elements include nitrogen, phosphorus, arsenic, antimony and bismuth. Nitrogen is the real constituent of the air and records for 78% of it by volume. It is the primary member of this group and happens in a free state as a diatomic gas, N_2.

Phosphorus is a fundamental constituent of animal and plant matter. Phosphate groups are constituents of nucleic acids, that is, DNA and RNA. Around 60% of bones and teeth are made out of phosphates. Phosphoproteins are available in egg yolk, milk, and bone marrow. The rest of the elements of the group, that is, arsenic, antimony, and bismuth, mostly happen as sulfides. For example, Stibnite, Arsenopyrite, and bismuth glance.

Learn more about the Characteristics of Group 16 Elements.

Trends in Group 15 Elements

Let us now look at some of the trends of atomic properties of these elements.

1) Atomic Radii

Moving down the group, the ionic radii, and atomic radii increases. This is because of the expansion of another main energy level in each progressive element.

2) Ionization Enthalpy

These elements demonstrate higher values of ionization enthalpy when contrasted with group 14 elements. As we move down the group, the ionization enthalpy values keep on decreasing. This is because of the progressive increase in the size of the nucleus.

Learn more about Some Important Compounds of Carbon and Silicon here in detail

3) Electronegativity

Electronegativity is the inclination of a particle to pull in a shared pair of electrons more towards itself. The electronegativity decreases gradually on moving down the group. The reason behind this is the increase in atomic radius.

4) Physical Properties

Physical properties incorporate physical state, boiling and melting points, metallic character, allotropy, and density. Nitrogen is a diatomic gas, while the rest of the elements are solids in nature.

Moving down a group, metallic character increases. On the other hand, the ionization enthalpy of the elements decreases due to an increase in their nuclear size.

5) Trends in Melting and Boiling Points

The melting point increments from nitrogen to arsenic because of the continuous increment in nuclear size. The low melting point of nitrogen is because of its discrete diatomic particles.

In spite of the fact that the nuclear size increments from arsenic to antimony, there is a decrease in their melting points. Despite the fact that antimony has a layered structure, it has a low melting point than arsenic on account of the generally free pressing of particles. Furthermore, the melting point of bismuth is not as much as antimony because of the packing of atoms loosely by metallic holding. Then again, the boiling point step by step increments from nitrogen to bismuth.

The density of these elements increases from nitrogen to bismuth.

6) Allotropy

All group fifteen elements, aside from bismuth, indicate allotropy. Nitrogen is found in two allotropic structures, that is, alpha nitrogen and beta nitrogen. Phosphorus exists in numerous allotropic structures. Of these, the two critical allotropic structures are red phosphorus and white phosphorus.

Arsenic exists in three essential allotropic structures – black, grey, and yellow. Antimony additionally has three essential allotropic structures, to be specific, yellow, metallic and explosive.

Solved Example for You

Q: Write a note on the oxidation states of the group 15 elements.

Ans: Every one of the elements of group 15 has 5 electrons in their outermost circle. They require just 3 electrons to finish their octet setup. The octet can be accomplished either by picking up 3 electrons or by sharing 3 electrons by a method for covalent bonds.

Accordingly, the basic negative oxidation state of these elements is – 3. Moving down the group, the propensity to display – 3 oxidation state diminishes. This is because of the expansion in nuclear size and metallic character.

Group 15 elements additionally indicate positive oxidation states of +3 and +5 by developing covalent bonds. Because of the inert pair effect, the +5 oxidation state stability diminishes down the group, while that of +3 oxidation state increments.

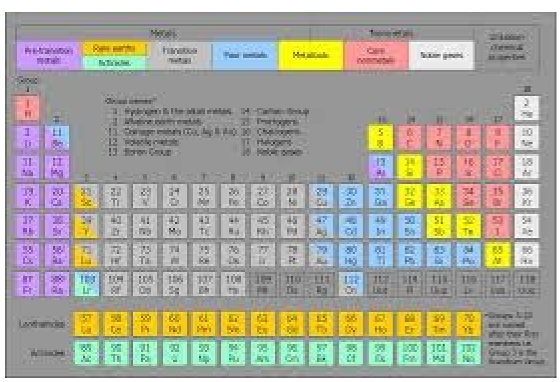
Nitrogen has just s and p-orbitals, yet no d-orbitals in its valence shell. Therefore, nitrogen can demonstrate a most extreme covalency of 4. A covalency of four is reached by the sharing of its lone pair of electrons to another iota or particle.

Group 16 elements

Oxygen is one of the most important elements in the universe. Don't you agree? However, do you know that there are many other elements that behave like oxygen? In this chapter, we will read about the Group 16 elements. Let us see what is similar between oxygen and the other members. We will also see why oxygen is the most important member of that group. So let's begin.

Occurrence and General Characteristics

The elements oxygen, sulfur, selenium, tellurium and polonium comprise the 16th vertical column or VI A group elements in the currently used long type of periodic table.



The initial four elements of the group are together termed as chalcogens or ore-forming elements. This is because an extensive number of metal ores are found in the earth's crust as sulfides or oxides. Oxygen is the most plenteous element that is accessible in nature. It shapes 20.946% of air by volume and 46.6% of the world's mass generally as silicates and different compounds like carbonates, oxides, and sulfates.

The vast majority of the oxygen in the air is delivered by photosynthesis in plants. It additionally occurs as ozone. Sulfur is the sixteenth most inexhaustible element. Sulfur in its combined state is found in ores.

Electronic Configuration of Group 16 Elements

Group 16 elements have 6 electrons in their valence shell and their general electronic configuration is ns^2np^4.

Element	Electronic Configuration
Oxygen	[He] $2s^2$ $2p^4$

Sulphur	[Ne] 3s2 3p4
Selenium	[Ar] 3d10 4s2 4p4
Tellurium	[Kr] 4d10 5s2 5p4
Polonium	[Xe] 4f14 5d10 6s2 6p4

Atomic and Physical Properties and the Trends of Group 16 elements

- Atomic and Ionic Radii: The atomic and ionic radius increases as we move from Oxygen to Polonium.
- Ionization Enthalpy: Ionization enthalpy decreases with increase in the size of the central atom. Therefore, it decreases as we move from Oxygen to Polonium since the size of the atom increases as we move down.
- Electron Gain Enthalpy: The electron gain enthalpy decreases with increase in the size of the central atom moving down the group. Oxygen molecule has a less negative electron gain enthalpy than sulfur. This is on the grounds that Oxygen,

because of its compressed nature encounter more repulsion between the electrons effectively present and the approaching electron.

- Electronegativity: The electronegativity decreases as we move down the group. Therefore, it decreases as we move from oxygen to polonium due to increase in nuclear size.Learn about Electronegativity here in detail.
- Nature of the Group 16 Elements: Oxygen and Sulfur are non-metals, Selenium and Tellurium are metalloids and Polonium is a metal under typical conditions. Polonium is a radioactive element.
- Allotropy: Each one of the element of group 16 displays allotropy. Oxygen has two allotropes: Oxygen and Ozone. Sulphur exists as many allotropic forms but only two of them are stable, which are: Rhombic Sulphur and Monoclinic Sulphur. Selenium and Tellurium are found in both amorphous and crystalline forms.
- The Melting and Boiling Points: As the atomic size increases from oxygen to tellurium, the melting and boiling points also increase. The huge distinction between the melting and boiling points of oxygen and sulfur might be clarified on the premise

that oxygen exists as a diatomic atom while sulfur exists as a polyatomic particle.

- Oxidation States: The group 16 elements have a configuration of ns^2 np^4 in their outer shell, they may accomplish noble gas configuration either by the gain of two electrons, framing M^{-2} or by sharing two electrons, in this manner shaping two covalent bonds.

Thus, these elements indicate both negative and positive oxidation states. The regular oxidation states showed by the elements of group 16 incorporate -2, +2, +4 and + 6.

Chemical Properties

The group sixteen elements react with hydrogen to form hydrides of the sort H_2E, where E could be any element- oxygen, sulfur, selenium, tellurium or polonium.

H_2	+	E	→	H_2E
Hydrogen		Group 16 Element		Hydride

The Physical States of Hydrides of Group 16 Elements

Water is an odourless and colourless liquid but the hydrides of the various elements of this group are poisonous gases which are colourless with disagreeable smells.

The boiling point of these hydrides extraordinarily diminishes from water to hydrogen sulfide, and after that increases. Water has an anomalously high boiling point since its particles are bonded with each other by the hydrogen bonds in both its liquid as well as solid states.

Learn more about the s Block elements here.

Solved Example for You

Q: Write a note on the acidic character of hydrides of Group 16 elements.

Ans: There is an expansion in acidic nature of hydrides from H_2O to H_2Te. The development in acidic character is a result of the decrease in the H-E bond separation enthalpy from H_2O to H_2Te. Except for water, the different hydrides go about as reducing agents. The reducing property of these hydrides increments from H_2S to H_2Te.

Every one of the elements of group 16 reacts with oxygen to shape dioxides and trioxides. Both dioxides and trioxides are acidic in nature. Sulfur trioxide is the primary basic triode in this group. Sulfur trioxide at room temperature is a solid and exists in three specific structures – alpha, beta and gamma.

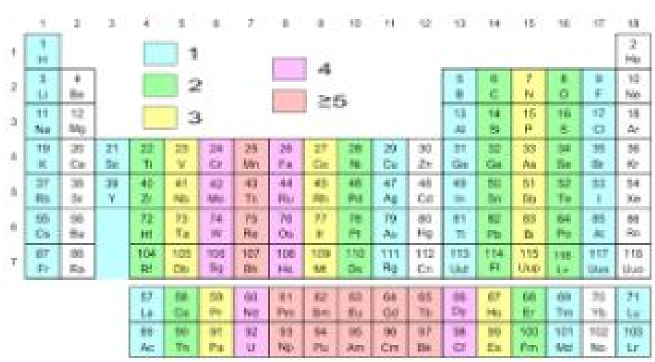
Elements of group 16 accommodate an enormous assortment of halides of the sort EX_6, EX_2, and EX_4, where E is the element of group 16 elements and X is a halogen. Among all hexahalides, just hexafluorides are latent. They encounter sp^3d^2 hybridization, and along these lines, have octahedral geometry.

Group 17 Elements

Now, you have quite some idea about the periodic table. Have you heard about the 'halogen' lights? Yes, the ones that you use in your garden, maybe? But, do you know that the Group 17 elements in the periodic table are called "halogens". Why? Well, we will find out soon! In this chapter, we will cover all about halogens, their properties, and uses.

Group 17 Elements

The group 17 elements include fluorine(F), chlorine(Cl), bromine(Br), iodine(I) and astatine(At) from the top to the bottom. They are called "halogens" because they give salts when they react with metals. So, now you know what halogens are! Let's now look at the electronic configuration of these elements.



Electronic Configuration of Group 17 Elements

The valence shell electronic configuration of these electrons is ns2np5. Thus, there are 7 electrons in the outermost shell of these elements. The element misses out on the octet configuration by one electron. Thus, these elements look out to either lose one electron and form a covalent bond or gain one electron and form an ionic bond. Therefore, these are very reactive non-metals.

Atomic Properties

Let us now look at the various atomic properties of the group 17 elements. We will speak about the ionic and atomic radii, ionization enthalpy and more.

1) Ionic and Atomic Radii

The nuclear and atomic radii of these elements keep on increasing as we move down the group. This happens because of the addition of an extra energy level. They have the minimal atomic radii compared to the other elements in the related periods. This can be attributed to the fact that their atomic charge is quite powerful.

2) Ionisation Enthalpy

These elements have higher ionization enthalpy. This value keeps on diminishing as we move down the group. This happens because of the increase in the size of the nucleus. However, it is interesting to note that fluorine has the highest ionization enthalpy than any other halogen, thanks to its minute size!

3) Electron Gain Enthalpy

The electron gain enthalpy of these elements becomes less negative upon moving down the group. Fluorine has lesser enthalpy than chlorine. We can attribute it to the small size and the smaller 2p sub-shell of the atom of fluorine.

4) Electro-Negativity

The halogens exhibit high electro-negativity values. However, it diminishes slowly on moving down the group from fluorine to iodine. this can be attributed to the increase in nuclear radii upon moving down the group.

Learn more about Group 16 Elements here.

Physical Properties

Let us now look at the various physical properties of these halogens.

- Physical state: The group 17 elements are found in diverse physical states. For example, Fluorine and Chlorine are gases. On the other hand, Bromine is a liquid and Iodine is solid.
- Colour: These elements have a variety of colours. For example, while Fluorine is pale yellow in colour, Iodine is dark violet in colour.
- Solubility: Florine and Chlorine are soluble in water. On the other hand, Bromine and Iodine are very less soluble in water.
- Melting and boiling points: Melting and boiling points of these elements increase as we move down the group from Fluorine to Iodine. Thus, Fluorine has the lowest boiling and melting points.

Chemical Properties

We will now have a look at some of the chemical properties of these elements.

1) Oxidising Power

All the halogens are great oxidizing agents. Of the list, fluorine is the most powerful oxidizing agent. It is capable of oxidizing all the halide particles to halogen. The oxidizing power reduces as we move down the group. The halide particles also act as reducing agents. However, their reducing capacity decreases down the group as well.

2) Reaction with Hydrogen

All halogens react with hydrogen and produce acidic hydrogen halides. The acidity of these hydrogen halides reduces from HF to HI. Fluorine reacts violently and chlorine requires the sunlight. On the other hand, bromine reacts upon heating and iodine needs a catalyst.

3) Reaction with Oxygen

Halogens react with oxygen to form oxides. However, it has been found that the oxides are not steady. Beside oxides, halogens also form a number of halogen oxoacids and oxoanions.

4) Reaction with Metals

As halogens are very reactive, they react with most of the metals instantly and form the resulting metal halides. For example, sodium reacts with chlorine gas and forms sodium chloride. This process is an exothermic one and gives out a bright yellow light and a lot of heat energy.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Metal halides are ionic in nature. This is because of the high electronegative nature of the halogens and high electropositivity of the metals. This ionic character of the halides reduces from fluorine to iodine.

Learn more about s-Block Elements here.

Solved Example for You

Q: Mention some uses of halogens.

Ans: The uses of halogens are:

- Fluorine compounds constitute an important ingredient in toothpaste. This is because fluoride compounds react with the enamel of the teeth and take care of teeth rotting.
- Chlorine majorly used as a bleach. It is also used in the metallurgy of elements like platinum and gold.

- Iodine is used as an antiseptic because it kills the germs on the skin.

Group 18 Elements

How many of you have played with the helium balloons when you were kids? Well, some of you do that even today! However, are you aware of what helium is? Why is it used in balloons? Why don't the other group 18 elements find their use in these balloons? We will read about all this in this chapter. We will look at the physical properties, chemical properties and uses of the group 18 elements.

Group 18 Elements

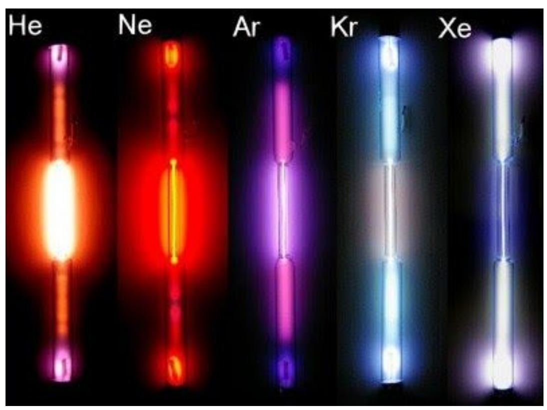
The Group 18 elements include Helium(He), Neon(Ne), Argon(Ar), Krypton(Kr), Xenon(Xe), and Radon(Rn). They are referred to as noble gases or inert gases. This means that these elements are chemically inert and do not take part in any reaction.

Electronic Configuration of Group 18 Elements

The general configuration of the valence shell is ns^2np^6. The exception to this is helium. It has the configuration of $1s^2$. As they have the octet configuration in their valence shells already, they are quite chemically inert. they have the valency of zero.

The Occurrence of These Elements

All of these elements occur in a free state in the atmosphere. Apart from Radon, every other noble gas exists in the atmosphere. Argon alone constitutes 0.93% of the total atmosphere. We can prepare this element by the fractional distillation of liquid air. We can find neon, helium and argon in certain water springs as disintegrated gasses. Also, we can obtain Radon by the decay of radium and thorium minerals.



(Source: io9.gizmodo)

Trends in the Atomic Properties

- Atomic Radius: The nuclear radii increment on moving down the group with increasing nuclear number. This is a result of the expansion of another shell at each progressive element on moving down the group.

- Electron Gain Enthalpy: Group 18 elements exhibit very stable electronic configurations. They do not have any tendency to accept an electron.
- Ionisation Potential: They have high ionisation potentials, thanks to their closed electronic configurations. This value decreases on moving down the group due to an expansion in the nuclear size.

Physical Properties

- Because of their stable nature, we find these elements as monatomic gases in a free state.
- They are colourless, tasteless and odourless gases. The particles of these elements have weak Van der Waals forces. This force increases on moving down the group. This is due to an expansion in the polarising capacity of the molecules.
- They exhibit low melting and boiling points. We can attribute this to the weak Van der Waals forces. The melting and boiling points increase as we move down the group.

- We can condense these elements at extremely low temperatures. As the size of the atoms increases down the group, the ease of liquefaction also increases.

Chemical Properties

- These elements are chemically latent because of their stable electronic configuration.
- Group 18 elements have high positive electron gain enthalpy and high ionization enthalpy.
- In 1962, Neil Bartlett anticipated that xenon ought to react with platinum hexafluoride. He was the first to set up a compound of xenon, called xenon hexafluoroplatinate(V). Later, many compounds of xenon were integrated, including fluorides, oxyfluorides, and oxides.

$$Xe + PtF_6 \rightarrow Xe[PtF_6]$$

Xenon	Platinum Hexafluoride	Xenon Hexafluoroplatinate(V)

- The chemical movement of group eighteen elements increments with a diminishment in the ionization enthalpy on moving down the group.
- The ionization enthalpies of helium, argon, and neon are too high for them to shape compounds.
- Krypton only forms krypton difluoride, since its ionization enthalpy is marginally higher than that of xenon.
- Although radon has less ionization enthalpy than xenon, it shapes just a few compounds like radon difluoride, and a few complexes, since radon has no steady isotopes. In any case, xenon shapes an especially more prominent number of compounds.

Uses of Helium

- Helium finds its use in filing air balloons and aircraft. This is because it is flammable and has a very low density. Because it has a very low boiling point of only 4.2 K, we use fluid helium as a cryogenic agent to perform tests at extremely low temperatures.

- Liquid helium is common in the process of cryoscopy that we need for superconductivity. Fluid helium is a common ingredient to cool the superconducting magnets in the atomic magnetic resonance spectrometers.
- We also use it as the cooling gas in gas-cooled atomic reactors, and also as a stream gas in gas-fluid chromatography.
- It is also a major component in the oxygen cylinders that we use in deep sea diving.

Solved Example for You

Q: Give some uses of Neon, Argon, Krypton and Xenon.

Ans: The various uses of Neon, Argon, Krypton and Xenon include:

- We use Neon in minute amounts in various release tubes and fluorescent light bulbs. It gives the reddish-orange shine to these bulbs and the “Neon signs”.

- We use argon in various metallurgical processes like welding stainless steel, aluminium etc. Also, we use it for giving a stable atmosphere in research centres.
- We use Krypton in fibre lights and release tubes. On the other hand, we use Xenon in release tubes to deliver the rapid flash that is essential for photography.
- Radon finds its use in the treatment of cancer.

Hydrogen Chloride

Have you done any experiments with hydrogen chloride yet? Did your chemistry teacher not show you the reaction between sodium hydroxide and hydrogen chloride? Well, in this chapter, we will look at the compound of hydrogen chloride and look at its properties and more. In the end, we will also see some of the important reactions of hydrochloric acid and its uses in various industries. Let's begin.

Manufacture of Hydrogen Chloride

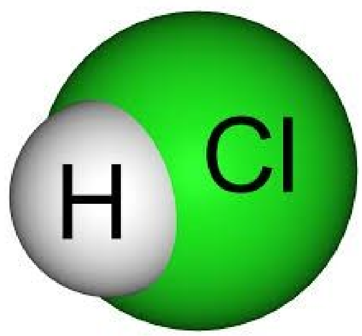
We produce Hydrogen Chloride in the laboratory by treating sodium chloride with concentrated sulphuric acid. We, then, heat this mixture up to 420K.

$$NaCl + H_2SO_4 \xrightarrow{420K} NaHSO_4 + HCl$$

We get Sodium bisulphate as a by-product which is insoluble. Therefore, we further mix it with more sodium chloride. This mixture has to be further warmed to a higher temperature of around 823K. It gives dissolvable sodium sulfate and HCl gas.

$$NaHSO_4 + NaCl \xrightarrow{823K} Na_2SO_4 + HCl$$

We dry this HCl by treating it with concentrated sulphuric acid. HCl is not dried over phosphorus pentoxide or brisk lime. This is because it reacts with both of these compounds.



Properties of Hydrogen Chloride

- Hydrogen Chloride is a vapid gas. It has a very sharp and pungent odour.
- It can melt to a colourless fluid at 189K. HCl forms a white solid at 159K upon freezing.

- It is very soluble in water. An aqueous solution of Hydrogen Chloride is what we know as the hydrochloric acid.
- Hydrochloric acid has a higher dissociation constant and is, therefore, a strong acid.
- It reacts with metals and salts to give various chlorides. For example, it reacts with zinc to form zinc chloride.
- Hydrochloric acid reacts with iron to form ferrous chloride.

$$Fe + 2HCl \rightarrow FeCl_2 + H_2$$

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- Trend and Anomalous Properties of Carbon
- Trends and Properties of Boron and Aluminium
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- Chlorine
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- Dioxygen
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- Sulphur – Allotropic Forms
- Sulphuric Acid
- Sulphuric Dioxide

Some Important Reactions of HCl

- Noble metals react with a mixture of nitric acid and hydrochloric acid in the proportion of 1:3. This arrangement what we call aquaregia.
- Hydrochloric acid reacts with the salts of weaker acids. Some of them include sodium carbonate, sodium bicarbonate etc.
- Hydrochloric acid also reacts with sodium carbonate and sodium bicarbonate and forms sodium chloride, carbon dioxide, and water.
- It also breaks down sodium sulphate to sodium chloride, sulfur dioxide, and water.
- Reaction with Sodium Carbonate:

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2\uparrow + H_2O$$

- Reaction with Sodium Bicarbonate:

$$NaHCO_3 + HCl \rightarrow NaCl + CO_2\uparrow + H_2O$$

- Reaction with Sodium Sulphate:

$$Na_2SO_3 + 2HCl \rightarrow 2NaCl + SO_2\uparrow + H_2O$$

Solved Example for You

Q: Give some uses of Hydrochloric acid.

Ans: The uses of the Hydrochloric acid include:

- We use hydrochloric acid in the process of fabrication of chlorine, and chlorides like ammonium chloride. It is also a common ingredient in the fabrication of glucose from corn starch.
- We also use hydrochloric acid as a research facility reagent and in medicines.
- We use a saturated arrangement of zinc chloride in dilute hydrochloric acid in removing the contaminations on a metal surface prior to welding or electroplating.

Interhalogen Compounds

Did you know that there are combinations of halogen compounds as well? They are the interhalogen compounds. They consist of two halogens. In this chapter, we will talk about these compounds, look at their properties and uses. Let us start with what interhalogen compounds are.

Interhalogen Compounds

We can refer to the Interhalogen Compounds as the subordinates of halogens. These are the compounds having two unique sorts of halogens. For example, the common interhalogen compounds include Chlorine monofluoride, bromine trifluoride, iodine pentafluoride, iodine heptafluoride, etc.

Types of Interhalogen Compounds

We can divide interhalogen compounds into four types, depending on the number of atoms in the particle. They are as follows:

- XY Compounds
- Compounds XY_3

- Compounds XY_5
- XY_7 Compounds

In these notations, we must understand that "X" is the bigger (or) less electronegative halogen. On the other hand, "Y" represents the smaller (or) more electronegative halogen. We can calculate the number of particles in the atom by the concept of the radius ratio. The formula for the same is as follows:

Radius Ratio = Radius of Bigger Halogen Particle/Radius of Smaller Halogen Molecule

With an increase in the radius proportion, we see that the number of atoms per molecule also increases. Therefore, we can make out that Iodine heptafluoride possesses the greatest number of particles per atom. This is because it has a magnificent radius proportion.

Preparation of Interhalogen Compounds

We can manufacture these interhalogen compounds by two main methods. One of them includes the direct mixing of halogens and the other includes a reaction of halogens with the lower Interhalogen compounds.

- The halogen atoms react to form an interhalogen compound. One example includes the reaction when a volume of chlorine reacts with an equal volume of fluorine at 473K. The resultant product is chlorine monofluoride.
- In other cases, a halogen atom acts with another lower interhalogen to form an interhalogen compound. For example, fluorine reacts with iodine pentafluoride at 543K. This gives rise to the compound of Iodine Heptafluoride.

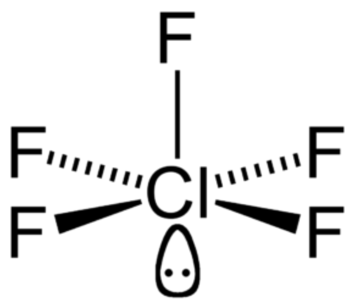
Properties of Interhalogen Compounds

- We can find Interhalogen compounds in vapour, solid or fluid state. A lot of these compounds are unstable solids or fluids at 298K. a few other compounds are gases as well. As an example, chlorine monofluoride is a gas. On the other hand, bromine trifluoride and iodine trifluoride are solid and liquid respectively.

- These compounds are covalent in nature. We can attribute it to the lesser electronegativity between the bonded molecules. Examples include Chlorine monofluoride, Bromine trifluoride and Iodine heptafluoride. These compounds are covalent in nature.
- These interhalogen compounds are diamagnetic in nature. This is because they have bond pairs and lone pairs.
- Interhalogen compounds are very reactive. One exception to this is fluorine. This is because the A-X bond in interhalogens is much weaker than the X-X bond in halogens, except for the F-F bond.
- We can use the VSEPR theory to explain the unique structure of these interhalogens. In chlorine trifluoride, the central atom is that of chlorine. It has seven electrons in its outermost valence shell. Three of these electrons form three bond pairs with three fluorine molecules leaving four electrons.

Common Shapes of these Compounds

Applying the VSEPR theory, we can see that it forms a trigonal bipyramid. The lone pairs take up the tropical positions. On the other hand, bond pairs take up the other three positions. The axial bond pairs bend towards the tropical position. This happens in order to minimize the repulsions that happen due to lone pair-lone pair bonds. Thus, it has the shape of a bowed T.



Chlorine Pentafluoride

Let us now take the case of Iodine Pentafluoride. The central atom in Iodine pentafluoride is the iodine atom. It has one lone pair and five bond pairs. This is the reason it has a square pyramidal shape. Similarly, let us consider the case of Iodine heptafluoride. It has seven bond pairs and has the shape of pentagonal bipyramid.

Uses of Interhalogen Compounds

- We use interhalogen compounds as non-watery solvents.

- Also, we use these compounds as a catalyst in a number of reactions.
- We use UF_6 in the enrichment of ^{235}U. We can produce this by using ClF_3 and BrF_3.

$$U(s) + 3ClF_3(l) \rightarrow UF_6(g) + 3ClF(g)$$

- We use these compounds as fluorinating compounds.

Solved Examples for You

Q1: Can fluorine ever be a central atom?

Ans: Fluorine cannot be a central particle in the inter-halogen compounds. This is because it is an element from the 2nd period in the periodic table. Since it has 7 valence electrons, it can form only one bond.

Q2: Why can't hydrogen be the central atom?

Ans: Hydrogen can't be the central atom. We can attribute this to the fact that an atom will always attempt and get to the condition of most minimal energy. In case of Hydrogen, this means it can form only a single bond. It also has a very small size and does not fit into the other molecules present around it.

Oxides of Nitrogen and Nitric Acid

Oxides of Nitrogen are a mixture of seven different gases and compounds which are form from nitrogen and oxygen. Gases in this group are Nitrous Oxide (N_2O), Nitrogen Monoxide (NO), Dinitrogen Trioxide (N_2O_3), Nitrogen Dioxide (NO_2), Dinitrogen Pentoxide (N_2O_5), Dinitrogen Tetroxide (N_2O_4). Nitric Oxide and Nitrogen Dioxide are common and hazardous. Without nitrogen, life would hardly sustain. In this chapter, we will look at oxides of nitrogen and Nitric acid.

Nitric Acid

Nitric acid is the most popular and helpful oxoacid of nitrogen. It has a molecular formula of HNO_3 and its molar mass is 63.01 g mol^{-1}. Now, let us see the various properties of the nitric acid.

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- Sulphur – Allotropic Forms

- Sulphuric Acid
- Sulphuric Dioxide

Physical Properties of Nitric Acid

- Pure nitric acid is nothing but a colourless fuming fluid. It has a typical and strong odour.
- Upon standing, it creates a yellow shading. This is because of the presence of various oxides of nitrogen that are dissolved in it. The primary oxide is NO_2.
- The acid is completely soluble in water.
- The thickness of the pure acid is 1.54 g/mL.
- We can see that anhydrous nitric acid bubbles at 355.6 K (83.6°C). It is capable of developing a white solid at 231.4 K (-41.7°C).
- It can corrode the skin and causes yellow rashes.

Chemical Properties of Nitric Acid

In this section, we will cover some of the most important chemical properties of Nitric Acid.

1) Stability

In its pure form, Nitric acid is not totally stable. Even at normal temperatures, it slightly decays when we expose it to daylight. Upon strong heating, it breaks down and gives nitrogen dioxide, oxygen and water.

$$\underset{\text{Nitric acid}}{4HNO_3\,(aq)} \longrightarrow \underset{\text{Water}}{2H_2O\,(l)} + \underset{\text{Nitrogen oxide}}{4NO_2\,(g)} + \underset{\text{Oxygen}}{O_2\,(g)}$$

Nitrogen dioxide has a reddish brown colour. However, it might further break down in the undecomposed acid and produce its yellowish chestnut shade.

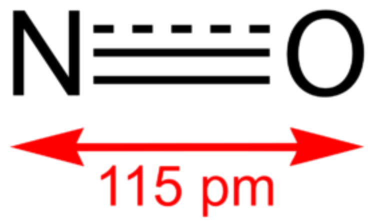
2) Acidic Nature

Nitric acid ionises easily in water and is a strong monobasic acid.

$$HNO_3\,(l) + H_2O(l) \longrightarrow H_3O^+\,(aq) + NO_3^-\,(aq)$$

- Reactions with Basic Oxides: Nitric acid reacts with basic oxides to give a nitrate salt and water. For example, it reacts with calcium oxide to give calcium nitrate and water.
- Reaction with Bases (Hydroxides): Nitric acid reacts with hydroxides to give nitrate salt and water. For example, it reacts with sodium hydroxide to give sodium nitrate and water.
- Reaction with Carbonates and Hydrogen Carbonates: Nitric acid reacts with carbonates and hydrogen carbonates to give nitrate salt, carbon dioxide and water. For example, it reacts with sodium carbonate to give sodium nitrate, water and carbon dioxide.
- Reaction with Metals: Nitric acid does not behave like an acid with metals to form a salt and free hydrogen. However, magnesium and manganese are the two metals that react with cold and extremely dilute (1%) nitric acid to give out hydrogen.

Let us now look at the concept of various oxides of nitrogen.



Oxides of Nitrogen

Nitrogen reacts with oxygen and results in a number of nitrogen oxides. The oxidation states of all these oxides are pretty different. They are in the range of +1 to +5. We will look at some of the important oxides below.

1) Dinitrogen Oxide, N_2O

This is a colourless and non-flammable gaseous compound that has neutral properties. We know it by the common name, laughing gas. We can prepare it by the decomposing ammonium nitrate under high temperature.

$$NH_4NO_3 \rightarrow N_2O + 2H_2O$$

2) Nitrogen Monoxide, NO

This is a colourless and gaseous compound. We can usually prepare it by reducing dilute nitric acid with copper.

$$2NaNO_2 + 2FeSO_4 + 3H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 2NaHSO_4 + 2H_2O + 2NO$$

3) Dinitrogen Trioxide, N_2O_3

Dinitrogen trioxide is a deep blue solid. It has acidic properties. It is prepared by mixing equal parts of nitric oxide and nitrogen dioxide and by further cooling the mixture below −21 °C (−6 °F).

$$NO + NO_2 \rightarrow N_2O_3$$

4) Nitrogen Dioxide, NO_2

Nitrogen dioxide is a common oxide of nitrogen. It is a reddish-brown toxic gas. We can know its presence with a sharp odour.

5) Dinitrogen Tetroxide, N_2O_4

Dinitrogen tetroxide is a colourless solid that we can find in equilibrium with nitrogen dioxide. It is a powerful oxidizer and is a common reagent in the production of many chemical compounds.

$$N_2O_4 \rightleftharpoons 2NO_2$$

6) Dinitrogen Pentoxide, N_2O_5

Dinitrogen pentoxide is a colourless solid. Its characteristic property is that it sublimes slightly above room temperature. It is unstable. It is a potentially dangerous oxidizer. We can prepare it by dehydrating nitric acid (HNO3) with phosphorus (V) oxide:

$$P_4O_{10} + 12HNO_3 \rightarrow 4H_3PO_4 + 6N_2O_5$$

Solved Example for You

Q: Give some uses of Nitric acid.

Ans: The essential uses of nitric acid are as follows:

- It plays the role of an important ingredient in explosives such as trinitrotoluene (T.N.T.), gun cotton etc.
- It is also important in the manufacture of fertilizers.
- We use it in the manufacturing of dyes, scents, drugs etc.

- Also, it is important for the cleaning of precious metals like gold, silver, platinum etc.

Oxoacids of Halogens

By now, you know a lot about the group 17 elements. Do you remember what they are commonly called as? Yes! You are right! It's halogens. So, what are we going to read about today? In this chapter, we will cover the concept of oxoacids of these elements, halogens. It is an important chapter as these oxoacids are very important compounds. Before we begin, let us have a brief recap of what we know about these group 17 elements.

Group 17 Elements: Halogens

Group 17 elements are fluorine, chlorine, bromine, iodine, and astatine from the top to the bottom of the group. We know them by "halogens" because they are salt producers. The members of this group are highly similar to each other. They exhibit a regular pattern in the physical and chemical properties. Did you know that astatine is the only radioactive element in the group? All of these elements have seven electrons in their valence shell. Their electronic configuration is ns^2 np^5

We can see that they are one electron less from the nearest noble gas or octet configuration. These elements have a small size because of their effective nuclear charge. Hence, these do not possess the tendency to lose electrons. As a matter of fact, they gain an electron easily and complete their octet configuration. Halogens produce several oxoacids. These are nothing but the acids containing oxygen in the acidic group.

Oxoacids of Halogens

You may ask, what are oxoacids? An oxoacid is a compound having hydrogen, oxygen, and no less than one other element. These do not have any lesser than one hydrogen molecule bound to oxygen. This hydrogen is capable of separating into the H^+ cation and the anion of the acid.

The fluorine atom is extremely small and thus, it is highly electronegative. Therefore, it can form a single oxoacid, HOF which is fluoric(I) acid or hypofluorous acid. The other elements of the halogen family produce several oxoacids.

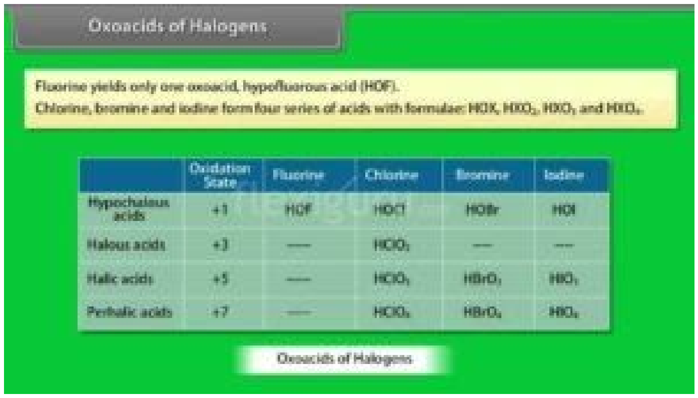
We cannot isolate them in the pure state. They are stable in aqueous solution. They are also very stable in their salt forms. Halogens generally form four series of oxoacids namely hypohalous acids (+1 oxidation state), halous acids (+3 oxidation state), halic acids (+5 oxidation state) and perhalic acids (+7 oxidation state).

Structures of the Oxoacids of Halogens

We can see that the focal halogen molecule is sp^3 hybridised in these oxoacids. We can find an X-OH bond in each oxoacid. In the majority of these oxoacids, "X = O" bonds are available. Hypohalous acids incorporate hypofluorous acid, hypochlorous acid, hypobromous acid and hypoiodous acid. The halogen has the oxidation condition of +1 in hypohalous acids.

Some More Examples

Chlorine is capable of forming four types of oxoacids. They are HOCl (hypochlorous acid), HOClO (chlorous acid), $HOClO_2$(chloric acid) and lastly $HOClO_3$ (perchloric acid). Bromine forms HOBr (hypobromous acid), $HOBrO_2$(bromic acid) and $HOBrO_3$ (perbromic acid). Iodine forms HOI (hypoiodous acid), $HOIO_2$ (iodic acid) and $HOIO_3$ (periodic acid).



Oxoacids of Halogens

Fluorine yields only one oxoacid, hypofluorous acid (HOF).
Chlorine, bromine and iodine form four series of acids with formulae: HOX, HXO_2, HXO_3 and HXO_4.

	Oxidation State	Fluorine	Chlorine	Bromine	Iodine
Hypochalous acids	+1	HOF	HOCl	HOBr	HOI
Halous acids	+3	—	$HClO_2$	—	—
Halic acids	+5	—	$HClO_3$	$HBrO_3$	HIO_3
Perhalic acids	+7	—	$HClO_4$	$HBrO_4$	HIO_4

Oxoacids of Halogens

(Source: YouTube)

The central atom in the oxoacids is sp3 hybridized. Every oxoacid has essentially one X-OH bond. Whereas most oxoacids have X=O bonds present in them.

Solved Example for You

Q: Is hydrochloric acid an oxoacid?

Ans: No. Hydrochloric acid is a hydro acid.

Oxoacids of Phosphorus

Oxoacids of Phosphorus are Hypophosphoric acid(H_3PO_4), Metaphosphoric acid(HPO_2), Pyrophosphoric acid ($H_4P_2O_7$), Hypophosphorous acid(H_3PO_2), Phosphorous acid (H_3PO_3), Peroxophosphoric acid (H_3PO_5), Orthophosphoric acid (H_3PO_5). Oxoacids are acids containing oxygen. Let us Learn more about Oxoacids and Oxoacids of Phosphorus.

What are Oxoacids?

In simple terms, oxoacids are the acids containing oxygen. Phosphorus is one such element that forms a number of oxoacids. A few common oxyacids include H_3PO_4, H_3PO_3, etc.

In oxoacids of phosphorus, we see that the phosphorus is tetrahedrally surrounded by other atoms. Generally, it is clear that there are at least one P=O bond and one P–OH bond in these acids. P–P or P–H bonds are also present besides the P=O bonds and P–OH bonds in oxoacids of phosphorus. In these cases, the oxidation state of phosphorus is less than +5.

These acids are generally seen to jump to higher and lower oxidation states. For example, upon heating, phosphorous acid disproportionates to result in phosphoric acid and phosphine.

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Oxoacids of Phosphorus

The P-H bonds in oxoacids are not ionisable to give H^+ ions. On the other hand, the H atoms attached to oxygen in P-OH form are ionisable. Hence, we can say that basicity is the property exhibited by the H atoms that are attached to oxygen.

As a result, phosphorous acid, H_3PO_3 is dibasic as it has two P-OH bonds. Similarly, phosphoric acid, H_3PO_4 is tribasic as it has three

P-OH bonds. The oxoacids of phosphorus that have P-H bonds exhibit strong reducing properties. For example, hypophosphorous acid containing two P-H bonds is a very good reducing agent.

$$4AgNO_3 + 2H_2O + H_3PO_2 \rightarrow 4Ag + 4HNO_3 + H_3PO_4$$

Few Popular Oxoacids of Phosphorus

In this section, we will look at some of the most important and popular oxoacids of phosphorus.

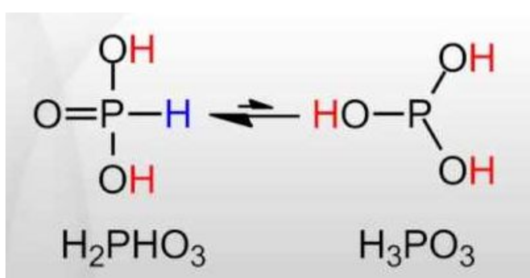
Name	Formula
Hypophosphoric acid	H4P2O6
Metaphosphoric acid	HPO2
Pyrophosphoric acid	H4P2O7
Hypophosphorous acid	H3PO2

Phosphorous acid	H3PO3
Peroxophosphoric acid	H3PO5
Orthophosphoric acid	H3PO4

1) Phosphorus acid, H_3PO_3

Phosphorous acid is a diprotic acid. This means that it ionizes two protons. We can describe it in a better manner by the structural formula $HPO(OH)_2$. We can prepare phosphorous acid by the hydrolysis of phosphorus trichloride with acid or steam.

$$PCl3 \quad + \quad 3H_2O \quad \rightarrow \quad HPO(OH)_2 \quad + \quad 3HCl$$



2) Phosphoric acid, H_3PO_4

Phosphoric acid is a triprotic acid. This means that it can ionise three protons. It is a non-toxic acid, when pure. It is solid at the room temperature and pressure. We can prepare phosphoric acid by adding sulfuric acid to tricalcium phosphate rock:

$$Ca_5(PO_4)_3X + 5H_2SO_4 + 10H_2O \rightarrow 3H_3PO_4 + 5CaSO_4.2H_2O + HX$$

(X can be F, Cl, Br, and OH).

Learn more about P Block Elements here in detail

3) Polymetaphosphoric Acid $(HPO_3)_n$

We can obtain it by warming orthophosphoric acid to around 850 K. It does not exist as a monomer. It exists as a cyclic trimer, cyclic tetramer or polymer.

Solved Example for You

Q: How do we obtain Hypophosphoric Acid ($H_4P_2O_6$)?

Ans: This acid can be set up by the controlled oxidation of red phosphorus with sodium chlorite. When we get the disodium salt of the hypophosphoric acid, we pass it through a cation exchanger to yield hypophosphoric acid. It is tetrabasic in nature.

Oxoacids of Sulphur

What do you know about sulphur? In the previous chapters, we have looked at its configuration and properties. However, are you aware of what oxyacids of sulphur are? Okay, don't be nervous! It's nothing that is complicated. Well, in this chapter we will help you with the basics of oxyacids of sulphur. It is an interesting concept and we are sure you are going to enjoy looking at the various properties of these oxoacids.

Oxoacids of Sulphur

As we have already mentioned, oxoacids are the acids that contain oxygen. A lot of experiments have shown sulphur to form many oxoacids. For example, these oxoacids are H_2SO_4, H_2SO_3, etc. You already knew about these acids, didn't you? Yes! See, we told you it wasn't that difficult!

In oxoacids, sulphur shows a tetrahedral structure with respect to oxygen. Generally, these oxoacids have a minimum of one S=O bond and one S-OH bond. We also observe terminal peroxide groups, terminal S=S, terminal and bridging oxygen atoms in these oxoacids.

Let us now look at some of the most popular oxoacids and their properties.

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1) Sulphuric Acid, H_2SO_4

Sulphuric acid is possibly the most common oxoacid. We have seen a lot of applications of this acid, haven't we? It is a diprotic acid. It signifies its property of ionising into two protons. In sulphuric acid, one atom of sulphur bonds to two hydroxyl groups. The other two oxygen atoms form pie bonds with the atom.

Thus, sulphuric acid exhibits tetrahedral geometry. As the bond length of the sulphur-oxygen bond (S=O) is quite less as compared to the bond length of S-OH, the oxygen atoms repel the OH groups. Hence, the bond angle of O=S=O bond is greater than the HO-S-OH bond angle. We produce it industrially by the contact process.

2) Sulphurous Acid, H_2SO_3

Sulphurous acid is a diprotic acid and thus, gives rise to two protons. In sulphurous acid, one atom of sulphur bonds with two hydroxyl groups. Also, one oxygen atom forms a pie bond with the sulphur atom. We prepare this by dissolving sulfur dioxide in water. As of this date, we do not have any evidence of the existence of sulphurous acid in solution phase. However, the molecule is capable to be isolated in its gaseous phase.

Watch Videos on Percent Free SO_3 –

3) Peroxodisulphuric Acid, $H_2S_2O_8$

Peroxodisulphuric acid contains sulphur in +6 oxidation state. Thus, it acts as a very strong oxidizing agent. It is, also, very explosive in nature. We know it by the common name, Marshall's acid.

It contains one peroxide group that acts as a bridge between the two sulphur atoms. Each atom connects to one hydroxyl group (S-OH bond) and two oxygen atoms (S=O bond) other than the peroxide group. We can prepare it by the reaction of the chlorosulfuric acid with hydrogen peroxide. The reaction is given as follows:

$$2ClSO_3H + H_2O_2 \rightarrow H_2S_2O_8 + 2HCl$$

4) Pyrosulphuric Acid, $H_2S_2O_7$

Pyrosulphuric acid is also known as oleum. Its molar mass is 178.13 g/mol. It is an anhydride of sulphuric acid, is a colourless, crystalline solid and has a melting point of 36°C. It can be prepared by reacting excess sulphur trioxide with sulphuric acid. The reaction goes like this:

$$H_2SO_4 + SO_3 \rightarrow H_2S_2O_7$$

It reacts with bases to form salts which are called pyrosulphates. We use it in the manufacturing of explosives and dyes. it is also used in petroleum refining. It is a strong dehydrating agent and is corrosive in nature. It can cause burns on the skin and irritation to the eyes. Long exposure can prove to be fatal.

Solved Example for You

Q: What are thioacids and peroxy acids?

Ans: Oxoacids having S-S linkages are thioacids, while those with Peroxy linkages (O-O) are Peroxo Acids.

Ozone

Ozone is important for us. Isn't it? You know how it protects us from the harmful UV rays of the sun. We all use sunscreen with UV protection, don't you? So, ozone gives you a natural sunscreen! In this chapter, we will look at ozone in greater detail. We will have a look at its properties and other uses as well.

What Exactly is Ozone?

Ozone is an allotrope of oxygen. It is highly unstable in nature. We can find its traces for about 20 kilometres above the sea level. So, how is ozone present at such heights? It is formed by the reaction of oxygen with the ultraviolet rays of the sun. The major role of this gas is protecting the earth's surface from the harmful ultraviolet radiations from the sun.

You have heard of the depletion of the ozone layer. You know why is that? We use chlorofluorocarbons in refrigerators and other aerosols also that release harmful things into the air leading to gaps in the layer. Due to this, we get the UV light and it causes a lot of skin problems and malignancy in people.

Apart from this, the various oxides of nitrogen, particularly nitric oxide, react very rapidly with ozone to produce oxygen and nitrogen dioxide. Hence, the various nitrogen oxides of nitrogen that appear from the fumes frameworks of supersonic fly planes lead to depleting the layer.

Preparation

We can prepare ozone by passing a silent electric discharge through dry, unadulterated, and cold oxygen in an extraordinary device. This device is what we know as the ozoniser. In this process, we obtain the gas of up to 10% concentration. This is an endothermic process and we must take care to complete it at high temperatures. That is why we use a silent electric discharge.

If we want to produce higher concentrations of ozone, we can do so by the process of fractional liquefaction of an oxygen and ozone mixture.

Physical Properties

- Ozone is a gas with a light blueish colour.
- It has a fishy smell.
- It condenses at – 120°C to give a dull blue fluid. On further cooling, it hardens to give dark violet crystals.
- Thermodynamically, it is very unstable and disintegrates to oxygen. This is an exothermic process and is catalyzed by numerous materials. However, we must know that high concentration of the gas can be very dangerous.

Oxidizing Action

- Ozone is taken as a very strong oxidizing agent. This is mainly because of the ease with which it gives out atoms of nascent oxygen. It helps in oxidising:

- Lead sulphide to lead sulphate:

$$4O_3 + PbS \rightarrow 4O_2 + PbSO_4$$

- Iodide ions to iodine:

$$2KI + H_2O + O_3 \rightarrow 2KOH + I_2 + O_2$$

- We use this particular reaction in the process of quantitative estimation of the gas. When we bring ozone in contact with potassium iodide solution with a borate buffer (pH 9.2), it liberates iodine. We can titrate this liberated iodine against a standard solution of sodium thiosulphate using starch as an indicator. The reactions that involve in this process are:

$$\underset{\text{Iodide}}{2I^-} + H_2O + \underset{\text{Ozone}}{O_3} \longrightarrow 2OH^- + \underset{\text{Iodine}}{I_2} + O_2$$

$$I_2 + \underset{\text{Sodium thiosulphate}}{2Na_2S_2O_3} \longrightarrow \underset{\text{Sodium tetrathionate}}{Na_2S_4O_6} + 2NaI$$

- It also helps in oxidising nitrogen dioxide to dinitrogen pentoxide.

$$2NO_2 + O_3 \rightarrow N_2O_5 + O_2$$

Solved Example For You

Q: Mention some uses of ozone.

Ans: The common uses of ozone include:

- We use it as an antiseptic and also as a disinfectant.
- We also use it as a delicate dying agent for fading oils, starch etc.

Phosphine

Phosphine (PH_3) is a chemical compound which is prepared by heating phosphorous acid or also by reacting calcium phosphide with water. Phosphine finds its place in the group of organophosphorus compounds with the chemical formula of PH_3. Philippe Gengembre discovered it in 1783. It can be found in human tissues, blood, urine, saliva, etc. Let us look at this chemical in greater detail in this chapter. We will see its physical properties, chemical properties and uses as well. Let's begin.

What is Phosphine?

Phosphine is a chemical that finds its place in the group of organophosphorus compounds. Philippe Gengembre discovered or acquired this chemical in the year 1783. He was the one who acquired phosphine by heating phosphorous in an aqueous solution of potassium carbonate.

Phosphine carries the chemical formula of PH_3. The concentration of this compound constantly alters in our environment. As we already mentioned, this chemical plays an important role in the phosphorous biochemical cycle.

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Preparation of Phosphine

(i) From Phosphide

We can obtain the compound by reacting phosphides with water. For example, calcium phosphide reacts with water to give calcium hydroxide and phosphine.

$$\underset{\text{Calcium Phosphide}}{Ca_3P_2} + 6H_2O \longrightarrow 2PH_3 + 3Ca(OH)_2$$

(ii) From Phosphorus Acid

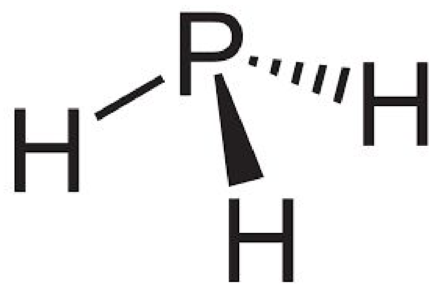
We can also obtain the compound by heating phosphorus acid. It decomposes to give a pure sample of phosphine.

$$\underset{\text{Phosphorous acid}}{4H_3PO_3} \longrightarrow \underset{\text{Phosphoric acid}}{H_3PO_4} + \underset{\text{Phosphine}}{PH_3}$$

Structure of Phosphine

You can see that the electronic configuration of phosphine resembles ammonia. It has the structure of a pyramid. The bond angle H-P-H is = 93°. On the other hand, ammonia has pyramidal geometry with a bond angle of 107.8$_0$. Thus, we see that both these chemicals have a comparative bond angle. Phosphorus is less electronegative than nitrogen.

The cloud of electrons around the central atom, phosphorus is less concentrated as compared to the nitrogen present in ammonia. Therefore, the lone pair of electrons causes significantly more contortion in PH_3. Thus, we note the decline in the bond angle in PH_3 to 93.5°.



Physical and Chemical Properties of Phosphine

- Phosphines are colourless gas.
- It has a characteristic smell like that of a spoiled fish.
- It is an exceedingly noxious gas. PH_3 is sparingly dissolvable in water. However, it can dissolve in natural solvents.
- PH_3 acts as a Lewis base by giving away its lone pair of electrons by reacting with hydrogen iodide.
- Under typical conditions, it is a non-ignitable gas. But, when you warm it, it bursts into flames, forming phosphoric acid.
- It explodes violently when we expose it to oxidising agents.

Effects of PH_3

Phosphine is an extremely dangerous gas. Exposure to even little quantities of the gas can lead to dizziness, loose bowels, cough, cerebral pain, and chest tightness, just to name a few. Upon a greater exposure, you have the risk of suffering from convulsions, coma, damage to the kidney and liver and irregular heartbeat.

Properties of PH_3 Ligands

- Tertiary phosphines or PR_3, are an important class of ligands. This is mainly due to the fact that we can change their electronic and steric properties in a very orderly path over a wide range by just shifting the R group(s).
- These are capable of stabilising a completely wider array of metal complexes that might interest organometallic scientific expert as their phosphine complexes (R_3P) nM−L.
- Phosphines are usually spectator ligands and not performer ligands.
- Like NR_3, phosphines also possess a lone pair on the focal particle which it gives to a metal.
- For alkyl phosphines, the π acidity is very weak.

Solved Example for You

Q: Write down the various uses of phosphine.

Ans: The various uses of phosphine include:

- We use phosphine as a dopant, in the semiconductor industries.
- PH_3 plays an important part in the Holme's signal. This is because it can rapidly ignite. The compartments having calcium carbide and calcium phosphide are penetrated and dropped into the ocean when the gasses develop smoulder. They help in serving as a signal to the Mariners.
- It is also an essential part of smoke screens. Containers having a punctured base and a gap at the top are recorded with calcium phosphide and calcium carbide. These are, then, let out into the ocean. Water enters the containers through the base. It reacts and produces acetylene and phosphine.
- Phosphine gets ignited immediately as it interacts with air. As a further step, it also incites the ignition in acetylene. Due to these reactions, we get a bright and nice red fire. This has a huge cloud of dense smoke due to the smouldering of phosphine. This helps in giving a signal to the approaching boats.
- Phosphine fumigants are a common ingredient in the households to control the bugs, rodents and rabbit invasion in a

huge range of stored grains. Human beings can recognise the scent of PH_3 easily as it is quite strong. However, the rodents and other bugs or insects can't know its presence easily. Thus, it helps in driving or killing them away. We transport PH_3 as compressed liquefied gas. A few solids (phosphides) discharge PH_3 gas.

Phosphorus – Allotropic Forms

Have you ever seen someone look similar to you? Too often, we come across people who look "exactly like" someone you know. However, they are NOT the same person! Such is the case with allotropes. Did you know about this? In this chapter, we will cover the various allotropes of phosphorus and look at their forms and properties. However, let us start with the basics of what allotropes are.

What is an Allotrope?

Allotropy is the science of occurrence of an element in multiple or more than one physical shapes. Allotropes are the various physical types of the same element. A lot of elements exhibit the property of allotropy. These allotropes show different physical properties. However, their chemical properties are comparable.

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- Sulphuric Dioxide

Allotropes of Phosphorus

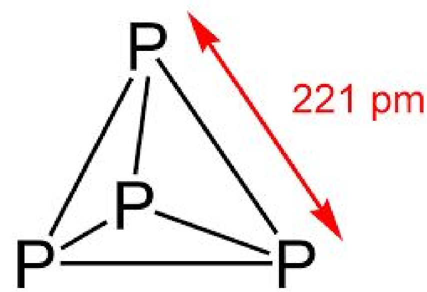
In this chapter, we will look at the allotropes of phosphorus. This element exists in a few allotropic forms. The main allotropes of phosphorus include the white phosphorus, red phosphorus and black phosphorus. In addition to these, there also exists a violet phosphorus. However, that is not a significant allotrope. So, let us start with the various allotropes of phosphorus now.

White Phosphorus

It is a common allotrope of phosphorus. White phosphorus is a waxy and translucent solid. It is very delicate and needs proper handling. It is insoluble in water. However, it dissolves in carbon disulphide or carbon tetrachloride. It breaks down in boiling caustic soda in a latent air and produces sodium hypophosphite and phosphine.

P_4	+	$3\,NaOH + 3H_2O$	$\rightarrow$ PH_3	+	$3NaH_2PO_2$
White		Caustic	Phosphine		Sodium
phosphorus		soda			hydrophosphite

Structure of White Phosphorus



In the above diagram, we see the structure of white phosphorus. As we can see, it has a tetrahedral shape. Every phosphorus particle has a covalent bond with three different atoms of phosphorus. There exist weak Van Der Waals forces of attraction between these particles. We must remember that this element is very reactive and too harmful.

- The melting point is quite low at 44°C.
- As we see, the bond angle in a P_4 particle is 60°. This is comparatively very less as compared to a normal bond angle or a hypothetical bond angle. Therefore, it has a strain in itself. This is why white phosphorus is highly unstable and reactive.

- White phosphorus catches fire suddenly in the air at around 35◦C. As you can notice, this temperature is marginally higher than the normal room temperature. This is the reason why it is kept in water. After combustion, it produces phosphorus pentoxide.

$$P_4 + 5O_2 \rightarrow 2P_2O_5 \text{ or } P_4O_{10}$$

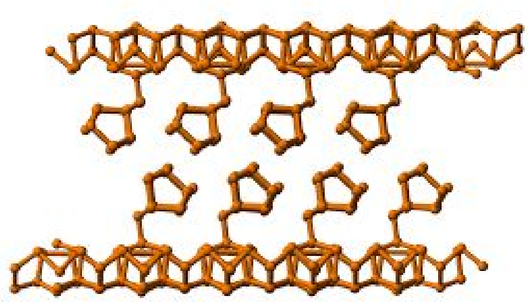
- When it comes in contact with moist air, white phosphorus undergoes an oxidation reaction. This reaction leads to a sparkling discharge of light. As an outcome, it sparkles oblivious.
- White phosphorus displays chemiluminescence.

Red Phosphorus

- We can obtain red phosphorus by heating white phosphorus to around 250◦Celsius within the sight of daylight.

White phosphorous $\xrightarrow[\text{(Inert atmosphere)}]{540\text{ K}}$ Red phosphorus

- Red phosphorus is iron-grey in colour. It is a radiant and bright crystalline solid.
- It is non-poisonous and does not have any odour. Red Phosphorus does not dissolve I water and also in carbon tetrachloride.
- It doesn't break up in boiling caustic soda-like white phosphorus. In fact, it disintegrates in alcoholic potash.
- We can find it in the state of a polymeric solid.



- It is steady under ordinary conditions and doesn't catch fire in the air.
- However, it experiences burning when we warm it to around 400◦C.

- Red phosphorus doesn't show chemiluminescence.

Black Phosphorus

- We can prepare black phosphorus from white phosphorus by heating it to 470K at inert temperature.
- $$\text{White phosphorus} \xrightarrow[\text{(Inert atmosphere)}]{470\text{ K}} \text{Black phosphorus}$$
- Black phosphorus is the most stable allotrope of phosphorus. It has a layered structure. It is a very highly polymerised form of the element.
- We can find black phosphorus in two main forms. They are alpha black phosphorus and beta black phosphorus.
- While beta black phosphorus forms when white phosphorus is heated at 473K, alpha black phosphorus forms when we heat red phosphorus at 803K.
- Beta black phosphorus conducts electricity while alpha black phosphorus doesn't conduct electricity.

Uses of Phosphorus

Phosphorus compounds assume a vital part of life forms. Phosphorus forms a basic constituent in the animal and plant matter.We find it present in blood, bones and the brain of all the animals and also, in living cells. A few of its compounds find applications in industries. The most essential of these chemicals are orthophosphoric acid and phosphatic composts.

Solved Examples For You

Q1: What is yellow phosphorus?

Ans: We use the term yellow phosphorus to denote white phosphorus. It is an allotrope of phosphorus.

Q2: Describe the structure of alpha black phosphorus and beta black phosphorus.

Ans: We find alpha black phosphorus in the shape of opaque monoclinic crystals. We might also see rhombohedral crystals for the same. On the other hand, beta black phosphorus is available as corrugated sheets. This is why they have a structure consisting of flaky and layered crystals. The difference in their structure is also a reason why beta black phosphorus can conduct electricity while the other can not.

Phosphorus Halides

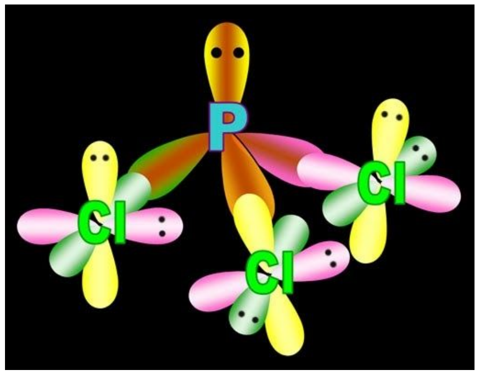
You probably know that chlorine finds its use in your swimming pools. However, are you aware that a phosphorus halide is also used as a cleansing and bleaching agent? First of all, do you know what a phosphorus halide is? Well, in this chapter, let us look at the halides of phosphorus and their properties. In the end, we will also look at its uses.

What is a Phosphorus Halide?

A phosphorus halide is a compound that a phosphorus forms with a halogen. A phosphorus halide is of two types. They are PX_3 and PX_5. Here, we refer X to a halogen. It could be anything from fluorine, chlorine, bromine or iodine. However, the most common phosphorus halide is that of chloride. These chlorides are usually covalent in their nature.

1) Phosphorus Trichloride

This is an oily and sleek fluid. It is very lethal in nature. The shape of this compound is that of a triangular pyramid. The atom of phosphorus exhibits sp^3 hybridization.



As we can see in the above diagram, phosphorus has its SP^3 orbitals. It has only one electron and it gives that electron to a p orbital electron from 3 chlorine atoms. The fourth sp^3 orbital is full. It is a solitary lone pair. Thus, it cannot form a bond. However, it repels alternate bonds. This creates a state of the shape of trigonal pyramidal.

Preparation

- We obtain phosphorus trichloride by passing dry chlorine overheated white phosphorus. The reaction that takes place is as follows:

$$P_4 + 6Cl_2 \rightarrow 4PCl_3$$

- We can also obtain this compound by the reaction of thionyl chloride with white phosphorus. Below is the reaction to it.

$$P + 8SOCl_2 \rightarrow 4PCl_3 + 4SO_2 + 2S_2Cl_2$$

Chemical Properties

- Phosphorus trichloride hydrolyses when we dampen it.

$$PCl_3 + 3H_2O \rightarrow H_3PO_3 + 3HCl$$

- It reacts with natural compounds having a –OH group and gives their 'chloro' subsidiaries as products.

$$3C_2H_5OH + PCl_3 \rightarrow 3C_2H_5Cl + H_3PO_3$$

Structure of PCl_3

The phosphorus particle in the centre of PCl_3 exhibits sp^3 hybridisation. It has 3 bond sets and 1 lone pair of electrons. Due to this reason, it has a pyramidal shape. It acts as a Lewis base because it has the capacity to donate its lone pair of electrons to other electron-lacking particles or atoms.

2) Phosphorus Pentachloride

It is yellowish-white in colour. Phosphorus Pentachloride is a very water delicate solid. It dissolves in organic solvents like carbon tetrachloride, benzene, carbon disulphide, diethyl ether.

Its structure is that of a trigonal bi-pyramid. We find this structure primarily in vaporous and fluid stages. In the solid state, we can find it as an ionic solid, $[PCl_4]^+[PCl_6]^-$. Here, the cation, $[PCl_4]^+$ is tetrahedral and the anion, $[PCl_6]^-$ is octahedral.

We must know that the molecule has three tropical P-Cl bonds and two pivotal P-Cl bonds. Due to the more prominent repulsion at hub positions as compared to the central positions, we see that the two axial bonds are longer than tropical bonds.

Preparation

- We can produce pentachloride by the reaction with an excess of dry chlorine.

$$P_4 + 10Cl_2 \rightarrow 4PCl_5$$

- We can also produce it by the reaction of SO_2Cl_2 and phosphorus.

$$P_4 + 10SO_2Cl_2 \rightarrow 4PCl_5 + 10SO_2$$

Chemical Properties

- In the presence of damp air, phosphorus pentachloride hydrolyses to $POCl_3$. This compound changes over to phosphoric acid over a period of time.

$$PCl_5 + H_2O \rightarrow POCl_3 + 2HCl$$

$$POCl_3 + 3H_2O \rightarrow H_3PO_4 + 3HCl$$

- When we heat it, it sublimes and further disintegrates to phosphorus trichloride.

$$PCl_5 \rightarrow PCl_3 + Cl_2$$

- It reacts with finely partitioned metals under the influence of heat to create metal chlorides.

$$2Ag + PCl_5 \rightarrow 2AgCl + PCl_3$$

- It reacts with natural compounds containing –OH group and produces their 'chloro' subordinates.

$$C_2H_5OH + PCl_5 \rightarrow C_2H_5Cl + POCl_3 + HCl$$

Structure of PCl_5

As we discussed earlier, the central phosphorus atom in phosphorus pentachloride experiences SP^3d hybridisation. All the five electrons combine in these hybrid orbitals as bond sets. The molecular shape of the particle is trigonal bipyramidal.

After the five electrons get hybridised, we get five electrons of equivalent size and shape. Three of them frame a triangle (120° partition) in the centre. One bond is above and one is underneath those three.

However, you must remember that we notice trigonal bipyramidal geometry in phosphorus pentachloride only in its fluid and vaporous state. In its solid state, it exists as a salt.

Solved Example for You

Q: Mention some of the common uses of phosphorus halides.

Ans: The various uses of Phosphorus Halides are as follows:

- We use phosphorus halides as a chlorinating agent. Specifically, it is used in cleaning water bodies.
- Also, we utilise them for making water treatment agents. Too often, we use them to make organophosphorus pesticides.

- They are also an important component in the lube oil and paint added substances. They act as an intermediate in these substances.
- We use phosphorus halides in the preparation of phosphorus acid. We also use them to prepare chloroanhydrides and phosphoric acid subsidiaries as an intermediate.

Simple Oxides

Did you ever wonder an oxide could be simple and complex? Well, as students of chemistry, we have to look into the depths of everything. So, while studying oxides, we have to look at simple oxide. In this chapter, we will read all about simple oxides, their types, and properties. However, do you first know what an oxide is?

What is an Oxide?

An oxide is a binary compound that we obtain upon the reaction of oxygen with other elements. Depending on the oxygen content, we can extensively arrange them into mixed and simple oxides. An oxide of a nonmetal generally has a tendency to be acidic.

On the other hand, an oxide a of metal shows a basic tendency. The oxides of elements in or close to the corner band of semimetals are by and large amphoteric. Let us now look at the different types of oxides.

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Simple Oxides

A simple oxide is one carrying a number of oxygen atoms that the normal valency of its metal allows. Examples of this include H_2O, MgO, and Al_2O_3.

Mixed Oxides

We get a mixed oxide upon the combination of two simple oxides. Examples of these include: Lead dioxide (PbO_2) and lead monoxide (PbO) combine to form the mixed oxide Red lead (Pb_3O_4). In another example, we can see that Ferric oxide (Fe_2O_3) and ferrous oxide (FeO) combine and form the mixed oxide Ferro-ferric oxide (Fe_3O_4).

Classification of Simple Oxides

On the basis of their chemical behaviour, there are acidic, basic, amphoteric and neutral oxides.

1) Acidic Oxide

An acidic oxide reacts with water and produces an acid. Usually, it is the oxide of non-metals. Examples include SO_2, CO_2, SO_3, Cl_2O_7, P_2O_5, and N_2O_5. It could also be the oxide of metals with high oxidation states, such as CrO_3, Mn_2O_7, and V_2O_5.

- Sulphur dioxide reacts with water and gives sulphurous acid.

$$SO_2 + H2O \rightarrow H_2SO_3$$

- Chromic anhydride reacts with water and results in chromic acid.

$$Cr_2O_3 + H_2O \rightarrow H_2Cr_2O_4$$

2) Basic Oxide

A basic oxide reacts with water to give a base. Examples include the oxide of most metals, such as Na_2O, CaO, BaO. These are basic in nature.

- Calcium oxide reacts with water and produces calcium hydroxide, a base.

$$CaO + H_2O \rightarrow Ca(OH)_2$$

3) Amphoteric Oxide

An amphoteric oxide is that metallic oxide displaying a dual behaviour. It shows the characteristics of both an acid as well as a base. It reacts with both alkalis as well as acids.

- For example, zinc oxide acts as an acidic oxide when it reacts with concentrated sodium hydroxide. However, it acts as a basic oxide while reacting with hydrochloric acid.

$$ZnO + 2H_2O + 2NaOH \rightarrow Na_3Zn[OH]_4 + H_2$$

$$ZnO + 2HCl \rightarrow ZnCl_2 + H_2O$$

- Aluminium oxide is another example that reacts with alkalis as well as acids.

$$Al_2O_3(s) + 6NaOH(aq) + 3H_2O(l) \rightarrow 2Na_3[Al(OH)_6](aq)$$

$$Al_2O_3(s) + 6HCl(aq) + 9H_2O(l) \rightarrow 2[Al(H_2O)_6]^{3+}(aq) + 6Cl^{-}(aq)$$

Solved Example for You

Q: What is a neutral oxide?

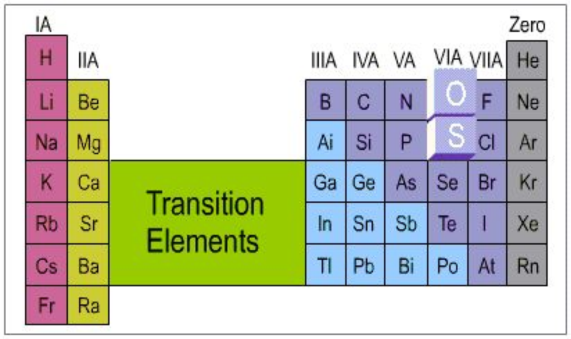
Ans: A neutral oxide is one that does not exhibit any tendency to form salts either with acids or bases. The examples of neutral oxides include nitrous oxide and carbon monoxide.

Sulphur – Allotropic Forms

You must have noticed that a lot of ointments contain sulphur. However, are you aware “which” sulphur it is talking about? Did you know that there are more than one type of sulphur, even if you don’t see it on the periodic table? They are the allotropic forms of sulphur. So, what are these allotropic forms exactly? Let us read about their types, properties and uses. However, before we begin, let’s take a quick recap of the properties of sulphur.

What is Sulphur?

Sulphur is a chemical element having atomic number 16. It is easily accessible at room temperature. It is basically a splendid yellow crystalline solid. Sulphur is a non-metal, as we obviously know! The position of Sulphur in the periodic table is as follows:



Properties of Sulphur

The following are the most common properties of Sulphur. We will look at its chemical and physical properties separately.

Physical Properties

- Sulphur is yellow in colour.
- It is insoluble in water. However, it is very soluble in toluene (methylbenzene) and carbon disulphide.
- It is a non-metal and therefore, a poor conductor of electricity and heat.
- At the point, when we consolidate Sulphur vapour, we get a fine powder, which shapes a pattern resembling a flower. This is the 'Flower of Sulphur'.

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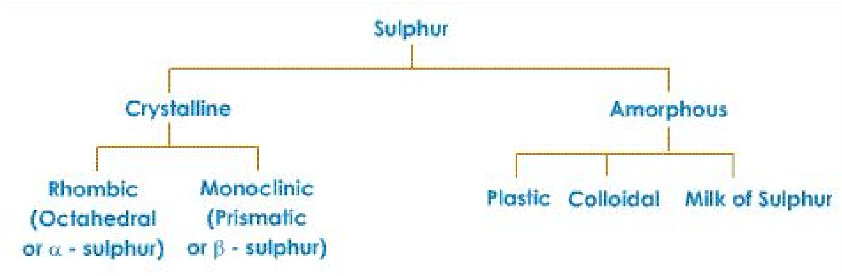
Chemical Properties

- Most metals and non-metals react with Sulphur, under specific conditions.
- Sulphur burns in excess of air with a bright blue fire and forms Sulphur (IV) oxide and a little amount of Sulphur (VI) oxide.
- It reacts with Hydrogen at high temperature and forms hydrogen sulphide.
- Sulphur vapour reacts with hot coke to produce a fluid, carbon disulphide.

Allotropic Forms of Sulphur



We can find sulphur in a number of structures in the same physical state. However, the most important crystalline structures are rhombic or octahedral (α – sulphur), and monoclinic sulfur (β – sulphur). We find the shape of rhombic sulphur at a temperature beneath $96^{o}C$. On the other hand, monoclinic sulphur occurs at a temperature over $96^{o}C$.

This temperature of $96^{o}C$ is the transitional temperature between the two crystalline structures. There is another allotrope of sulphur, polymeric sulfur (S_8). It is an eight-part ring particle. This is insoluble in organic media, synthetic and natural rubber. It also does not dissolve in carbon disulphide.

Let us now look at the properties of the two main allotropes of sulphur: rhombic and monoclinic sulphur.

1) Rhombic Sulphur

- We find them as yellow and translucent crystals.

- Rhombic Sulphur has a melting point of 114^{o} C.
- The density of rhombic Sulphur is 2.08 g/cm^3
- It is stable at temperatures below 96^{o}C.

2) Monoclinic Sulphur

- These are transparent and amber crystals.
- They have a melting point of 119^{o}C.
- The density of monoclinic sulphur is 1.98 gcm^3
- It is unstable at temperatures below 96^{o}C and changes into rhombic form.
- We must remember that at a temperature of 96^{o}C or above, rhombic sulphur changes to kaleidoscopic or prismatic sulphur. At 96^{o}C or beneath, kaleidoscopic or prismatic sulphur changes to rhombic sulphur.
- These allotropes that alter their configuration from one form to another by a change in the temperature are Enantiotropic Allotropes.

3) Colloidal Sulphur

- We can produce this sulphur by passing hydrogen sulphide through a saturated and cooled solution of sulphur dioxide in water. Another method is by including a solution of alcohol and sulphur in the water.
- It acts as a solvent in carbon disulfide.
- We utilise it as a part of medicines.

4) Milk of Sulphur

- We can produce this by the action of weak hydrochloric acid on ammonium sulphide. In a similar fashion, this milk of sulphur forms by the boiling of sulphur with calcium hydroxide (aqueous solution). We filter this mixture and add weak hydrochloric acid to get the milk of sulphur.
- This compound is non-crystalline and white in colour.
- It is soluble in carbon disulphide.
- At the point when we heat it, it changes to the conventional yellow colour of sulphur that we use as a part of medicines.

Solved Example for You

Q: Give some important uses of sulphur.

Ans: The most common uses of sulphur are:

- We use sulphur to develop specific sorts of fungus in vines.
- Sulphur is a common ingredient in the production of tetraoxosulphate(VI) acid. We can say, this is the most important use of sulphur.
- We use sulphur in the making of calcium hydrogen trioxosulphate(IV), $Ca(HSO_3)_2$. Here, this compound finds its use as a bleacher of wood pulp in the paper manufacturing industry.
- Sulphur is a common and important ingredient in the vulcanization of rubber. This process involves making the rubber tough and hard by binding the rubber molecules close to each other.
- We use sulphur in dye manufacturing.
- Sulphur is common in the fabrication of sulphur compounds, for example, carbon disulfide, CS_2 and sulfur monochloride, S_2Cl_2.
- It finds its significant use in ointments.

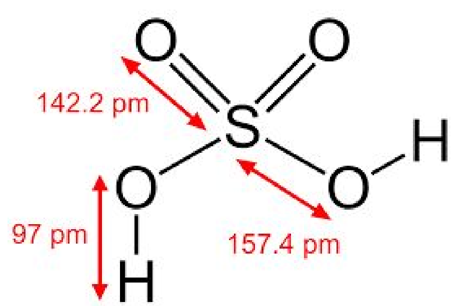
- Sulphur is an important ingredient in sulphides like phosphorus sulphide. We use this as a part of making firecrackers, gunpowder and matches.

Sulphuric Acid

You surely must have come across a lot of experiments with sulphuric acid. Haven't you? But, have you ever put it directly in water? NO! Never ever do that! Why? Let's find out. In this chapter, we will read all about sulphuric acid and its properties. In the end, we will also look at its uses. Let us first start with what it is.

What is Sulphuric Acid?

Sulphuric acid or as we write it, H_2SO_4, is an odourless and colourless, oily liquid that is very corrosive. People named it Oil of Vitriol. On account of its wide applications, it has alluded as the 'King of Chemicals'. We can find it in both combined and free state.



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Manufacturing Process

How do you think we can manufacture sulphuric acid? Well, we can commercially produce this by two techniques. They are:

- Lead chamber process
- Contact process

Let us now look at these processes closely.

1) Contact Process

The contact process has three major steps:

- Step – I: Production of Sulphur Dioxide

We can make this by heating sulphur or sulphide ores, for example, iron pyrites in excess of air.

S (Sulphur) + O_2(Oxygen) + Δ(Heating) → SO_2(Sulphur dioxide)

4FeS(Iron pyrites) + $7O_2$(Oxygen) + Δ(heating) → $2Fe_2O_3$(Ferric Oxide) + $4SO_2$(Sulphur dioxide)

- Step -II: Formation of Sulphur Trioxide

We can oxidize Sulphur dioxide to Sulphur trioxide with atmospheric oxygen by using V_2O_5 as a catalyst.

$2SO_2$(Sulphur dioxide) + O_2(Oxygen) + V_2O_5(Catalyst) → SO_3(Sulphur trioxide)

- Step -III: Conversion of Sulphur Trioxide into Sulphuric Acid

We break down the sulphur trioxide from the above step in 98% sulphuric acid to give pyrosulphuric acid or oleum. We, then, dilute the Oleum with water to give sulphuric acid of the desired concentration.

$$SO_3\text{(Sulphur trioxide)} + H_2SO_4\text{(Sulphuric acid-98\%)} \rightarrow H_2S_2O_7\text{(Pyrosulphuric acid/Oleum)}$$

$$H_2S_2O_7\text{(Pyrosulphuric acid/Oleum)} + H_2O\text{(Dilution)} \rightarrow 2H_2SO_4\text{(Sulphuric acid)}$$

2) Lead Chamber Process

Lead Chamber process is one of the most common manufacturing strategies that results in around 50-60 B grade acids. In this process, we use the wet SO_2 (Sulphur Dioxide) in the presence of nitrogenous oxides (dynamic impetus). It gets oxidised and forms sulphur trioxide with the oxygen exhibit in the air. We then react Sulphur trioxide with water to get H_2SO_4. The reactions are:

$$2SO_2 + O_2 \rightarrow 2SO_3$$

$$SO_3 + H_2O \rightarrow H_2SO_4$$

Physical Properties

- Sulphuric acid is a thick, colourless and an oily fluid.
- It has a specific gravity of 1.84 at 298 K.
- The boiling point of the acid is 611 K. It attributes its higher boiling point and thickness to hydrogen bonding.
- The strong acid reacts with water vigorously releasing quite a lot of heat. Therefore, you must never add water to sulphuric acid. Instead, you should add the acid to water, slowly with proper stirring.

Chemical Properties

- Sulphuric acid is a strong dibasic acid. It is diprotic and ionises in two stages in the aqueous solution.
- It is highly corrosive and reactive and is soluble in water. Sulphuric acid has a very high oxidising power and thus, acts as a strong oxidising agent. It has very low volatility.
- We use this acid as a part of the preparation of more volatile acids from their comparing salts because of its low volatility.
- Concentrated sulphuric acid is a very strong dehydrating agent. This property is utilized as a part of drying many wet gases which do not react with the acid.

- It additionally expels water from natural mixes like starches.
- As it is a good oxidising agent, it can oxidise both non-metals as well as metals. It itself reduces to sulphur dioxide.

Some Common Reactions

- Hot concentrated sulphuric acid oxidizes copper to copper sulphate.

$$Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + H_2O$$

- Concentrated sulphuric acid gives out hydrogen chloride from sodium chloride and hydrogen fluoride from calcium fluoride.

$$CaF_2 + H_2SO_4 \rightarrow CaSO_4 + 2HF$$

- It burns glucose, sugar, and starch to carbon.

$$C_{12}H_{22}O_{11} + (H_2SO_4) \rightarrow 12C + 11H_2O$$

Solved Example for You

Q: Write down the main uses of sulphuric acid.

Ans: The uses of sulphuric acid are:

- It is a common ingredient in the preparation of fertilisers like ammonium sulphate and superphosphate
- We use it in the manufacture of dyes, shades, and paints.
- It is a common ingredient in the manufacture of explosives, for example, TNT.
- Other imperative chemicals like hydrochloric acid, phosphoric acid, nitric acid, and sodium carbonate need the presence of sulphuric acid. Without sulphuric acid, we cannot obtain the chemicals.
- We utilise it as a part of the refining of petroleum.
- It acts as a pickling agent.
- It is common as a laboratory agent, an oxidizing and dehydrating agent.

Sulphur Dioxide

What happens when sulphur and oxygen, which are in the same group of the periodic table, react? They form sulphur dioxide. But how is this possible? How is this compound formed? What are its properties and uses? Do you want to find out? Then keep reading the article further.

What is Sulphur Dioxide?

Sulphur dioxide is a common gas that has a pungent smell. We will look at the significance and methods of preparation of this gas in this chapter.

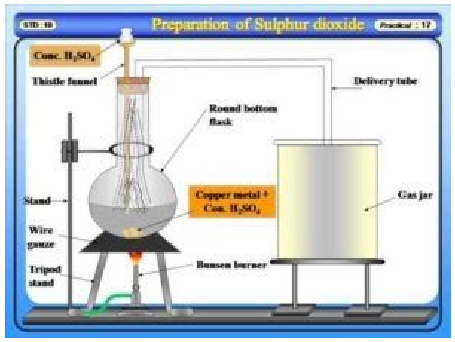
Methods of Preparation of Sulphur Dioxide Gas

We can create sulphur dioxide in the laboratory by the action of dilute sulphuric acid on sulphites

$$Na_2SO_3 + H_2SO_4 \rightarrow Na_2SO_4 + H_2O + SO_2\uparrow$$

Sodium sulphite   Sulphuric acid   Sodium sulphate   water

Sulphur dioxide



(Source: YouTube)

Commercially, chemists produce vast volumes of sulphur dioxide by heating a sulphide ore, for example, let's say iron sulphide. They, then, liquefy this gas subsequent to drying under 25 atm pressure. It is usually stored in steel barrels. The process of roasting involves the following reaction:

$$Fe_2S_3 + 4O_2 \rightarrow FeO + 3SO_2$$

Liquefaction at 25 atm gives takes place as given below:

$$So_2(g) \rightarrow SO_2(l)$$

We can obtain Sulphur dioxide likewise, on an extensive scale by blazing sulphur in the air.

$$S + O_2 \rightarrow SO_2$$

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Physical Properties of Sulphur Dioxide

- Sulphur dioxide is a dull gas.
- It has a very pungent smell. Its odour resembles smoulder sulphur.
- It is one of the most straightforward gases to melt. This is because it consolidates at room temperature under a pressure of 2 atm.

Chemical Properties of Sulphur Dioxide

- It is an acidic oxide.
- It is readily dissolvable in water.
- Sulphur dioxide breaks up in water and gives out sulphurous acid.

$$SO_2 + H_2O \rightarrow H_2SO_3$$

Sulphur dioxide   Water   Sulphurous acid

- It reacts vigorously with sodium hydroxide solution and forms sodium sulphite.

$$SO_2 + 2NaOH \rightarrow Na_2SO_3 + H_2O$$

- In the cases that we pass more sulphur dioxide into this arrangement, we get sodium hydrogen sulphite.

$$SO_2 + Na_2SO_3 \rightarrow H_2O + 2NaHSO_3$$

- The sulphur particle in a sulphur dioxide atom is tetravalent. Subsequently, it can improve it's covalency to six by specifically reacting with elements like O_2 and Cl_2 to shape the comparing addition compounds. For example, it reacts with chlorine under the influence of charcoal as a catalyst to give sulphuryl chloride (SO_2Cl_2).

$$SO_2 + Cl_2 \rightarrow SO_2Cl_2$$

- Within the sight of vanadium pentoxide(V_2O_5) as an impetus, it gives sulphur trioxide.

$$2SO_2 + O_2 \rightarrow SO_3$$

- In the presence of moisture, it can start giving nascent oxygen, and, along these lines, go about as a reducing agent. For example, it reduces ferric salts to ferrous salt, and halogens to halogen acids.

$$2Fe^{3+} + SO_2 + 2H_2O \rightarrow 2Fe^{2+} + SO_4^{2-} + 4H^+$$

$$X_2 + SO_2 + 2H_2O \rightarrow SO_4^{2-} + 2X^- + 4H^+$$

Identifying Tests for Sulphur Dioxide Gas

How do we test for the presence of sulphur dioxide gas? Well, it's easy! We can perform some simple steps to identify this gas. We will talk about them below.

1) It decolorizes acidified $KMnO_4$ solution

It reacts with potassium permanganate to give potassium sulphate, manganese sulphate and sulphuric acid.

$$5SO_2 + 2KMnO_4 + 2H_2O \rightarrow K_2SO_4 + 2MnSO_4 + 2H_2SO_4$$

2) It turns a filter paper moistened with acidified $K_2Cr_2O_7$ solution green

It reacts with potassium dichromate and sulphuric acid to give potassium sulphate and chromium sulphate.

$$3SO_2 + K_2Cr_2O_7 + H_2SO_4 \rightarrow K_2SO_4 + 2Cr_2(SO_4)_3 + H_2O$$

3) It turns starch iodate paper blue

It reacts with potassium iodate to give potassium hydrogen sulphate and iodine and sulphuric acid.

$$5SO_2 + 2KIO_3 + 4H_2O \rightarrow 2KHSO_4 + 3H_2SO_4 + I_2$$

Structure of Sulphur Dioxide

It has an angular shape with an O-S-O bond edge of 119.50. We must note the fact that sulphur dioxide has two unique sorts of pi bonds, i.e. p pi-p pi and d pi-p pi. The two sulphur-oxygen bond lengths are similar or equivalent. This signifies that sulphur dioxide is a resonance hybrid of two canonical structures.

Solved Example for You

Q: Mention some common uses of sulphur dioxide gas.

Ans: The different uses of sulphur dioxide include:

- We use it in the assembling of sulphites, sulphuric acid, and hydrogen sulphite.
- It is a common ingredient in the sugar industry. It finds its use in the refining and decolourising of sugar.
- We use it to refine lamp oil and other petroleum items.
- It is a common disinfectant and a popular fumigant.
- We use it for dying fragile articles.
- As an antichlor, we use it to expel the overabundance chlorine from substances those have been faded by chlorine.
- It is a glue solvent.
- We use it as a refrigerant in refrigerators.
- It is used as an additive for wines, meat, dry natural products etc.