

Oxygen Family.

Oxygen is the first member of group 16 or VIA of the periodic table. It consists of five elements Oxygen (O), sulphur (S), selenium (Se), tellurium (Te) and polonium (Po). These (except polonium) are the ore forming elements and thus called **chalcogens**.

Electronic configuration

Elements	Electronic configuration ($ns^2 np^4$)
${}_8\text{O}$	$1s^2, 2s^2 2p^4$ or $[\text{He}]2s^2 2p^4$
${}_{16}\text{S}$	$1s^2, 2s^2 2p^6, 3s^2 3p^4$ or $[\text{Ne}]3s^2 3p^4$
${}_{34}\text{Se}$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^4$ or $[\text{Ar}]3d^{10} 4s^2 4p^4$
${}_{52}\text{Te}$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6 4d^{10}, 5s^2 5p^4$ or $[\text{Kr}]4d^{10} 5s^2 5p^4$
${}_{84}\text{Po}$	$1s^2, 2s^2 2p^6, 3s^2 3p^6 3d^{10}, 4s^2 4p^6 4d^{10} 4f^{14}, 5s^2 5p^6 5d^{10}, 6s^2 6p^4$ or $[\text{Xe}]4f^{14} 5d^{10} 6s^2 6p^4$

Physical properties

(1) **Physical state:** Oxygen is gas while all other are solids.

(2) **Atomic radii:** Down the group atomic radii increases because the increases in the number of inner shells overweighs the increase in nuclear charge.

(3) **Ionization energy:** Down the group the ionization energy decrease due to increase in their atomic radii and shielding effect.

(4) **Electronegativity:** Down the group electronegativity decreases due to increase in atomic size.

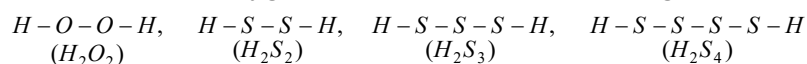
(5) **Electron affinity:** Element of this group have high electron affinity electron affinity decreases down the group.

(6) **Non – metallic and metallic character:** These have very little metallic character because of their higher ionization energies.

(7) **Nature of bonding:** Compound of oxygen with nonmetals are predominantly covalent. S, Se, and Te because of low electronegativities show more covalent character.

(8) **Melting and boiling points:** The m. pt. and b. pt increases on moving down the group.

(9) **Catenation:** Oxygen has some but sulphur has greater tendency for catenation.



(10) **Allotropy**

Oxygen	-	O_2 and O_3
Sulphur	-	Rhombic, monoclinic, plastic sulphur
Selenium	-	Red (non-metallic) grey (metallic)
Tellurium	-	Non-metallic and metallic (more stable)
Polonium	-	α and β (both metallic)

(11) **Oxidation states:** Oxygen \rightarrow -2 and -1 oxidation state. These element shows +2, +4 and +6 oxidation state.

Chemical properties

(1) **Hydrides:** H_2O, H_2S, H_2Se, H_2Te and H_2Po, H_2O - colorless and odorless.

H_2S, H_2Se, H_2Te and H_2Po - colorless, unpleasant smell.

Increasing order of reducing power of hydrides: $H_2O < H_2S < H_2Se < H_2Te$

Increasing order of bond angles in hydrides: $H_2Te < H_2Se < H_2S < H_2O$

The order of stability of hydrides: $H_2O > H_2S > H_2Se > H_2Te$

The order of increasing acidic nature of hydrides: $H_2O < H_2S < H_2Se < H_2Te$

(2) **Oxides:** These elements form monoxides (MO), dioxides (MO_2) increasing order of acidic nature of oxides is $TeO_3 < SeO_3 < SO_3$.

(3) **Oxyacids:** $H_2SO_3, H_2SO_4, H_2S_2O_3, H_2SO_5, H_2S_2O_8, H_2S_2O_7, H_2S_2O_6$

(4) **Halides:** Oxygen : OF_2, Cl_2O, Br_2O

Sulphur: $S_2F_2, S_2Cl_2, SF_2, SCl_2, SBr_2, SF_4, SCl_4$ and SF_6

Selenium and tellurium: SeF_6 and TeF_6

Oxygen and its compounds

Oxygen is the most abundant element in the earth crust (46.5%). It was discovered by **Karl Scheele** and **Joseph Priestley**. It occurs in three isotopic forms:

${}_8O^{16}$
(Abundance : 99.76%)

${}_8O^{17}$
(Abundance : 0.037%)

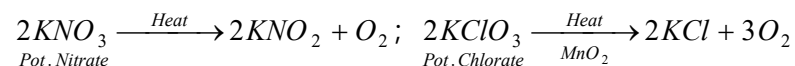
${}_8O^{18}$
(Abundance : 0.204%)

Out of the three isotopes, ${}_8O^{18}$ is radioactive.

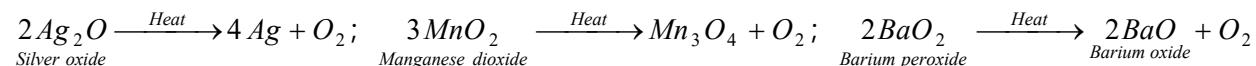
(1) **Occurrence:** In free state, it occurs in air and constitutes 21% by volume of air.

(2) **Preparation of Dioxygen:** Oxygen is prepared by the following methods.

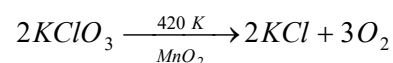
(i) **By the decomposition of oxygen rich compounds:** e.g.



(ii) **By heating dioxides, Peroxides and higher oxides:** e.g.

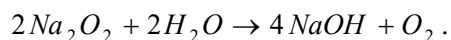


(iii) **Laboratory Method:** In the laboratory, O_2 is prepared by thermal decomposition of potassium chlorate.



Note: In the absence of MnO_2 catalyst, the decomposition takes place at 670-720 K. Therefore, MnO_2 acts as a catalyst and also lowers the temperature for the decomposition of $KClO_3$.

(iv) O_2 can also be prepared by the action of water on sodium peroxide as,



(v) **Industrial preparation:** The main sources for the industrial preparation of dioxygen are air and water.

(a) From air: O_2 is prepared by fractional distillation of air. During this process, N_2 with less boiling point (78 K) distills as vapour while O_2 with higher boiling point (90 K) remains in the liquid state and can be separated.

(b) From water: O_2 can also be obtained by the electrolysis of water containing a small amount of acid or alkali,

$$2H_2O \xrightarrow{\text{Electrolysis}} 2H_2(g) + O_2(g).$$

(3) **Physical properties of O_2 :** It is a colorless, tasteless and odorless gas. It is slightly soluble in water and its solubility is about 30 cm^3 per litre of water at 298 K. Physical properties of atomic and molecular oxygen

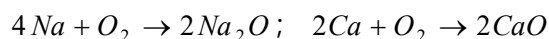
Atomic properties	Molecular properties
Atomic radius (pm) – 73	Bond length (pm) – 120.7
Ionic radius O^{2-} (pm) – 140	Bond energy (kJ mol^{-1}) – 493
Electronegativity – 3.5	Density at S.T.P. (g cm^{-3}) – 1.429
Ionization energy (kJ mol^{-1}) – 1310	Melting point (K) – 54.4
Electron affinity (kJ mol^{-1}) – 140	Boiling point (K) – 90.2

(4) **Chemical properties of O_2 :** It does not burn itself but helps in burning. It is quite stable in nature and its bond dissociation energy is very high. Therefore, it is not very reactive as such,
 $O_2 \rightarrow O + O.$

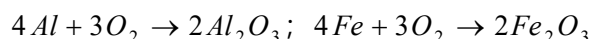
Therefore, dioxygen reacts at higher temperatures. However, once the reaction starts, it proceeds of its own. This is because the chemical reactions of dioxygen are exothermic and the heat produced during the reaction is sufficient to sustain the reactions.

(i) **Action with litmus:** Like dihydrogen, it is also neutral and has no action on blue or red litmus.

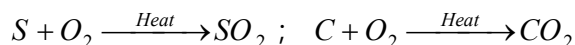
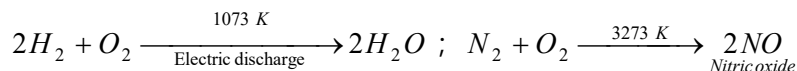
(ii) **Reaction with metals:** Active metals like Na, Ca react at room temp. to form their respective oxides.



It reacts with Fe, Al, Cu etc. metals at high temperature

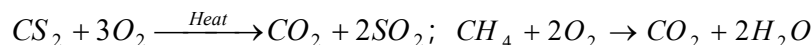
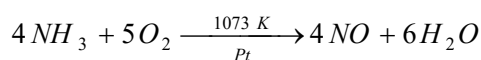


(iii) **Action with Non-metals:** It form oxides.



(iv) **Reaction with compounds:** Dioxygen is an oxidizing agent and it oxidizes many

compounds under specific conditions. e.g. $4HCl + O_2 \xrightarrow[CuCl_2]{700 K} 2H_2O + 2Cl_2;$



(5) Uses of dioxygen

(i) It is used in the oxy-hydrogen or oxy-acetylene torches which are used for welding and cutting of metals.

(ii) It is used as an oxidizing and bleaching agent,

(iii) Liquid O_2 is used as rocket fuel.

(iv) It is used in metallurgical processes to remove the impurities of metals by oxidation.

(6) Compounds of Oxygen

Oxides: A binary compound of oxygen with another element is called oxide. On the basis of acid-base characteristics, the oxides may be classified into the following four types,

(a) **Basic oxides:** Alkali, alkaline earth and transition metals form basic oxides - Na_2O, MgO, Fe_2O_3 etc. their relative basic character decreases in the order: alkali metal oxides > alkaline earth metal oxides > transition metal oxides.

(b) **Acidic oxides:** Non-metal oxides are generally acidic - $CO_2, SO_2, SO_3, NO_2, N_2O_5, P_4O_{10}, Cl_2O_7$ etc.

(c) **Amphoteric oxides:** Al_2O_3, SnO_2 etc.

(d) **Neutral oxides:** H_2O, CO, N_2O, NO etc.

Trends of oxides in the periodic Table : On moving from left to the right in periodic table, the nature of the oxides change from basic to amphoteric and then to acidic. For example, the oxides of third period has the following behavior,


Na_2O strongly basic	MgO basic	Al_2O_3 amphoteric	SiO_2 weakly acidic	P_4O_{10} acidic	SO_2 strongly acidic	Cl_2O_7 very strongly acidic
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Basic to acidic character increase



However, on moving down a group, acidic character of the oxides decreases. For example in the third group, the acidic character of oxides decreases as:

B_2O_3 acidic	Al_2O_3 amphoteric	Ga_2O_3 (weakly basic)	In_2O_3, Tl_2O_3 basic
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On the basis of oxygen content the oxides may be classified into the following types,

Normal oxides: These contain oxygen atoms according to the normal oxidation number i.e. -2 . For example, $MgO, H_2O, CaO, Li_2O, Al_2O_3$ etc.

Polyoxides: These contain oxygen's atoms more than permitted by the normal valency. Therefore, these contain oxygen atoms in oxidation state different than -2 .

Peroxides: These contains O_2^{2-} ion having oxidation number of oxygen as -1 . For example, $H_2O_2, Na_2O_2, BaO_2, PbO_2$ etc.

Superoxides: These contains O_2^- ion having oxidation number of oxygen as $-1/2$. For example, KO_2, PbO_2 , etc.

Suboxides: These oxides contain less oxygen than expected from the normal valency. For example, N_2O .

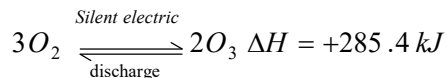
Mixed oxides: These oxides are made up of two simple oxides. For example, red lead $Pb_3O_4 (2PbO_2 + PbO_2)$, magnetic oxide of iron, $Fe_3O_4 (FeO + Fe_2O_3)$ and mixed oxide of manganese, $Mn_3O_4 (MnO_2 + 2MnO)$.

Ozone or trioxygen

Ozone is an allotrope of oxygen. It is present in the upper atmosphere, where it is formed by the action of U. V. radiations on O_2 , $3O_2 \xrightarrow[\text{Ozone}]{\text{U.V. radiation}} 2O_3$.

O_3 Protects us from the harmful U. V. radiations which causes skin cancer. Now a days, ozone layer in the stratosphere is depleting due to NO released by supersonic aircrafts and Chloro fluoro carbons (CFC'S) i.e. freon which is increasingly being used in aerosols and as a refrigerant.

Preparation: Ozone is prepared by passing silent electric discharge through pure, cold and dry oxygen in a specially designed apparatus called ozonize. The formation of ozone from oxygen is an endothermic reaction.



Ozone is prepared in the laboratory by the following two types of ozonizers,

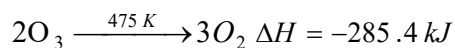
- (a) Siemen's ozoniser
- (b) Brodie's ozoniser

For the better yield of ozone: (a) Only pure and dry oxygen should be used. (b) The ozoniser must be perfectly dry. (c) A fairly low temperature ($\approx 273 \text{ K}$) must be maintained. (d) The electric discharge must be sparkless.

Physical properties: Ozone is a light blue colored gas, having pungent odour. It is heavier than air. Its vapour density is 24. It is slightly soluble in water.

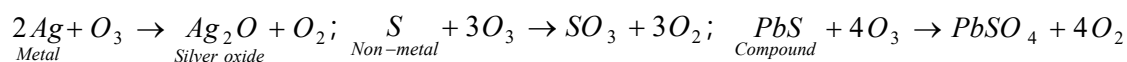
Chemical properties: The important chemical properties of ozone are discussed below,

- (1) **Decomposition:** Pure ozone decomposes on heating above 475 K to form O_2 gas.



- (2) **Oxidising agent:** Ozone is one of the most powerful oxidizing agents with the liberation of dioxygen. In fact, ozone is a stronger oxidising agent than molecular oxygen because ozone has higher energy content and decomposes to give atomic oxygen as: $O_3 \rightarrow O_2 + \underset{\text{Atomic oxygen}}{O}$

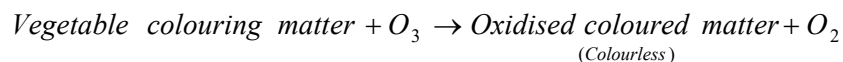
Therefore, ozone oxidises a number of non-metals and other reducing agents. e.g.



Mercury is oxidised to mercurous oxide, $2Hg + O_3 \rightarrow \underset{\text{Mercurous oxide}}{Hg_2O} + O_2$

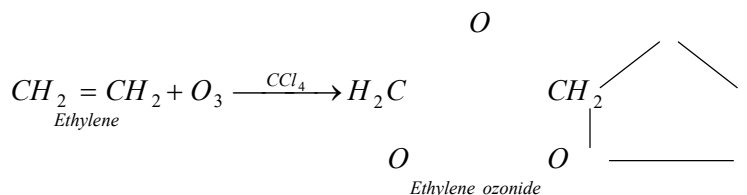
During this reaction mercury loses its meniscus and starts sticking to the sides of the glass. This is known as tailing of mercury. Mercurous oxide formed in this reaction dissolves in mercury and starts sticking to the glass surface.

- (3) **Bleaching agent:** Due to the oxidizing action of ozone, it acts as a mild bleaching agent as well as a sterilizing agent. It acts as a bleaching agent for vegetable coloring matter.

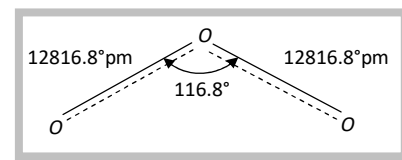


For example, ozone bleaches indigo, ivory, litmus, delicate fabrics etc.

(4) **Formation of ozonides:** Ozone reacts with alkenes in the presence of CCl_4 to form an ozonide. e.g.



Structure of O_3 : The structure of O_3 molecule is angular as shown in fig. The $\text{O}-\text{O}-\text{O}$ bond angle is 116.8° and $\text{O}-\text{O}$ bond length is 128 pm.



Uses of ozone

- (1) O_3 is used for disinfecting water for drinking purposes because ozone has germicidal properties.
- (2) It is used for purifying air of crowded places such as cinemas, underground railway, auditoriums, tunnels, mines etc.
- (3) It is used in industry for the manufacture of KMnO_4 , artificial silk, synthetic camphor etc.

Sulphur and its compounds

Sulphur is the second member of oxygen family and belongs to group-16 (VI A) of the periodic table.

(1) **Occurrence:** Sulphur occurs in the earth's crust to the extent of 0.05%. It occurs in the Free State as well as in combined state. Sulphure occurs mainly as sulphides and sulphates. eg.

Sulphide Ores		Sulphate Ores	
Iron pyrites (fool's gold)	– FeS_2	Gypsum	– $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$
Galena	– PbS	Epsom salt	– $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$

Copper pyrites	– $CuFeS_2$	Barytes	– $BaSO_4$
Cinnabar	– HgS	Zinc blende	– ZnS

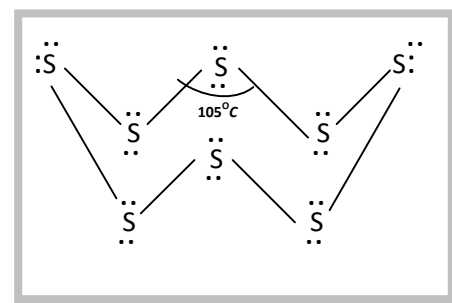
(2) **Extraction of sulphur** (Frasch process): Sulphur is generally extracted from underground deposits by drilling three concentric pipes up to the beds of sulphur (700 – 1200 feet deep).

(3) **Allotropy in sulphur**: Sulphur exists in four allotropic forms,

(i) **Rhombic or octahedral or α -sulphur**: It is a bright yellow solid, soluble in CS_2 and stable at room temp. All other varieties of sulphur gradually change into this form on standing.

(ii) **Monoclinic sulphur or prismatic or β -sulphur**: It is prepared by melting the sulphur and then cooling it till a crust is formed. On removing the crust, needle shaped crystals of monoclinic sulphur separate out. It is dull yellow in colour, soluble in CS_2 and stable only above 369K. Below this temperature it changes into rhombic form.

Thus, at 369K both these varieties co-exist. This temperature is called transition temperature and the two sulphurs are called enantiotropic substances. It also exist as molecules similar to that of rhombic sulphur but the symmetry of the crystals is different.



(iii) **Plastic or amorphous or γ -sulphur**: It is a super cooled liquid insoluble in CS_2 , soft and amorphous. It consists of long zig-zag chains of S-atoms.

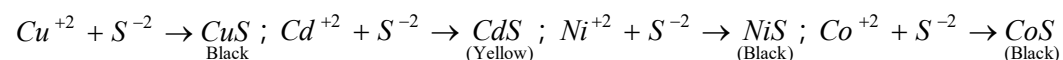
(iv) **Colloidal or δ -sulphur**: It is prepared by passing H_2S through a solution of an oxidizing agent or water or by treating sodium thiosulphate with dil. HCl.

(4) **Properties of sulphur:** It burns in air with, a blue flame forming SO_2 , gives sulphur hexafluoride with F_2 and sulphur mono chloride with Cl_2 , sulphides with metals like Na, Ca, Zn, Hg, Fe, Cu etc., reduces HNO_3 to NO_2 and H_2SO_4 to SO_2 . With NaOH solution on heating, $S_8 + 12NaOH \longrightarrow 4Na_2S + 2Na_2S_2O_3 + 6H_2O$. It gives sodium sulphide and sodium thiosulphate, with excess of sulphur, $2Na_2S + S_8 \longrightarrow 2Na_2S_5$.

(5) **Uses of sulphur:** It is used in the manufacture of matches, gun powder (mixture of charcoal, sulphur and potassium nitrate), explosives and fireworks SO_2 , H_2SO_4 , CS_2 and dyes, sulphur drugs and ointment for curing skin diseases and in the vulcanization of rubber.

Compounds of Sulphur

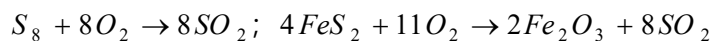
(1) **Hydrogen Sulphide:** It is prepared in the laboratory by the action of dil. H_2SO_4 on ferrous sulphide in kipp's apparatus, $FeS + H_2SO_4 \rightarrow FeSO_4 + H_2S$. It is colorless gas having foul smell resembling that of rotten eggs. It reacts with many cations (of group II and IV) to give colored sulphides,



The solubility of sulphides can be controlled by the H^+ ions concentration and therefore, H_2S finds extensive use in qualitative analysis of cation radicals.

(2) **Oxides of sulphur:** Sulphur forms several oxides of which sulphur dioxide (SO_2) and sulphur trioxide (SO_3) are most important.

(i) **Sulphur dioxide** (SO_2): It is prepared by burning sulphur or iron pyrites in air.

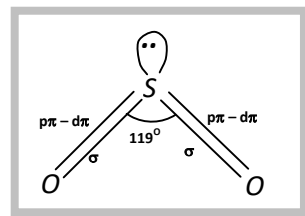


In laboratory, it is prepared by heating copper turnings with conc. H_2SO_4

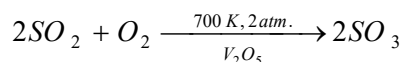


It is a colorless gas with irritating and suffocating smell.

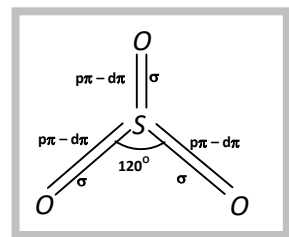
SO_2 Molecule has a bent structure with a O – S – O bond angle of 119° . Sulphur is sp^2 hybridized.



(ii) **Sulphur trioxide** (SO_3): It is formed by the oxidation of SO_2 .



In the gaseous phase, it exists as planar triangular molecular species involving hybridization of the S-atom. It has three S–O σ bonds and three S–O π bonds. The O–S–O bond angle is of 120° .



(3) **Oxyacids of sulphur:** Sulphur forms many oxyacids. Some of these are

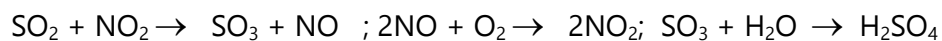
Formula	Name	Important properties	Structural formula
H_2SO_3 (+4)	Sulphurous acid	Free acid does not exist diprotic, strong reducing agent	$\begin{array}{c} \text{O} \\ \vdots \\ \text{O} = \ddot{\text{S}} - \text{OH} \\ \\ \text{OH} \end{array}$
H_2SO_4 (+6) (Oil of vitriol)	Sulphuric acid	Stable diprotic, dehydrating agent	$\begin{array}{c} \text{O} \\ \\ \text{O} = \ddot{\text{S}} - \text{OH} \\ \\ \text{OH} \end{array}$

$H_2S_2O_3$ (-2 and +6)	Thiosulphuric acid	Free acid does not exist but its salts e.g. $Na_2S_2O_3$ All quite stable reducing agent	$\begin{array}{c} S \\ \\ O = S - OH \\ \\ OH \end{array}$
$H_2S_2O_4$ (+3)	Dithionous acid		$\begin{array}{c} O \quad O \\ \quad \\ HO - S - S - OH \end{array}$
$H_2S_2O_6$ (+5)	Dithionic acid	Free acid is moderately stable but its salts are quite stable.	$\begin{array}{c} O \quad O \\ \quad \\ O = S - S = O \\ \quad \\ OH \quad OH \end{array}$
$H_2S_2O_7$ (+6) (Oleum)	Disulphuric acid (Pyrosulphuric acid)	Strong oxidising agent	$\begin{array}{c} O \quad O \\ \quad \\ O = S - O - S = OH \\ \quad \\ OH \quad OH \end{array}$
H_2SO_5 (+6) (Caro's acid)	Peroxomonosulphuric acid (Its salts known as persulphates)	Stable crystalline solid, powerful oxidising agent	$\begin{array}{c} O \\ \\ HO = S - OOH \\ \\ O \end{array}$
$H_2S_2O_8$ (+6) (Marshals acid)	Peroxodisulphuric acid (its salts are known as disulphates)	Strong oxidising agent.	$\begin{array}{c} O \quad O \\ \quad \\ O = S - O - O - S = O \\ \quad \\ OH \quad OH \end{array}$

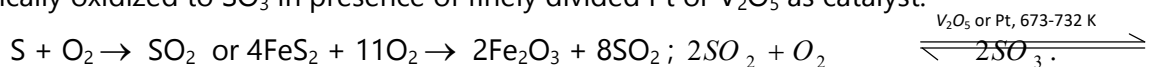
(4) **Sulphuric acid** (H_2SO_4): H_2SO_4 is a very stable oxyacid of sulphur. It is often called king of chemicals, since it is one of the most useful chemicals in industry.

Manufacture of sulphuric acid: H_2SO_4 can be manufactured by following process,

(1) **Lead chamber process:** In this process, SO_2 is oxidized to SO_3 by the oxides of nitrogen and the SO_3 thus formed is dissolved in steam to form H_2SO_4 .



(2) **Contact process:** In the contact process, SO_2 obtained by burning of S or iron pyrites is catalytically oxidized to SO_3 in presence of finely divided Pt or V_2O_5 as catalyst.

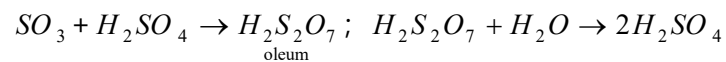


V_2O_5 is, however, preferred since is much cheaper than Pt and is also not poisoned by arsenic impurities.

The favorable conditions for maximum yield of SO_3 are,

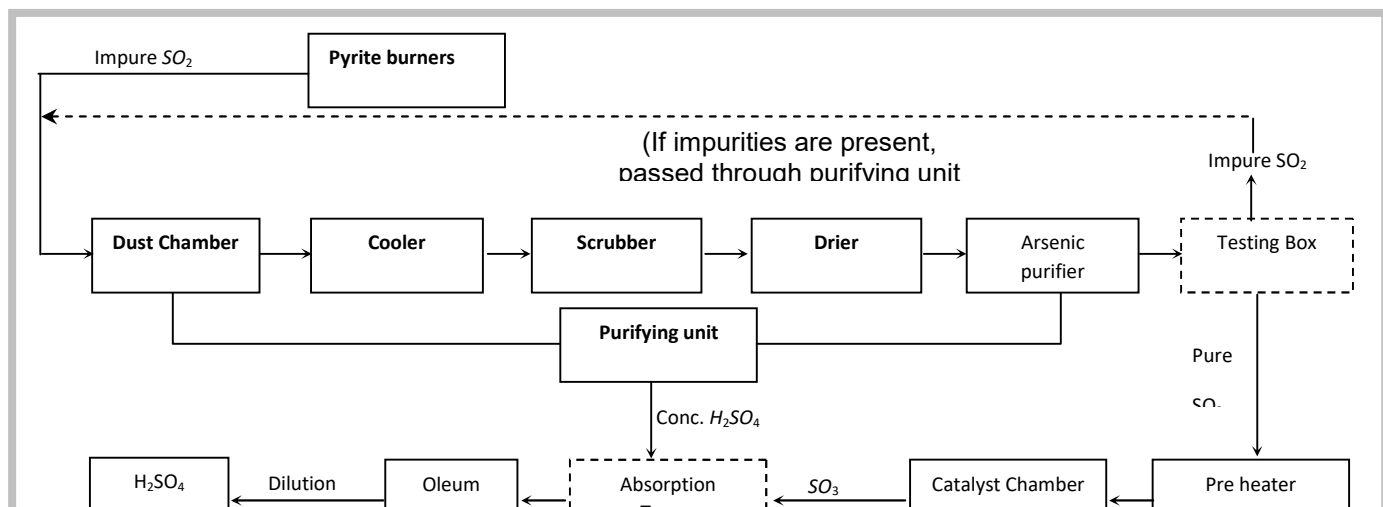
- (i) High concentration of SO_2 and O_2 .
- (ii) Low temperature of 673 to 723 K,
- (iii) High pressure about 2 atmospheres.

SO_3 thus obtained is absorbed in 98% H_2SO_4 to form oleum which on dilution with water gives H_2SO_4 of desired concentration.



Contact process is preferred over lead chamber process (gives 98% pure H_2SO_4) since it gives H_2SO_4 of greater purity (100%).

Flow sheet diagram of its preparation is as follows

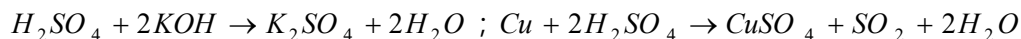


Structure: H_2SO_4 is a covalent compound and has tetrahedral (S is sp^3 - hybridized) structure.

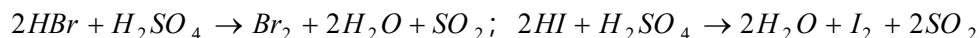
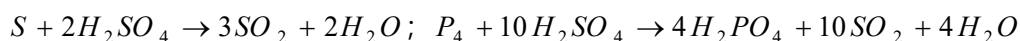
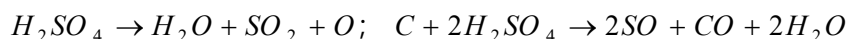
Properties: H_2SO_4 has high b.p. (611K) and is also highly viscous due to H-bonding. It has strong affinity for H_2O and a large amount of heat is evolved when it is mixed with water.

(1) H_2SO_4 is a strong dibasic acid. It neutralizes alkalies, liberates CO_2 from carbonates and bicarbonates.

(2) It reacts with more electropositive (than hydrogen) metals to evolve H_2 and produces SO_2 on heating with less electropositive metals than hydrogen .eg.,



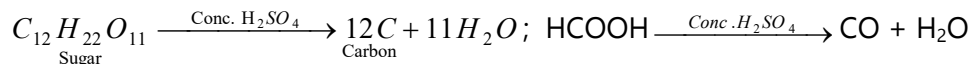
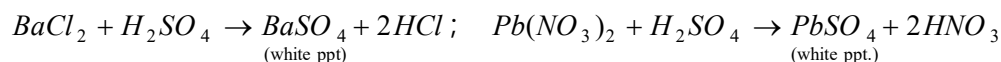
(3) It is a strong oxidizing agent and oxidizes as follows,



(4) It reacts with number of salts. It liberates HCl from chlorides, H_2S from sulphides, HNO_3 from nitrates.

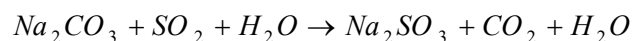
(5) It acts as a strong dehydrating agent, as it dehydrates, sugar to sugar charcoal (carbon), and formic acid to CO, oxalic acid to $CO + CO_2$ and ethyl alcohol to ethylene.

(6) It is also a good sulphonating agent and used for sulphonation of aromatic compounds. eg.,



Uses : H_2SO_4 is used (i) in the preparation of fertilizers like $(NH_4)_2SO_4$ and super phosphate of lime, (ii) in lead storage batteries (iii) in preparation of dyes, paints and explosives (iv) in textile and paper industry (v) for tanning (vi) as a dehydrating agent.

(5) **Sodium thiosulphate** $Na_2S_2O_3 \cdot 5H_2O$: It is manufactured by saturating a solution of sodium carbonate with SO_2 which gives a solution of sodium sulphite,



The resulting solution is boiled with powdered sulphur as, $Na_2SO_3 + S \xrightarrow{373 K} Na_2S_2O_3$

The solution is then cooled to get crystals of sodium thiosulphate.

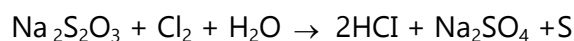
Physical properties:

- (1) Sodium thiosulphate is a colourless crystalline solid. In the hydrated form, it is called hypo.
- (2) It melts at 320 K and loses its water molecules of crystallization on heating to 490K.

Chemical properties

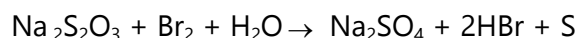
(1) **Action with halogens:** It reacts with halogens as,

(i) Chlorine water oxidizes sodium thiosulphate to sodium sulphate and sulphur is precipitated,

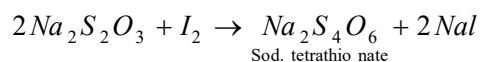


This property enables it to act as an antichlor in bleaching i.e. it destroys the unreacted chlorine in the process of bleaching.

(ii) Bromine water also oxidizes sodium thiosulphate to sodium sulphate and sulphur,



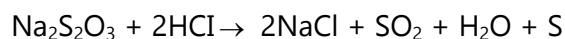
(iii) With iodine it forms a soluble compound called sodium tetrathionate,



Therefore, hypo is commonly used to remove iodine stains from the clothes.

(2) **Action of heat:** Upon heating, sodium thiosulphate decomposes to form sodium sulphate and sodium pentasulphide, $4 Na_2S_2O_3 \xrightarrow{\text{Heat}} 3 Na_2SO_4 + Na_2S_5$
Sodium pentasulphide

(3) **Action with acids:** Sodium thiosulphate reacts with dilute hydrochloric acid or Sulphuric acid forming sulphur dioxide and sulphur. The solution turns milky yellow due to sulphur.



(4) **Action with silver halides:** Sodium thiosulphate forms soluble complex when treated with silver chloride or silver bromide, $2 Na_2S_2O_3 + 2 AgBr \rightarrow Na_3Ag(S_2O_3)_2 + NaBr$.
Sodium dithiosulphate
 argentate (I) complex

This property of hypo is made use in photography.

Uses of sodium thiosulphate

- (1) It is largely used in photography as a fixing agent.
- (2) It is used as a preservative for fruit products such as jams and squashes.
- (3) It is used as an antichlor in bleaching.
- (4) It is used as a volumetric agent for the estimation of iodine.
- (5) It is used in medicine.