Sodium

Contents

1 - Introduction :

Sodium is a metallic element with a symbol **Na** (from Latin *natrium* or Arabic *natrun*) and atomic number 11 . It is a soft, silvery - white, highly reactive metal and is a member of the alkali metals within " group 1 " (formerly known as ‘ group IA ’) . It has only one stable isotope , ^{23}Na.

Elemental sodium was first isolated by Sir Humphry Davy in 1806 by passing an electric current through molten sodium hydroxide. Elemental sodium does not occur naturally on Earth, but quickly oxidizes in air and is violently reactive with water, so it must be stored in an inert medium, such as a liquid hydro carbon. The free metal is used for some chemical synthesis and heat transfer applications.

Sodium ion is soluble in water in nearly all of its compounds, and is thus present in great quantities in the Earth's oceans and other stagnant bodies of water. In these bodies it is mostly counterbalanced by the chloride ion, causing evaporated ocean water solids to consist mostly of sodium chloride, or common table salt. Sodium ion is also a component of many minerals.

Sodium is an essential element for all animal life and for some plant species. In animals, sodium ions are used in opposition to potassium ions, to allow the organism to build up an electrostatic charge on cell membranes, and thus allow transmission of nerve impulses when the charge is allowed to dissipate by a moving wave of voltage change. Sodium is thus classified as a “dietary inorganic macro-mineral” for animals. Sodium's relative rarity on land is due to its solubility in water, thus causing it to be leached into bodies of long-standing water by rainfall. Such is its relatively large requirement in animals, in contrast to its relative scarcity in many inland soils, that herbivorous land animals have developed a special taste receptor for sodium ion.

Appearance	Silvery white metallic
Name , Symbol , Number	Sodium , Na , 11
Element category	Alkali metal
Group, period, block	1 , 3 , s
Standard atomic weight	23 g ·mol $^{-1}$
Electron configuration	[Ne] $3s^1$
Electrons per shell	2,8,1
Phase	solid
Density	0.968 g ·cm^{-3}
Liquid density at m.p.	0.927 g ·cm^{-3}
Melting point	97.72 ° C
Boiling point	883 ° C
Critical point	(extrapolated) 2573 K , 35 MPa
Heat of fusion	2.60 kJ · mol^{-1}

Heat of vaporization	97.42 kJ · mol^{-1}
Specific heat capacity	(25 °C) 28.230 J · mol^{-1}·K^{-1}
Oxidation states	+1 , -1 (strongly basic oxide)
Electro negativity	0.93 (Pauling scale)
Ionization energies	1st : 495.8 kJ ·mol^{-1} 2nd : 4562 kJ ·mol^{-1} 3rd : 6910.3 kJ ·mol^{-1}
Atomic radius	186 pm
Covalent radius	166 ± 9 pm
Van der Waals radius	227 pm
Crystal structure	Body - centered cubic
Magnetic ordering	paramagnetic
Electrical resistivity	(20 °C) 47.7 nΩ ·m
Thermal conductivity	(300 K) 142 W·m^{-1} · K^{-1}
Thermal expansion	(25 °C) 71 µm·m^{-1} · K^{-1}
Speed of sound	(20 °C) 3200 m / s
Mohs hardness	0.5

2 . Characteristics

At room temperature, sodium metal is soft enough that it can be cut with a knife. In air, the bright silvery luster of freshly exposed sodium will rapidly tarnish. The density of alkali metals generally increases with increasing atomic number, but sodium is denser than potassium.

2 . 1 . Chemical properties

Compared with other alkali metals, sodium is generally less reactive than potassium and more reactive than lithium,[2] in accordance with " periodic law ": for example, their reaction in water, chlorine gas, etc.;

Sodium reacts exothermically with water : small pea - sized pieces will bounce across the surface of the water until they are consumed by it, whereas large pieces will explode. While sodium reacts with water at room temperature, the sodium piece melts with the heat of the reaction to form a sphere, if the reacting sodium piece is large enough. The reaction with water produces very caustic sodium hydroxide (lye) and highly flammable hydrogen gas. These are extreme hazards . When burned in air, sodium forms sodium peroxide Na_2O_2, or with limited oxygen, the oxide Na_2O (unlike lithium , the nitride is not formed) . If burned in oxygen under pressure , sodium superoxide NaO_2 will be produced. In chemistry, most sodium compounds are considered soluble but nature provides examples of many insoluble sodium compounds such as the feldspars. There are other insoluble sodium salts such as sodium bismuthate $NaBiO_3$, sodium octamolybdate $Na_2Mo_8O_{25} \bullet 4H_2O$, sodium thioplatinate $Na_4Pt_3S_6$, sodium uranate Na_2UO_4. Sodium meta-antimonate's $2NaSbO_3 \bullet 7H_2O$ solubility is 0.3g/L as is the pyro form $Na_2H_2Sb_2O_7 \bullet H_2O$ of this salt. Sodium metaphosphate $NaPO_3$ has a soluble and an insoluble form .

2 . 2 . Compounds

Sodium compounds are important to the chemical, glass, metal, paper, petroleum, soap, and textile industries. Hard soaps are generally sodium salt of certain fatty acids (potassium produces softer or liquid soaps) .

The sodium compounds that are the most important to industries are common salt (NaCl) , soda ash (Na_2CO_3) , baking soda (Na HCO_3) , caustic soda (Na OH) , sodium nitrate (Na NO_3) , di - and tri - sodium phosphates, sodium thio sulfate (hypo , $Na_2 S_2O_3 \cdot 5H_2O$) , and borax ($Na_2 B_4O_7 \cdot 10H_2O$) .

2 . 3 . Spectroscopy

When sodium or its compounds are introduced into a flame , they turn the flame a bright yellow color.

One notable atomic spectral line of sodium vapor is the so - called D - line, which may be observed directly as the sodium flame - test line and also the major light output of low-pressure sodium lamps (these produce an unnatural yellow, rather than the peach - colored glow of high pressure lamps). The D - line is one of the classified Fraunhofer lines observed in the visible spectrum of the Sun's electromagnetic radiation. Sodium vapor in the upper layers of the Sun creates a dark line in the emitted spectrum of electromagnetic radiation by absorbing visible light in a band of wavelengths around 589.5 nm. This wavelength corresponds to transitions in atomic sodium in which the valence-electron transitions from a 3p to 3s electronic state. Closer examination of the visible spectrum of atomic sodium reveals that the D - line actually consists of two lines called the D_1 and D_2 lines at 589.6 nm and 589.0 nm, respectively. This fine structure results from a spin-orbit interaction of the valence electron in the 3p electronic state. The spin - orbit interaction couples the spin angular momentum and orbital angular momentum of a 3p electron to form two states that are respectively notated as and in the LS coupling scheme. The 3s state of the electron gives rise to a single state which is notated as $3s\ (^2S_{1/2})$ in the LS coupling scheme. The D_1 - line results from an electronic transition between $3s\ (^2S_{1/2})$ lower state and upper state. The D_2 - line results from an electronic transition between $3s\ (^2S_{1/2})$ lower state and upper state. Even closer examination of the visible spectrum of atomic sodium would reveal that the D - line actually consists of a lot more than two lines. These lines are associated with hyperfine structure of the 3p upper states and 3s lower states. Many different transitions involving visible light near 589.5 nm may occur between the different upper and lower hyperfine levels .

A practical use for lasers which work at the sodium D-line transition is to create artificial laser guide stars , which assist in the adaptive optics for large land - based visible light telescopes.

2 . Isotopes of Sodium :

Thirteen isotopes of sodium have been recognized. The only stable isotope is ^{23}Na . Sodium has two radioactive cosmogenic isotopes which are also the two isotopes with longest half life , ^{22}Na, with a half-life of 2.6 years and ^{24}Na with a half - life of 15 hours. All other isotopes have a half life of less than one minute .

Acute neutron radiation exposure (e.g., from a nuclear criticality acciden t) converts some of the stable ^{23}Na in human blood plasma to ^{24}Na. By measuring the concentration of this isotope , the neutron radiation dosage to the victim can be computed .

3 . History of Sodium :

Salt has been an important commodity in human activities, as testified by the English word *salary*, referring to *salarium*, the wafers of salt sometimes given to Roman soldiers along with their other wages.

In medieval Europe a compound of sodium with the Latin name of *sodanum* was used as a headache remedy. The name sodium probably originates from the Arabic word suda meaning headache as the headache - alleviating properties of sodium carbonate or soda were well known in early times .

Sodium's chemical abbreviation *Na* was first published by Jöns Jakob Berzelius in his system of atomic symbols (Thomas Thomson, *Annals of Philosophy*) and is a contraction of the element's new Latin name *natrium* which refers to the Egyptian *natron* , the word for a natural mineral salt whose primary ingredient is hydrated sodium carbonate. Hydrated sodium carbonate historically had several important industrial and household uses later eclipsed by soda ash, baking soda and other sodium compounds .

Although sodium (some times called " soda " in English) has long been recognized in compounds, it was not isolated until 1807 by Sir Humphry Davy through the electrolysis of caustic soda .

Sodium imparts an intense yellow color to flames. As early as 1860, Kirchhoff and Bunsen noted the high sensitivity that a flame test for sodium could give .

4 . Occurrence of Sodium :

Owing to its high reactivity, sodium is found in nature only as a compound and never as the free element . Sodium makes up about

2.6 % by weight of the Earth's crust, making it the sixth most abundant element overall and the most abundant alkali metal. Sodium is found in many different minerals, of which the most common is ordinary salt (sodium chloride) , which occurs in vast quantities dissolved in seawater, as well as in solid deposits (halite) . Others include amphibole, cryolite, soda niter and zeolite.

Sodium is relatively abundant in stars and the D spectral lines of this element are among the most prominent in star light. Though elemental sodium has a rather high vaporization temperature, its relatively high abundance and very intense spectral lines have allowed its presence to be detected by ground telescopes and confirmed by spacecraft .

5 . Commercial production of Sodium :

Sodium was first produced commercially in 1855 by thermal reduction of sodium carbonate with carbon at 1100 °C, in what is known as the Deville process .

$$Na_2CO_3 \text{ (liquid) } + 2\,C \text{ (solid) } \rightarrow 2\,Na \text{ (vapor) } + 3\,CO \text{ (gas).}$$

A process based on the reduction of sodium hydroxide was developed in 1886 .

Sodium is now produced commercially through the electrolysis of liquid sodium chloride, based on a process patented in 1924 . This is done in a Downs Cell in which the NaCl is mixed with calcium chloride to lower the melting point below 700 °C. As calcium is less electropositive than sodium, no calcium will be formed at the anode. This method is less expensive than the previous Castner process of electrolyzing sodium hydroxide .Very pure sodium can be isolated by the thermal decomposition of sodium azide .

6 . Applications

Sodium in its metallic form can be used to refine some reactive metals, such as zirconium and potassium , from their compounds . This alkali metal as the Na^+ ion is vital to animal life . Other uses :

- In certain alloys to improve their structure.
- In soap, in combination with fatty acids. Sodium soaps are harder (higher melting) soaps than potassium soaps.
- To descale metal (make its surface smooth) .
- To purify molten metals.
- In some medicine formulations, the salt form of the active ingredient usually with sodium or potassium is a common modification to improve bioavailability.
- In sodium vapor lamps, an efficient means of producing light from electricity , often used for street lighting in cities. Low-pressure sodium lamps give a distinctive yellow-orange light which consists primarily of the twin sodium D lines. High - pressure sodium lamps give a more natural peach - colored light, composed of wavelengths spread much more widely across the spectrum.
- As a heat transfer fluid in some types of nuclear reactors and inside the hollow valves of high - performance internal combustion engines.
- Sodium chloride (NaCl) , a compound of sodium ions and chloride ions, is an important heat transfer material.
- In organic synthesis, sodium is used as a reducing agent, for example in the Birch reduction.
- In chemistry, sodium is often used either alone or with potassium in an alloy, NaK as a desiccant for drying solvents. Used with benzophenone, it forms an intense blue coloration when the solvent is dry and oxygen-free.

7 . Biological role

7 . 1 . In maintenance of body fluid volume in animals

The serum sodium and urine sodium play important roles in medicine, both in the maintenance of sodium and total body fluid homeostasis, and in the diagnosis of disorders causing homeostatic disruption of salt/sodium and water balance.

In mammals, decreases in blood pressure and decreases in sodium concentration sensed within the kidney result in the production of renin, a hormone which acts in a number of ways, one of them being to act indirectly to cause the generation of aldosterone, a hormone which decreases the excretion of sodium in the urine. As the body of the mammal retains more sodium, other osmoregulation systems which sense osmotic pressure in part from the concentration of sodium and water in the blood, act to generate antidiuretic hormone. This, in turn, which causes the body to retain water, thus helping to restoring the body's total amount of fluid.

There is also a counterbalancing system, which senses volume. As fluid is retained, receptors in the heart and vessels which sense distension and pressure, cause production of atrial natriuretic peptide, which is named in part for the Latin word for sodium. This hormone acts in various ways to cause the body to lose sodium in the urine. This causes the body's osmotic balance to drop (as low concentration of sodium is sensed directly), which in turn causes the osmoregulation system to excrete the "excess" water. The net effect is to return the body's total fluid levels back toward normal.

7 . 2 . In maintenance of resting electrical potential in excitable tissues in animals :

Sodium cations are important in neuron (brain and nerve) function , and in influencing osmotic balance between cells and the interstitial fluid, with their distribution mediated in all animals (but not in all plants) by the so - called Na+ / K+-AT Pase pump . Sodium is the chief cation in fluid residing outside cells in the mammalian body (the so - called extra cellular compartment) , with relatively little sodium residing inside cells. The volume of extracellular fluid is typically 15 litres in a 70 kg human, and the 50 grams of sodium it contains is about 90 % of the body's total sodium content.

8 . Dietary uses

The most common sodium salt , sodium chloride (table salt) , is used for seasoning and warm - climate food preservation, such as pickling and making jerky (the high osmotic content of salt inhibits bacterial and fungal growth) . The human requirement for sodium in the diet is about 500 mg per day , which is typically less than a tenth as much as many diets "seasoned to taste." Most people consume far more sodium than is physiologically needed. For certain people with salt-sensitive blood pressure, this extra intake may cause a harmful effect on health. However, low sodium intake may lead to sodium deficiency .

9 . Precautions

Extreme care is required in handling elemental/metallic sodium. Sodium is potentially explosive in water (depending on quantity) and is a corrosive substance, since it is rapidly converted to sodium hydroxide on contact with moisture. The powdered form may combust spontaneously in air or oxygen. Sodium must be stored either in an inert (oxygen and moisture free) atmosphere (such as nitrogen or argon) , or under a liquid hydrocarbon such as mineral oil or kerosene.

The reaction of sodium and water is a familiar one in chemistry labs, and is reasonably safe if amounts of sodium smaller than a pencil eraser are used and the reaction is done behind a plastic shield by people wearing eye protection. However, the sodium-water reaction does not scale up well, and is treacherous when larger amounts of sodium are used. Larger pieces of sodium melt under the heat of the reaction, and the molten ball of metal is buoyed up by hydrogen and may appear to be stably reacting with water, until splashing covers more of the reaction mass, causing thermal runaway and an explosion which scatters molten sodium, lye solution, and sometimes flame . (18.5 g explosion) This behavior is unpredictable, and among the alkali metals it is usually sodium which invites this surprise phenomenon, because lithium is not reactive enough to do it, and potassium is so reactive that chemistry students are not tempted to try the reaction with larger potassium pieces.

Sodium is much more reactive than magnesium; a reactivity which can be further enhanced due to sodium's much lower melting point. When sodium catches fire in air (as opposed to just the hydrogen gas generated from water by means of its reaction with sodium) it more easily produces temperatures high enough to melt the sodium, exposing more of its surface to the air and spreading the fire.

Few common fire extinguishers work on sodium fires. Water, of course, exacerbates sodium fires, as do water-based foams . CO_2 and Halon are often ineffective on sodium fires, which reignite when the extinguisher dissipates. Among the very few materials effective on a sodium fire are Pyromet and Met- L - X . Pyromet is a NaCl / $(NH_4)_2$ HPO_4 mix , with flow/anti - clump agents. It smothers the fire, drains away heat, and melts to form an impermeable crust. This is the standard dry - powder canister fire extinguisher for all classes of fires. Met - L - X is mostly sodium chloride, NaCl, with approximately 5 % Saran plastic as a crust-former, and flow / anti - clumping agents. It is most commonly hand - applied, with a scoop. Other extreme fire extinguishing materials include Lith+ , a graphite based dry powder with an organophosphate flame retardant; and Na+ , a $Na_2 CO_3$ - based material.

Because of the reaction scale problems discussed above, disposing of large quantities of sodium (more than 10 to 100 grams) must be done through a licensed hazardous materials disposer. Smaller quantities may be broken up and neutralized carefully with ethanol (which has a much slower reaction than water), or even methanol (where the reaction is more rapid than ethanol's but still less than in water) , but care should nevertheless be taken, as the caustic products from the ethanol or methanol reaction are just as hazardous to eyes and skin as those from water. After the alcohol reaction appears complete, and all pieces of reaction debris have been broken up or dissolved, a mixture of alcohol and water, then pure water, may then be carefully used for a final cleaning. This should be allowed to stand a few minutes until the reaction products are diluted more thoroughly and flushed down the drain. The purpose of the final water soaking and washing of any reaction mass or container which may contain sodium, is to ensure that alcohol does not carry unreacted sodium into the sink trap, where a water reaction may generate hydrogen in the trap space which can then be potentially ignited, causing a confined sink trap explosion .

Sodium Aluminum Fluoride (Cryolite)

Cryolite from Ivigtut Greenland

1 – Introduction :

Cryolite ($Na_3 Al F_6$, sodium hexa fluoro aluminate) is an uncommon mineral identified with the once large deposit at Ivigtût on the west coast of Greenland, depleted by 1987.

The cryolite mine Ivigtut, Greenland, summer 1940

It was historically used as an ore of aluminium and later in the electrolytic processing of the aluminium - rich oxide ore bauxite (itself a combination of aluminium oxide minerals such as gibbsite, boehmite and diaspore) . The difficulty of separating aluminium from oxygen in the oxide ores was overcome by the use of cryolite as a flux to dissolve the oxide mineral (s). Pure cryolite itself melts at 1012 °C (1285 K), and it can dissolve the aluminium oxides sufficiently well to allow easy extraction of the aluminium by electrolysis. Considerable energy is still required for both heating the materials and the electrolysis, but it is much more energy - efficient than melting the oxides themselves. Now, as natural cryolite is too rare to be used for this purpose, synthetic sodium aluminium fluoride is produced from the common mineral fluorite.

Cryolite occurs as glassy, colorless, white-reddish to gray-black prismatic monoclinic crystals. It has a Mohs hardness of 2.5 to 3 and a specific gravity of about 2.95 to 3.0. It is translucent to transparent with a very low refractive index of about 1.34, which is very close to that of water; thus if immersed in water, cryolite becomes essentially invisible .

Cryolite has also been reported at Pikes Peak, Colorado; Mont Saint - Hilaire , Quebec; and at Miass , Russia. It is also known in small quantities in Brazil, the Czech Republic, Namibia, Norway, Ukraine, and several American states.

Cryolite was first described in 1799 from a deposit of it in Ivigtut and Arksukfiord, West Greenland. The name is derived from the Greek language words *cryò* = chill, and *lithòs* = stone.[2] The Pennsylvania Salt Manufacturing Company used large amounts of cryolite to make caustic soda at its Natrona, Pennsylvania works during the 19th and 20th centuries.

Category	Halide Mineral
Chemical Formula	$Na_3 Al F_6$
Crystal Symmetry	Monoclinic 2 / m
Unit cell	a = 7.7564 (3) Å , b = 5.5959 (2) Å , c = 5.4024 (2) Å ; β = 90.18° ; Z = 2
Molar Mass	210 g mol^{-1}
Color	Colorless to white, also brownish, reddish and rarely black
Crystal habit	Usually massive , coarsely granular . The rare crystals are equant and pseudo cubic
Crystal system	Monoclinic 2/m
Twinning	Very common, often repeated or poly synthetic with simultaneous occurrence of several twin laws
Cleavage	None observed

Fracture	Uneven
Tenacity	Brittle
Mohs scale hardness	2.5 to 3
Streak	White
Diaphaneity	Transparent to translucent
Specific gravity	2.95 to 3.0
Optical properties	Biaxial (+)
Refractive index	n_α = 1.3385 –1.339 , n_β = 1.3389 –1.339 , n_γ = 1.3396 –1.34
Birefringence	δ = 0.001
2V angle	43°
Dispersion	r < v
Melting point	1012 °C
Solubility	Easily soluble in $AlCl_3$ solution, soluble in H_2SO_4 with the evolution of HF, which is poisonous. Slightly soluble in water.
Other characteristics	Weakly thermo lumine scent . Small clear fragments become nearly invisible when placed in water, since its refractive index is close to that of water. May fluoresce intense yellow under SWUV, with yellow phosphorescence, and pale yellow phosphorescence under LWUV. Not radioactive.

2 - Pesticide and insecticide :

Cryolite is used as a insecticide and a pesticide.

Sodium Bi Carbonate

Contents

1 . Introduction :

Sodium bicarbonate or **sodium hydrogen carbonate** is the chemical compound with the formula $NaHCO_3$. Sodium bicarbonate is a white solid that is crystalline but often appears as a fine powder. It has a slight alkaline taste resembling that of washing soda (sodium carbonate) . It is a component of the mineral natron and is found dissolved in many mineral springs. The natural mineral form is known as **nahcolite**. It is also produced artificially.

Since it has long been known and is widely used, the salt has many related names such as **baking soda** , **bread soda**, **cooking soda** , **bicarbonate of soda** . Colloquially, its name is shortened to ***sodium bicarb*** , ***bicarb soda*** , or simply ***bicarb*** . The word *saleratus*, from Latin *sal æratus* meaning "aerated salt", was widely used in the 19th century for both sodium bicarbonate and potassium bicarbonate. The term has now fallen out of common usage.

Other Names	Sodium bicarbonate Bicarbonate of soda Baking soda Sodium hydrogen carbonate Nahcolite
Molecular Formula	$Na\ HCO_3$
Molar Mass	84 g / mol
Appearance	White crystalline solid
Odor	Odorless
Density	2.159 g / cm^3
Melting Point	270 °C (decomp)
Solubility in Water	7.8 g / 100 mL (18 ° C) 10 g / 100 mL (20 ° C)
Solubility	Insoluble in alcohol
Acidity (pK_a)	10.3
Refractive index (n_D)	1.500
Flash point	Non - flammable
LD_{50}	4220 mg / kg

2 . History

The ancient Egyptians used natural deposits of natron, a mixture consisting mostly of sodium carbonate deca hydrate and sodium bicarbonate. The natron was used as a cleansing agent like soap.

In 1791, a French chemist, Nicolas Leblanc, produced sodium bicarbonate as we know it today. In 1846 two New York bakers, John Dwight and Austin Church, established the first factory to develop baking soda from sodium carbonate and carbon dioxide.

3 . Production (Solvay process) .

$NaHCO_3$ is mainly prepared by the Solvay process, which is the reaction of calcium carbonate , sodium chloride , ammonia , and carbon dioxide in water . It is produced on the scale of about 100,000 ton / year (as of 2001) .

$NaHCO_3$ may be obtained by the reaction of carbon dioxide with an aqueous solution of sodium hydroxide. The initial reaction produces sodium carbonate:

$$CO_2 + 2\ Na\ OH \rightarrow Na_2\ CO_3 + H_2O$$

Further addition of carbon dioxide produces sodium bicarbonate which at sufficiently high concentration will precipitate out of solution :

$$Na_2\ CO_3 + CO_2 + H_2O \rightarrow 2\ Na\ HCO_3$$

Commercial quantities of baking soda are also produced by a similar method: soda ash, mined in the form of the ore trona, is dissolved in water and treated with carbon dioxide. Sodium bicarbonate precipitates as a solid from this method:

$$Na_2\ CO_3 + CO_2 + H_2O \rightarrow 2\ Na\ HCO_3$$

4 . Mining

Naturally occurring deposits of nahcolite ($Na\ HCO_3$) are found in the Eocene - age (55.8 ± 0.2 - 33.9 ± 0.1 Ma) Green River Formation , Piceance Basin in Colorado. Nahcolite was deposited as beds during periods of high evaporation in the basin. It is commercially mined using in-situ leach techniques involving dissolution of the nahcolite by heated water which is pumped through the nahcolite beds and reconstituted through a natural cooling crystallization process. Natural Soda, Inc., the only North American processor of natural sodium bicarbonate is owned by AmerAlia (AALA) and The Sentient Group. The operation is located 54 miles northwest of Rifle , Colorado . The Natural Soda facility has changed owner ship several times over the last decade, proving difficult to turn a profit . All other North American producers base their production sodium bicarbonate on trona found in the Green River .

5 . Chemistry

Sodium bicarbonate is an amphoteric compound. Aqueous solutions are mildly alkaline due to the formation of carbonic acid and hydroxide ion:

$$HCO\ {-3} + H_2O \rightarrow H_2CO_3 + OH^-$$

Sodium bicarbonate can be used as a wash to remove any acidic impurities from a " crude " liquid, producing a purer sample. Reaction of sodium bicarbonate and an acid to give a salt and carbonic acid, which readily decomposes to carbon dioxide and water:

$$Na\ HCO_3 + H\ Cl \rightarrow Na\ Cl + H_2\ CO_3$$
$$H_2\ CO_3 \rightarrow H_2O + CO_2\ (g)$$

Sodium bicarbonate reacts with acetic acid (CH_3COOH) to form sodium acetate:

$$Na\ HCO_3 + CH_3COOH \rightarrow CH_3COONa + H_2O + CO_2\ (g)$$

Sodium bicarbonate reacts with bases such as sodium hydroxide to form carbonates:

$$NaHCO_3 + Na\ OH \rightarrow Na_2CO_3 + H_2O$$

Sodium bicarbonate reacts with carboxyl groups in proteins to give a brisk effervescence from the formation of CO_2. This reaction is used to test for the presence of carboxylic groups in protein.

5 . 1 . Thermal decomposition

Above 70 °C, sodium bicarbonate gradually decomposes into sodium carbonate , water and carbon dioxide . The conversion is fast at 250 °C :

$$2\ Na\ HCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

Most bicarbonates undergo this dehydration reaction. Further heating converts the carbonate into the oxide (at around 1000 ° C) :

$$Na_2CO_3 \rightarrow Na_2O + CO_2$$

These conversions are relevant to the use of Na HCO_3 as a fire-suppression agent (" BC powder ") in some dry powder fire extinguishers .

6 . Applications

6 . 1 . Cooking

Sodium bicarbonate is primarily used in cooking (baking) where it reacts with other components to release carbon dioxide, that helps dough " rise " . The acidic compounds that induce this reaction include phosphates, cream of tartar, lemon juice, yogurt, buttermilk, cocoa, vinegar, etc. Sodium bicarbonate can be substituted for baking powder provided sufficient acid reagent is also added to the recipe.[3] Many forms of baking powder contain sodium bicarbonate combined with one or more acidic phosphates (especially good) or cream of tartar.

6 . 2 . Neutralization of acids and bases

Many laboratories keep a bottle of sodium bicarbonate powder within easy reach, because sodium bicarbonate is amphoteric, reacting with acids and bases. Furthermore, as it is relatively innocuous in most situations, there is no harm in using excess sodium bicarbonate. Lastly, sodium bicarbonate powder may be used to smother a small fire .

A wide variety of applications follows from its neutralization properties, including ameliorating the effects of white phosphorus in incendiary bullets from spreading inside an afflicted soldier's wounds. Sodium bicarbonate can be added as a simple solution for raising the pH balance of water (increasing total alkalinity) where high levels of chlorine (2 - 5 ppm) are present as in swimming pools and aquariums .

6 . 3 . Medical uses

Sodium bicarbonate is used in an aqueous solution as an antacid taken orally to treat acid indigestion and heart burn .

An aqueous solution is sometimes administered intravenously for cases of acidosis, or when there are insufficient sodium or bicarbonate ions in the blood . It is used as well for treatment of hyperkalemia . Since sodium bicarbonate can cause alkalosis, it is sometimes used in to treat aspirin overdoses. Aspirin requires an acidic environment for proper absorption, and the basic environment diminishes aspirin absorption in the case of an overdose. Sodium bicarbonate has also been used in the treatment of tricyclic antidepressant overdose . It can also be applied topically as a paste, with three parts baking soda to one part water, to relieve insect bites .

Adverse reactions to the administration of sodium bicarbonate can include metabolic alkalosis, edema due to sodium overload, congestive heart failure , hyper osmolar syndrome, hypervolemic hyper natremia , and hypertension due to increased sodium. In patients who consume a high calcium or dairy-rich diet, calcium supplements, or calcium - containing antacids such as calcium carbonate (e.g., Tums) , the use of sodium bicarbonate can cause milk - alkali syndrome, which can result in metastatic calcification, kidney stones, and kidney failure.

Sodium bicarbonate is also used as an ingredient in some mouthwashes. It works as a mechanical cleanser on the teeth and gums, neutralizes the production of acid in the mouth and also as an antiseptic to help prevent infections occurring.

6 . 4 . Personal hygiene

A paste made from sodium bicarbonate and a 3 % hydrogen peroxide solution can be used as an alternative to commercial tooth pastes, and sodium bicarbonate in combination with other ingredients can be used to make a dry or wet deodorant. Sodium bicarbonate is a common ingredient in alternative and natural brands of toothpaste and deodorant. It can be used as a treatment for dandruff due to its anti - fungal properties.

6 . 5 . Soda loading

Soda loading is a dangerous practice in which an athlete drinks sodium bicarbonate (baking soda) dissolved in water in the belief that it will enhance athletic performance and endurance. In reality, the opposite occurs, and the resulting alkalosis can leave an athlete unable to compete .

This practice is also extremely dangerous due to the risk of an overdose of sodium bicarbonate and potentially life - threatening alkalosis. Symptoms include constipation, convulsions, diarrhea, feeling of being full, frequent urination, irritability, muscle spasms, muscle weakness and vomiting. If an overdose of baking soda is suspected emergency medical treatment should be sought immediately. The individual should not be forced to vomit the contents of the stomach unless directed to by a medical professional or a poison control center .

6 . 6 . As a cleaning agent

A paste from baking soda can be very effective when used in cleaning and scrubbing . For cleaning aluminium objects, sodium bicarbonate is to be discouraged as it attacks the thin unreactive protective oxide layer of this otherwise very reactive metal. A solution in warm water will remove the tarnish from silver when the silver is in contact with a piece of aluminium foil .

It is more effective than vinegar , salt , or hot water alone in removing pesticides from vegetables .

Baking soda is commonly added to the rinse cycles of washing machines (together with the detergent) as a replacement for softener and also to remove odors. Sodium bicarbonate is also effective in removing heavy tea and coffee stains from cups when diluted with warm water.

6 . 7 . Miscellaneous

Sodium bicarbonate can be used to extinguish small grease or electrical fires by being poured or dumped over the fire . However, it should not be poured or dumped onto fires in deep fryers as it may cause the grease to splatter . Sodium bicarbonate is used in BC dry chemical fire extinguishers as an alternative to the more corrosive ammonium phosphate in ABC extinguishers. The alkali nature of sodium bicarbonate makes it the only dry chemical agent, besides Purple - K , that was used in large scale fire suppression systems installed in commercial kitchens. The agent has a mild saponification effect on grease, which forms a smothering soapy foam. Dry chemicals have since fallen out of favor for kitchen fires as they have no cooling effect compared to the extremely effective wet chemical agents specifically designed for such hazards .

Sodium bicarbonate is used in a process for cleaning paint called soda blasting. It can be administered to pools, spas, and garden ponds to raise pH levels . It has disinfectant and antiseptic properties , and it may be an effective fungicide against some organisms .

Because it can be used to absorb odors, it is a tried-and-true method of used booksellers. The baking soda will absorb the musty smell , leaving the books less odorous .

Sodium Bi Sulfate

1 . Sodium bisulfate, also **sodium hydrogen sulfate**, has the chemical formula Na HSO_4 . The anhydrous form is hygroscopic. Its melting point is poorly defined because it begins to decompose into sodium pyro sulfate and water before it reaches its melting points .

Solutions of sodium bisulfate are acidic , with a 1M solution having pH of 1.4. In some applications, such solutions can be used instead of sulfuric acid solution. For example, from a solution of sodium bisulfate and sodium acetate it is possible to distill acetic acid Sodium bisulfate solutions will also liberate CO_2 from most carbonates . Sodium bisulfate behaves, to some degree , as if it were a complex of sodium sulfate with sulfuric acid. This is evident if either the anhydrous form or the monohydrate come in contact with ethanol, which causes them to separate into those two components .

Other Names	Sodium acid sulfate Bisulfate of soda
Molecular Formula	Na HSO_4
Molar Mass	120 g / mol (anhydrous) 138 g / mol (monohydrate)
Appearance	White solid
Density	2.742 g / cm^3 (anhydrous) 1.8 g / cm^3 (mono hydrate)
Melting Point	58.5°C (mono hydrate) 315°C (anhydrous)
Boiling Point	decomposes to $Na_2S_2O_7$ (+ H_2O)
Solubility in Water	50 g / 100 mL (0° C) 100 g / 100 mL (100° C)
Solubility	Insoluble in ammonia decomposed by alcohol
Acidity (pK_a)	1.9
EU Index	Corrosive (**C**)

2 . Production

Sodium bisulfate is produced by two methods. One method involves mixing stoichiometric quantities of sodium hydroxide and sulfuric acid which react to form sodium bisulfate and water.

$$\mathrm{Na\,OH} + \mathrm{H_2SO_4} \rightarrow \mathrm{NaHSO_4} + \mathrm{H_2O}$$

A second production method involves reacting sodium chloride (salt) and sulfuric acid at elevated temperatures to produce sodium bisulfate and hydrogen chloride gas.

$$\mathrm{Na\,Cl} + \mathrm{H_2SO_4} \rightarrow \mathrm{NaHSO_4} + \mathrm{H\,Cl}$$

The liquid sodium bisulfate is sprayed and cooled so that it forms a solid bead. The hydrogen chloride gas is dissolved in water to produce hydrochloric acid as a useful byproduct of the reaction.

There are only two producers in the USA , Jones Hamilton Company, uses the sulfuric acid/sodium chloride process, which produces the anhydrous form. Jost Chemical uses the sodium hydroxide/sulfuric acid method, which produces the monohydrate.

3 . Uses

- Household cleaners .
- Silver pickling .
- To reduce alkalinity (lowering the pH) in swimming pools
- In pet foods
- As a preservative for soil and water samples in analytical laboratory analysis

Sodium Bromide

Contents

1 – Introduction :

Sodium bromide is an inorganic compound with the formula Na Br. It is a high-melting white, crystalline solid that resembles sodium chloride. It is a widely used source of the bromide ion and has many applications.

IUPAC name : Sodium Bromide	
Molecular formula	Na Br
Molar mass	103 g / mol
Appearance	White powder
Density	3.21 g/cm^3 (anhydrous) 2.18 g/cm^3 (di hydrate)
Melting point	747 °C (anhydrous) 36 °C (dihydrate)
Boiling point	1396 °C
Solubility in water	90.5 g / 100 mL (20 °C) 121 g / 100 mL (100 °C)
Solubility in methanol	16.7 g / 100 mL
Refractive index (n_D)	1.6459
EU Index	Not listed

Flash point	800 °C
LD_{50}	3500 mg/kg

2 - Synthesis , structure , reactions :

Na Br crystallizes in the same cubic motif as Na Cl , Na F and Na I . It is produced by treating sodium hydroxide with hydrogen bromide. The anhydrous salt crystallizes above 50.7 °C.

Sodium bromide can be used as a source of the chemical element bromine. This can be accomplished by treating an aqueous solution of Na Br with chlorine gas :

$$2\ NaBr + Cl_2 \rightarrow Br_2 + 2\ NaCl$$

3 – Applications :

Sodium bromide is the most useful inorganic bromide in industry .

3 – 1 – Medicine :

Also known as Sedoneural , sodium bromide has been used as a hypnotic , anticonvulsant, and sedative in medicine , widely used as an anticonvulsant and a sedative in the late 19th and early 20th centuries. Its action is due to the bromide ion, and for this reason potassium bromide is equally effective.

3 – 2 – Preparation of other bromine compounds :

Sodium bromide is widely used for the preparation of other bromides in organic synthesis and other areas. It is a source of the bromide nucleophile to convert alkyl chlorides to more reactive alkyl bromides by the Finkelstein reaction :

$$NaBr + RCl \rightarrow RBr + NaCl\ (R = \text{alkyl})$$

Once a large need in photography, but now shrinking, the photosensitive salt silver bromide is prepared using NaBr.

3 – 3 – Disinfectant :

Na Br is used in conjunction with chlorine as a disinfectant for swimming pools.

3 – 4 – Petroleum industry :

Sodium bromide is used to prepare dense fluids used in oil wells.

4 – Safety :

Na Br has a very low toxicity with an oral LD_{50} estimated at 5 g/kg for rats.

Sodium Carbonate

Contents

1 . Introduction :

Sodium carbonate (also known as **washing soda**, **soda crystals** or **soda ash**), Na_2CO_3, is a sodium salt of carbonic acid. It most commonly occurs as a crystalline heptahydrate, which readily effloresces to form a white powder, the monohydrate; and is domestically well known for its everyday use as a water softener. It has a cooling alkaline taste, and can be extracted from the ashes of many plants. It is synthetically produced in large quantities from table salt in a process known as the Solvay process.

Other Names	Soda ash Washing soda Soda crystals
Molecular formula	$Na_2 CO_3$
Molar Mass	106 g / mol (anhydrous) 124 g / mol (monohydrate) 286 g / mol (deca hydrate)
Appearance	White solid , hygroscopic
Density	2.54 g / cm^3 (anhydrous) 2.25 g / cm^3 (monohydrate) 1.46 g / cm^3 (deca hydrate)

Melting point	851 ° C (anhydrous) 100 ° C (decomp , mono hydrate) 34 ° C (decomp , deca hydrate)
Boiling Point	1600 ° C (anhydrous)
Solubility in Water	30 g / 100 ml (20 °C)
Solubility	Insoluble in alcohol , ethanol
Basicity (pK_b)	3.67
Refractive index (n_D)	1.495 (anhydrous) 1.420 (mono hydrate)
Flash point	Non - flammable

2 . Uses

The manufacture of glass is the most important use of sodium carbonate. When combined with sand (SiO_2) and calcium carbonate ($CaCO_3$) and heated to very high temperatures , then cooled very rapidly, glass is produced.

Sodium carbonate is also used as a relatively strong base in various settings. For example, sodium carbonate is used as a pH regulator to maintain stable alkaline conditions necessary for the action of the majority of developing agents. It is a common additive in municipal pools used to neutralize the acidic effects of chlorine and raise pH . In cooking, it is sometimes used in place of sodium hydroxide for lying, especially with German pretzels and lye rolls. These dishes are treated with a solution of an alkaline substance in order to change the pH of the surface of the food and thus improve browning.

In taxidermy, sodium carbonate added to boiling water will remove flesh from the skull or bones of trophies to create the " European skull mount " or for educational display in biological and historical studies.

In chemistry, it is often used as an electrolyte. This is because electrolytes are usually salt - based, and sodium carbonate acts as a very good conductor in the process of electrolysis. Additionally, unlike chloride ions which form chlorine gas, carbonate ions are not corrosive to the anodes. It is also used as a primary standard for acid - base titrations because it is solid and air - stable, making it easy to weigh accurately.

In domestic use, it is used as a water softener during laundry. It competes with the ions magnesium and calcium in hard water and prevents them from bonding with the detergent being used. Without using washing soda, additional detergent is needed to soak up the magnesium and calcium ions . Called **Washing Soda** , **Soda crystals** or **Sal Soda** in the detergent section of stores, it effectively removes oil, grease, and alcohol stains. Sodium carbonate is also used as a descaling agent in boilers such as found in coffee pots , espresso machines, etc.

In dyeing with fiber - reactive dyes, sodium carbonate (often under a name such as soda ash fixative or soda ash activator) is used to ensure proper chemical bonding of the dye with the fibers, typically before dyeing (for tie dyes) , mixed with the dye (for dye painting) , or after dyeing (for immersion dyeing)

2 . 1 . Other applications

Sodium carbonate is a food additive (E500) used as an acidity regulator, anti - caking agent, raising agent and stabilizer. It is one of the components of *kansui*, a solution of alkaline salts used to give ramen noodles their characteristic flavor and texture. Sodium carbonate is also used in the production of sherbet powder. The cooling and fizzing sensation results from the endothermic reaction between sodium carbonate and a weak acid, commonly citric acid, releasing carbon dioxide gas, which occurs when the sherbet is moistened by saliva.

Sodium carbonate is used by the brick industry as a wetting agent to reduce the amount of water needed to extrude the clay .

In casting, it is referred to as " bonding agent " and is used to allow wet alginate to adhere to gelled alginate .

Sodium carbonate is used to encapsulate and kill mold. When mixed with water and put in a spray bottle, it is sold for its anti-mold cleaning ability. It is also used to blast off mold from wood or other materials.

Sodium carbonate is used in toothpastes, where it acts as a foaming agent, an abrasive, and to temporarily increase mouth pH .

The crystalline form of washing soda can be used to induce vomiting in dogs. A tablespoon for large breeds is sufficient to force the animal to empty the contents of its stomach .

Sodium carbonate may be used for safely cleaning silver. First, aluminium foil is added to a glass or ceramic container, and covered with very hot water and some sodium carbonate. Silver items are dipped into this "bath" to clean them, making sure the silver makes contact with the aluminium foil. Finally, the silver is rinsed in water and let to dry.

3 . Occurrence

Sodium carbonate is soluble in water, but can occur naturally in arid regions, especially in the mineral deposits (*evaporites*) formed when seasonal lakes evaporate. Deposits of the mineral natron, natural sodium carbonate decahydrate, have been mined from dry lake bottoms in Egypt since ancient times, when natron was used in the preparation of mummies and in the early manufacture of glass. Sodium carbonate has three known forms of hydrates: sodium carbonate deca hydrate (natron), sodium carbonate hepta hydrate (not known in mineral form) and sodium carbonate mono hydrate (mineral thermo natrite). The anhydrous mineral form of sodium carbonate is quite rare and called natrite . Sodium carbonate also erupts from Tanzania's unique volcano Ol Doinyo Lengai, and probably erupted from other volcanoes in the past. All three mineralogical forms of sodium carbonate, as well as trona (tri sodium hydrogen di carbonate di hydrate) , are also known from ultra - alkaline pegmatitic rocks, i.e. from the Kola Peninsula.

4 . Production

4 . 1 . Mining

Trona, tri sodium hydrogen di carbonate di hydrate (Na_3 HCO_3 CO_3 · $2H_2O$) , is mined in several areas of the United States and provides nearly all the domestic sodium carbonate. Large natural deposits found in 1938, such as the one near Green River, Wyoming, have made mining more economical than industrial production in North America.

It is also mined out of certain alkaline lakes such as Lake Magadi in Kenya by using a basic dredging process and it is also self - regenerating so will never run out in its natural source.

4 . 2 . Barilla and kelp

Several "halophyte" (salt tolerant) plant species and seaweed species can be processed to yield an impure form of sodium carbonate, and these sources predominated in Europe and elsewhere until the early 19th Century. The land plants (typically glassworts or saltworts) or the seaweed (typically *Fucus* species) were harvested, dried, and burned. The ashes were then "lixiviated" (washed with water) to form an alkali solution. This solution was boiled dry to create the final product, which was termed "soda ash"; this very old name refers to the archetypal plant source for soda ash, which was the small annual shrub *Salsola soda* ("barilla plant").

The sodium carbonate concentration in soda ash varied very widely, from 2 - 3 % for the seaweed - derived form ("kelp"), to 30 % for the best barilla produced from saltwort plants in Spain. Plant and seaweed sources for soda ash, and also for the related alkali "potash," became increasingly inadequate by the end of the 18 th Century, and the search for commercially - viable routes to synthesizing soda ash from salt and other chemicals intensified .

4 . 3 . Leblanc process

In 1791, the French chemist Nicolas Leblanc patented a process for producing sodium carbonate from salt, sulfuric acid, limestone, and coal. First, sea salt (sodium chloride) was boiled in sulfuric acid to yield sodium sulfate and hydrogen chloride gas, according to the chemical equation

$$2\ NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2\ HCl$$

Next, the sodium sulfate was blended with crushed limestone (calcium carbonate) and coal, and the mixture was burnt, producing calcium sulfide.

$$Na_2SO_4 + CaCO_3 + 2\ C \rightarrow Na_2CO_3 + 2\ CO_2 + CaS$$

The sodium carbonate was extracted from the ashes with water, and then collected by allowing the water to evaporate.

The hydrochloric acid produced by the Leblanc process was a major source of air pollution, and the calcium sulfide byproduct also presented waste disposal issues. However, it remained the major production method for sodium carbonate until the late 1880s .

4 . 4 . Solvay process

In 1861, the Belgian industrial chemist Ernest Solvay developed a method to convert sodium chloride to sodium carbonate using ammonia. The Solvay process centered around a large hollow tower. At the bottom, calcium carbonate (limestone) was heated to release carbon dioxide:

$$CaCO_3 \rightarrow CaO + CO_2$$

At the top, a concentrated solution of sodium chloride and ammonia entered the tower. As the carbon dioxide bubbled up through it, sodium bicarbonate precipitated:

$$NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$$

The sodium bicarbonate was then converted to sodium carbonate by heating it, releasing water and carbon dioxide:

$$2\ NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

Meanwhile, the ammonia was regenerated from the ammonium chloride byproduct by treating it with the lime (calcium hydroxide) left over from carbon dioxide generation:

$$CaO + H_2O \rightarrow Ca\ (OH)_2$$
$$Ca\ (OH)_2 + 2\ NH_4Cl \rightarrow CaCl_2 + 2\ NH_3 + 2\ H_2O$$

Because the Solvay process recycles its ammonia, it consumes only brine and limestone, and has calcium chloride as its only waste product. This made it substantially more economical than the Leblanc process, and it soon came to dominate world sodium carbonate production. By 1900 , 90 % of sodium carbonate was produced by the Solvay process, and the last Leblanc process plant closed in the early 1920s.

4 . 5 . Hou's process

Developed by a Chinese chemist Hou Debang in 1930s, It is the same as the Solvay process in the first few steps. But, instead of treating the remaining solution with lime, carbon dioxide and ammonia is pumped into the solution, and sodium chloride is added until it is saturated at 40 °C. Then the solution is cooled down to 10 °C. Ammonium chloride precipitates and is removed by filtration, the solution is recycled to produce more sodium bicarbonate. Hou's Process eliminates the production of calcium chloride and the byproduct ammonium chloride can be used as a fertilizer.

Sodium chlorate

Contents

1 . Introduction :

Sodium chlorate is a chemical compound with the chemical formula ($NaClO_3$) . When pure, it is a white crystalline powder that is readily soluble in water. It is hygroscopic. It decomposes above 250 °C to release oxygen and leave sodium chloride.

Other Names	Sodium chlorate (V)
Molecular Formula	$Na\,Cl\,O_3$
Molar Mass	106 .5 g / mol
Appearance	White solid
Odor	Odorless
Density	2.5 g / cm^3
Melting Point	248 ° C
Boiling Point	~300 °C *decomp.*
Solubility in Water	101.0 g / 100 ml (20 °C)
Refractive index (n_D)	1.572
EU classification	Oxidant (**O**) Harmful (**Xn**) Dangerous for the environment (**N**)
Flash point	Non - flash

2 . Synthesis

Industrially, sodium chlorate is synthesized from the electrolysis of hot sodium chloride solution in a mixed electrode tank :

$$Na\,Cl + 3H_2O \rightarrow Na\,ClO_3 + 3H_2$$

It can also be synthesized by passing chlorine gas to a hot sodium hydroxide solution. It is then purified by crystallization.

3 . Uses

3 . 1 . Herbicides

Sodium chlorate is used as a non-selective herbicide. It is considered phytotoxic to all green plant parts . It can also kill through root absorption.

The herbicide is mainly used on non - crop land for spot treatment and for total vegetation control on roadsides, fenceways, ditches and suchlike.

Sodium chlorate is also used as a defoliant and desiccant for :

Cotton , safflower , corn , flax , peppers , soybeans , grain sorghum , southern peas , dry beans , rice and sunflowers .

If used in combination with atrazine, it increases the persistence of the effect. If used in combination with 2 , 4 - D , it improves performance of the material. Sodium chlorate has a soil-sterilant effect. Mixing with other herbicides in aqueous solution is possible to some extent, so long as they are not susceptible to oxidation.

3 . 2 . Chemical oxygen generation

Chemical oxygen generator for example in commercial aircraft provide emergency oxygen to passengers to protect them from drops in cabin pressure by catalytic decomposition of sodium chlorate. The catalyst is normally some iron powder . Barium peroxide (BaO_2) is used to absorb the chlorine which is a minor product in the decomposition .

Sodium chlorate is used in some aircraft as a source of supplemental oxygen. Iron powder is mixed with sodium chlorate and ignited by a charge which is activated by pulling on the emergency mask. The reaction produces more oxygen than is required for combustion.

4 . Toxicity in humans

Due to its oxidative nature, sodium chlorate can be very toxic if ingested. The oxidative effect on hemoglobin leads to methaemo globin formation, which is followed by denaturation of the globin protein and a cross-linking of erythrocyte membrane proteins with resultant damage to the membrane enzymes. This leads to increased permeability of the membrane, and severe hemolysis . The denaturation of hemoglobin overwhelms the capacity of the G6 P D metabolic pathway. In addition, this enzyme is directly denatured by chlorate reducing its activity.

Therapy with ascorbic acid and methylene blue may be effective. However, since methylene blue requires the presence of NADPH that requires normal functioning of G6 P D system, it is less effective than in other conditions characterized by hemoglobin oxidation. Acute severe hemolysis results, with multi-organ failure, including DIC and renal failure.

5 . Formulations

Sodium chlorate comes in dust, spray and granule formulations. There is a risk of fire and explosion in dry mixtures with other substances, especially organic materials, that is other herbicides, sulfur, phosphorus, powdered metals, strong acids. Particularly when mixed with sugar it has explosive properties. If accidentally mixed with one of these substances do not store inside your home, garage etc .

Marketed formulations contain a fire depressant, but this has little effect if deliberately ignited. Most commercially available chlorate weed killers contain approximately 53 % sodium chlorate with the balance being a fire depressant such as sodium meta borate or ammonium phosphates.

6 . Trade Names

The active ingredient sodium chlorate is found in a variety of commercial herbicides . Some trade names for products containing sodium chlorate include :

Atlacide , Defol , De - Fol-Ate , Drop – Leaf , Fall , Harvest – Aid , Kusatol , Leafex , and Tumbleaf .

The compound may be used in combination with other herbicides such as :

atrazine, 2 , 4 – D , bromacil , diuron , and sodium meta borate .

In the United Kingdom there are many brands of Sodium Chlorate Weed killers . Two well known ones are Doff and Wilkinsons . In Italy it is found under the Zapi brand , sold as " Grass killer Liquid " .

Sodium chloride

Contents

1 . Introduction :

Sodium chloride, also known as **common salt** , **table salt** , or halite, is an ionic compound with the formula $NaCl$. Sodium chloride is the salt most responsible for the salinity of the ocean and of the extracellular fluid of many multi cellular organisms. As the major ingredient in edible salt, it is commonly used as a condiment and food preservative.

Other Names	Common salt halite , table salt , rock salt
Molecular Formula	$NaCl$
Molar Mass	58 .5 g / mol
Appearance	Colorless / white crystalline solid
Odor	Odorless
Density	2.165 g / cm^3
Melting Point	801 °C
Boiling Point	1465 °C

Solubility in Water	35.6 g /100 mL (0 ° C) 35.9 g / 100 mL (25 ° C) 39.1 g / 100 mL (100 ° C)
Solubility	Soluble in glycerol , ethylene glycol , formic acid , Insoluble in HCl
Refractive index (n_D)	1.5442 (589 nm)
Crystal structure	Cubic , cF8
Flash point	Non - flammable
LD_{50}	3000 – 8000 mg / kg (oral in rats , mice , rabbits)
Thermodynamic data	Phase behaviour : Solid , liquid , gas

Solubility of NaCl in various solvents (g NaCl / 100 g of solvent at 25 ° C)	
H_2O	36
Liquid Ammonia	3.02
Methanol	1.4
Formic acid	5.2
Formamide	9.4

2 . Production and use :

Salt is currently mass - produced by evaporation of seawater or brine from other sources, such as brine wells and salt lakes, and by mining **rock salt**, called halite. In 2002, world production was estimated at 210 million metric tones , the top five producers (in million tonnes) being the United States (40.3) , China (32.9), Germany (17.7), India (14.5) and Canada (12.3).[2]

As well as the familiar uses of salt in cooking, salt is used in many applications, from manufacturing pulp and paper, to setting dyes in textiles and fabric, to producing soaps, detergents, and other bath products. It is the major source of industrial chlorine and sodium hydroxide, and used in almost every industry.

Sodium chloride is some times used as a cheap and safe desiccant because it appears to have hygroscopic properties, making salting an effective method of food preservation historically. Even though more effective desiccants are available , few are safe for humans to ingest .

2 . 1 . Synthetic uses

Sodium chloride is also the raw material used to produce chlorine which itself is required for the production of many modern materials including PVC and pesticides. Industrially, elemental chlorine is usually produced by the electrolysis of sodium chloride dissolved in water. Along with chlorine, this chloralkali process yields hydrogen gas and sodium hydroxide, according to the chemical equation

$$2NaCl + 2H_2O \rightarrow Cl_2 + H_2 + 2NaOH$$

Sodium metal is produced commercially through the electrolysis of liquid sodium chloride. This is now done in a Down's cell in which sodium chloride is mixed with calcium chloride to lower the melting point below 700 °C. As calcium is more electropositive than sodium, no calcium will be formed at the cathode. This method is less expensive than the previous method of electrolyzing sodium hydroxide.

Sodium chloride is used in other chemical processes for the large-scale production of compounds containing sodium or chlorine. In the Solvay process, sodium chloride is used for producing sodium carbonate and calcium chloride. In the Mannheim process and in the Hargreaves process, it is used for the production of sodium sulfate and hydrochloric acid.

2 . 2 . Biological uses

Many microorganisms cannot live in an overly salty environment: water is drawn out of their cells by osmosis. For this reason salt is used to preserve some foods, such as smoked bacon or fish. It can also be used to detach leeches that have attached themselves to feed. It is also used to disinfect wounds.

2 . 3 . Optical uses

Pure NaCl crystal is an optical compound with a wide transmission range from 200 nm to 20 um. It was often used in the infrared spectrum range and it is still used sometimes.

NaCl crystal is soft, hygroscopic and inexpensive. This limits its application to protected environment or for short term uses (proto typing) . Exposed to free air Na Cl optics will " rot ".

Today tougher crystals like Zn Se are used instead of Na Cl (for the IR spectral range) .

2 . 4 . House hold uses

Since at least medieval times, people have used salt as a cleansing agent rubbed on household surfaces. It is also used in many brands of shampoo, and popularly to de - ice drive ways and patches of ice.

3 . Biological functions

In humans, a high - salt intake has long been known to generally raise blood pressure, especially in certain individuals. More recently, it was demonstrated to attenuate nitric oxide production. Nitric oxide (NO) contributes to vessel home ostasis by inhibiting vascular smooth muscle contraction and growth, platelet aggregation, and leukocyte adhesion to the endothelium .

4 . Crystal structure

Sodium chloride forms crystals with face - centered cubic symmetry. In these, the larger chloride ions, shown to the right as green spheres, are arranged in a cubic close-packing, while the smaller sodium ions, shown to the right as silver spheres, fill all the cubic gaps between them. Each ion is surrounded by six ions of the other kind; the surrounding ions are located at the vertices of a perfect octahedron.

This same basic structure is found in many other minerals and is commonly known as the halite or rock-salt crystal structure. It can be represented as a face - centered cubic (fcc) lattice with a two atom basis. The first atom is located at each lattice point, and the second atom is located half way between lattice points along the fcc unit cell edge .

It is held together with an ionic bond and electrostatic forces.

5 . Road salt

While salt was once a scarce commodity in history, industrialized production has now made salt plentiful. Approximately 51 % of world output is now used by cold countries to de - ice roads in winter, both in grit bins and spread by winter service vehicles. Calcium chloride is preferred over sodium chloride, since $Ca\ Cl_2$ releases energy upon forming a solution with water, heating any ice or snow it is in contact with. It also lowers the freezing point, depending on the concentration. Na Cl does not release heat upon solution; however, it does lower the freezing point. It is also more readily available and does not have any special handling or storage requirements, unlike calcium chloride. The salinity (S) of water is measured as grams salt per kilogram (1000g) water , and the freezing temperatures are as follows.

S (g / kg)	0	10	20	24.7	30	35
T (freezing) (C)	0	- 0.5	- 1.08	- 1.33	- 1.63	- 1.91

5 . 1 . Additives

Most table salt sold for consumption today is not pure sodium chloride. In 1911 magnesium carbonate was first added to salt to make it flow more freely . In 1924 trace amounts of iodine in form of sodium iodide, potassium iodide or potassium iodate were first added, to reduce the incidence of simple goiter .

Salt for de-icing in the UK typically contains sodium hexa cyano ferrate (II) at less than 100 ppm as an anti- caking agent. In recent years this additive has also been used in table salt .

5 . 2 . Common chemicals

Chemicals used in de - icing salts are mostly found to be sodium chloride (Na Cl) or calcium chloride ($CaCl_2$) . Both are similar and are effective in de-icing roads. When these chemicals are produced, they are mined/made, crushed to fine granules, then treated with an anti - caking agent. Adding salt lowers the freezing point of the water, which allows the liquid to be stable at lower temperatures and allows the ice to melt. Alternative de-icing chemicals have also been used. Chemicals such as calcium magnesium acetate and potassium formate are being produced. These chemicals have few of the negative chemical effects on the environment commonly associated with NaCl and $CaCl_2$.

Sodium chlorite

1 . Introduction :

Sodium chlorite :

Is a chemical compound used in the manufacture of paper .

Other Names	Chlorous acid sodium salt , Textone
Molecular Formula	$Na\ ClO_2$
Molar Mass	90.5 g / mol
Appearance	White solid
Density	2.5 g / cm^3 , solid
Melting Point	180 – 200 ° C *decomp* .
Solubility in Water	39 g / 100 ml (17 °C)
Flash point	Non - flammable

2 . **Manufacture**

The free acid, chlorous acid , $H\ ClO_2$, is only stable at low concentrations. Since it cannot be concentrated, it is not a commercial product. However, the corresponding sodium salt, sodium chlorite, $Na\ ClO_2$ is stable and inexpensive enough to be commercially available. The corresponding salts of heavy metals (Ag^+ , Hg^+ , Tl^+, Pb^{2+}, and also Cu^{2+} and NH_4^+) decompose explosively with heat or shock .

Sodium chlorite is derived indirectly from sodium chlorate , $Na\ ClO_3$. First, the explosively unstable gas chlorine dioxide, ClO_2 is produced by reducing sodium chlorate in a strong acid solution with a suitable reducing agent (for example, sodium chloride, sulfur dioxide, or hydrochloric acid) . The chlorine dioxide is then absorbed into an alkaline solution and reduced with hydrogen peroxide, $H_2\ O_2$ yielding sodium chlorite .

3 . Usage

The main application of sodium chlorite is the generation of chlorine dioxide for bleaching and stripping of textiles, pulp, and paper. It is also used for disinfection of a few municipal water treatment plants after conversion to chlorine dioxide. An advantage in this application, as compared to the more commonly used chlorine, is that tri halo methanes are not produced from organic contaminants. Sodium chlorite, $NaClO_2$ also finds application as a component in therapeutic rinses, mouthwashes, toothpastes and gels, mouth sprays, chewing gums and lozenges, and also in contact lens cleaning solution under the trade name purite. Under the brand name Oxine it is used for sanitizing air ducts and HVAC / R systems and animal containment areas (walls , floors , and other surfaces) .

In organic synthesis, sodium chlorite is frequently used for the oxidation of aldehydes to carboxylic acids. The reaction is usually performed in the presence of a chlorine scavenger.

Sodium chlorite , like many oxidizing agents, should be protected from inadvertent contamination by organic materials to avoid the formation of an explosive mixture .

Recently, sodium chlorite has been used as an oxidizing agent to convert alkyl furans to the corresponding 4 – oxo – 2 - alkenoic acids in a simple one pot synthesis .

Sodium Di Hydrogen Phosphate

Introduction :

Mono Sodium Phosphate (NaH_2PO_4) , also known as anhydrous monobasic sodium phosphate, is used as a laxative and, in combination with other sodium phosphates, as a pH buffer.

The pKa is 7.20. The sodium chloride equivalent value, or E-Value, is 0.49. It is soluble in 4.5 parts water. It is most commonly available as colorless crystals or a white powder. It is insoluble in alcohol.

IUPAC name	Sodium Di Hydrogen Phosphate
Molecular formula	Na H_2PO_4
Molar mass	120 g / mol (anhydrous) 246 g / mol (hepta hydrate)
Appearance	White powder or crystals

Sodium fluoride

Contents

1 – Introduction :

Sodium fluoride is an inorganic chemical compound with the formula **NaF.** A colorless solid, it is a source of the fluoride ion in diverse applications. Sodium fluoride is less expensive and less hygroscopic than the related salt potassium fluoride.

IUPAC Name : Sodium Fluoride	
Other Names : Florocid	
UN number	1690
Molecular Formula	Na F
Molar mass	42 g / mol
Appearance	White solid
Odor	Odorless
Density	2.558 g / cm^3
Melting point	993 °C
Boiling point	1695 °C
Solubility in water	0.5 M (20 °C)
Solubility	Soluble in HF
EU classification	Toxic (**T**) Irritant (**Xi**)
Flash point	Non - flammable
LD_{50}	52–200 mg/kg (oral in rats, mice, rabbits)

2 - Structure , general properties , occurrence

Sodium fluoride is an ionic compound, dissolving to give separated Na^+ and F^- ions. Like sodium chloride, it crystallizes in a cubic motif where both Na^+ and F^- occupy octahedral coordination sites.

The mineral form of Na F , villiaumite , is moderately rare. It is known from plutonic nepheline syenite rocks.

3 – Production :

Na F is prepared by neutralizing hydrofluoric acid or hexa fluoro silicic acid (H_2SiF_6) , byproducts of the production of superphosphate fertilizer. Neutralizing agents include sodium hydroxide and sodium carbonate. Alcohols are sometimes used to precipitate the Na F :

$$HF + NaOH \rightarrow NaF + H_2O$$

From solutions containing HF, sodium fluoride precipitates as the bi fluoride salt $NaHF_2$. Heating the latter releases HF and gives Na F .

$$HF + NaF \rightleftharpoons NaHF_2$$

In a 1986 report, the annual worldwide consumption of NaF was estimated to be several million tonnes.[6]

4 - Applications :

Fluoride salts are used to enhance the strength of teeth by the formation of Fluor apatite , a naturally occurring component of tooth enamel . Although sodium fluoride is also used to fluoridate water and, indeed, is the standard by which other water-fluoridation compounds are gauged, hexa fluoro silicic acid (H_2SiF_6) and its salt sodium hexa fluoro silicate (Na_2SiF_6) are more commonly used additives in the U.S . Toothpaste often contains sodium fluoride to prevent cavities . Alternatively, sodium fluoride is used as a cleaning agent, e.g. as a "laundry scour". A variety of specialty chemical applications exist in synthesis and extractive metallurgy. It reacts with electrophilic chlorides including acyl chlorides, sulfur chlorides, and phosphorus chloride. Like other fluorides, sodium fluoride finds use in desilylation in organic synthesis. The fluoride is the reagent for the synthesis of fluorocarbons.

Sodium fluoride is sold in tablets for cavity prevention.

In medical imaging, fluorine -18 - labelled sodium fluoride is used in positron emission tomography (PET). Relative to conventional bone scintigraphy carried out with gamma cameras or SPECT systems, PET offers more sensitivity and spatial resolution. A disadvantage of PET is that fluorine -18 labelled sodium fluoride is less widely available than conventional technetium – 99 m - labelled radiopharmaceuticals.

5 - Safety

The lethal dose for a 70 kg human is estimated at 5 – 10 g . Sodium fluoride is classed as toxic by both inhalation (of dusts or aerosols) and ingestion. In high enough doses, it has been shown to affect the heart and circulatory system. In the higher doses used to treat osteoporosis, plain sodium fluoride can cause pain in the legs and incomplete stress fractures when the doses are too high; it also irritates the stomach, sometimes so severely as to cause ulcers. Slow-release and enteric-coated versions of sodium fluoride do not have gastric side effects in any significant way, and have milder and less frequent complications in the bones.[13] In the lower doses used for water fluoridation, the only clear adverse effect is dental fluorosis , which can alter the appearance of children's teeth during tooth development; this is mostly mild and is unlikely to represent any real effect on aesthetic appearance or on public health.

Sodium hydrogen Sulfide

Contents

1 – Introduction :

Sodium hydrosulfide is the chemical compound with the formula Na HS . This compound is the product of the half neutralization of hydrogen sulfide with sodium hydroxide. NaHS is a useful reagent for the synthesis of organic and inorganic sulfur compounds. It is a colorless solid that typically smells like H_2S due to hydrolysis by atmospheric moisture. In contrast with sodium sulfide, Na_2S, which is insoluble in organic solvents, Na HS , being a 1:1 electrolyte, is more soluble. Alternatively, in place of Na HS , H_2S can be treated with an organic amine to generate an ammonium salt. Solutions of HS^- are sensitive to oxygen, converting mainly to poly sulfides , indicated by the appearance of yellow.

IUPAC name :	Sodium hydrosulfide
Other Name	Sodium bi sulfide , Sodium hydrogen sulfide
Molecular Formula	Na HS
Molar Mass	56 g / mol
Appearance	white solid , deliquescent
Density	1.79 g / cm^3
Melting point	52 – 54 ºC (hydrate) 350 ºC (anhydrous)
EU Index	Toxic
Main hazards	Flammable solid , stench , releases hydrogen sulfide
Flash point	90 ºC

2 - Structure and properties :

Crystalline Na HS undergoes two phase transitions. At temperatures above 360 K, Na HS adopts the Na Cl structure, which implies that the HS^- behaves as a spherical anion due to its rapid rotation leading to equal occupancy of eight equivalent positions. Below 360 K, a rhombohedral structure forms, and the HS^- sweeps out a discoidal shape. Below 114 K, the structure becomes monoclinic. The analogous rubidium and potassium compounds behave similarly.

Na HS has a relatively low melting point of 350 °C. In addition to the aforementioned anhydrous forms, it can be obtained as two different hydrates, $NaHS \cdot 2H_2O$ and $NaHS \cdot 3H_2O$. These three species are all colorless and behave similarly, but not identically.

3 – Preparation :

The usual laboratory synthesis entails treatment of sodium methoxide , Na O Me , with hydrogen sulfide :

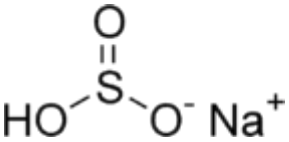
$$\text{Na O Me} + H_2S \rightarrow \text{Na HS} + \text{Me OH}$$

Industrially, Na OH is employed as the base. The quality of the NaHS can be assayed by iodo metric titration, exploiting the ability of HS^- to reduce I_2 .

4 - Applications ;

Thousands of tons of Na HS are produced annually. Its main uses are in paper manufacture as a make up chemical for sulfur used in the Kraft process , as a flotation agent in copper mining where it is used to activate oxide mineral species, and in the leather industry for the removal of hair from hides.

Sodium Hydrogen Sulfite



Contents

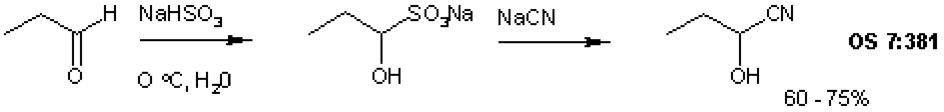
1 – Introduction :

Sodium Bi Sulfite (**Sodium Hydrogen Sulfite**) is a chemical compound with the chemical formula $NaHSO_3$. Sodium bi sulfite is a food additive with E number E222. This salt of bisulfite can be prepared by bubbling sulfur dioxide in a solution of sodium carbonate in water. Sodium bi sulfite in contact with chlorine bleach (aqueous solution of sodium hypochlorite) will release harmful fumes.

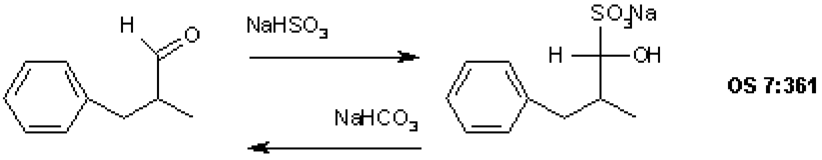
IUPAC Name : Sodium hydrogen sulfite	
Other Names : E222	
Molecular formula	$Na\ H\ SO_3$
Molar mass	104 g / mol
Appearance	White solid
Density	1.48 g / cm^3
Melting point	150 °C
Solubility in water	42 g / 100 mL
EU classification	Harmful (**Xn**)
Flash point	Non - flammable

2 - Uses in chemistry :

In organic chemistry sodium bi sulfite has several uses. It forms a bi sulfite adduct with aldehyde groups and with certain cyclic ketones to a sulfonic acid.



This reaction has limited synthetic value (s) but it is used in purification procedures. Contaminated aldehydes in a solution precipitate as the bi sulfite adduct which can be isolated by filtration. The reverse reaction takes place in presence of a base such as sodium bicarbonate or sodium hydroxide and the bisulfite is liberated as sulfur dioxide.



Examples of such procedures are described for benzaldehyde , tetralone , citral , the ethyl ester of pyruvic acid and glyoxal. In the ring - expansion reaction of cyclo hexanone with diazald , the bi sulfite reaction is reported to be able to differentiate between the primary reaction product cyclo heptanone and the main contaminant cyclo octanone.

The other main use of sodium bisulfite is as a mild reducing agent in organic synthesis in particular in purification procedures. It can efficiently remove traces or excess amounts of chlorine, bromine, iodine, hypochlorite salts, osmate esters, chromium trioxide and potassium permanganate.

A third use of sodium bi sulfite is as a decoloration agent in purification procedures because it can reduce strongly coloured oxidizing agents, conjugated alkenes and carbonyl compounds.

Sodium bi sulfite is also the key ingredient in the Bucherer reaction. In this reaction an aromatic hydroxyl group is replaced by an aromatic amine group and vice versa because it is a reversible reaction. The first step in this reaction is an addition reaction of sodium bi sulfite to an aromatic double bond. The Bucherer carbazole synthesis is a related organic reaction.

3 - Uses in food :

While the related compound, sodium meta bi sulfite , is used in almost all commercial wines to prevent oxidation and preserve flavor, sodium bi sulfite is sold by some home winemaking suppliers for the same purpose . In fruit canning, sodium bi sulfite is used to prevent browning (caused by oxidation) and to kill microbes.

In the case of wine making, sodium bi sulfite releases sulfur dioxide gas when added to water or products containing water. The sulfur dioxide kills yeasts, fungi, and bacteria in the grape juice before fermentation. When the sulfur dioxide levels have subsided (about 24 hours), fresh yeast is added for fermentation.

It is later added to bottled wine to prevent the formation of vinegar if bacteria are present, and to protect the color, aroma and flavor of the wine from oxidation, which causes browning and other chemical changes. The sulfur dioxide quickly reacts with oxidation by-products and prevents them from causing further deterioration.

Sodium bi sulfite is also added to leafy green vegetables in salad bars and elsewhere, to preserve apparent freshness, under names like Leaf Green . The concentration is sometimes high enough to cause severe allergic reactions .

In the 1980s, sodium bi sulfite was banned from use on raw fruits and vegetables in the United States following the deaths of 13 people who unknowingly consumed produce treated with excessive amounts of the substance .

4 - Industrial uses :

Sodium bi sulfite is a common reducing agent in the chemical industries. As it readily reacts with dissolved oxygen :

$$2\ NaHSO_3 + O_2 \rightarrow 2\ NaHSO_4$$

It is usually added to large piping systems to prevent oxidative corrosion. In biochemical engineering applications, it is helpful to maintain anaerobic conditions within a reactor. Sodium bi sulfite should not be confused with sodium bisulfate which is used as a pH lowering chemical for swimming pools.

In waste water treatment , sodium bi sulfite is often added following disinfection with a chlorine solution to neutralize the residual chlorine before discharging the treated effluent.

5 - Bi sulfite DNA sequencing

The chemical reaction that underlies the bi sulfite - mediated conversion of cytosine to uracil.

Sodium bi sulfite is used in the analysis of methylation status of cytosines in DNA. In this technique, sodium bi sulfite deaminates cytosine into uracil , but does not affect 5- methyl cytosine , a methylated form of cytosine with a methyl group attached to carbon 5.

When the bi sulfite - treated DNA is amplified via polymerase chain reaction, the uracil is amplified as thymine and the methylated cytosines are amplified as cytosine. DNA sequencing techniques are then used to read the sequence of the bi sulfite - treated DNA. Those cytosines that are read as cytosines after sequencing represent methylated cytosines, while those that are read as thymines represent un methylated cytosines in the genomic DNA.

Sodium hydroxide

Contents :

1 . Introduction :

Sodium hydroxide (**Na OH**) , also known as **lye** and **caustic soda** , is a caustic metallic base. Sodium hydroxide forms a strong alkaline solution when dissolved in a solvent such as water. However, only the hydroxide ion is basic. It is used in many industries, mostly as a strong chemical base in the manufacture of pulp and paper, textiles, drinking water, soaps and detergents and as a drain cleaner. Worldwide production in 1998 was around 45 million tones. Sodium hydroxide is a common base in chemical laboratories.

Pure sodium hydroxide is a white solid ; available in pellets, flakes, granules and as a 50 % saturated solution. It is hygroscopic and readily absorbs water from the air, so it should be stored in an airtight container. It is very soluble in water with liberation of heat. It also dissolves in ethanol and methanol, though it exhibits lower solubility in these solvents than potassium hydroxide. It is insoluble in ether and other non-polar solvents. A sodium hydroxide solution will leave a yellow stain on fabric and paper.

Other Names	Caustic soda Lye
Molecular Formula	Na OH
Molar Mass	40 g / mol
Appearance	White solid , hygroscopic
Density	2.13 g / cm^3
Melting Point	323 °C
Boiling Point	1388 °C
Solubility in Water	111 g / 100 ml (20 °C)
Solubility in Ethanol	13.9 g / 100 mL
Solubility in Methanol	23.8 g / 100 mL
Solubility in glycerol	Soluble
Refractive index (n_D)	1.412
EU classification	Corrosive (**C**)
Flash point	Non - Flammable

2 . Physical properties

Δ H° dissolution for diluted aqueous - 44.45 kJ / mol ;

From aqueous solutions at 12.3 - 61.8° C, it crystallizes in mono hydrate , with a melting point 65.1 °C and density of 1.829 g / cm 3 ;

Δ H° form - 734.96 kJ / mol ;

Mono hydrate from - 28 to - 24°C ;

Hepta hydrate from - 24 to - 17.7°C;

Penta hydrate from - 17.7 to - 5.4°C;

Tetra hydrate (α - changed) , at -5 , 4 - 12.3° C Also know meta stable β – Na OH 4* H_2O. Which above 61.8° C are crystallized.

3 . Chemical properties

Sodium hydroxide is completely ionic, containing sodium cations and hydroxide anions. The hydroxide anion makes sodium hydroxide a strong base which reacts with acids to form water and the corresponding salts, e.g., with hydrochloric acid, sodium chloride is formed:

$$\text{Na OH (aq)} + \text{H Cl (aq)} \rightarrow \text{Na Cl (aq)} + H_2O\text{ (l)}$$

In general such neutralization reactions are represented by one simple net ionic equation:

$$OH^- \text{(aq)} + H_3O^+ \text{(aq)} \rightarrow 2H_2O$$

This type of reaction with a strong acid, releases heat, and hence is referred to as exothermic. Such acid - base reactions can also be used for titrations, which is a common method to determine the concentration of acids. Another type of reaction that sodium hydroxide is involved in is with acidic oxides. The reaction of carbon dioxide has already been mentioned, but other acidic oxides such as sulfur dioxide (SO_2) also react completely. Such reactions are often used to "scrub" harmful acidic gases (like SO_2 and H_2S) and prevent their release into the atmosphere.

$$2\text{Na OH} + CO_2 \rightarrow Na_2\,CO_3 + H_2O$$

Sodium hydroxide slowly reacts with glass to form sodium silicate, so glass joints and stopcocks exposed to NaOH have a tendency to " freeze ". Flasks and glass-lined chemical reactors are damaged by long exposure to hot sodium hydroxide, and the glass becomes frosted. Sodium hydroxide does not attack iron since iron does not have amphoteric properties. A few transition metals, however, may react with sodium hydroxide in a vigorous way .

In 1986 an aluminium road tanker in the UK was mistakenly used to transport 25 % sodium hydroxide solution, causing pressurization of the contents and damage to the tanker. The pressurization was due to the hydrogen gas which is produced in the reaction between sodium hydroxide and aluminium :

$$2Al_{(s)} + 6Na\,OH_{(aq)} \rightarrow 3H_{2\,(g)} + 2Na_3\,AlO_{3\,(aq)}$$

Unlike Na OH , the hydroxides of most metals are insoluble, and therefore sodium hydroxide can be used to precipitate metal hydroxides. One such hydroxide is aluminium hydroxide, used as a gelatinous floc to filter out particulate matter in water treatment. Aluminium hydroxide is prepared at the treatment plant from aluminium sulfate by reacting with Na OH . This reaction is highly profitable, and is hence an important synthesis reaction.

Sodium hydroxide reacts readily with carboxylic acids to form their salts and is even a strong enough base to form salts with phenols. Na OH can be used for the base - driven hydrolysis of esters (as in saponification) , amides and alkyl halides. However, the limited solubility of Na OH in organic solvents means that the more soluble KOH is often preferred.

4 . **Manufacture**

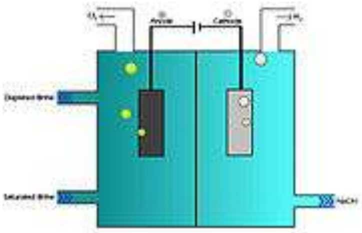
In 1998, total world production was around 45 million tonnes. North America and Asia collectively contributed around 14 million tonnes, while Europe produced around 10 million tones. In the United States, the major producer of sodium hydroxide is the Dow Chemical Company, which has annual production around 3.7 million tonnes from sites at Freeport , Texas, and Plaquemine, Louisiana. Other major US producers include Oxychem , PPG , Olin , Pioneer Companies (which was purchased by Olin) , Inc. (PIONA) , and Formosa. All of these companies use the chlor alkali process .

4 . 1 Methods of production

Sodium hydroxide is produced (along with chlorine and hydrogen) via the chlor alkali process. This involves the electrolysis of an aqueous solution of sodium chloride. The sodium hydroxide builds up at the cathode, where water is reduced to hydrogen gas and hydroxide ion :

$$2Na^+ + 2H_2O + 2e^- \rightarrow H_2 + 2NaOH$$

To produce Na OH it is necessary to prevent reaction of the NaOH with the chlorine. This is typically done in one of three ways, of which the membrane cell process is economically the most viable.



Basic membrane cell used in the electrolysis of brine.

- **Mercury cell process** (also called the Castner - Kellner process) – Sodium ions are reduced to sodium metal, which forms an amalgam with a mercury cathode; this sodium is then reacted with water to produce Na OH . There have been concerns about mercury releases, although modern plants claim to be safe in this regard .
- **Diaphragm cell process** – uses a steel cathode, and the reaction of Na OH with Cl_2 is prevented using a porous diaphragm, often made of asbestos fibers. In the diaphragm cell process the anode area is separated from the cathode area by a permeable diaphragm. The brine is introduced into the anode compartment and flows through the diaphragm into the cathode compartment. A diluted caustic brine leaves the cell. The sodium hydroxide must usually be concentrated to 50 % and the salt removed. This is done using an evaporative process with about three tones of steam per tone of sodium hydroxide. The salt separated from the caustic brine can be used to saturate diluted brine. The chlorine contains oxygen and is purified by liquefaction and evaporation .
- **Membrane cell process** – similar to the diaphragm cell process, with a Nafion membrane to separate the cathode and anode reactions. Only sodium ions and a little water pass

through the membrane. It produces a higher quality of NaOH. Of the three processes, the membrane cell process requires the lowest consumption of electric energy and the amount of steam needed for concentration of the caustic is relatively small (less than one tone per tone of sodium hydroxide) .

An older method for sodium hydroxide production was the Leblanc process, which produced sodium carbonate, followed by roasting to create carbon dioxide and sodium oxide. This method is still occasionally used. It helped establish sodium hydroxide as an important commodity chemical.

The LeBlanc process was superseded by the Solvay process in the late 19th century .

5 . Uses

Sodium hydroxide is the principal strong base used in the chemical industry. In bulk it is most often handled as an aqueous solution, since solutions are cheaper and easier to handle. It is used to drive chemical reactions and also for the neutralization of acidic materials. It can be used also as a neutralizing agent in petroleum refining. It is sometimes used as a cleaner.

5 . 1 . Alumina production

In the Bayer process, sodium hydroxide is used in the refining of alumina containing ore (bauxite) to produce alumina (aluminium oxide) which is the raw material used to produce aluminium metal via the smelting process .

5 . 2 . Soap production

Sodium hydroxide was traditionally used in soap making (cold process soap , saponification) .

5 . 3 . Oil drilling :

Sodium hydroxide is a component of drilling mud, used to increase viscosity which prevents the weighted material from settling

5 . 4 . Paper making :

Sodium hydroxide was also widely used in making paper. Along with sodium sulfide, Na OH is a key component of the white liquor solution used to separate lignin from cellulose fibers in the Kraft process. It also plays a key role in several later stages of the process of bleaching the brown pulp resulting from the pulping process. These stages include oxygen delignification, oxidative extraction, and simple extraction, all of which require a strong alkaline environment with a pH > 10.5 at the end of the stages.

5 . 5 . Biodiesel

For the manufacture of biodiesel, sodium hydroxide is used as a catalyst for the trans esterification of methanol and tri glycerides. This only works with anhydrous sodium hydroxide, because combined with water the fat would turn into soap, which would be tainted with methanol. It is used more often than potassium hydroxide because it is cheaper and a smaller quantity is needed.

5 . 6 . Aluminium etching

Strong bases attack aluminium . Sodium hydroxide reacts with aluminium and water to create hydrogen gas. The aluminium takes the oxygen atom from sodium hydroxide (Na OH) , which in turn takes the oxygen atom from the water, and releases the two hydrogen atoms. In this reaction, sodium hydroxide acts as an active catalyst. This reaction can be useful in etching through a resist, removing anodizing, or converting a polished surface to a satin-like finish, but without further passivation such as anodizing or alodining the surface may become degraded, either under normal use or in severe atmospheric conditions.

5 . 7 . Cleaning agent

Sodium hydroxide is frequently used as an industrial cleaning agent where it is often called " caustic ". It is added to water, heated, and then used to clean the process equipent , storage tanks, etc. It can dissolve grease, oils, fats and protein based deposits. The sodium hydroxide solution might also be added surfactants to stabilize dissolved substances to prevent redeposition . A sodium hydroxide soak solution is used as a powerful degreaser on stainless and glass bakeware. It is also the most common ingredient in oven cleaners .

5 . 8 . Paint stripper :

A solution of sodium hydroxide in water was traditionally used as the most common paint stripper on wooden objects. Due to its caustic nature and the fact that it can damage the wood surface raising the grain and staining the color, its use has become less common.

5 . 9 . Tissue digestion :

This is a process that was used with farm animals at one time. This process involves the placing of a carcass into a sealed chamber, which then puts the carcass in a mixture of sodium hydroxide and water, which breaks chemical bonds keeping the body intact. This eventually turns the body into a coffee-like liquid, and the only solid that remains are bone hulls, which could be crushed between one's fingertips. It is also of note that sodium hydroxide is frequently used in the process of decomposing road kill dumped in landfills by animal disposal contractors .

Sodium hydroxide has also been used by serial killers and other criminals to dispose of their victims' bodies .

5 . 10 . Food preparation :

Food uses of sodium hydroxide include washing or chemical peeling of fruits and vegetables, chocolate and cocoa processing, caramel coloring production, poultry scalding, soft drink processing, and thickening ice cream. Olives are often soaked in sodium hydroxide to soften them, while pretzels and German lye rolls are glazed with a sodium hydroxide solution before baking to make them crisp. Due to the difficulty in obtaining food grade sodium hydroxide in small quantities for home use, Sodium carbonate is often used in place of sodium hydroxide .

Specific foods processed with sodium hydroxide include :

- The Scandinavian delicacy known as lutefisk (from *lutfisk* , " lye fish") .
- Hominy is dried maize (corn) kernels reconstituted by soaking in lye - water. These expand considerably in size and may be further processed by frying to make corn nuts or by drying and grinding to make grits. Nixtamal is similar, but uses calcium hydroxide instead of sodium hydroxide.
- Sodium hydroxide is also the chemical that causes gelling of egg whites in the production of Century eggs.
- German pretzels are poached in a boiling sodium carbonate solution or cold sodium hydroxide solution before baking, which contributes to their unique crust.
- Most yellow coloured Chinese noodles are made with lye - water but are commonly mistaken for containing egg.

5 . 11 . Domestic uses

Sodium hydroxide is used in the home as a drain cleaning agent for clearing clogged drains. It is distributed as a dry crystal or as a thick liquid gel. The chemical mechanism employed is the conversion of grease to a form of soap. Soap is water-soluble, and can be dissolved by flushing with water. Sodium hydroxide also decomposes complex molecules such as the protein that composes hair. Such **drain cleaners** (and their acidic versions) are highly caustic and should be handled with care.

Sodium hydroxide has been used as a relaxer to straighten hair. However, because of the high incidence and intensity of chemical burns, chemical relaxer manufacturers have now switched to other alkaline chemicals, although sodium hydroxide relaxers are still available, used mostly by professionals .

5 . 12 . Use in chemical analysis

In analytical chemistry, sodium hydroxide solutions are often used to measure the concentration of acids by titration. Since Na OH is not a primary standard, solutions must first be standardised by titration against a standard such as KHP. Burettes exposed to NaOH should be rinsed out immediately after use to prevent " freezing " of the stopcock . Sodium hydroxide was traditionally used to test for cations in Qualitative Inorganic Analysis, as well as to provide alkaline media for some reactions that need it, such as the Biuret test.

6 . Safety

Solid sodium hydroxide and solutions of more than 2 % by weight (0.5 M) should be labelled as corrosive .

Solid sodium hydroxide or solutions of sodium hydroxide will cause chemical burns, permanent injury or scarring, and blindness if it contacts unprotected human or animal tissue. Protective equipment such as rubber gloves, safety clothing and eye protection should always be used when handling the material or its solutions.

Dissolution of sodium hydroxide is highly exothermic, and the resulting heat may cause heat burns or ignite flammables.

Sodium hypo chlorite

Contents

1 . Introduction :

Sodium hypochlorite is a chemical compound with the formula Na ClO . Sodium hypochlorite solution, commonly known as bleach, is frequently used as a disinfectant or a bleaching agent.

Other Names	Sodium chlorate
Molecular Formula	Na Cl O
Molar Mass	74.5 g / mol
Appearance	White solid
Density	1.11 g / cm^3
Melting Point	18 °C (penta hydrate)
Boiling Point	101 °C (decomp)
Solubility in Water	29.3 g / 100 ml (0 °C)
EU classification	Corrosive (**C**) Dangerous for the environment (**N**)

2 . Production :

Hypo chlorite was first produced in 1789 in Javel , France , by passing chlorine gas through a solution of sodium carbonate . The resulting liquid, known as "Eau de Javel" or "Javel water" was a weak solution of sodium hypochlorite. However, this process was not very efficient and alternate production methods were sought. One such method involved the extraction of chlorinated lime (known as bleaching powder) with sodium carbonate to yield low levels of available chlorine. This method was commonly used to produce hypochlorite solutions for use as a hospital antiseptic which was sold under the trade names "Eusol" and "Dakin's solution." Near the end of the nineteenth century, E. S. Smith patented a method of hypochlorite production involving hydrolysis of brine to produce caustic soda and chlorine gas which then mix to form hypochlorite.[*citation needed*] Both electric power and brine solution were in cheap supply at this time and various enterprising marketers took advantage of this situation to satisfy the market's demand for hypochlorite. Bottled solutions of hypochlorite were sold under numerous trade names; one such early brand produced by this method was called Parozone. Today, an improved version of this method, known as the Hooker process, is the only large scale industrial method of sodium hypochlorite production. In this process sodium hypochlorite (Na OCl) and sodium chloride (NaCl) is formed when chlorine is passed into cold and dilute sodium hydroxide solution. It is prepared industrially by electrolysis minimal separation between the anode and the cathode. The solution must be kept below 40 °C (by cooling coils) to prevent the undesired formation of sodium chlorate.

$$Cl_2 + 2\ Na\,OH \rightarrow Na\,Cl + Na\,Cl\,O + H_2O$$

Sodium hydroxide and chlorine are commercially produced by the chlor alkali process, and there is no need to isolate them to prepare sodium hypo chlorite . The commercial solutions always contain significant amounts of sodium chloride as the main by - product , as seen in the equation above .

3 . Packaging and sale :

House hold bleach sold for use in laundering clothes is a 3 – 6 % solution of sodium hypo chlorite at the time of manufacture. Strength varies from one formulation to another and gradually decreases with long storage.

A 12 % solution is widely used in water works for the chlorination of water and a 15 % solution is more commonly used for disinfection of waste water in treatment plants. High-test hypochlorite (HTH) is sold for chlorination of swimming pools and contains approximately 30 % calcium hypochlorite. The crystalline salt is also sold for the same use; this salt usually contains less than 50 % of calcium hypochlorite. However, the level of " active chlorine " may be much higher.

It can also be found on store shelves in " Daily Sanitizing Sprays " , as the sole active ingredient at 0.0095 %.

4 . Uses

4 . 1 . Bleaching

In household bleach form , sodium hypochlorite is used for removal of stains from laundry. It is particularly effective on cotton fiber, which stains easily but bleaches well. 50 to 250 ml per load is usually recommended for a standard-size washer. The properties of household bleach that make it effective for removing stains also result in cumulative damage to organic fibers such as cotton, and the useful lifespan of these materials will be shortened with regular bleaching. The sodium hydroxide (Na OH) that is also found in household bleach (as noted later) causes fiber degradation as well. It is not volatile, and residual amounts of Na OH not rinsed out will continue slowly degrading organic fibers in the presence of humidity. For these reasons, if stains are localized, spot treatments should be considered whenever possible. With safety precautions, post-treatment with weak organic acids such as acetic acid (vinegar) will neutralize the Na OH and volatilize the chlorine from residual hypochlorite. Old t - shirts and cotton sheets that rip easily demonstrate the costs of laundering with household bleach. Hot water increases the activity of the bleach, owing to the thermal decomposition of hypochlorite which ultimately generates environmentally-undesirable chlorate.

4 . 2 . Disinfection

A weak solution of 1 % house hold bleach in warm water is used to sanitize smooth surfaces prior to brewing of beer or wine. Surfaces must be rinsed to avoid imparting flavors to the brew; these chlorinated by products of sanitizing surfaces are also harmful.

US Government regulations (21 CFR Part 178) allow food processing equipment and food contact surfaces to be sanitized with solutions containing bleach provided the solution is allowed to drain adequately before contact with food, and the solutions do not exceed 200 parts per million (ppm) available chlorine (for example, one tablespoon of typical household bleach containing 5.25 % sodium hypochlorite, per gallon of water) . If higher concentrations are used, the surface must be rinsed with potable water after sanitizing.

A 1 in 5 dilution of household bleach with water (1 part bleach to 4 parts water) is effective against many bacteria and some viruses, and is often the disinfectant of choice in cleaning surfaces in hospitals (Primarily in the United States). The solution is corrosive, and needs to be thoroughly removed afterwards, so the bleach disinfection is sometimes followed by an ethanol disinfection. Even "scientific grade", commercially produced disinfection solutions such as Virocidin -X usually have sodium hypochlorite as their sole **active** ingredient, though they also contain surfactants (to prevent beading) and fragrances (to conceal the bleach smell) .

4 . 3 . Water treatment

For shock chlorination of wells or water systems, a 2 % solution of household bleach is used. For larger systems, HTH is more practical because lower rates can be used. The alkalinity of the sodium hypochlorite solution also causes the precipitation of minerals such as calcium carbonate, so that the shock chlorination is often accompanied by a clogging effect. The precipitate also preserves bacteria, making this practice somewhat less effective.

Sodium hypochlorite has been used for the disinfection of drinking water. A concentration equivalent to about 1 liter of household bleach per 4000 liters of water is used. The exact amount required depends on the water chemistry, temperature, contact time, and presence or absence of sediment. In large -scale applications, residual chlorine is measured to titrate the proper dosing rate. For *emergency* disinfection, the United States Environmental Protection Agency recommends the use of 2 drops of 5 % ac house hold bleach per quart of water. If the treated water doesn't smell of bleach , 2 more drops are to be added.

The use of chlorine - based disinfectants in domestic water, although widespread, has led to some controversy due to the formation of small quantities of harmful byproducts such as chloroform.

An alkaline solution (pH 11.0) of sodium hypochlorite is used to treat dilute (< 1 g / L) cyanide wastewater, e.g. rinse water from an electroplating shop. In batch treatment operations, sodium hypochlorite has been used to treat more concentrated cyanide wastes, such as silver cyanide plating solutions. A well - mixed solution is fully treated when an excess of chlorine is detected.

Sodium hypochlorite in the form of household bleach is often used to oxidize foul - smelling thiol wastes generated in a chemistry laboratory .

4 . 4 . Endodontics

Sodium hypochlorite is now used in endodontics during root canal treatments. It is the medicament of choice due to its efficacy against pathogenic organisms and pulp digestion. Historically, Henry Drysdale Dakin's solution (0.5 %) had been used. Its concentration for use in endodontics today varies from 0.5 % to 5.25 %. At low concentrations it will dissolve mainly necrotic tissue; whereas at higher concentrations tissue dissolution is better but it also dissolves vital tissue, a generally undesirable effect. It has been shown clinical effectiveness does not increase conclusively for concentrations higher than 1 % .

4 . 5 . Oxidation

House hold bleach , with a phase - transfer catalyst , has been reported to oxidize alcohols to the corresponding carbonyl compound

5 . **Safety**

Sodium hypochlorite is a strong oxidizer. Products of the oxidation reactions are corrosive. Solutions burn skin and cause eye damage, particularly when used in concentrated forms. However, as recognized by the NFPA, only solutions containing more than 40% sodium hypochlorite by weight are considered hazardous oxidizers. Solutions less than 40 % are classified as a moderate oxidizing hazard (NFPA 430 , 2000).

House hold bleach and pool chlorinator solutions are typically stabilized by a significant concentration of lye (caustic soda, Na OH) as part of the manufacturing reaction. Skin contact will produce caustic irritation or burns due to defatting and saponification of skin oils and destruction of tissue. The slippery feel of bleach on skin is due to this process . Sodium thio sulfate (hypo) is an effective chlorine neutralizer. Rinsing with a 5 mg / L solution, followed by washing with soap and water, quickly removes chlorine odor from the hands . Chlorination of drinking water can oxidize organic contaminants, producing tri halo methanes (also called halo forms) , which are carcinogenic. The extent of the hazard thus created is a subject of disagreement . Mixing bleach with some household cleaners can be hazardous. For example, mixing an acid cleaner with sodium hypochlorite bleach generates chlorine gas. Mixing with ammonia solutions (including urine) produces chloramines. Both chlorine gas and chloramine gas are toxic. Bleach can react violently with hydrogen peroxide and produce oxygen gas :

$$H_2O_2\,(aq) + Na\,O\,Cl\,(aq) \rightarrow Na\,Cl\,(aq) + H_2O\,(l) + O_2\,(g)$$

It is estimated that there are about 3300 accidents needing hospital treatment caused by sodium hypo chlorite solutions each year in British homes .

A recent European study indicated that sodium hypochlorite and organic chemicals (e.g., surfactants , fragrances) contained in several household cleaning products can react to generate chlorinated volatile organic compounds (VOCs) . These chlorinated compounds are emitted during cleaning applications, some of which are toxic and probable human carcinogens. The study showed that indoor air concentrations significantly increase (8 - 52 times for chloroform and 1- 1170 times for carbon tetrachloride, respectively, above baseline quantities in the household) during the use of bleach containing products. The increase in chlorinated volatile organic compound concentrations was the lowest for plain bleach and the highest for the products in the form of “thick liquid and gel”. The significant increases observed in indoor air concentrations of several chlorinated VOCs (especially carbon tetrachloride and chloroform) indicate that the bleach use may be a source that could be important in terms of inhalation exposure to these compounds. While the authors suggested that using these cleaning products may significantly increase the cancer risk, this conclusion appears to be hypothetical :

* The highest level cited for concentration of carbon tetrachloride (seemingly of highest concern) is 459 micrograms per cubic meter, translating to 0.073 ppm (part per million) , or 73 ppb (part per billion). The OSHA - allowable time - weighted average concentration over an eight - hour period is 10 ppm , almost 140 times higher;

* The OSHA highest allowable peak concentration (5 minute exposure for five minutes in a 4-hour period) is 200 ppm , twice as high as the reported highest peak level (from the headspace of a bottle of a sample of bleach plus detergent) .

Further studies of the use of these products and other possible exposure routes (i.e., dermal) may reveal other risks. Though the author further cited ozone depletion greenhouse effects for these gases, the very low amount of such gases, generated as prescribed, should minimize their contribution relative to other sources

Sodium iodide

Contents

1 – Introduction :

Sodium iodide is a white, crystalline salt with chemical formula Na I used in radiation detection, treatment of iodine deficiency, and as a reactant in the Finkelstein reaction.

Molecular formula	Na I
Molar mass	150 g / mol
Appearance	white solid deliquescent
Density	3.67 g/cm^3
Melting point	661 °C,
Boiling point	1304 °C,
Solubility in water	178.8 g /100 mL (20 °C) 184 g / 100 mL (25 °C) 294 g / 100 mL (70 °C)
Solubility	soluble in ethanol and acetone (39.9 g/100 mL)
Coordination geometry	Octahedral
EU Index	Not listed
Main hazards	Irritant , can harm the unborn child
Flash point	Non-flammable

2 - Uses :

Sodium iodide is commonly used to treat and prevent iodine deficiency.

Sodium iodide is used in polymerase chain reactions, and also (as an acetone solution) in the Finkelstein reaction, for conversion of an alkyl chloride into an alkyl iodide. This relies on the insolubility of sodium chloride in acetone to drive the reaction.

$$R\text{-}Cl + NaI \rightarrow R\text{-}I + NaCl$$

Sodium iodide activated with thallium, Na I (Tl) , when subjected to ionizing radiation, emits photons (i.e., scintillate) and is used in scintillation detectors, traditionally in nuclear medicine, geophysics, nuclear physics , and environmental measurements . Na I (Tl) is the most widely used scintillation material and has the highest light yield of the commonly used scintillators . The crystals are usually coupled with a photomultiplier tube, in a hermetically sealed assembly, as sodium iodide is hygroscopic. Fine-tuning of some parameters (i.e., radiation hardness, afterglow, transparency) can be achieved by varying the conditions of the crystal growth. Crystals with a higher level of doping are used in X - ray detectors with high spectrometric quality. Sodium iodide can be used both as single crystals and as polycrystals for this purpose. The radioactive iodide salt of sodium, $Na^{131}I$, is used for the treatment of thyroid cancer and hyperthyroidism.

3 - Solubility data :

Solubility of NaI in various solvents
(g NaI / 100g of solvent at 25°C)

H_2O	184
Liquid ammonia	162
Liquid sulfur dioxide	15
Methanol	62.5 - 83.0
Formic acid	61.8
Acetonitrile	24.9
Acetone	28.0
Formamide	57 - 85
Acetamide	32.3
Dimethyl formamide	3.7 - 6.4

Sodium meta bi sulfite

Contents

1 – Introduction :

Sodium meta bi sulfite or **sodium pyro sulfite** (IUPAC spelling; Br. E . **sodium meta bi sulphite** or **sodium pyro sulphite**) is an inorganic compound of chemical formula $Na_2 S_2 O_5$. The name is sometimes referred to as *disodium* (meta bi sulfate , etc). It is used as a disinfectant, antioxidant and preservative agent.

Other Names	Sodium Pyro Sulfite , Sodium Di Sulfite
Molecular Formula	$Na_2 S_2 O_5$
Molar Mass	190 g / mol
Appearance	white powder
Density	1.48 g / cm^3
Melting point	>170 °C (begins at 150)
Solubility in water	65 (25 °C) ; 47 (20 °C)

2 - Uses :

2 – 1 - Food additive :

It is used as a food additive, mainly as a preservative and is sometimes identified as E223 .

It may cause allergic reactions in those who are sensitive to sulfites.

It is present in many dilute able squashes (i.e. fruit juice concentrates) and in candy bars such as Mounds . Sodium meta bi sulfite and potassium meta bi sulfite are the primary ingredients in Camp den tablets , used for wine and beer making.

The acceptable daily intake is up to 0.7 mg per kg of body weight. Sodium meta bi sulfite has no side effects; it is oxidized in the liver to harmless sulfate and excreted in the urine.

2 – 2 - Sanitization / Cleaning agent :

It is commonly used in home brewing preparations to sanitize equipment. It is used as a cleaning agent for potable water reverse osmosis membranes in desalination systems. It is also used to remove chloramine from drinking water after treatment.

2 – 3 - Other uses :

Concentrated sodium meta bi sulfite can be used to remove tree stumps. The *Stump - Out* brand has 98 % sodium meta bi sulfite, and causes degradation of lignin in the stumps, facilitating removal. It is also used as an excipient in some tablets, such as paracetamol . A very important health related aspect of this substance is that it can be added to a blood smear in a test for sickle cell anaemia (and other similar forms of haemoglobin mutation). The substances causes defunct cells to sickle (through a complex poly merisation) hence confirming disease.

3 - Chemical properties :

When mixed with water, sodium meta bi sulfite releases sulfur dioxide (SO_2), a pungent, unpleasant smelling gas that can also cause breathing difficulties in some people. For this reason, sodium meta bi sulfite has fallen from common use in recent times, with agents such as hydrogen peroxide becoming more popular for effective and odorless sterilization of equipment. Released sulfur dioxide however makes the water a strong reducing agent.

Sodium Mono Hydrogen Phosphate

1 – Introduction :

Disodium hydrogen phosphate (Na_2HPO_4) is a sodium salt of phosphoric acid. It is a white powder that is highly hygroscopic and water soluble . It is therefore used commercially as an anti - caking additive in powdered products. It is also known as disodium hydrogen orthophosphate, sodium hydrogen phosphate or sodium phosphate dibasic. It is commercially available in both the hydrated and anhydrous forms.

2 - Other uses :

- Disodium phosphate can be used in cream of wheat to quicken cook time , as described on the ingredients panel of the product package .
- Disodium phosphate is used in conjunction with tri sodium phosphate in many steam boiler applications. It supplies the inventory of free phosphates to retard calcium scale formation.

Sodium Nitrate

1 . Introduction :

Sodium nitrate is the chemical compound with the formula $NaNO_3$. This salt, also known as " Chile salt peter " or " Peru salt peter " (to distinguish it from ordinary saltpeter , potassium nitrate) , is a white solid which is very soluble in water. The mineral form is also known as **nitratine** or **soda niter**.

Sodium nitrate is used as an ingredient in fertilizers, pyrotechnics, as a food preservative, and as a solid rocket propellant, as well as in glass and pottery enamels; the compound has been mined extensively for those purposes.

The mining of South American saltpeter was such a profitable business that Chile fought against the allies Peru and Bolivia and took over the richest deposits in the War of the Pacific. The world's largest natural deposits of caliche ore were in the Atacama desert of Chile, and many deposits were mined for over a century, until the 1940s, when its value declined dramatically in the first decades of the twentieth century .

Chile still has the largest reserves of caliche , with active mines in such locations as Pedro de Valdivia, Maria Elena and Pampa Blanca, and there it used to be called "white gold". Sodium nitrate, potassium nitrate, sodium sulfate and iodine are all obtained by the processing of caliche . The former Chilean saltpeter mining communities of Humber stone and Santa Laura were declared Unesco World Heritage sites in 2005.

Sodium nitrate is also synthesized industrially by neutralizing nitric acid with soda ash .

Other Names	Caliche Chile salt peter Nitrate of soda Nitratine

	Peru saltpeter Soda niter
Molecular Formula	Na NO_3
Molar Mass	85 g / mol
Appearance	White powder or colorless crystals
Density	2.257 g / cm^3 , solid
Melting Point	308 ° C
Boiling Point	380 ° C decomp .
Solubility in Water	92.1 g / 100 ml (25 °C) 180 g / 100mL (100 °C)
Solubility	Very soluble in ammonia ; soluble in alcohol
Refractive index (n_D)	1.587 (trigonal) 1.336 (rhomobohedral)
Crystal structure	trigonal and rhombohedral
Std enthalpy of formation $\Delta_f H^{\circ}_{298}$	− 468 kJ / mol
Main hazards	Oxidant , irritant
Flash point	Non - flammable
LD_{50}	3236 mg / kg

2 . Applications :

Sodium nitrate was used extensively as a fertilizer and a raw material for the manufacture of gunpowder in the late nineteenth century. Sodium nitrate has antimicrobial properties when used as a food preservative. It is found naturally in leafy green vegetables. it can also be combined with iron hydroxide to make a resin.

Sodium nitrate should not be confused with the related compound , sodium nitrite.

It can be used in the production of nitric acid by combining it with sulfuric acid and subsequent separation through fractional distillation of the nitric acid, leaving behind a residue of sodium bisulfate. Hobbyist gold refiners use sodium nitrate to make a hybrid aqua regia that dissolves gold and other metals.

Less common applications include its use as a substitute oxidizer used in fireworks as a replacement for potassium nitrate commonly found in black powder and as a component in instant cold packs .

Because sodium nitrate can be used as a Phase Change Material it may be used for heat transfer in solar power plants .

Sodium nitrite

Contents

1 . Introduction :

Sodium nitrite, with chemical formula Na NO_2 , is used as a color fixative and preservative in meats and fish . When pure, it is a white to slight yellowish crystalline powder. It is very soluble in water and is hygroscopic. It is also slowly oxidized by oxygen in the air to sodium nitrate , Na NO_3. The compound is a strong oxidizing agent.

It is also used in manufacturing diazo dyes , nitroso compounds, and other organic compounds; in dyeing and printing textile fabrics and bleaching fibers; in photography; as a laboratory reagent and a corrosion inhibitor; in metal coatings for phosphatizing and detinning; and in the manufacture of rubber chemicals . It may also be used as an electrolyte in electro chemical grinding manufacturing processes, typically diluted to about 10 % concentration in water. Sodium nitrite also has been used in human and veterinary medicine as a vasodilator, a bronchodilator, and an antidote for cyanide poisoning.

Molecular Formula	Na NO_2
Molar Mass	69 g / mol
Appearance	white solid
Density	2.168 g / cm^3
Melting Point	271 ° C decomp.
Solubility in Water	82 g / 100 ml (20 °C)
Crystal structure	Trigonal

EU Classification	Oxidant (**O**) Toxic (**T**) Dangerous for the environment (**N**)
Autoignition temperature	489 °C
LD_{50}	85 mg / kg

2 . Uses :

2 . 1 . Food additive

As a food additive, it serves a dual purpose in the food industry since it both alters the color of preserved fish and meats and also prevents growth of *Clostridium botulinum*, the bacteria which causes botulism. In the European Union it may be used only as a mixture with salt containing at most 0.6 % sodium nitrite. It has the E number E 250 . Potassium nitrite (E249) is used in the same way .

While this chemical will prevent the growth of bacteria, it can be toxic in high amounts for animals, including humans. Sodium nitrite's LD_{50} in rats is 180 mg / kg and its human LD_{Lo} is 71 mg / kg , meaning a 65 kg person would likely have to consume at least 4.615 g to result in toxicity . To prevent toxicity, sodium nitrite sold as a food additive is dyed bright pink to avoid mistaking it for something else.

A principal concern of sodium nitrite is the formation of carcinogenic nitrosamines in meats containing sodium nitrite when exposed to high temperatures. Sodium nitrite's usage is carefully regulated in the production of cured products in the United States as the concentration in finished products is limited to 200 ppm, and is usually lower. In about 1970, it was found that ascorbic acid (vitamin C) , an antioxidant, inhibited nitrosamine formation . Consequently, the addition of at least 550 ppm of ascorbic acid is required in meats manufactured in the United States. Manufacturers sometimes instead use erythorbic acid , a cheaper but equally effective isomer of ascorbic acid. Additionally, manufactures may include alpha-tocopherol (vitamin E) to further inhibit nitrosamine production . Alpha – tocopherol , ascorbic acid , and erythorbic acid all inhibit nitrosamine production by their oxidation-reduction properties. Ascorbic acid for example, forms de hydro ascorbic acid when oxidized, which when in the presence of nitrous anhydride, a potent nitro sating agent formed from sodium nitrate, reduces the nitrous anhydride into the nitric oxide gas .

Sodium nitrite consumption has also been linked to triggering migraines in individuals who already suffer from them .

A recent study has found a link between frequent ingestion of cured meats and the COPD form of lung disease. The study's researchers suggest that the high amount of nitrites in the meats was responsible; however, the team did not prove the nitrite theory. Additionally, the study doesn't prove that nitrites or cured meat caused higher rates of COPD , merely a link. The researchers did adjust for many of COPD ' s risk factors, but they commented they cannot rule out all possible un measurable causes or risks for COPD .

2 . 2 . Disease treatment :

Recently, sodium nitrite has been found to be an effective means to increase blood flow by dilating blood vessels , acting as a vasodilator. Research is ongoing to investigate its applicability towards treatments for sickle cell anemia , cyanide poisoning, heart attacks , brain aneurysms , and pulmonary hyper tension in infants .

2 . 3 . Synthetic reagent :

Sodium nitrite is used to convert amines into diazo compounds. The synthetic utility of such a reaction is to render the amino group labile for nucleophilic substitution, as the N_2 group is a better leaving group.

In the laboratory, sodium nitrite is also used to destroy excess sodium azide .

$$NaNO_2 + H_2SO_4 \rightarrow HNO_2 + NaHSO_4$$
$$2\,NaN_3 + 2\,HNO_2 \rightarrow 3\,N_2 + 2\,NO + 2\,NaOH$$

Sodium Per Borate

Contents

1 – Introduction :

Sodium per borate (**PBS**) is a white, odorless, water-soluble chemical compound with chemical formula $NaBO_3$. It crystallizes as the monohydrate, $NaBO_3.H_2O$, tetra hydrate, $NaBO_3.4H_2O$ and tri hydrate , $NaBO_3.3H_2 O$. The monohydrate and tetra hydrate are the commercially important forms.

Other Names	PBS - 1 (mono) , PBS - 4 (tetra)
Molecular Formula	$Na BO_3 . nH_2O$
Molar mass	100 g / mol (mono hydrate) ; 154 g / mol (tetra hydrate)
Appearance	white powders
Flash point	non-flammable

2 - Preparation and chemistry :

Sodium per borate is manufactured by reaction of disodium tetra borate penta hydrate , hydrogen peroxide , and sodium hydroxide. The mono hydrate form dissolves better than the tetra hydrate and has higher heat stability; it is prepared by heating the tetra hydrate. Sodium per borate undergoes hydrolysis in contact with water, producing hydrogen peroxide and borate.

3 - Structure :

Unlike sodium per carbonate and per phosphate , the sodium per borate are not simply addition compounds of hydrogen peroxide, but contain a cyclic anion with peroxo bridges and do not contain the BO_3^- ion . This makes the material more stable, safer for handling and storage. The formulae of the mono and tetra hydrates can therefore be written as $Na_2B_2O_8H_4$ (anhydrous) and $Na_2B_2O_8H_4.6H_2O$ respectively

4 - Uses :

It serves as a source of active oxygen in many detergents, laundry detergents, cleaning products, and laundry bleaches. It is also present in some tooth bleaching formulas. It is used as a bleaching agent for internal bleaching of a tooth that has had root canal treatment. The sodium per borate is placed inside the tooth and left in place for an extended period of time to allow it to diffuse into the tooth and bleach stains from the inside out. It has antiseptic properties and can act as a disinfectant. It is also used as a "disappearing" preservative in some brands of eye drops.

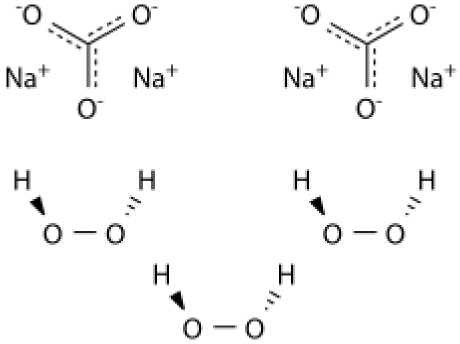
Sodium per borate is a less aggressive bleach than sodium hypochlorite, causing less degradation to dyes and textiles. Borates also have some non-oxidative bleaching properties.

Sodium per borate releases oxygen rapidly at temperatures over 60°C. To make it active at lower temperatures (40 - 60 °C) , it has to be mixed with a suitable activator, typically tetra acetyl ethylene di amine (TAED).

5 – Hazards :

It is a skin irritant.

Sodium per carbonate



Contents

1 – Introduction :

Sodium per carbonate is a chemical, an adduct of sodium carbonate and hydrogen peroxide (a per hydrate) , with formula $Na_2\ CO_3 \cdot 1.5\ H_2O_2$. It is a colorless, crystalline, hygroscopic and water-soluble solid. It is used in some eco - friendly cleaning products and as a laboratory source of anhydrous hydrogen peroxide.

This product contains the carbonate anion, and should not be confused with sodium per oxo carbonate Na_2CO_4 or per oxo di carbonate $Na_2C_2O_6$, which contain different anions.

2 – Structure :

At room temperature, solid sodium percarbonate has the orthorhombic crystal structure, with the *Cmca* crystallographic space group. The structure changes to *Pbca* as the crystals are cooled below about −30 °C.

3 - Production :

Sodium per carbonate is produced industrially by reaction of sodium carbonate and hydrogen peroxide, followed by crystallization. Also, dry sodium carbonate may be reacted directly with concentrated hydrogen peroxide solution. World production capacity of this compound was estimated at several hundred thousand tonnes for 2004.[3] It can be obtained in the laboratory by reacting the two substances in aqueous solution with proper control of the pH[4] or concentrations.

4 – Uses :

Sodium per carbonate is an oxidizing agent and ingredient in a number of home and laundry cleaning products, including bleach products such as Oxi Clean and Tide laundry detergent . It contains no phosphorus or nitrogen. Dissolved in water, it yields a mixture of hydrogen peroxide (which eventually decomposes to water and oxygen) and sodium carbonate ("soda ash") .

Sodium per carbonate can be used in organic synthesis as a convenient source of anhydrous H_2O_2, in particular in solvents that cannot dissolve the carbonate but can leach the H_2O_2 out of it.

Sodium Per Chlorate

1- Introduction :

Sodium per chlorate is a per chlorate of sodium and has the formula $NaClO_4$. Sodium per chlorate melts with decomposition at 480 °C.[1] Its heat of formation is - 382.75 kJ mol^{-1}. It is a white crystalline solid. It is hygroscopic. It is soluble in water and in alcohol. It usually comes as the monohydrate, which has a rhombic crystal structure.

Other names	Sodium Chlorate (VII) , Sodium Hyper chlorate , Perchloric Acid Sodium Salt
Molecular Formula	$Na\ Cl\ O_4$
Molar Mass	122.5 g / mol
Appearance	White Crystalline Solid
Melting point	400 °C (Decomposes)
Solubility in water	Soluble , Hygroscopic in Air

2 - Uses :

The present major use of per chlorate salts is as oxidizers in solid propellants such as NASA's solid rocket boosters. The potassium salt was first used and quickly followed by what is now the more important salt, ammonium per chlorate. Lithium per chlorate, which has the highest weight percentage of oxygen of all compounds, has been tested as an oxidizer in solid propellants, but has not found favor with propellant manufacturers due to its hygroscopicity.

Sodium per chlorate itself finds only minimal use in pyrotechnics because it is hygroscopic; ammonium and potassium per chlorates are preferred. They are made by double decomposition from a solution of sodium per chlorate and potassium or ammonium chlorides.

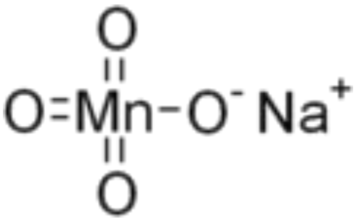
It is also used in standard DNA extraction and hybridization reactions in molecular biology.

3 - Synthesis :

$$ClO_3^- + H_2O \rightarrow ClO_4^- + 2H^+ + 2e^-$$

Sodium per chlorate is manufactured by anodic oxidation of sodium chloride or sodium chlorate at high current density, with platinum (or in some cases , lead dioxide , manganese dioxide , and possibly magnetite and cobalt oxide) anodes and graphite, steel, nickel, or titanium cathodes .

Sodium permanganate



1 – Introduction :

Sodium permanganate is the inorganic compound with the formula Na Mn O_4 . It is closely related to the more commonly encountered potassium permanganate, but it is generally less desirable, because it is more expensive, absorbs water from the atmosphere, and has a low melting point. Being about 15 times more soluble than K Mn O_4 , sodium permanganate finds some applications where very high concentrations of Mn O_4^- are sought.

IUPAC Name : Sodium manganate (VII)

Other names :

Sodium permanganate , permanganate of sodium

Molecular Formula	Na Mn O_4
Molar Mass	142 g / mol 160 g / mol (mono hydrate)
Appearance	Red solid
Density	1.972 g / cm^3 (mono hydrate)
Melting point	36 °C
Boiling point	100 °C
Solubility in water	900 g / L
Main hazards	Oxidizer

2 - Preparation and properties :

Sodium permanganate cannot be prepared analogously to the route to K Mn O_4 because the required intermediate manganate salt,

Na_2MnO_4 , does not form. Thus less direct routes are used including conversion from K Mn O_4.

Sodium permanganate behaves similarly to potassium permanganate . It dissolves readily in water to give deep purple solutions, evaporation of which gives prismatic purple-black glistening crystals of the monohydrate Na Mn $O_4 \cdot H_2O$. The potassium salt does not form a hydrate. Because of its hygroscopic nature, it is less useful in analytical chemistry than its potassium counterpart.

It can be prepared by the reaction of manganese dioxide with sodium hypochlorite:

$$2\ MnO_2 + 3\ Na\,Cl\,O + 2\ Na\,OH \rightarrow 2\ NaMnO_4 + 3\ NaCl + H_2O$$

3 - Applications :

Because of its high solubility, its aqueous solutions are used as etchants in printed circuitry.

Sodium peroxide

Contents

1 – Introduction ;

Sodium peroxide is the inorganic compound with the formula Na_2O_2. This solid is the product when sodium is burned with oxygen. It is a strong base and a potent oxidizing agent. It exists in several hydrates and per oxy hydrates including $Na_2O_2 \cdot 2H_2O_2 \cdot 4H_2O$, $Na_2O_2 \cdot 2H_2O$, $Na_2O_2 \cdot 2H_2O_2$, and $Na_2O_2 \cdot 8H_2O$.

Other Names : Sodium dioxide , Flo cool , Solo zone , Disodium peroxide	
Molecular Formula	$Na_2 O_2$
Molar Mass	78 g / mol
Appearance	yellow to white powder
Density	2.805 g / cm^3
Melting point	675 °C
Boiling point	decomp.
Solubility in water	reacts violently
Crystal structure	Hexagonal
Std enthalpy of formation $\Delta_f H^\circ_{298}$	−513 kJ/mol
Standard molar entropy S°_{298}	95 J K^{-1} mol^{-1}
EU classification	Oxidant (O) Corrosive (C)
Flash point	Non-flammable

2 - Properties :

Sodium peroxide crystallizes with hexagonal symmetry . Upon heating, the hexagonal form undergoes a transition into a phase of unknown symmetry at 512 °C. With further heating above the 675 °C melting point, the compound decomposes, releasing O_2, before reaching a boiling point.[5]

Sodium peroxide is hydrolyzed to give sodium hydroxide and hydrogen peroxide according to the reaction:

$$2Na_2O_2 + 2\ H_2O \rightarrow 4\ Na\ OH + O_2$$

3 - Preparation

The synthesis is no longer of commercial significance since the development of efficient routes to hydrogen peroxide . Sodium peroxide formerly was prepared on a large scale by the reaction with sodium with oxygen at 130 –200 °C , a process that generates sodium oxide, which in a separate stage absorbs oxygen :

$$4\ Na + O_2 \rightarrow 2\ Na_2O$$
$$2\ Na_2O + O_2 \rightarrow 2\ Na_2O_2$$

More specialized routes have been developed. At ambient temperatures (0–20 °C), O_2 reacts with a dilute (0.1–5.0 mole percent) sodium amalgam. It may also be produced by passing ozone gas over solid sodium iodide inside a platinum or palladium tube. The ozone oxidizes the sodium to form sodium peroxide. The iodine is freed into iodine crystals, which can be sublimed by mild heating. The platinum or palladium catalyzes the reaction and is not attacked by the sodium peroxide.

4 - Uses

Sodium peroxide was used to bleach wood pulp for the production of paper and textiles. Presently it is mainly used for specialized laboratory operations, e.g. the extraction of minerals from various ores. Sodium peroxide may go by the commercial names of *Solozone* .

Sodium per sulfate

1 – Introduction :

Sodium per sulfate ($Na_2S_2O_8$) is a chemical compound. It is a strong oxidizer. It is a severe irritant of skin, eyes, and respiratory system. It is almost non-hygroscopic and has particularly good ability to be stored for long time. It is easy and safe to handle. It is not combustible, but releases oxygen easily and assists combustion of other materials.

It is used as a bleach, both standalone (particularly in hair cosmetics) and as a detergent component. It is a replacement for ammonium per sulfate in etching mixtures for zinc and printed circuit boards, and is used for pickling of copper and some other metals. It is a source of free radicals, making it useful as a radical initiator for emulsion polymerization reactions and for accelerated curing of low formaldehyde adhesives. It is also used as a soil conditioner and in manufacture of dyestuffs, modification of starch, bleach activator, desizing agent for oxidative desizing , etc.

Other Names	sodium per oxodi sulfate sodium per oxodi sulphate sodium per oxydi sulfate sodium per oxydi sulphate
Molecular Formula	$Na_2 S_2 O_8$
Molar mass	238 g / mol
Appearance	white powder
Density	2.40 g / cm³ , solid
Melting point	180°C (decomposes)
Solubility in water	700 g / L (20°C)
Acidity (pK_a)	6.0 (1 % solution)

Conditions/substances to avoid mixing per sulfates with are: moisture, heat, flame, ignition sources, shock, friction, reducing agents, organic material, sodium peroxide, water, aluminium and powdered metals.

Sodium per sulfate is a oxidizing agent in chemistry, for example in a synthesis of di apocynin from apocynin where iron (II) sulfate is the radical initiator :

The sulfate radical formed in situ has a standard electrode potential of 2.7V.

Sodium Sesqui Carbonate

1 – Introduction :

Sodium sesqui carbonate (Systematic name tri sodium hydrogen di carbonate) $Na_3\ H(CO_3)_2$, which is a double salt of sodium bicarbonate and sodium carbonate, and has a needle-like crystal structure. However, the term is also applied to an equimolar mixture of those two salts, with what ever water of hydration the sodium carbonate includes, supplied as a powder.

The di hydrate , $Na_3H(CO_3)_2 \cdot 2H_2O$, occurs in nature as the evaporate mineral trona

2 – Uses :

Sodium sesqui carbonate is used in bath salts, swimming pools and as an alkalinity source for water treatment.

Sodium sesqui carbonate is used in the archaeological conservation of objects made of copper and copper alloys that have been corroded by salt water. Treatment with sodium sesqui carbonate removes copper chlorides from the corroded layer ; if these were not removed the object would continue to corrode ("copper disease") while exposed to air.

It is also used as a precipitating water softener, which combines with hard water minerals (calcium - and magnesium - based minerals) to form an insoluble precipitate, removing these hardness minerals from the water. It is the carbonate moiety which forms the precipitate, the bicarbonate being included to moderate the material's alkalinity.

Sodium Sulfate

Contents

1 . Introduction :

Sodium sulfate is the sodium salt of sulfuric acid. Anhydrous, it is a white crystalline solid of formula $Na_2 SO_4$ known as the mineral thenardite ; the deca hydrate $Na_2 SO_4 \cdot 10H_2O$ has been known as **Glauber 's salt** or , historically , **sal mirabilis** since the 17th century. Other solid is the hepta hydrate, which transforms to mirabilite when cooled . With an annual production of 6 million tones, it is one of the world's major commodity chemicals and one of the most damaging salts in structure conservation: when it grows in the pores of stones it can achieve high levels of pressure, causing structures to crack.

Sodium sulfate is mainly used for the manufacture of detergents and in the Kraft process of paper pulping. About two-thirds of the world's production is from mirabilite , the natural mineral form of the decahydrate, and the remainder from by - products of chemical processes such as hydrochloric acid production .

Other Names	Thenardite (mineral) Glauber 's salt (deca hydrate) Salt mirabilis (deca hydrate) Mirabilite (deca hydrate)

Molecular Formula	$Na_2 SO_4$
Molar Mass	142 g / mol (anhydrous) 322 g / mol (deca hydrate)
Appearance	white crystalline solid hygroscopic
Density	2.664 g / cm^3 (anhydrous) 1.464 g / cm^3 (deca hydrate)
Melting point	884 °C (anhydrous) 32.4 °C (decahydrate)
Boiling point	1429 °C (anhydrous)
Solubility in Water	4.76 g / 100 ml (0 °C) 42.7 g / 100 ml (100 °C)
Solubility	Insoluble in ethanol
Refractive index (n_D)	1.468 (anhydrous) 1.394 (deca hydrate)
Crystal structure	orthorhombic or hexagonal (anhydrous) monoclinic (deca hydrate)
Main hazards	Irritant
Flash Point	Non - flammable

2 . History :

The hydrate of sodium sulfate is known as Glauber 's Salt after the Dutch / German chemist and apothecary Johann Rudolf Glauber (1604 – 1670) , who discovered it 1625 in Austrian spring water. He named it *sal mirabilis* (miraculous salt), because of its medicinal properties: the crystals were used as a general purpose laxative, until more sophisticated alternatives came about in the 1900s .

In the 18th century, Glauber 's salt began to be used as a raw material for the industrial production of soda ash (sodium carbonate) , by reaction with potash (potassium carbonate) . Demand for soda ash increased and supply of sodium sulfate had to increase in line. Therefore, in the nineteenth century, the Leblanc process, producing synthetic sodium sulfate as a key intermediate, became the principal method of soda ash production .

3 . Physical and chemical properties

Sodium sulfate is chemically very stable, being unreactive toward most oxidizing or reducing agents at normal temperatures. At high temperatures, it can be reduced to sodium sulfide . It is a neutral salt, which forms aqueous solutions with pH of 7. The neutrality of such solutions reflects the fact that $Na_2\ SO_4$ is derived, formally speaking, from the strong acid sulfuric acid and a strong base sodium hydroxide. Sodium sulfate reacts with an equivalent amount of sulfuric acid to give an equilibrium concentration of the acid salt sodium bisulfate :

$$Na_2 SO_4 \text{ (aq) } + H_2 SO_4 \text{ (aq) } \rightleftharpoons 2 \text{ Na } HSO_4 \text{ (aq)}$$

In fact, the equilibrium is very complex, depending on concentration and temperature, with other acid salts being present.

Sodium sulfate is a typical ionic sulfate, containing Na^+ ions and SO_4^{2-} ions. Aqueous solutions can produce precipitates when combined with salts of Ba^{2+} or Pb^{2+} , which form insoluble sulfates

$$Na_2 SO_4 \text{ (aq) } + Ba Cl_2 \text{ (aq) } \rightarrow 2 \text{ Na Cl (aq) } + Ba SO_4 \text{ (s)}$$

Sodium sulfate has unusual solubility characteristics in water.[7] Its solubility rises more than tenfold between 0 °C to 32.4 °C , where it reaches a maximum of 49.7 g $Na_2\ SO_4$ per 100 g water. At this point the solubility curve changes slope, and the solubility becomes almost independent of temperature. In the presence of NaCl, the solubility of sodium sulfate is markedly diminished. Such changes provide the basis for the use of sodium sulfate in passive solar heating systems, as well is in the preparation and purification of sodium sulfate. This nonconformity can be explained in terms of hydration, since 32.4 °C corresponds with the temperature at which the crystalline deca hydrate (Glauber 's salt) changes to give a sulfate liquid phase and an anhydrous solid phase.

Sodium sulfate deca hydrate is also unusual among hydrated salts in having a measureable residual entropy (entropy at absolute zero) of 6.32 $J \cdot K^{-1} \cdot mol^{-1}$. This is ascribed to its ability to distribute water much more rapidly compared to most hydrates .

Sodium sulfate displays a moderate tendency to form double salts. The only alums formed with common trivalent metals are $NaAl(SO_4)_2$ (unstable above 39 °C) and $NaCr(SO_4)_2$, in contrast to potassium sulfate and ammonium sulfate which form many stable alums . Double salts with some other alkali metal sulfates are known , including $Na_2SO_4 \cdot 3K_2SO_4$ which occurs naturally as the mineral glaserite . Formation of glaserite by reaction of sodium sulfate with potassium chloride has been used as the basis of a method for producing potassium sulfate , a fertilizer . Other double salts include $3Na_2SO_4 \cdot CaSO_4$, $3Na_2SO_4 \cdot MgSO_4$ (vanthoffite) and $NaF \cdot Na_2SO_4$.

4 . Production

The world production of sodium sulfate, mostly in the form of the deca hydrate amounts to approximately 5.5 to 6 million tonnes annually (Mt / a) . In 1985 , production was 4.5 Mt/a, half from natural sources, and half from chemical production. After 2000, at a stable level until 2006, natural production had increased to 4 Mt/a, and chemical production decreased to 1.5 to 2 Mt/a, with a total of 5.5 to 6 Mt / a . For all applications, naturally produced and chemically produced sodium sulfate are practically inter changeable.

4 . 1 . Natural sources

Two thirds of the world's production of the deca hydrate (Glauber 's salt) is from the natural mineral form mirabilite , for example as found in lake beds in southern Saskatchewan. In 1990, Mexico and Spain were the world's main producers of natural sodium sulfate (each around 500,000 tonnes) , with Russia, USA and Canada around 350,000 tones each . Estimatedly , natural resources amount to over 1 billion tones .

Major producers of 200,000 – 1,500,000 tonnes/a in 2006 include Searles Valley Minerals (California, USA), Airborne Industrial Minerals (Saskatchewan , Canada), Química del Rey (Coahuila, Mexico) , Criaderos Minerales Y Derivados and Minera de Santa Marta , also known as Grupo Crimidesa (Burgos , Spain), FMC Foret (Toledo , Spain) , Sulquisa (Madrid, Spain), and in China Chengdu Sanlian Tianquan Chemical (Sichuan) , Hongze Yinzhu Chemical Group (Jiangsu), Nafine Chemical Industry Group (Shanxi), and Sichuan Province Chuanmei Mirabilite (Sichuan), and Kuchuksulphat JSC (Altai Krai , Siberia , Russia) .

Anhydrous sodium sulfate occurs in arid environments as the mineral thenardite. It slowly turns to mirabilite in damp air. Sodium sulfate is also found as glauberite , a calcium sodium sulfate mineral. Both minerals are less common than mirabilite.

4 . 2 . Chemical industry

About one third of the world 's sodium sulfate is produced as by-product of other processes in chemical industry. Most of this production is chemically inherent to the primary process, and only marginally economical. By effort of the industry, therefore, sodium sulfate production as by - product is declining.

The most important chemical sodium sulfate production is during hydrochloric acid production, either from sodium chloride (salt) and sulfuric acid, in the Mannheim process, or from sulfur dioxide in the Hargreaves process. The resulting sodium sulfate from these processes are known as *salt cake* .

Mannheim : $2\ NaCl + H_2SO_4 \rightarrow 2\ HCl + Na_2SO_4$
Hargreaves : $4\ NaCl + 2\ SO_2 + O_2 + 2\ H_2O \rightarrow 4\ HCl + 2\ Na_2SO_4$

The second major production of sodium sulfate are the processes where surplus sulfuric acid is neutralized by sodium hydroxide , as applied on a large scale in the production of rayon. This method is also a regularly applied and convenient laboratory preparation.

$$2\ NaOH\ (aq) + H_2SO_4\ (aq) \rightarrow Na_2SO_4\ (aq) + 2\ H_2O\ (l)$$

Formerly, sodium sulfate was also a by - product of the manufacture of sodium dichromate, where sulfuric acid is added to sodium chromate solution forming sodium dichromate, or subsequently chromic acid. Alternatively, sodium sulfate is or was formed in the production of lithium carbonate, chelating agents, resorcinol, ascorbic acid, silica pigments, nitric acid, and phenol .

Bulk sodium sulfate is usually purified via the decahydrate form, since the anhydrous form tends to attract iron compounds and organic compounds. The anhydrous form is easily produced from the hydrated form by gentle warming .

Major sodium sulfate by-product producers of 50 – 80 Mt/a in 2006 include Elementis Chromium (chromium industry , Castle Hayne, NC, USA) , Lenzing AG (200 Mt/a, rayon industry, Lenzing, Austria), Addiseo (formerly Rhodia, methionine industry, Les Roches-Roussillon, France), Elementis (chromium industry, Stockton – on – Tees , UK), Shikoku Chemicals (Tokushima, Japan) and Visko - R (rayon industry, Russia) .

5 . Applications

Sodium sulfate used to dry an organic liquid.

5 . 1 . Commodity industries

sodium sulfate is a very cheap material . The largest use is as filler in powdered home laundry detergents, consuming approx. 50 % of world production. This use is waning as domestic consumers are increasingly switching to compact or liquid detergents that do not include sodium sulfate .

Another formerly major use for sodium sulfate, notably in the USA and Canada, is in the Kraft process for the manufacture of wood pulp. Organics present in the "black liquor" from this process are burnt to produce heat, needed to drive the reduction of sodium sulfate to sodium sulfide. However, this process is being replaced by newer processes; use of sodium sulfate in the USA and Canadian pulp industry declined from 1.4 Mt/a in 1970 to only approx. 150,000 tones in 2006 .

The glass industry provides another significant application for sodium sulfate, as second largest application in Europe. Sodium sulfate is used as a fining agent, to help remove small air bubbles from molten glass. It fluxes the glass, and prevents scum formation of the glass melt during refining. The glass industry in Europe has been consuming from 1970 to 2006 a stable 110,000 tonnes annually .

Sodium sulfate is important in the manufacture of textiles, particularly in Japan, where it is the largest application. Sodium sulfate helps in "levelling", reducing negative charges on fibres so that dyes can penetrate evenly. Unlike the alternative sodium chloride, it does not corrode the stainless steel vessels used in dyeing. This application in Japan and USA consumed in 2006 approximately 100,000 tones .

5 . 2 . Thermal storage

The high heat storage capacity in the phase change from solid to liquid, and the advantageous phase change temperature of 32 degrees Celsius (90 degrees Fahrenheit) makes this material especially appropriate for storing low grade solar heat for later release in space heating applications. In some applications the material is incorporated into thermal tiles that are placed in an attic space while in other applications the salt is incorporated into cells surrounded by solar–heated water. The phase change allows a substantial reduction in the mass of the material required for effective heat storage (83 calories per gram stored across the phase change, versus one calorie per gram per degree Celsius using only water) , with the further advantage of a consistency of temperature as long as sufficient material in the appropriate phase is available.

5 . 3 . Small - scale applications

In the laboratory, anhydrous sodium sulfate is widely used as an inert drying agent, for removing traces of water from organic solutions.[18] It is more efficient, but slower - acting, than the similar agent magnesium sulfate. It is only effective below about 30 °C, but it can used with a variety of materials since it is chemically fairly inert. Sodium sulfate is added to the solution until the crystals no longer clump together; the two video clips (see above) demonstrate how the crystals clump when still wet, but some crystals flow freely once a sample is dry.

Glauber's salt , the deca hydrate , was historically used as a laxative. It is effective for the removal of certain drugs such as acetaminophen from the body, for example, after an overdose .

In 1953, sodium sulfate was proposed for heat storage in passive solar heating systems. This takes advantage of its unusual solubility properties, and the high heat of crystallization (78.2 kJ / mol) .

Other uses for sodium sulfate include de - frosting windows, in carpet fresheners, starch manufacture, and as an additive to cattle feed.

Lately, sodium sulfate has been found effective in dissolving very finely electroplated micrometer gold that is found in gold electroplated hardware on electronic products such as pins, and other connectors and switches. It is safer and cheaper than other reagents used for gold recovery, with little concern for adverse reactions or health effects .

At least one company makes a laptop computer chill mat using sodium sulfate deca hydrate inside a quilted plastic pad. The material slowly turns to liquid as the heat from the laptop is transferred.

6 . Safety

Although sodium sulfate is generally regarded as non-toxic , it should be handled with care. The dust can cause temporary asthma or eye irritation; this risk can be prevented by using eye protection and a paper mask. Transport is not limited, and no Risk Phrase or Safety Phrase apply .

Sodium sulfide

Contents

1 – Introduction :

Sodium sulfide is the name used to refer to the chemical compound Na_2S , but more commonly it refers to the hydrate $Na_2S \cdot 9H_2O$. Both are colorless water - soluble salts that give strongly alkaline solutions. When exposed to moist air, Na_2S and its hydrates emit hydrogen sulfide, which smells much like rotten eggs or flatus. Generally, commercially available **sodium sulfide** is not a unique chemical entity, but it is specified as $Na_2S \cdot xH_2O$, where a weight percentage of Na_2S is specified. Commonly available grades have around 60 % Na_2S by weight, which means that x is around 3. Such technical grades of sodium sulfide have a yellow appearance. These grades of sodium sulfide are marketed as 'sodium sulfide flakes'. Although the solid is yellow, solutions of it are colorless.

Other names : Disodium sulfide	
UN number	1385 (anhydrous) 1849 (hydrate)
Molecular formula	$Na_2 S$
Molar mass	78 g / mol (anhydrous) 240 g / mol (nono hydrate)
Appearance	Colorless , hygroscopic solid
Density	1.85 g / cm^3 (anhydrous) 1.58 g / cm^3 (penta hydrate) 1.43 g / cm^3 (nono hydrate)
Melting Point	1176 °C (anhydrous)

	100 °C (penta hydrate) 50 °C (nono hydrate)
Solubility in Water	186 g / L (20 °C) 390 g / L (50 °C)
Solubility	Insoluble in ether Slightly soluble in alcohol
Crystal structure	Anti fluorite (cubic)
EU classification	Corrosive (**C**) Dangerous for the environment (**N**)
Auto ignition temperature	> 480 ºC

2 - Structure

Na_2S adopts the anti fluorite structure , which means that the Na^+ centers occupy sites of the fluoride in the CaF_2 framework, and the larger S^{2-} occupy the sites for Ca^{2+}. In solution, the salt, by definition, dissociates. The di anion S^{2-} does not, however, exist in appreciable amounts in water. Sulfide is too strong a base to coexist with water. Thus, the dissolution process can be described as follows:

$$Na_2S(s) + H_2O(l) \rightarrow 2Na^+(aq) + HS^- + OH^-$$

Sodium sulfide can oxidize when heated to sodium carbonate and sulfur dioxide :

$$2\ Na_2S + 3\ O_2 + 2\ CO_2 \rightarrow 2\ Na_2CO_3 + 2\ SO_2$$

3 - Production

Industrially Na_2S is produced by reduction of Na_2SO_4 with carbon, in the form of coal :

$$Na_2SO_4 + 4\ C \rightarrow Na_2S + 4\ CO$$

In the laboratory, the anhydrous salt can be prepared by reduction of sulfur with sodium in anhydrous ammonia. Alternatively, sulfur can be reduced by sodium in dry THF with a catalytic amount of naphthalene :

$$2\,Na + S \rightarrow Na_2S$$

4 - Uses

It is primarily used in pulp and paper industry in the kraft process . It is used in water treatment as an oxygen scavenger agent, in the photographic industry to protect developer solutions from oxidation, in textile industry as a bleaching, as a desulfurising and as a dechlorinating agent and in leather trade for the sulfitisation of tanning extracts. It is used in chemical manufacturing as a sulfonation and sulfo methylation agent. It is used in the production of rubber chemicals, sulfur dyes and other chemical compounds. It is used in other applications including ore flotation, oil recovery, food preservative, making dyes, and detergent.

5 - Safety

Like sodium hydroxide, sodium sulfide is strongly alkaline and can cause skin burns. Acids react with it to rapidly produce hydrogen sulfide, which is a toxic and foul - smelling gas.

Sodium Sulfite

Contents

1 – Introduction :

Sodium sulfite (sodium sulphite) is a soluble compound of sodium. Its chemical formula is Na_2SO_3. It has a molecular weight of 126 . It is a product of SO_2 scrubbing, a part of the flue gas desulfurization process. It is also used as a preservative to prevent dried fruit from discoloring, and for preserving meats, and is used in the same way as sodium thiosulfate to convert elemental halides to their respective acids, in photography and for reducing chlorine levels in pools.

IUPAC Name	Sodium Sulfite
Other Names	Sodium Sulphite Hypo clear (photography) E221
Molecular Formula	$Na_2 SO_3$
Molar mass	126.0418 g/mol
Appearance	white, granular
Density	2.633 g / cm³ (anhydrous) 1.561 g / cm³ (hepta hydrate)
Melting point	251.8°C (525 K) (anhydrous) hepta hydrate loses $7H_2O$ at 3.4°C
Boiling point	decomposes
Solubility in water	67.8 g /100 ml of hepta hydrate (18°C)
Solubility in other solvents	3 g /100 ml (?°C) (hepta hydrate)
Crystal structure	hexagonal (anhydrous) monoclinic (hepta hydrate)

2 - Applications :

Sodium sulfite is primarily used in the pulp and paper industry. It is used in water treatment as an oxygen scavenger agent , in the photographic industry to protect developer solutions from oxidation and (as hypo clear solution) to wash fixer (sodium thiosulfate) from film and photo-paper emulsions, in the textile industry as a bleaching, desulfurizing and dechlorinating agent and in the leather trade for the sulfitization of tanning extracts. It is used in the purification of TNT for military use. It is used in chemical manufacturing as a sulfonation and sulfo methylation agent. It is used in the production of sodium thio sulfate . It is used in other applications, including ore flotation, oil recovery, food preservatives, making dyes, and detergent. It forms a bi sulfite adduct with aldehydes, and with ketones forms a sulfonic acid. It is used to purify or isolate aldehydes and ketones.

3 - Descriptive chemistry :

Sodium sulfite is decomposed by even weak acids, giving up sulfur dioxide gas.

$$Na_2SO_3 + 2\ H^+ \rightarrow 2\ Na^+ + H_2O + SO_2$$

A saturated aqueous solution has pH of ~9. Solutions exposed to air are eventually oxidized to sodium sulfate. If sodium sulfite is allowed to crystallize from aqueous solution at room temperature or below, it does so as a hepta hydrate. The hepta hydrate crystals effloresce in warm dry air. Hepta hydrate crystals also oxidize in air to form the sulfate. The anhydrous form is much more stable against oxidation by air.

4 - Hazards :

Parenteral injection containing sodium sulfite may cause life - threatening allergic reactions .

Sodium Thio Sulfate

Contents

1 – Introduction :

Sodium thio sulfate ($Na_2S_2O_3$) is a colorless crystalline compound that is more familiar as the penta hydrate , $Na_2S_2O_3 \bullet 5H_2O$ an efflorescent, monoclinic crystalline substance also called sodium hyposulfite or “hypo.”

The thio sulfate anion is tetrahedral in shape and is notionally derived by replacing one of the oxygen atoms by a sulfur atom in a sulfate anion. The S-S distance indicates a single bond, implying that the sulfur bears significant negative charge and the S - O interactions have more double bond character. The first protonation of thio sulfate occurs at sulfur.

IUPAC Name	Sodium thio sulfate
Other Names	Sodium hyposulfite , Hypo sulphite of soda
Appearance	White crystals
Density	1.667 g / cm³, solid
Melting point	48.3 °C
Boiling point	Decomposes
Solubility in water	Very Soluble
EU classification	Non - toxic.
Flash point	Non flammable

2 - Industrial production and laboratory synthesis :

On an industrial scale, sodium thio sulfate is produced chiefly from liquid waste products of sodium sulfide or sulfur dye manufacture. In the laboratory, this salt can be prepared by heating an aqueous solution of sodium sulfite with sulfur.

3 - Principal reactions and applications :

Thio sulfate anion characteristically reacts with dilute acids to produce sulfur , sulfur dioxide and water :

$$Na_2S_2O_3 + 2HCl \rightarrow 2NaCl + S + SO_2 + H_2O$$

This reaction has been employed to generate colloidal sulfur. When the protonation is conducted at low temperatures, $H_2S_2O_3$ (thio sulfuric acid) can be obtained. It is a strong acid $pK_a = 0.6, 1.7$.

3 – 1 - Iodometry :

In analytical chemistry, the most important use comes from the fact that the thio sulfate anion reacts stoichio metrically with iodine, reducing it to iodide as it is oxidized to tetra thionate :

$$2\ S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2\ I^-(aq)$$

Due to the quantitative nature of this reaction, as well as the fact that $Na_2S_2O_3 \bullet 5H_2O$ has an excellent shelf – life , it is used as a titration in iodometry. $Na_2S_2O_3 \bullet 5H_2O$ is also a component of *iodine clock* experiments.

This particular use can be set up to measure the oxygen content of water through a long series of reactions. It is also used in estimating volumetrically, the concentrations of certain compounds in solution (hydrogen peroxide, for instance), and in estimating the chlorine content in commercial bleaching powder and water.

3 - 2 - Photographic processing :

The terminal sulfur atom in $S_2O_3^{2-}$ binds to soft metals with high affinity. Thus, silver halides, e.g. Ag Br , typical components of photographic emulsions, dissolve upon treatment with aqueous thio sulfate:

$$2\ S_2O_3^{2-} + Ag\ Br \rightarrow [\ Ag\ (S_2O_3)_2\ ]^{3-} + Br^{-}$$

In this application to photographic processing, discovered by John Herschel and used for both film and photographic paper processing, the sodium thio sulfate is known as a *photographic fixer*, and is often referred to as *hypo*, from the original chemical name, *hypo sulphite of soda.*

3 - 3 - Gold extraction :

Sodium thio sulfate is one component of an alternative lixiviant to cyanide for extraction of gold.[3] It forms a strong complex with gold(I) ions, $[Au(S_2O_3)_2]^{3-}$. The advantage of this approach is that thio sulfate is essentially non-toxic and that ore types that are refractory to gold cyanidation (e.g. carbonaceous or Carlin type ores) can be leached by thio sulfate. Some problems with this alternative process include the high consumption of thio sulfate, and the lack of a suitable recovery technique, since $[Au(S_2O_3)_2]^{3-}$ does not adsorb to activated carbon, which is the standard technique used in gold cyanidation to separate the gold complex from the ore slurry.

4 – Medical :

It is used as an antidote to cyanide poisoning . Thio sulfate acts as a sulfur donor for the conversion for cyanide to thio cyanate (which can then be safely excreted in the urine), catalyzed by the enzyme rhodanase .

It is also used as treatment of calciphylaxis in hemodialysis patients with end-stage renal disease .

5 - Other uses :

Sodium thio sulfate is also used :

- As a component in hand warmers and other chemical heating pads that produce heat by exothermic crystallization of a super cooled solution.
- In Bleach
- In pH testing of bleach substances. The universal indicator and any other liquid pH indicator are destroyed by bleach, rendering them useless for testing the pH. If one first adds sodium thio sulfate to such solutions, it will neutralize the color-removing effects of bleach and allow one to test the pH of bleach solutions with liquid indicators. The relevant reaction is akin to the iodine reaction: thio sulfate reduces the hypochlorite (active ingredient in bleach) and in so doing becomes oxidized to sulfate. The complete reaction is:

$$4\ NaClO + Na_2S_2O_3 + 2\ NaOH \rightarrow 4\ NaCl + 2\ Na_2SO_4 + H_2O$$

- To dechlorinate tap water for aquariums or treat effluent from waste water treatments prior to release into rivers. The reduction reaction is analogous to the iodine reduction reaction. Treatment of tap water requires between 0.1 grams and 0.3 grams of penta hydrated (crystalline) sodium thiosulfate per 10 liters of water.
- To lower chlorine levels in swimming pools and spas following super chlorination.
- To remove iodine stains, e.g. after the explosion of nitrogen tri iodide .
- In bacteriological water assessment.
- In the tanning of leather.
- To demonstrate the concept of reaction rate in chemistry classes. The thio sulfate ion can decompose into the sulfite ion and a colloidal suspension of sulfur, which is opaque. The equation for this acid- catalyzed reaction is as follows :

$$S_2O_3^{2-}(aq) \rightarrow SO_3^{2-}(aq) + S(s)$$

- To demonstrate the concept of supercooling in physics classes. Melted sodium thio sulfate is very easy to overcool to room temperature and when crystallization is forced, the sudden temperature jump to 48.3°C can be experienced by touch.
- As part of patina recipes for copper alloys.
- Often used in pharmaceutical preparations as an anionic surfactant to aid in dispersion.

Sodium Tri Poly Phosphate

Contents

1 – Introduction :

Sodium tri poly phosphate (STPP , some times STP or sodium tri phosphate or TPP) , with formula $Na_5\ P_3\ O_{10}$, is a polyphosphate of sodium . It is the sodium salt of tri phosphoric acid.

IUPAC Name	sodium tri phosphate
Molecular Formula	$Na_5\ O_{10}\ P_3$
Molar Mass	368 g / mol

2 - Preparation :

Industrially sodium tri poly phosphate is prepared by heating a stoichiometric mixture of di sodium phosphate, Na_2HPO_4 and mono sodium phosphate , NaH_2PO_4 in carefully controlled conditions.

$$2Na_2HPO_4 + NaH_2PO_4 \rightarrow Na_5P_3O_{10} + 2H_2O$$

3 - Uses :

It is used in various applications such as a preservative for seafood, meats, poultry and pet foods. It is also used in toothpaste and as a builder in soaps and detergents, improving their cleansing ability.

The United States Food and Drug Administration lists STPP as "generally recognized as safe", along with salt, vinegar, and baking powder.

STPP is a solid inorganic compound used in a large variety of house hold cleaning products, mainly as a builder, but also in human foodstuffs, animal feeds, industrial cleaning processes and ceramics manufacture. STPP is widely used in regular and compact laundry detergents and automatic dishwashing detergents (in powder, liquid, gel and/or tablet form), toilet cleaners, surface cleaners, and coffee urn cleaners . It also provides a number of chemical functions including: sequestration of "water hardness", enabling surfactants to function effectively; pH buffering; dirt emulsification and prevention of deposition; hydrolysis of grease; and dissolving-dispersing dirt particles.

3 – 1 - Food Applications :

In foods , STPP is used to retain moisture. Many governments regulate the quantities allowed in foods, as it can substantially increase the sale weight of seafood in particular.

Many people find STPP to add an unpleasant taste to food, particularly delicate seafood. The taste tends to be slightly sharp and soapy and is particularly detectable in mild-tasting foods. The increased water holding properties can also lead to a more diluted flavor in the food.

4 - Health effects :

4 – 1 - Toxic kinetics and acute toxicity :

Poly phosphates are hydrolyzed into smaller units (ortho phosphates) in the gut before absorption , which may induce a metabolic acidosis. The acute toxicity of poly phosphonates is low as the lowest LD50 after oral administration is > 1,000 mg/kg body weight.

4 – 2 - Skin and eye irritation :

Poly phosphates are moderately irritating to skin and mucous membrane because of their alkalinity.

4 – 3 - Mutagenicity and carcinogenicity :

No mutagenic potential was observed when TTP was tested in a Salmonella / micro some assay (Ames test) and in a chromosomal aberration assay in vitro using a Chinese hamster fibroblast cell line (Ishidate et al. 1984). Tetra sodium pyrophosphate was not mutagenic in an in vitro assay using S. cerevisiae strains and S. typhimurium strains with and without the addition of mammalian metabolic activation preparations (IPCS 1982).

4 – 4 - Reproductive toxicity :

Sodium tri poly phosphate showed no maternal toxicity or teratogenic effects at dose levels up to 238 mg/kg body weight in mice and 40 mg/kg in rats (IPCS 1982). Reproduction studies in three generations of rats on diets with 0.5 % TTP were performed. TTP had no effects on fertility or litter size, or on growth or survival on offspring (Hodge 1964). Tetrasodi

5 - Environmental Effects :

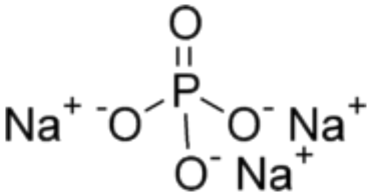
In 2000, the total consumption of STPP in these applications was estimated to be approximately 300,000 tones in Western Europe and is estimated to represent 90 - 95 % of STPP use in Europe. Due to its physico - chemical properties, STPP is not distributed or transported to the atmosphere, and thus is not expected to end up in soil via atmospheric deposition. Because it is very water-soluble, it is not significantly transferred to sewage sludge, and therefore to soil by sludge spreading. No environmental risk related to STPP use in detergents is indicated in soil or air. As an ingredient of household cleaning products, STPP present in domestic waste waters is mainly discharged to the aquatic compartment, directly, via waste water treatment plants, via septic tanks, infiltration or other autonomous waste water systems.

As STPP is an inorganic substance, biodegradation studies are not applicable. However, STPP can be hydrolyzed, finally to ortho phosphate, which can be assimilated by algae and / or by micro - organisms. STPP thus ends up being assimilated into the natural phosphorus cycle. Reliable published studies confirm biochemical understanding, showing that STPP is progressively hydrolyzed by biochemical activity in contact with waste waters (in sewerage pipes and within sewage works) and also in the natural aquatic environment. This information enabled the calculation of “worst case” PEC (Predicted Environmental Concentrations) using the EUSES model and the HERA detergent scenario. A default regional release of 10 % was applied instead of the 7 % regional release indicated in the HERA detergent scenario. Reliable acute aquatic ecotoxicity studies are available which show that STPP is not toxic to aquatic organisms: all EC/LC50 are above 100 mg / l (*Daphnia*, fish, algae). Because of this, and because of the only temporary presence of STPP in the aquatic environment (due to hydrolysis), no studies have been carried out to date concerning the chronic effects of STPP on these aquatic organisms.

5 – 1 - Effects of waste water containing phosphorus :

Detergents containing phosphorus contribute together with other sources of phosphorus to the eutrophication of many fresh waters. Eutrophication is an increase in chemical nutrients -- typically compounds containing nitrogen or phosphorus -- in an ecosystem. It may occur on land or in water. The term is however often used to mean the resultant increase in the ecosystem's primary productivity (excessive plant growth and decay), and further effects including lack of oxygen and severe reductions in water quality, fish, and other animal populations . Phosphorus can theoretically generate its weight 500 times in algae (Wetzel 1983) . Whereas the primary production in marine waters is mainly nitrogen limited, freshwaters are considered to be phosphorus limited. A large part of the sewage effluents in many countries is released untreated into freshwater recipients, and here the use of phosphorus as complexing agents is still an environmental concern.

Tri Sodium Phosphate



Contents

1 – Introduction :

Tri sodium phosphate (TSP , E339) is a cleaning agent, food additive, stain remover and degreaser. It is a white, granular or crystalline solid, highly soluble in water producing an alkaline solution. The item of commerce is often partially hydrated and may range from anhydrous tri sodium phosphate, Na_3PO_4, to the dodeca hydrate, $Na_3PO_4 \cdot 12H_2O$. Most often found in white powder form, it can also be called **tri sodium ortho phosphate** or just plain **sodium phosphate**. Tri sodium phosphate was at one time extensively used in formulations for a wide variety of consumer grade soaps and detergents, but ecological problems have largely ended that practice, at least in the western world. Substitutes are not as effective, but the raw chemical can be bought in bulk to add to underpowered detergents.

IUPAC Name : Tri sodium phosphate
Other Names : Sodium phosphate tri basic
Molecular Formula $Na_3\,PO_4$

Molar Mass	164 g / mol
Density	1.620 g / cm^3 (dodeca hydrate)
Melting point	73.5 °C decomp. (dodeca hydrate)
Solubility in water	1.5 g/100 mL (0 °C) 8.8 g/100 mL (25 °C)
Basicity (pK_b)	2.23
Crystal structure	Trigonal
Flash point	Non - flammable

2 - Uses :

2 – 1- Cleaning :

The major use for tri sodium phosphate is in cleaning agents. The pH of a 1 % solution is 12 , and the solution is sufficiently alkaline to saponify grease and oils . In combination with surfactants, TSP is an excellent agent for cleaning everything from laundry to concrete driveways. This versatility and low manufacturing price, made TSP the preferred basis for a plethora of cleaning products sold in the mid - 20th century. TSP is still sold, and used, as a cleaning agent, but during the late 1960s in the United States,[*citation needed*] government regulators in seventeen states determined that overuse led to a series of ecological problems .

By the end of the 20th century, many products that formerly contained TSP were manufactured with **TSP Substitutes**, which consist mainly of sodium carbonate along with various admixtures of nonionic surfactants and a limited percentage of sodium phosphates.

TSP is commonly used after cleaning with mineral spirits in order to remove hydrocarbon residues . TSP may be used with household chlorine bleach in the same solution without hazardous reactions . This mixture is particularly good for removing mildew.

Although it is still the active ingredient in some toilet bowl cleaning tablets, TSP is generally not good for cleaning bathrooms, because it can corrode metal and can damage grout .

2 – 2 - Flux :

In the United States, tri sodium phosphate is an approved flux for use in hard soldering joints in medical grade copper plumbing. The flux is applied as a concentrated water solution and dissolves copper oxides at the temperature used in copper brazing . Residues are fully water soluble and can be rinsed out of plumbing before it is put in service.

TSP is used as an ingredient in fluxes designed to deoxygenate nonferrous metals for casting.

TSP can be used in ceramic production to lower the flow point of glazes.

2 – 3 - Painting enhancement :

TSP is still in common use for the cleaning , degreasing and deglossing of walls prior to painting. TSP breaks the gloss of oil based paints and opens the pores of latex based paint providing a surface better suited for the adhesion of the subsequent layer of paint .

2 – 4 - Food additive :

Tri sodium phosphate is approved as a food additive in the EU and other countries. It has the E number E339 and is used as an acidity regulator in a variety of food products.

2 – 5 - Exercise performance enhancement :

TSP has gained a following as a nutritional supplement that can improve certain parameters of exercise performance . The basis of this belief is the fact that phosphate is required for the energy-producing Krebs cycle central to aerobic metabolism . Phosphates are available from a number of other sources that are much milder than TSP. While tri sodium phosphate is not toxic per se , it is severely irritating to gastric mucosa unless used as part of a buffered solution. Studies have shown that loading with TSP may increase an athlete's VO2 Max as well as reducing the accumulation of lactic acid in the muscles . Dosing is typically 3- 4 grammas per day for three days prior to competition.

3 - TSP Substitutes :

Products sold as TSP substitute, containing soda ash and zeolites are promoted as a direct substitute . How ever, sodium carbonate is not as strongly basic as tri sodium phosphate, making it less effective in demanding applications. Zeolites are added to laundry detergents as bulking agents that rapidly break down in water and are essentially nonpolluting. Even cleaning products labeled as TSP may contain other ingredients as well, and may, in fact, be less than half tri sodium phosphate.

4 - Phosphates

Many similar chemicals, including tri sodium phosphate, are included in the umbrella term, *phosphates*. Many phosphates are used as cleaners, pH buffers, surfactants, detergents, food-grade acidifiers, emulsifiers, and food preservatives. Some phosphates which are similar to TSP include:

- Sodium tri phosphate, a water softener used in dishwashing detergent and a food preservative.
- Sodium di hydrogen phosphate, commonly termed monosodium phosphate, (NaH_2PO_4), is a laxative a pH buffer
- Disodium hydrogen phosphate, commonly termed disodium phosphate, (Na_2HPO_4), is a laxative and anti-caking agent used in food.
- Sodium aluminium phosphate, ($Na_8Al_2(OH)_2(PO_4)_4$) is a pH buffer and an emulsifying agent used in food.