and is imported into Europe in enormous quantities as cubic nitre for the preparation of saltpetre. Cryolite, a fluoride of aluminium and sodium, is extensively mined in Greenland and elsewhere for industrial purposes. These form the principal natural sources of sodium compounds—the chloride as rock salt and in sea-water being of such predominating importance as quite to outweigh all the others. But it is questionable whether, taken altogether, the mass of sodium they represent is as much as that disseminated throughout the rocky crust in the form of soda felspar (*i.e.* as silicate of soda) and in other soda-contain­ing rocks. From this source all soils contain small proportions of sodium in soluble forms, hence the ashes of plants, although they preferably imbibe potassium salts, contain traces and sometimes notable quantities of sodium salts. Sodium salts also form essential ingredients in all animal juices.

Although many sodium compounds have been known from very remote times, the element was not isolated until 1807, when Sir H. Davy obtained it by electrolysing caustic soda. This method was followed by that proposed by Gay-Lussac and Thénard, who decomposed molten caustic soda with red-hot iron; and this in turn was succeeded by Brunner’s process of igniting sodium carbonate with charcoal. Deville made many improvements, but the method remained wasteful and uneconomical, and in 1872 the metal cost 4s. a pound. In 1886, however, Castner replaced the carbonate by caustic soda, and materially cheapened the cost of production; but this method was discarded for an electrolytic one, patented by Castner in 1890. Electrolytic processes had, in fact, been considered since 1851, when Charles Watt patented his method for the production of sodium and potassium from fused chlorides. Among the difficulties here to be contended with are the destructive action of fused chlorides and of the reduced alkali metals upon most non-metallic sub­stances available for the containing vessel and its partition, and also of the anode chlorine upon metals; also the low fusing-point (95° C. for sodium, and 62° C. for potassium) and the low specific gravity of the metals, so that the separated metal floats as a fused layer upon the top of the melted salt. Again, pure sodium chloride melts at about 775° C., while sodium boils at 877° C., so that the margin of safety is but small if loss by vaporization is to be prevented. Borchers endeavoured to con­tend against the first difficulty by employing an iron cathode vessel and a chamotte (fire-clay) anode chamber united by a specially constructed water-cooled joint. The other difficulty is to some extent met by using mixed chlorides *(e.g.* sodium, potassium and strontium chlorides for sodium extraction), as these melt at a lower temperature than the pure chloride. In Castner's process (as employed at Oldbury and Niagara Falls and in Germany) fused caustic soda is electrolysed. The apparatus described in the patent specification is an iron »cylinder heated by gas rings below, with a narrower cylinder beneath, through which passes upwards a stout iron cathode rod cemented in place by caustic soda solidified in the narrower vessel. Iron anodes are suspended around the cathode, and between the two is a cylinder of iron gauze at the bottom with a sheet-iron continuation above, the latter being provided with a movable cover. During electrolysis, oxygen is evolved at the anode and escapes from the outer vessel, while the sodium deposited in globules on the cathode floats upwards into the iron cylinder, within which it accumulates, and from which it may be re­moved at intervals by means of a perforated iron ladle, the fused salt, but not the metal, being able to pass freely through the perforations. The sodium is then cast into moulds. Sodium hydroxide has certain advantages compared with chloride, although it is more costly; its fusing-point is only 320° C., and no anode chlorine is produced, so that both containing vessel and anode may be of iron, and no porous partition is necessary.

Metallic sodium possesses a silvery lustre, but on exposure to moist air the surface is rapidly dulled by a layer of the hydroxide. It may be obtained crystallized in the quadratic system by melting in a sealed tube containing hydrogen, allowed to cool partially, and then pouring off the still liquid portion by inverting the tube. The specific gravity is 0∙9735 at 13∙5° (Baumhauer). At ordinary temperatures the metal has the consistency of wax and can be readily cut; on cooling it hardens. On heating it melts at 95∙6° (Bunsen) to a liquid resembling mercury, and boils at 877∙5° (Ruff and Johannsen, *Ber.,* 1905, 38, p. 3601), yielding a vapour, colourless in thin layers but a peculiar purple, with a greenish fluorescence, when viewed through thick layers. (For the optics of sodium vapour see R. W. Wood, *Physical Optics.)* According to A. Matthiessen, sodium ranks fourth to silver, copper and gold as a conductor of electricity and heat, and according to Bunsen it is the most electropositive metal with the exception of caesium, rubidium and potassium.

The metal is very reactive chemically. Exposed to moist air it rapidly oxidizes to the hydroxide; and it burns on heating in air with a yellow flame, yielding the monoxide and dioxide. A fragment thrown on the surface of water rapidly disengages hydrogen, which gas, however, does not inflame, as happens with potassium; but inflammation occurs if hot water be used, or if the metal be dropped on moist filter paper. Sodium also combines directly, sometimes very energetically, with most non-metallic elements. It also combines with dry ammonia at 300-400° to form sodamide, NaNH2, a white waxy mass when pure, which melts at 155°. Heated in a current of carbon dioxide sodamide yields caustic soda and cyanamide, and with nitrous oxide it gives sodium azoimide; it deflagrates with lead or silver nitrate and explodes with potassium chlorate. Sodamide was introduced by Claisen (Ber., 1905, 38, p. 693) as a condensing agent in organic chemistry, and has since been applied in many directions. Sodium is largely employed in the manufacture of cyanides and in reduction processes leading to the isolation of such elements as magnesium, silicon, boron, aluminium (formerly), &c.; it also finds application in organic chemistry. With potassium it forms a liquid alloy resembling mercury, which has been employed in high temperature thermometers (see Thermometry).

*Compounds.*

In its chemical combinations sodium is usually monovalent; its salts arc generally soluble in water, the least soluble being the metantimonate.

Sodium hydride, NaH, is a crystalline substance obtained directly from sodium and hydrogen. at about 400°. It burns when heated in dry air, and ignites in moist air; it is decomposed by water, giving caustic soda and hydrogen. Dry carbon dioxide is decomposed by it, free carbon being produced; moist carbon dioxide, on the other hand, gives sodium formate.

Several oxides are known. A suboxide, Na3O, appears to be formed as a grey mass when a clean surface of the metal is exposed to air, or when pure air is passed through the metal just above its melting point (De Forcrand, *Compt. rend.,* 1898, 127, pp. 364, 514). The monoxide, Na2O,.is obtained by heating the metal above 180°in a limited amount of slightly moist oxygen (Holt and Sims, *Journ. Chem. Soc.,* 1894., i. 442); it may also be prepared by heating the nitrate or nitrite with metallic sodium, free nitrogen being eliminated (German patent, 142467, 1902). It forms a grey mass, which melts at a red heat and violently combines with water to give the hydroxide. The hydroxide or caustic soda, NaOH, is usually manufactured from the carbonate or by electrolysis of salt solution (see Alkali Manufacture). When anhydrous it is a colourless opaque solid which melts at 310°, and decomposes at about 1100°. It is very soluble in water, yielding a strongly alkaline solution; it also dissolves in alcohol. It absorbs moisture and carbon dioxide from the atmosphere. Several hydrates are known: 2NaOH∙7H2O is obtained as large monoclinic crystals by cooling a solution of specific gravity 1∙365 to -8°; Pickering *(Journ. Chem. Soc.,* 1893, 65, p. 890) obtained NaOH∙H2O from hot concentrated solutions and NaOH∙2H2O from a solution of the hydroxide in 96∙8% alcohol. (See also De Forcrand, *Compt. rend.,* 1901, 133, p. 223.)

Sodium dioxide, Na2O2, is formed when the metal is heated in an excess of air or oxygen. In practice the metal is placed on aluminium trays traversing an iron tube heated to 300°, through which a current of air, freed from moisture and carbon dioxide, is passed; the process is made continuous, and the product contains about 93% Na2O2. When pure, sodium dioxide has a taint yellowish tinge, but on exposure it whitens (W. R. Bousfield and T. M. Lowry, *Phil. Trans.,* 1905, A. 204, p. 253). When dissolved in water it yields some NaOH and H2O2; on crystallizing a cold solution Na2O2·8H2O separates as large tabular hexagonal crystals, which on drying over sulphuric acid give Na2O2·2H2O; the former is also obtained by precipitating a mixture of caustic soda and hydrogen peroxide solutions with alcohol. Acids yield a sodium salt and free oxygen or hydrogen peroxide; with carbon dioxide it gives sodium carbonate