

A TEXT-BOOK OF CHEMISTRY

INTENDED FOR THE USE OF PHARMA-
CEUTICAL AND MEDICAL STUDENTS

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FIFTH EDITION REVISED AND REWRITTEN

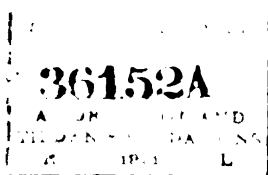


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PREFACE TO FIFTH EDITION

THE present edition of this pharmaceutical text-book has been in course of preparation for several years, having been begun before the appearance of the 9th Revision of the U. S. Pharmacopœia.

As is well known, chemical manufacturing has undergone quite important changes in recent years, and in particular since the revolution in trade incident to the outbreak of the present great world war. The Editors have sought to bring the book abreast of this changed condition in the chemical world.

The sections on Inorganic Chemistry have been therefore largely rewritten and much that is new introduced, and older statements withdrawn in numerous sections.

This part of the work has been attended to by Professor Coblenz, who has been of recent years in close touch with the manufacturing side of chemistry, and Professor Jeannot Hostmann of the Department of Pharmacy of Columbia University, who is actively engaged in chemical laboratory work.

The section on Elementary Physics has also been largely rewritten by Professor Hostmann.

The section on Organic Chemistry has been brought up to date with the introduction of considerable new matter by Professor Sadtler, who has also contributed the chapter on Electrolysis and its new applications.

Some new and useful tables have, with the permission of the Trustees of the U. S. Pharmacopœia, been inserted for the greater convenience of the pharmaceutical student and for convenient reference for pharmacists using the book.

We believe that the book in its present form is adapted for use in the classes of our Pharmaceutical Schools as a text-book and at the same time is sufficiently complete in its scheme of treatment to make it valuable for the reference libraries of graduates and others who might turn to it as a reference book on chemistry and chemical products.

SAMUEL P. SADTLER,
VIRGIL COBLENTZ.

September, 1918.

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PREFACE TO THE FIRST EDITION.

AN experience of some seventeen years as teachers in a professional school, where Chemistry is considered one of the foundation-stones upon which young men must build a technical education, has convinced the authors that the subject must be taught in a way that, while thoroughly scientific, is yet distinctly adapted to the ultimate uses of the students.

The professions of Pharmacy and Medicine are closely related, and their uses for Chemistry are very much alike. The members of these professions must understand the composition and chemical properties and possibilities of the whole *materia medica*, both inorganic and organic, in order safely to maintain the important trusts committed to them,—trusts involving the continued health and life of their fellow-men. The importance of thorough chemical knowledge for these related professions is now recognized in the fullest way in their official publication, the U. S. Pharmacopœia, the last edition of which has embodied throughout the most detailed and exact statements of the chemical nature, properties, and reactions of medicinal substances, together with methods of analysis or assay and tests for purity.

We have sought in the following pages to offer a text-book on Chemistry which should combine scientific accuracy and completeness with that special reference to the needs of the pharmaceutical and medical student, as well as those in active professional practice, which will make it of value for both study and reference. Compounds recognized in the U. S. Pharmacopœia are specially indicated, but they are not taken out of their proper position for such discussion.

A brief outline of Elementary Physics, much of which is absolutely essential for the understanding of chemical methods, forms Part I. The convenient division of the elements into Non-Metals and Metals is followed in the main in Parts II. and III., subject to the limitations made necessary by the chemical relationship as indicated by the valence groups. Part IV. deals with Organic Chemistry. While the arrangement here followed is strictly scientific, it has been sought to give the subject a distinctly practical bearing, and to this end the industrial applications of organic compounds have been noted, and in a number of cases illustrated.

In Part V. a brief outline of Qualitative and Quantitative Analysis is given, the methods being such as have proved of value in our own prac-

tice. A number of the Pharmaceutical Assays of the U. S. Pharmacopœia are also added.

The Appendix contains, besides a list of the chemical elements with atomic weights and valences, thermometric scales and the specific gravity tables in most general use.

The authors have made free use of many chemical works, such as Flückiger's "Pharmaceutische Chemie," Schmidt's "Pharmaceutische Chemie," Meyer and Jacobson's "Organische Chemie," Bernthsen's "Organische Chemie," Weyl's "Organische Chemie für Mediciner," Fischer's "Neuere Arzneimittel," Gamgee's "Physiological Chemistry," Pictet's "Alcaloides Végétaux," Power's "Essential Oils," Allen's "Commercial Organic Analysis," Prescott's "Organic Analysis," and Sadtler's "Industrial Organic Chemistry."

Our thanks are due to J. C. Peacock for valuable assistance in the preparation of Parts III. and V., to S. S. Sadtler and H. Blount Hunter for work on the index, and to Edwin Thorpe for the execution of many original drawings for the illustration of the work.

S. P. S.
H. T.

PHILADELPHIA, February, 1895.

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CHAPTER III

THE MAGNESIUM GROUP

MAGNESIUM

Symbol, Mg.

Atomic Weight, 24.32.

Valence, II.

History.—Attention was first directed to the magnesium salts about the close of the seventeenth century by Grew, a physician of London. He found a peculiar salt in the mineral spring at Epsom, which salt soon became celebrated for its medicinal virtues. The metal was first isolated by Davy, in 1808, although he did not obtain it in a pure condition. Bussy was the first to obtain a tangible quantity of the metal, which he accomplished by fusing the chloride with potassium.

Occurrence.—The metal is not found in the uncombined state in nature. Its compounds, however, are almost as widely distributed as those of calcium, although not in such large quantities. The principal minerals containing it are *magnesite*, MgCO_3 ; *dolomite*, $\text{MgCa}(\text{CO}_3)_2$; *kieserite*, $\text{MgSO}_4 \cdot \text{H}_2\text{O}$; *carnallite*, a chloride of magnesium and potassium, KCl , $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$; and the following silicates: *talc*, $3\text{MgO} \cdot 4\text{SiO}_2 \cdot \text{H}_2\text{O}$; *meerschaum*, $2\text{MgO} \cdot 3\text{SiO}_2 \cdot 2\text{H}_2\text{O}$; and *serpentine asbestos*, $3\text{MgO} \cdot 2\text{SiO}_2 \cdot 2\text{H}_2\text{O}$. Magnesium salts are found in plants, in which they may to a certain extent replace the salts of calcium. As phosphate, magnesium is found in the bones and urine of animals.

Isolation.—Magnesium is now usually prepared by the electrolysis of the fused double chloride; that is, carnallite ($\text{MgCl}_2 \cdot \text{KCl}$) to which some fluorspar or sodium chloride is added as a flux. The vessel in which the decomposition is effected is made of especially prepared cast-steel, which acts as the negative pole at which the magnesium is liberated; the positive pole is made of gas-carbon. In order to prevent the oxidation of the metal, a stream of hydrogen or nitrogen is led through the vessel during the operation.

The metal may also be prepared by the reduction of fused magnesium chloride with sodium or aluminum.

The metal is furnished usually either in form of powder ("flash-powder") or ribbon.

Magnesium unites directly with nitrogen at 670° to magnesium nitride.

Properties.—Magnesium is a silver-white, highly lustrous metal, having a specific gravity of 1.75. It is unchangeable in dry air, but in the presence of moisture it gradually becomes covered with a film of oxide. When heated without access of air, magnesium melts at about 650° and boils at 1120° . Heated in the air it ignites, and burns with a blinding white light which is especially rich in actinic rays. Water is decomposed by magnesium but slightly at or above 100° , with evolution of hydrogen and the formation of magnesium oxide.

Uses.—On account of the comparatively low ignition-point (800°), high heat of combustion (134,000 cal.), and relatively low heat of vapor-

ization of magnesium and the actinic activity of the magnesium light, it is useful in photography, and enters into the composition of the numerous "flash-light" powders and illuminants employed for military purposes. Previous to the development of electric lighting it was used in making strong lights for signalling purposes. It is used in pyrotechny. Bengal fires are made as follows: Red—1 part of pulverized shellac, 5 parts of dried strontium nitrate, 2.5 parts of powdered magnesium. Green—1 part of pulverized shellac, 6 parts of dried barium nitrate, 2.5 parts of powdered magnesium. An alloy of magnesium (10 per cent.) with aluminum is known as "magnalium" and is of a silver-white appearance, very light, fuses at 633° , and is used in aviation machines. Magnesium also has some use in the laboratory, where, on account of its freedom from arsenic, it may replace zinc in Marsh's test. In metallurgy magnesium is employed for scavenging alloys; that is, for removing oxides of other metals, as aluminum, nickel, copper, brass, bronze, etc. Magnesium reduces many metallic oxides to their metals, as boric oxide, silica, etc.

MAGNESIUM AND THE HALOGENS

Magnesium Chloride, $\text{MgCl}_2 \cdot 6\text{H}_2\text{O}$.—Sea water and most mineral waters contain magnesium chloride. In the solid state it is found abundantly as carnallite, $\text{KCl} \cdot \text{MgCl}_2 \cdot 6\text{H}_2\text{O}$, at Stassfurt; also in a number of the double salts in that locality. It may be prepared by dissolving magnesite in hydrochloric acid; but, since it is a by-product of not much value in the preparation of potassium chloride from carnallite, it is usually derived from this source.

Properties.—Magnesium chloride occurs in colorless, very deliquescent, monoclinic crystals. It cannot be deprived of its water of crystallization by heat, since, when heated, it decomposes with evolution of hydrochloric acid and formation of magnesium oxychloride, $\text{MgCl} \cdot \text{OH}$. This decomposition may be prevented by adding 3 parts of ammonium chloride for every part of magnesia. A double salt, $\text{MgCl}_2 \cdot 2\text{NH}_4\text{Cl}$, forms, which may be evaporated to dryness without decomposition.

The anhydrous salt is obtained by heating the double chloride, $\text{MgCl}_2 \cdot \text{NH}_4\text{Cl} \cdot 6\text{H}_2\text{O}$, until a crystalline mass, which fuses at 708° , forms; this distills at red heat; or it may be prepared by drying the crystallized chloride in a stream of hydrochloric acid gas.

Anhydrous magnesium chloride occurs in transparent, laminated, pearly plates, which dissolve in water with the evolution of heat. It possesses a bitter, saline taste, and is readily soluble in water. With magnesium oxide it forms an oxychloride which is insoluble in water. Water containing magnesium chloride cannot be used in boilers owing to its liberation of hydrochloric acid which attacks the iron.

Uses.—Magnesium chloride is used in the preparation of a cement

prepared by mixing magnesium oxide and chloride with infusorial earth or sawdust and sufficient water to form a paste. It is also used as a finisher for cotton and woolen textiles.

Magnesium Ammonium Chloride, $\text{MgCl}_2 \cdot \text{NH}_4\text{Cl} \cdot 6\text{H}_2\text{O}$, is a double salt which is usually prepared by evaporating a solution of 3 parts of ammonium chloride with one part of magnesium chloride. The addition of ammonia water does not precipitate the magnesium.

Magnesium Oxychloride.—When an aqueous solution of magnesium chloride of 1.16 to 1.26 specific gravity is mixed with freshly-prepared magnesium oxide, the mass after some time hardens sufficiently to be capable of taking a high polish. It has a composition represented by the formula $\text{Mg}_{11}\text{Cl}_2(\text{OH})_{22} \cdot 4\text{H}_2\text{O}$. On exposure to the air for some time it absorbs carbon dioxide. Another oxychloride is formed when magnesium sulphate, containing ammonium hydroxide and chloride, is exposed to the air for some time: the crystalline precipitate, which is deposited, has the composition $\text{Mg}_2\text{Cl}(\text{OH})_{11} \cdot 4\text{H}_2\text{O}$.

Magnesium Bromide, $\text{MgBr}_2 \cdot 6\text{H}_2\text{O}$, occurs in sea water and in a number of salt springs. It is this salt more than any other that is the source of bromine. It may be prepared by heating magnesium in bromine vapor or dissolving freshly precipitated magnesium hydroxide in hydrobromic acid and evaporating over sulphuric acid.

The anhydrous salt forms in solid, white, crystalline masses, which are very deliquescent, and become hot in contact with water. When the salt contains water of crystallization, it loses hydrobromic acid on the application of heat, leaving a residue of magnesium oxybromide.

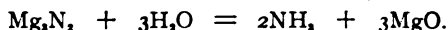
Magnesium Iodide, $\text{MgI}_2 \cdot 8\text{H}_2\text{O}$, occurs associated with the bromide in sea water. It may be prepared by adding iodine to an excess of magnesium powder suspended in water, digesting, filtering, and evaporating over sulphuric acid. It forms very deliquescent, hydrated crystals, which readily decompose with liberation of iodine.

Magnesium Nitride, Mg_3N_2 .—Obtained as a yellow powder when magnesium is heated in an atmosphere of nitrogen or a current of ammonia:

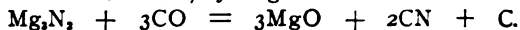


When magnesium is heated in a current of air at 670° , a mixture of oxide and nitride forms.

When the nitride is boiled with water, ammonia is given off:



When heated in a current of CO, cyanogen is obtained:



MAGNESIUM AND OXYGEN

Magnesium Oxide, MgO . **Magnesium Oxidum**, U.S.P.—Black, in 1755, appears to have been the first to prepare this oxide. He noted its solubility in diluted sulphuric acid, which served to distinguish it from lime.

Occurrence.—Magnesium oxide is found native in the rare mineral *periclase*, found near Naples.

Preparation.—Magnesia is prepared on a large scale by heating the carbonate, $(\text{MgCO}_3)_2 \cdot \text{Mg}(\text{OH})_2 \cdot 5\text{H}_2\text{O} = 5\text{MgO} + 6\text{H}_2\text{O} + 4\text{CO}_2$. When the light carbonate is ignited, the light or calcined magnesia, *magnesia usta*, results. When a denser carbonate is used, the result is heavy magnesia, **Magnesium Oxidum Ponderosum**, U.S.P.

The operation is conducted in iron or earthenware crucibles; at 300° to 350° the decomposition readily takes place, and it is known to be complete when a small quantity, taken from the middle of the vessel with an iron spoon, cooled and mixed with water, gives no effervescence with dilute hydrochloric acid.

Probably the most important feature in the successful manufacture of this compound is that of avoiding too high a temperature. In order to prevent parts of the powder from becoming too hot it should be stirred frequently, or it has been recommended to perform the ignition in shallow pans, so that the depth of the carbonate does not exceed 2 to 3 inches. Much also depends upon the purity and physical condition of the carbonate employed.

Properties.—Official “light” or “heavy” magnesia, which after ignition contains not less than 96 per cent. of MgO and not more than 2 per cent. of CaO , occurs as a loose, white, odorless powder; possessing an earthy but not saline taste. It is almost insoluble in water (1 to 55,000), insoluble in alcohol, but readily soluble in dilute acids without effervescence. Being a weak base, magnesium readily forms insoluble basic (oxy) salts with acids. When mixed with water magnesium oxide forms the hydroxide $\text{Mg}(\text{OH})_2$.

The presence of ammonium salts materially increases its solubility in water. In moist air it readily absorbs moisture and carbon dioxide, becoming converted into a basic carbonate.

Magnesia is not visibly affected by any temperature below that of the oxyhydrogen flame, but it will, at lower temperatures, become more dense and crystalline. When heated with distilled water it does not yield more than 2 per cent. of soluble salts (U.S.P.). The heavy magnesia is a denser and finer powder, but in most other characters resembles the light compound.

Magnesium Hydroxide, $\text{Mg}(\text{OH})_2$, occurs in nature as the mineral brucite. It may be prepared by allowing magnesium oxide which has not been overcalcined to stand in contact with water, or more quickly by heating the same compound with water to 150° . Sodium or potassium hydroxide precipitates magnesium hydroxide from solutions of the salts.

Magnesium hydroxide is a white, nearly insoluble powder. It is easily soluble in solutions of ammonium salts. Magnesium salts are not precipitated by the reagents for the heavy metals, as ammonium hydroxide or hydrosulphide, when ammonium salts as NH_4Cl are present due to the formation of complex magnesium ammonium compounds. At 100° magnesium hydroxide is not affected, but when the temperature approaches redness it is easily converted into oxide. It is sufficiently soluble in water to impart an alkaline reaction.

Magma Magnesiae, U. S. P. **Milk of Magnesia**.—A suspension of magnesium hydroxide in water, obtained by interaction between magnesium carbonate, 125 gm., and a solution of sodium hydroxide, 80 gm. This mixture yields not less than 6.5 per cent. nor more than 7.5 per cent. of $\text{Mg}(\text{OH})_2$.

SALTS OF MAGNESIUM

Magnesium Sulphite, $\text{MgSO}_3 \cdot 6\text{H}_2\text{O}$, may be made by adding sulphurous acid in excess to a mixture of 8 parts of magnesium oxide in 16 parts of water.

Properties.—Magnesium sulphite is a white, crystalline powder, becoming oxidized on exposure to air, odorless, having a slightly bitter, somewhat sulphurous taste, and a neutral or slightly alkaline reaction. It is soluble in 20 parts of water at 15° , and in 19 parts of boiling water; insoluble in alcohol. When heated to 200° the salt loses its water of crystallization, amounting to 50.9 per cent., and is converted into magnesia and anhydrous magnesium sulphate.

Magnesium Sulphate, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$. **Magnesii Sulphas**, U.S.P.—Epsom salt became known about the end of the seventeenth century,

because of its valuable purgative qualities, and was designated on the continent of Europe as bitter salt and *Sal Anglicum*.

Occurrence.—Magnesium sulphate occurs in many mineral waters, the most notable being those of Pullna and Seidlitz, in Bohemia, of Epsom, in England, of Friedrichshall, in Germany, and Hunyadi Janos, in Hungary. Under the name of *kieserite*, $\text{MgSO}_4 \cdot \text{H}_2\text{O}$, magnesium sulphate is found at Stassfurt. The upper layer of this salt deposit, known as *Abraumsalz*, is a mixture of magnesium and potassium salts, chiefly carnallite, kainite, and kieserite.

Preparation.—Formerly much magnesium sulphate was obtained by the concentration of natural bitter waters. At the present time large quantities are obtained as a by-product in the manufacture of carbon dioxide for use in mineral waters. Magnesite is used to furnish this gas, and sulphuric acid is used to effect its liberation, as follows:



After the escape of carbon dioxide, the residue is dissolved in hot water, neutralized with magnesite, some ammonium sulphide added, and the mixture allowed to stand for some time in order to separate iron; finally, the clear liquid is concentrated and allowed to crystallize. The *Abraumsalz* is also employed as a source of magnesium sulphate. This salt is usually suspended in baskets over open evaporators, so that it is brought in contact with the ascending steam; by this means the very soluble magnesium and potassium chlorides are dissolved, while the *kieserite* remains behind. *Kieserite* is not easily dissolved by water, but on contact with steam it disintegrates and is converted into Epsom salt, containing 7 molecules of water; in this form it is dissolved and recrystallized.

Properties.—Magnesium sulphate, "which contains not less than 48.59 nor more than 53.45 per cent. of the anhydrous salt," crystallizes from water in colorless, rhombic prisms, or in small acicular crystals; the latter form is obtained by allowing the crystallization to take place from hot concentrated solutions. The salt is without odor, and has a cooling, saline, bitter taste; the crystals slowly effloresce in dry air. It is soluble in 1 part of water at 25° , and in 0.2 part of boiling water; insoluble in alcohol. When heated to 52° the salt loses 1 molecule of water; from this temperature up to 132° it loses 5 additional molecules, and at somewhat above 200° the remaining seventh molecule of water is removed and the salt becomes anhydrous. The aqueous solution is neutral to litmus.

One molecule of water of crystallization (above 200°) of $\text{MgSO}_4 \cdot \text{H}_2\text{O} \cdot 6\text{H}_2\text{O}$ is readily replaced by sulphates of the alkali metals, yielding double isomorphous salts of the general formula $\text{MgSO}_4 \cdot \text{MeSO}_4 \cdot 6\text{H}_2\text{O}$, as $\text{MgSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$; $\text{MgSO}_4 \cdot (\text{NH}_4)_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$, etc. This may be extended to such other isomorphous sulphates as $\text{NiSO}_4 \cdot 7\text{H}_2\text{O}$, $\text{CoSO}_4 \cdot 7\text{H}_2\text{O}$, and $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$, which form similar double salts with the alkali sulphates.

When anhydrous magnesium sulphate is dissolved in hot concentrated sulphuric acid, six-sided tabular crystals separate on cooling, which have the composition $\text{MgSO}_4 \cdot \text{H}_2\text{SO}_4$.

Uses.—Magnesium sulphate is used in medicine as an active cathartic. It also has some use in dyeing and calico printing as a finisher. When this salt is used with aniline colors, the finished goods are found to withstand the action of soap.

Magnesium Nitrate, $\text{Mg}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$, is prepared by neutralizing nitric acid with magnesium carbonate. It forms deliquescent monoclinic prisms or acicular crystals. From hot solutions it separates with 2 molecules of water of crystallization. The crystals are soluble in 0.5 part of water at 15° , and in 9 parts of alcohol.

Normal Magnesium Phosphate, $\text{Mg}_3(\text{PO}_4)_2$, is found in bones, in the seeds of many plants, and in a number of minerals. It is precipitated when a solution of trisodium orthophosphate is added to one of magnesium sulphate. It is soluble in about 5000 parts of water.

Magnesium Hydrogen Phosphate, MgHPO_4 , is formed when a solution of ordinary sodium phosphate, Na_2HPO_4 , is added to one of magnesium sulphate. On standing in contact with water, this precipitate crystallizes with 7 molecules of water. One part of this crystalline precipitate is soluble in 322 parts of water. If the above precipitation takes place in presence of ammonia and an ammonium salt, the insoluble $\text{Mg}(\text{NH}_4)\text{PO}_4 \cdot 6\text{H}_2\text{O}$ forms.

Ammonium Magnesium Phosphate, $\text{Mg}(\text{NH}_4)\text{PO}_4 \cdot 6\text{H}_2\text{O}$.—This salt frequently separates from alkaline urine, and forms the triple phosphate calculi. It is formed when ammonium chloride followed by ammonium hydroxide is added to a solution of a magnesium salt followed by a solution of sodium phosphate (Na_2HPO_4). The crystals which are thus produced are transparent, quadratic prisms, soluble in 15,000 parts of water at 15° and in 44,000 parts of ammoniacal water. At 100° five molecules of water are given off without loss of ammonia, and at higher temperatures the remainder of the water and the ammonia escape, leaving magnesium pyrophosphate, $\text{Mg}_2\text{P}_2\text{O}_7$. Magnesium as well as phosphorus is determined quantitatively in this form.

Magnesium Arsenate and *Magnesium Ammonium Arsenate* ($\text{MgNH}_4\text{AsO}_4 \cdot 6\text{H}_2\text{O}$) are similar in their method of preparation and properties to the corresponding phosphates.

Magnesium Borate.—When solutions of borax and magnesium sulphate are mixed in the cold, no precipitation takes place, but on long boiling a precipitate forms having the composition $\text{Mg}(\text{BO}_2)_2 \cdot 2\text{Mg}(\text{OH})_2 \cdot 7\text{H}_2\text{O}$. This compound has had some use as an antiseptic.

Magnesium Carbonate, MgCO_3 .—This is the normal magnesium carbonate, and is found abundantly in nature as magnesite. In pharmacy and medicine we are concerned chiefly with the artificially prepared or basic magnesium carbonate, **Magnesii Carbonas**, U.S.P. The official salt is a mixture of hydrated carbonate and hydroxide, corresponding to not less than 39.2 per cent. of MgO containing not more than 0.8 per cent. CaO . This compound became known as a valuable medicinal agent about the beginning of the eighteenth century, under the name of *magnesia alba*.

Preparation.—The light carbonate may be prepared by precipitating magnesium sulphate with sodium carbonate and boiling the mixture, a slight excess of the sulphate being used. The precipitate is collected on a muslin strainer and washed with hot water until the washings cease to give a reaction with barium chloride for sulphuric acid. The heavy carbonate of magnesia is prepared as above, finally evaporating the resulting mixture to dryness. The

dry residue is then washed thoroughly with hot water until all sodium sulphate is removed. The product is a dense powder. The reaction which takes place in these processes is as follows:



The method of Pattinson consists in gently igniting dolomite, a magnesium and calcium carbonate, whereby magnesium oxide forms and the calcium remains as carbonate, then treating the product with water and carbon dioxide under a pressure of 5 to 6 atmospheres. The magnesium carbonate passes into solution as bicarbonate much more readily than the calcium carbonate. The clear solution is separated from the precipitate of the latter, and on heating deposits a precipitate of the official carbonate, with evolution of some carbon dioxide, which may be used again.

Properties.—The official magnesium carbonate comes into commerce in loose, white, easily-pulverized masses, or more usually as a fine, white, amorphous powder, without odor, having an earthy taste, and permanent in the air. It dissolves in about 2500 parts of cold water and about 9000 parts of hot water, and imparts to the solution an alkaline reaction; alcohol exerts no solvent action upon it. Dilute acids dissolve it with active effervescence, and without leaving any residue. When boiled with distilled water, magnesium carbonate does not yield more than 1 per cent. of soluble salts.

Solutions of ammonium salts have considerable solvent action on magnesium carbonate, and in the presence of ammonium chloride no precipitation will occur after the addition of ammonium carbonate.

On the application of heat, water is first given off, and at 300° the decomposition of the carbonate is complete, magnesium oxide remaining.

When the official magnesium carbonate is shaken with water until a uniform mixture results, and carbon dioxide is passed in, there results a clear solution of *acid magnesium carbonate*, $\text{MgH}_2(\text{CO}_3)_2$, which has not been separated in the solid state. This solution, on exposure to the air and warming to 50°, deposits crystals of the neutral carbonate with 3 molecules of water, $\text{MgCO}_3 \cdot 3\text{H}_2\text{O}$, or at a low temperature $\text{MgCO}_3 \cdot 5\text{H}_2\text{O}$ separates. Under ordinary atmospheric pressure 1 part of magnesium carbonate requires about 761 parts of carbonated water to dissolve it, under 3 atmospheres of pressure the solubility is 1 to 134, and at 6 atmospheres it is 1 to 76 of the carbonated water.

Uses.—Magnesium carbonate is used in medicine as an antacid, and in the presence of acids becomes cathartic in its action.

GLUCINUM (BERYLLIUM) (Gl)

Symbol, Gl.

Atomic Weight, 91.

Valence, II.

Occurrence.—Glucinum occurs in nature in a few rare minerals, its name being derived from *glukos* = *sweet*, referring to the sweet taste of its salts. Beryl is a silicate of glucinum and aluminum, $3\text{BeSiO}_3 \cdot \text{Al}_2(\text{SiO}_3)_3$, which occurs in beautiful hexagonal crystals of a green color; the transparent, green-colored varieties are known as emerald (which owes its color to the presence of a trace of chromium), and those of a bluish-green tint (due to FeO) are termed

aquamarine. Phenacite is a glucinum silicate of the composition Be_2SiO_4 , and chrysoberyl is a glucinum aluminate, $\text{BeO} \cdot \text{Al}_2\text{O}_3$.

Isolation.—Metallic glucinum is prepared by electrolysis of the fused double fluoride $\text{NaF} \cdot \text{BeF}_2$.

Properties.—Glucinum is a silver-white, lustrous metal; malleable, and having the specific gravity of 1.84. It is soluble in acids (excepting dilute HNO_3) and dilute alkalis and does not decompose water.

Glucinum Chloride, GlCl_3 , is obtained in the anhydrous condition by passing chlorine over a heated mixture of GlO and charcoal. It is easily fusible and volatile, and may be obtained with 4 molecules of water of crystallization from aqueous solution.

Glucinum Oxide and Hydroxide, GlO , yields with acid salts having a sweet taste. The solutions of these salts form precipitates of the hydroxide, $\text{Gl}(\text{OH})_3$, on the addition of sodium or potassium hydroxide; this hydroxide is distinguished from that of magnesium by being soluble in excess of the reagent, and from the hydroxide of aluminum by its solubility in excess of alkali carbonate. The hydroxide behaves as a weak acid towards strong bases, like aluminum. When heated, glucinum hydroxide yields GlO .

Glucinum Sulphate, $\text{GlSO}_4 \cdot 4\text{H}_2\text{O}$, separates from hot, neutral concentrated solutions, while $\text{GlSO}_4 \cdot 7\text{H}_2\text{O}$ crystallizes from acid solution; it is isomorphous with magnesium sulphate. The double salt $\text{GlSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 3\text{H}_2\text{O}$ is but slightly soluble.

Glucinum Nitrate, $\text{Gl}(\text{NO}_3)_3 \cdot 3\text{H}_2\text{O}$, is an extremely deliquescent salt, easily soluble in alcohol.

Glucinum Carbonate, GlCO_3 .—This is soluble in water and readily gives off CO_2 .

ZINC

Symbol, Zn.

Atomic Weight, 65.37.

Valence, II.

History.—Paracelsus in the sixteenth century first recognized zinc as a metal. It was known, however, long before that certain ores of zinc known as *cadmia* would, when smelted with copper ores, furnish a yellow metal (brass). Works for the preparation of metallic zinc were first established in the year 1743, at Bristol, England.

Occurrence.—Zinc occurs chiefly as a sulphide or *zinc blende* (ZnS), as carbonate or *smithsonite* (ZnCO_3), as a mixture of carbonate and silicate ($2\text{ZnO} \cdot \text{SiO}_2 \cdot \text{H}_2\text{O}$) or *calamine*. The most noted European localities are Belgium, Silesia, Saxony, and England. In the United States fully half the zinc is produced in Missouri, although there are important deposits in New Jersey, Virginia, Wisconsin, Kansas, Arkansas and Montana.

In the United States there are two ores peculiar to the localities named and are not found abroad—namely, the native oxide, called red oxide, or *zincite*, and a mixture of zinc oxide with manganese and iron oxide, known as *franklinite*. Both of these minerals occur at Franklin and Mine Hill, New Jersey.

Metallurgy.—Zinc ores are first roasted to remove carbon dioxide and water in the case of smithsonite or calamine, and in the blende to convert the sulphide into oxide. The roasted ore is then reduced by heating to redness in fireclay retorts with charcoal; the zinc, which boils at about 920° , distils over.

Modern methods consist in making an intimate mixture of the pulverized roasted ore with powdered coal; this is introduced into long, cylindrical clay retorts placed horizontally in rows slightly tipped toward the front. These are heated to 1200° to 1300° by means of natural or process gas; as soon as

the CO gas becomes luminous, long receivers (alembics) are attached to the ends of the retorts for condensation of the metal. Zinc powder or dust appears first, followed by the molten metal. This "zinc dust" is largely employed as a reducing agent in organic synthesis.

Commercial zinc often contains as impurities lead, cadmium, copper, silver, carbon, iron, arsenic, antimony, and sulphur. It, therefore, requires further purification. This is accomplished by melting in a small reverberatory furnace with an inclined hearth at a temperature somewhat above its melting-point; the Pb and Fe collect on the bottom, while the other impurities collect on the surface. Zinc purified in this manner is usually remelted and cast into rods, rolled into thin sheets, or granulated by pouring into cold water, in which form it is known as "mossy" zinc. In the United States it is customary to prepare pure (arsenic free) zinc from such ores as are free from foreign metals. Zinc is marketed in plates or ingots. Arsenic may be removed from impure zinc by adding 0.2 per cent. of sodium to the fused metal, whereby a sodium zinc arsenate rises to the surface, which may be removed by skimming.

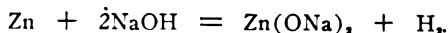
Electrolytic zinc of high purity is also produced.

Properties.—Zinc is a bluish-white metal, with a luster which is almost insignificant when compared with that of some other metals. When cast it exhibits a tabular structure. Its specific gravity ranges from 6.9, when it is cast, to 7.2, when it is rolled. It is very ductile, considering its hardness. Its malleability is considerable when pure, but in the impure condition it is quite brittle.

When heated to 100° and not above 150°, its malleability and ductility are materially increased; above 200° it becomes sufficiently brittle to be powdered in an iron mortar. Zinc melts at 418° to 419°, and at 918° it boils. Above this temperature zinc burns with a greenish-yellow flame, yielding a light white powder of zinc oxide. It is official under the name of **Zincum**, and this metal dissolves in dilute hydrochloric or sulphuric acid with evolution of hydrogen, and leaves a slight residue. "When 1 gm. of zinc is dissolved in 25 mls of diluted sulphuric acid the liberated hydrogen does not discolor a strip of paper moistened with lead acetate, T.S. (sulphur), or with silver nitrate, T.S. (As, Sb, P)." The purest metal, owing to the practical absence of foreign metals, more especially iron, reacts so slowly with diluted sulphuric acid that it is useless for the generation of hydrogen. An active reaction with the evolution of a steady current of this gas may be brought about by the addition of a little platinic chloride or copper sulphate or stannous chloride solution, which establishes galvanic action.

Zinc is not readily affected by dry air; in damp air it becomes coated with a thin film of oxide and carbonate, which prevents further action. For this reason zinc is used as a protective coating on iron (*galvanized iron*).

Sodium or potassium hydroxide dissolves zinc with evolution of hydrogen and formation of sodium and potassium zincate:



Zinc, such as is used as a reagent in arsenic testing, should be in mossy or globular (shot) form, the latter being known as No. 7 shot zinc. The "mossy" zinc is made by dropping the molten metal into warm water. Very

pure arsenic-free zinc will dissolve very slowly in diluted sulphuric acid. Rapid solution (gas evolution) indicates presence of much iron or added copper. The presence of much over 0.08 per cent. Fe as well as platinic chloride tends to retain some of the arsenic in the metal. Diluted sulphuric acid is used in the U.S.P. test for arsenic because it can be readily obtained free from this impurity, while the purest reagent hydrochloric acid (when kept in glass containers) always responds to the tests for arsenic.

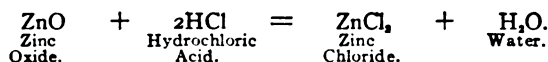
Uses.—Zinc is used in the manufacture of brass and galvanized iron in the form of wire, nails, and sheets. It is also employed in desilvering lead and in battery construction.

ZINC AND THE HALOGENS

Zinc Chloride, ZnCl_2 . *Zinci Chloridum*, U.S.P., which contains not less than 95 per cent. of ZnCl_2 , is obtained by evaporating a solution of zinc chloride.

Liquor Zinci Chloridi, U.S.P., is prepared by dissolving the metal in hydrochloric acid to saturation; a small quantity of nitric acid is then added to oxidize the iron present as impurity, and the solution evaporated to a dry mass at a temperature not exceeding 115° . The residue is dissolved in water and allowed to stand in contact with zinc carbonate for 24 hours (to remove Fe as hydrated oxide); the clear, supernatant liquid is then drained off and brought to a strength of 50 per cent.

Since zinc oxide is to be had in commerce of great purity and at a reasonable price, this solution can be economically prepared by saturating the acid with oxide as follows:



The dilute solution of zinc chloride may be filtered through paper, but when more concentrated it should be filtered through asbestos, or allowed to stand until clear and then decanted.

In order to obtain the dry salt the solution is evaporated in a porcelain dish on a sand-bath with constant stirring until brought to dryness. Some oxychloride (ZnOH.Cl) is formed during the evaporation, or larger quantities if overheated, which may be avoided by the cautious addition of hydrochloric acid from time to time. The presence of the basic salt (oxychloride) may be detected by mixing equal volumes of an aqueous solution of the salt (1 in 20) with alcohol, when a lesser or greater degree of turbidity results, according to the quantity present (see U.S.P. limit).

Properties.—Zinc chloride is found in commerce as a white, granular powder, or in irregular, porcelain-like masses, or it may be obtained moulded into sticks. It is odorless, intensely caustic, and, very deliquescent; soluble in 0.25 part of water and about 1.3 parts of alcohol at 25° . When heated to 115° the salt fuses to a clear liquid, and at a higher temperature it is volatilized with some decomposition, leaving a residue of oxychloride and oxide. Zinc chloride distils at about 600° .

The official salt forms a clear solution with water, from which on prolonged boiling a basic salt is deposited. The syrupy solution, on the addition of a few drops of hydrochloric acid, deposits octahedral crystals, containing one molecule of water of crystallization. The official solution has a specific gravity of 1.548 at 25°, and reacts acid with litmus paper.

The concentrated, aqueous solution dissolves paper and cotton, while the oxychloride dissolves wool and silk, hence the former cannot be filtered through paper. A concentrated solution of zinc chloride dissolves zinc oxide, forming zinc oxychloride ($\text{Zn}(\text{OH})\text{Cl}$), which, if sufficient oxide has been added, hardens quickly. This forms the base of dental enamel filling.

Uses.—Zinc chloride in solution is largely used as a caustic and, when diluted, as an antiseptic and disinfectant. It enters into the composition of some embalming fluids, and constitutes Burnett's disinfecting liquid. A solution of 100 parts of zinc chloride in 12 parts of water, saturated with 6 parts of potassium iodide, and then as much iodine added as the solution will dissolve, furnishes a valuable microscopic reagent for cellulose, with which it gives a blue color. A commercial solution of calamine, zinc blende, or the impure metal in hydrochloric acid, is used to impregnate wood to effect its preservation.

Zinc Bromide, ZnBr_2 .—This is prepared by interaction between zinc and bromine in the presence of water as detailed under zinc iodide.

Properties.—Zinc bromide occurs as a white, granular powder, odorless, and having a sharp, saline, and metallic taste. It is a very deliquescent salt, readily soluble in water and alcohol.

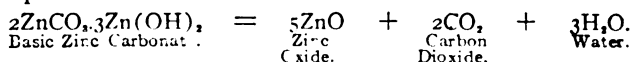
Zinc Iodide, ZnI_2 .—Prepared by a method similar to that employed in making ferrous iodide, in which powdered iodine is added to an excess of granulated zinc in the presence of water. The mixture is gently warmed until the color of iodine has disappeared and the solution has become quite concentrated, the clear solution of zinc bromide is drained off and evaporated at a low temperature and finally over sulphuric acid.

Properties.—Zinc iodide occurs as a white, granular powder, odorless, and having a sharp, saline, and metallic taste. It is very deliquescent, and, when exposed to light and air, it becomes brown, owing to gradual decomposition into iodine and zinc oxide. Zinc iodide is very soluble in water, alcohol and ether.

ZINC AND OXYGEN

Zinc Oxide, ZnO . Zinci Oxidum, U.S.P.—Zincite, or red oxide of zinc, associated with oxides of manganese and iron, occurs with other zinc minerals at Mine Hill, New Jersey.

Zinc oxide is prepared on a large scale by distilling the metal in earthenware retorts, through which a current of air is drawn, and conducting the vapors into large chambers. The product is the commercial zinc white; cadmium must be absent, since it dissolves the oxide. For pharmaceutical purposes a heavy oxide is prepared by heating the basic carbonate to low redness, when the following reaction takes place:



OXYGEN SALTS OF ZINC.

The oxide was formerly known as *Lana Philosophi*, *Alba*, and *Nix Alba*. Commercially it is known as *zinc oxide*.

ca, Nihilum Al-
white or Chinese

Properties.—Zinc oxide is a white, amorphous powder, without odor or taste. It is insoluble in water or alcohol, but is soluble, without water and effervescence, in dilute acids; it is also soluble in ammonia, but not in solution of ammonium carbonate. It turns yellow when heated, but gradually recovers its whiteness on cooling. On exposure to the air it gradually but slowly absorbs carbon dioxide.

Uses.—Zinc oxide, when ground with oil, is extensively used as a paint. It has the advantage of not darkening in the presence of sulphur compounds, and the disadvantage of possessing less power than white lead when used as paint. Lithopone, zinc sulphide and barium sulphate, is generally preferred because it possesses the advantages and none of the disadvantages of zinc oxide. Zinc oxide is used in making glass for purposes such as thermometers, since the glass has a high resistance to compression, and chemical action. It is one of the most important and satisfactory compounds from which to prepare the salts of zinc, and is considerably used in medicine.

Zinc Hydroxide, Zn(OH)_2 , is obtained as a white precipitate when potassium hydroxide is added to a solution of zinc sulphate soluble in excess of the reagent as well as ammonia, forming zincate.

When the above zincate solution is boiled the hydroxide is precipitated. Hydroxide deposits itself as a weak acid in the presence of a zinc salt, anions and H^+ cations.

On immersing a strip of zinc in ammonia water the hydroxide is slowly deposited in the form of rhombic prisms. Ammonia water does not precipitate zinc hydroxide from solutions containing ammonium salts, owing to the formation of double salts like those of magnesium.

Zinc Dioxide or Peroxide, ZnO_2 .—A mixture of zinc dioxide and zinc oxide or hydroxide containing not less than 45 per cent. of the dioxide, which may be obtained by interaction between barium dioxide and zinc sulphate, and is known under the commercial name of *Zinc Peroxide*. It contains from 7.4 to 10 per cent. of available oxygen. Zinc peroxide is a yellowish-white, voluminous, inodorous powder, insoluble in water, but on longer contact is gradually decomposed into oxygen and zinc hydroxide. It dissolves in dilute acids with formation of hydrogen peroxide.

OXYGEN SALTS OF ZINC

Zinc Sulphate, $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$. **Zinci Sulphas**, U.S.P.—Under the name of *white vitriol* this salt was known in the fifteenth century. It was prepared by lixiviation of zinc blende which has been roasted at low red heat ($\text{ZnS} + 2\text{O}_2 = \text{ZnSO}_4$), a process which is used on a large scale at the present time.

The sulphate is also made by dissolving the metal or zinc white in dilute sulphuric acid, and purifying by recrystallization.

Properties.—Zinc sulphate, containing 7 molecules of water of crystallization, occurs in colorless, transparent, rhombic crystals, without odor, and with an astringent, metallic taste. The crystals effloresce in

dry air. The salt is soluble in 0.6 part of water at 25° and in 0.25 part of boiling water, also soluble in 2.5 parts of glycerin.

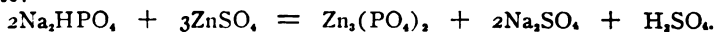
Zinc sulphate is isomorphous with and closely resembles magnesium sulphate, differing physically, however, in that it effloresces when exposed.

It melts in its water of crystallization when rapidly heated, and may be partly decomposed, losing H_2O and H_2SO_4 at higher temperatures. On applying heat gradually, 5 molecules of water are driven off at 50° (31.3 per cent.) without melting. At 100° the sixth is lost, and at 110°, in a current of dry air, the salt becomes anhydrous. It loses molecules of water in a vacuum. Monoclinic crystals, containing only molecules of water, may be obtained by having crystallization take place in concentrated solutions and above 40°. A salt having the composition $\text{ZnSO}_4 \cdot \text{H}_2\text{O}$ may also be obtained from the hot, concentrated solution. Like magnesium sulphate, one H_2O may be replaced by other sulphates, forming such double salts as $\text{ZnSO}_4 \cdot \text{K}_2\text{SO}_4 \cdot 6\text{H}_2\text{O}$. An aqueous solution of zinc sulphate has an acid reaction. Basic salts may be obtained by boiling concentrated solutions of zinc sulphate with metallic zinc or with zinc oxide.

Uses.—Zinc sulphate is used in medicine; in very small doses as a tonic and astringent; in larger doses as an emetic. It has the properties, common to all the other zinc salts, of acting as an irritant poison in large doses. This salt is also used as a wood and hide preservative, and for rendering fabrics fire-proof, and also as a disinfectant.

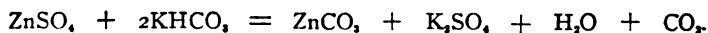
Zinc Nitrate. $\text{Zn}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}$, forms deliquescent crystals when 1 part of zinc oxide is dissolved in 6.5 parts of nitric acid of 25 per cent. strength. The process by which the metal is dissolved in nitric acid is not employed, since a part of the acid is decomposed and lost in the reaction. The crystals of zinc nitrate form colorless, striated, four-sided prisms, very soluble in water and in alcohol. When heated, the salt loses nitric acid and becomes converted into a basic nitrate.

Zinc Phosphate. $\text{Zn}_3(\text{PO}_4)_2 \cdot 4\text{H}_2\text{O}$, is obtained as a white, nearly insoluble precipitate, by the addition of sodium phosphate solution to a solution of zinc sulphate:



This precipitate is gelatinous when first formed, but becomes crystalline on standing.

Zinc Carbonate. **Zinci Carbonas Præcipitatus**, U.S.P.—This compound is officially defined as “an impalpable powder, of somewhat variable chemical composition, corresponding to not less than 68 per cent. of ZnO .” The true carbonate, ZnCO_3 , occurs native as zinc-spar, in hexagonal prisms, and massive as smithsonite, hardly to be distinguished from calamine. It may be prepared by precipitating a solution of zinc sulphate in the cold with a solution of acid potassium carbonate in excess:



The precipitate, when ammonium carbonate is used, is soluble in excess of the precipitant.

The official basic carbonate is not of constant composition, but is generally expressed as consisting of 2 molecules of normal zinc carbonate and 3 molecules of zinc hydroxide, $2\text{ZnCO}_3 \cdot 3\text{Zn}(\text{OH})_2$.

To prepare it, 1 part of sodium carbonate is dissolved in 10 parts of water and heated to the boiling point; to it is then added, in a thin stream so as not to interfere with the boiling, a solution of 1 part of zinc sulphate in 10 parts of water. The proportions should be such that, after all the zinc solution has been added, the mixture has a slight alkaline reaction. The following expresses the reaction which occurs:

$$5\text{ZnSO}_4 \cdot 7\text{H}_2\text{O} + 5\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O} = 2\text{ZnCO}_3 \cdot 3\text{Zn}(\text{OH})_2 + 5\text{Na}_2\text{SO}_4 + 3\text{CO}_2 + 82\text{H}_2\text{O}.$$

The precipitate is first washed with hot water by decantation, then collected on a filter and further washed with hot water until the washings cease to react with barium chloride.

Properties.—Basic zinc carbonate is an impalpable, amorphous powder, permanent in the air, and practically insoluble in water or alcohol; dilute acids, ammonia water, and solution of ammonium carbonate dissolve it freely. It is also soluble in water containing carbon dioxide. At about 300° the salt loses water and carbon dioxide, and is converted into zinc oxide, which is yellow while hot, but becomes white when cold.

Zinc Sulphide, ZnS .—Zinc blende is found native in crystals, belonging to the regular system. The color varies from a light brown to a deep black, depending on the amount of iron present.

When hydrogen sulphide is passed into a neutral solution of a zinc salt, the sulphide is formed as a white, amorphous precipitate. It is used as a pigment (lithopone) mixed with barium sulphate, and is made by mixing barium sulphide with a solution of zinc sulphate.

Zinc sulphide is soluble in the dilute mineral acids, but insoluble in acetic acid.

Zinc Permanganate, $\text{Zn}(\text{MnO}_4)_2 \cdot 6\text{H}_2\text{O}$.—Prepared like other permanganates. It occurs in dark brown, lustrous, deliquescent crystals, soluble in 3 parts of water. This salt gives up oxygen more readily than the potassium salt. Used in medicine like potassium permanganate, but is more astringent.

Zinc Phosphide, Zn_3P_2 , is obtained as a gray mass by heating powdered zinc in the vapor of phosphorus. It should be preserved in small, glass-stoppered vials. Zinc phosphide is a gritty powder of a dark-gray color, or crystalline fragments of a dark, metallic luster, and having a faint odor and taste of phosphorus. In contact with the air it slowly emits phosphorous vapor.

Uses.—Zinc phosphide is employed in medicine. Its physiological properties so closely resemble those of phosphorus that it is used as a convenient method of administering that element.

CADMIUM

Symbol, Cd.

Atomic Weight, 112.4.

Valence, II.

History.—Cadmium was discovered in 1817 by Stromeyer of Göttingen. Hermann of Schönebeck discovered it independently and at about the same time.

Occurrence.—Cadmium is found in small quantity in a number of zinc ores, and chiefly as sulphide. The fibrous zinc blende found at Przibram, Bohemia, contains from 2 to 3 per cent. of cadmium. The rare mineral greenockite consists almost entirely of cadmium sulphide. It is found at Bishoptown, in Scotland, near Friedensville, in Pennsylvania, and at Granby

and Joplin, Missouri. The presence of a brown flame at the orifice of the flame appears in the distillation of zinc.

Metallurgy.—In the preparation of zinc the latter metal distils first, and by saving coal and distilling it at the lowest possible temperature, a distillate is obtained very rich in cadmium, while the zinc remains behind, the boiling point of cadmium being 766° , while that of zinc is 918° . The cadmium is finally separated from the zinc by dissolving in hydrochloric acid, having the latter in slight excess, and passing in hydrogen sulphide; the cadmium sulphide is precipitated, while the zinc remains in solution in the presence of acid. The cadmium sulphide is dissolved in hot concentrated hydrochloric acid and the metal precipitated as carbonate by an excess of a solution of ammonium carbonate, by which means the copper and arsenic are retained in solution. The cadmium carbonate is washed, dried, and converted by heat into cadmium oxide, which, on distillation with coal, yields the metal.

Properties.—Cadmium is usually found in commerce in sticks. It is a white, lustrous metal, with a fibrous fracture. When pure, it is malleable and ductile. Its specific gravity is 8.6. It melts at 321° , and boils at 766° . When heated in the air it burns, giving off a brown vapor. Most of the acids dissolve it readily. Its salts are colorless. In many of their properties the salts of cadmium are identical with those of zinc.

COMPOUNDS OF CADMIUM

Cadmium Chloride, $\text{CdCl}_2 \cdot 2\text{H}_2\text{O}$, **Cadmium Bromide**, $\text{CdBr}_2 \cdot 2\text{H}_2\text{O}$, and **Cadmium Iodide**, CdI_2 , are colorless, soluble salts, which are prepared similarly to the corresponding zinc salts. The chloride is made by dissolving the metal or oxide in hydrochloric acid, while the bromide and iodide are prepared by digesting the metal with bromine and iodine respectively in the presence of water. These salts are readily soluble in alcohol and water, but not deliquescent. Cadmium potassium iodide (U. S. P. reagent), made by mixing equal weights of cadmium and potassium iodides, is a useful reagent for alkaloids. Cadmium iodide has some use in photography. The chloride and iodide are used as U. S. P. reagents.

Cadmium Oxide, CdO , forms as a brown, infusible, amorphous powder, when cadmium is burned in the air, or by ignition of the carbonate or nitrate.

Cadmium Hydroxide, $\text{Cd}(\text{OH})_2$, is prepared by precipitation of a soluble cadmium salt, by potassium or sodium hydroxide. It forms, when collected and dried, a white, amorphous powder, which at 300° loses water and is converted into oxide.

Cadmium Sulphate, CdSO_4 .—This salt was official in the Pharmacopœia of 1870, but since that time it and the other salts of cadmium have practically ceased to have any use in medicine. To prepare this salt, 10 parts of cadmium and 40 pieces are placed in a porcelain dish with 10 parts of pure sulphuric acid and 40 parts of water. The mixture is warmed until the solution is effected, and the solution is filtered and evaporated to crystallize. On account of the slowness with which sulphuric acid acts on cadmium, the metal is sometimes first dissolved in nitric acid, and the hydroxide is precipitated from the resulting nitrate with sodium hydroxide, and is then dissolved in diluted sulphuric acid.

Properties.—Cadmium sulphate crystallizes in large, colorless, monoclinic crystals, having the composition $3\text{CdSO}_4 \cdot 8\text{H}_2\text{O}$. It is permanent in the air, easily soluble in water, and the aqueous solution has an acid reaction. The sulphates of the alkalis and of the alkaline earths yield with cadmium sulphate well-crystallized double salts similar to those of zinc and magnesium. It has been used as an astringent in diseases of the eye.

Cadmium Nitrate, $\text{Cd}(\text{NO}_3)_2 \cdot 4\text{H}_2\text{O}$, is obtained by dissolving the metal in nitric acid. It crystallizes in fibrous needles, which are very deliquescent and are, therefore, very soluble in water; soluble also in alcohol.

Cadmium Carbonate.—A basic salt is obtained by precipitating a solution of cadmium sulphate with sodium carbonate. It has a variable composition, and in other respects resembles the corresponding zinc salt.

Cadmium Sulphide, CdS , is obtained as a yellow powder, insoluble in hydrochloric acid. It is used as a pigment under the names of *King's yellow* and *jaune brilliant*.