ESSENTIALS OF CHEMISTRY

AND

URINALYSIS.

WOLFF.

THE ESSENTIALS

OF

MEDICAL CHEMISTRY

AND

URINALYSIS.

BY

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PREFACE.

The rapid exhaustion of the first edition has encouraged the author and publishers to present this, the second edition, revised, enlarged, and illustrated. As the name implies, the aim of the work is to present the *essential* facts of a course of lectures on Medical Chemistry and Urinalysis, so that the student need not wade through the more exhaustive text-books. The author claims no originality for the facts, but has appropriated any thing of service wherever found. The selection of material and the plan of presentation is the outgrowth of his experience as a general practitioner and as a teacher of Medical Chemistry for the past ten years. The subjects treated are so numerous that the descriptions are as brief as is consistent with clearness, but the principles of the science and the application of the facts to medicine have been stated more fully. The writer hopes the book may prove useful to medical students, for whom it was designed, and that his labor may lessen theirs.

1506 W. Walnut Street, Louisville.

THE ESSENTIALS OF

MEDICAL CHEMISTRY.

INTRODUCTION.

"Chemistry is that branch of science which treats of the composition of substances, their changes in composition, and the laws governing such changes." (Webster.)

The distinctive characteristic of chemical action is change in *composition.* A bar of soft iron is the same in composition whether it be hot or cold, luminous or non-luminous, magnetized or unmagnetized. But when it undergoes chemical action a substance is formed which, though it contains iron, is entirely different from it in composition and properties.*

Matter is that of which the sensible universe is composed. It is *indestructible.* Although we may cause matter to assume a variety of forms, and render it under certain conditions invisible, yet in none of these changes is the smallest particle destroyed.†

When carbon is burned in an atmosphere of oxygen it disappears, and, so far as we can learn by the senses of sight and touch, it is lost; but the result of this burning is an invisible gas, whose weight is equal to that of the carbon which has disappeared, plus the weight of the oxygen required to burn it.‡

Fig. 1.

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* **Experiment.** Heat pieces of platinum and magnesium wire. Note that while the platinum is unaltered, the magnesium burns and is converted into a white powder of magnesium oxide.

† **Experiment.** Weigh a small porcelain crucible containing powdered iron. Heat it, and when combustion is complete, weigh again and note the increase in weight, and also that a new substance is formed, which, though it contains iron, is not iron.

‡ **Experiment.** Burn charcoal in a jar of oxygen or air. (Fig. 1.) Add lime-water and agitate. Note the white precipitate of carbonate of lime.

All matter has WEIGHT. *Apparent weight* is that which we determine by balances in the open air. The *absolute weight* of a body is its weight *in vacuo*.

By the *specific weight* or *specific gravity* of a substance we mean its weight as compared to that of some other substance taken as a standard. The standard for solids and liquids is water; for gases, air or hydrogen.

Matter exists in one of three states, solid, liquid, or gaseous. In the solid state the particles are held together so rigidly as to give the body a definite shape; while in the liquid state the attraction is so slight as to allow the particles to move freely upon each other and the body to take the shape of the vessel that contains it. In the gaseous state the mutual attraction of the particles is entirely overcome, and their distance from each other depends upon the pressure to which the gas is subjected. The term fluid is applied to any thing capable of flowing, whether liquid or gaseous. It is highly probable that all substances, which are not decomposed by heat or cold, are capable of existing in all three states. Heat is absorbed and cold produced wherever the particles are to be driven farther apart, as in the passage of a substance from the solid to the liquid or from the liquid to the gaseous state.

When the two solids ice and common salt are mixed, they form a liquid, and great cold is produced.* Perspiration in evaporating assumes the gaseous state, and absorbs in the change enough heat to keep the body at its normal temperature in spite of the hottest weather.†

On the other hand, when a substance passes from a rarer to a denser state it gives out again the heat absorbed in its passage in the opposite direction.

If we examine the infinite variety of substances upon our earth we find most of them are *compounds*, *i. e.* they can be decomposed into two or more other substances, distinct in their properties from the substance from which they were derived and

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*__Experiment.__ Fold tin foil into the shape of a little dish; add powdered ice and salt. Spill water on the table and set the dish in it. Note how quickly it is frozen fast to the table.

†__Experiment.__ Put a little water in a similar dish. Against the sides and bottom throw a spray of ether. Note that the evaporation of the ether is so rapid that the water is quickly frozen.

from each other. There are some substances which have never been decomposed. These are called *elements.* Only seventy elements are at present known; but, as our methods of investigation improve, this number may be increased by the discovery of other elements, or decreased by decomposing some of those now considered elements. Only about one half of these enter into the materia medica, and will be noticed in this work.

TABLE OF ELEMENTARY BODIES, WITH THEIR SYMBOLS AND ATOMIC WEIGHTS.

Name.	Symbol	Atomic Weight.	Name.	Symbol	Atomic Weight.
ALUMINIUM	Al	27	NICKEL	Ni	58.8
ANTIMONY (Stibium)	Sb	122	Niobium	Nb	94
ARSENIC	As	75	NITROGEN	N	14
BARIUM	Ba	137	Norwegium	Ng	214
Beryllium	Be	9.4	Osmium	Os	200
BISMUTH	Bi	210	OXYGEN	O	16
BORON	B	11	Palladium	Pd	106.6
BROMINE	Br	80	PHOSPHORUS	P	31
CADMIUM	Cd	112	PLATINUM	Pt	197.4
Cæsium	Cs	133	POTASSIUM (Kalium)	K	39.1
CALCIUM	Ca	40	Rhodium	Rh	104.4
CARBON	C	12	Rubidium	Rb	85.4
CERIUM	Ce	141	Ruthenium	Ru	104.4
CHLORINE	Cl	35.5	Samarium	Sm	150
CHROMIUM	Cr	52.2	Scandium	Sc	44
COBALT	Co	58.8	Selenium	Se	79.4
COPPER (Cuprum)	Cu	63.4	SILICON	Si	28
Didymium	D	145	SILVER (Argentum)	Ag	108
Erbium	E	168	SODIUM (Natrium)	Na	23
FLUORINE	F	19	STRONTIUM	Sr	87.6
Gallium	Ga	70	SULPHUR	S	32
Germanium	Ge	163	Tantalum	Ta	182
GOLD (Aurum)	Au	197	Tellurium	Te	128
HYDROGEN	H	1	Thallium	Tl	204
Indium	In	113.4	Thorinum	Th	235
IODINE	I	127	TIN (Stannum)	Sn	118
Iridium	Ir	198	Titanium	Ti	50
IRON (Ferrum)	Fe	56	Tungsten, or Wolfram	W	184
Lanthanum	La	139	Uralium	Ul	187
LEAD (Plumbum)	Pb	207	Uranium	U	240
LITHIUM	Li	7	Vanadium	V	51.2
MAGNESIUM	Mg	24	Ytterbium	Yb	173
MANGANESE	Mn	55	Yttrium	Y	92
MERCURY (Hydrargyrum)	Hg	200	ZINC	Zn	65.2
Molybdenum	Mo	96	Zirconium	Zr	89.6

To explain the laws governing chemical phenomena we adopt the old *atomic theory.*[*]

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[*] Democritus, 460 B. C., said: "The atoms are invisible by reason of their smallness—indivisible by reason of their solidity; impenetrable and unalterable."

We will take up the theories and laws, not in the order of their enunciation, but of their natural sequence.

It is assumed that matter is composed ultimately of infinitely small particles called *atoms;* that each element is composed of *atoms*, all of a certain size, weight, etc. Atoms do not exist alone, but in groups called *molecules.* In an element the molecule is composed of pairs of atoms of the same kind; in compounds they consist of two or more atoms of different kinds. It has been determined that equal volumes of all substances in the gaseous state, and under like conditions, contain the same number of molecules. So a gallon of hydrogen gas and one of oxygen gas containing the same number of molecules, and those molecules consisting of pairs of atoms, must contain the same number of atoms. Furthermore, it is found that the gallon of oxygen is sixteen times as heavy as the gallon of hydrogen. So each oxygen atom must also be sixteen times as heavy as the hydrogen atom. Hydrogen being the lightest substance known, its *atomic weight* is taken as 1, and consequently the atomic weight of oxygen is 16. The atomic weights of other elements are determined in a similar way. By "atomic weight" is not meant the absolute weight of atoms (for that is not known), but the weight of the atom as compared to the hydrogen atom. The atomic weight of carbon is 12. If carbon combines with oxygen, atom for atom, the new substance (CO) resulting from that action will consist of molecules, in each of which the carbon will weigh 12 and the oxygen 16, and, as the whole mass is composed of these molecules, the same proportion obtains throughout the new compound. So 12 is found to be the *combining weight* of carbon, and 16 of oxygen. If, however, the combination should occur in the proportion of one atom of carbon to two atoms of oxygen, then each molecule must consist of 12 by weight of carbon to 32 of oxygen, and that must be the proportion throughout the entire substance.

Between these two compounds no intermediate one can occur, for the carbon atom must take one or two, or more, oxygen atoms. It can not take a fraction of one, for atoms are indivisible. Hence, we deduce the following Law: *Substances combine in certain fixed proportions* (atomic weights), *or in multiples of these proportions.*

Symbols are abbreviations of the names of the elements. They consist of the initial letter of the Latin name; but if the names of several elements begin with the same letter, the single-letter symbol is reserved for the most common element, and for the others another letter is added. Thus, we have nine elements whose names begin with C; the most common is carbon, whose symbol is C; the others add other letters, as chlorine, Cl; cobalt, Co; copper (cuprum), Cu, etc. The symbol indicates just one atom. When more than one atom is to be represented, the number is written just after and below the symbol, thus, C_4.

Formulæ are to molecules what symbols are to elements. They indicate the kind and number of atoms composing the molecule. When more than one molecule is to be indicated, the number is placed in front of the formula, thus, $5H_2O$. A parenthesis inclosing several symbols or formulæ should be treated as a single symbol, thus, $2(NH_4)_2CO_3=N_4H_{16}C_2O_6$.

An *equation* is a combination of formulæ and algebraic signs to indicate a chemical reaction and its results. As no matter is ever lost or created in a reaction, the number of each kind of atoms before the equality sign must be the same as after it.

INORGANIC CHEMISTRY.

CLASSIFICATION OF THE ELEMENTS. The elements are usually divided into two great classes: (*a*) *Metals*, about fifty-five in number, possessing a peculiar luster, good conductors of heat and electricity, and whose oxides when combined with water form bases; (*b*) *Non-metals*, about fifteen in number, possessing little luster, poor conductors of heat and electricity, and whose oxides combine with water to form acids. A better classification, and the one we shall adopt, is the following, based upon chemical properties:

THE NON-METALLIC ELEMENTS.

I. The **Hydrogen and Oxygen Group.**

II. The **Chlorine Group**: Fluorine, Chlorine, Bromine, Iodine.

III. The **Sulphur Group**: (Oxygen) Sulphur, Selenium, Tellurium.

IV. The **Nitrogen Group**: Nitrogen, Phosphorus, Arsenic, Antimony, Bismuth.

V. Boron.

VI. The **Carbon Group**: Carbon, Silicon.

THE METALS.

I. The **Alkaline Group**: Lithium, Ammonium, Sodium, Potassium, Rubidium, and Cæsium.

II. The **Alkaline Earths Group**: Magnesium, Calcium, Strontium, Barium.

III. The **Earths Group**: Aluminium, Lanthanum, Cerium, Didymium.

IV. The **Zinc Group**: Zinc, Cadmium.

V. The **Iron Group**: Chromium, Manganese, Iron, Cobalt, Nickel.

VI. The **Lead Group**: Tin, Lead.

VII. The **Copper Group**: Copper, Mercury.

VIII. The **Silver Group**: Silver, Gold, Platinum.

THE NON-METALS.

I. **Hydrogen and Oxygen Group.** In a strict arrangement hydrogen would be placed in Group I of the metals and oxygen in the sulphur group. But we will consider them in a group to themselves, because (*a*) of all the elements hydrogen is taken as the standard for atomic weights, combining weights, valence, etc.; (*b*) oxygen plays a most important rôle in chemistry, and its deportment with the other elements forms the basis of our classification; (*c*) the chemistry of these two will best serve as an introduction to the study of the other elements.

Hydrogen (H — 1) *occurs* free in volcanoes, gas-wells, etc.; combined in water and all organized bodies. All acids are salts of hydrogen. *Prepared* in various ways from its compounds, the most convenient being to displace it from sulphuric acid by zinc, thus,

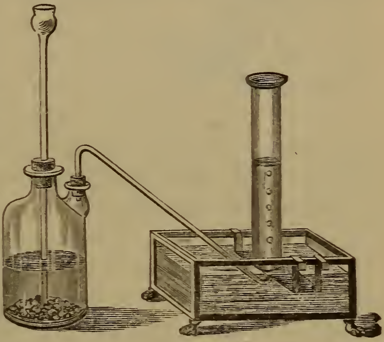

Fig. 2.

$$H_2SO_4 + Zn = ZnSO_4 + H_2. \quad \text{(Fig. 2.)}$$

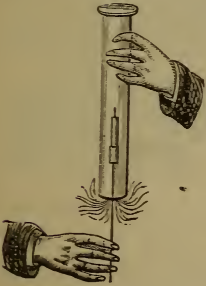

Fig. 3.

Physical Properties. Transparent, colorless, odorless, tasteless gas; the lightest substance known, fourteen and a half times as light as air; hence used in balloons. Long suspected to be a metal, because it displaces metals in chemical compounds, forms alloys with certain metals, and conducts electricity. This was proved in 1877, when Pictet condensed it under great cold and enormous pressure into a bluish metallic liquid.

Chemical Properties. Hydrogen does not support ordinary combustion or animal respiration. It burns in air with a pale but

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Experiment. By means of a wire gauze spoon hold some sodium beneath the water and under a cylinder. The hydrogen gas liberated by the sodium from the water will rise in bubbles, fill the cylinder, and displace the water.

very hot flame.[*][†] With pure oxygen it forms the oxyhydrogen flame. This is the hottest flame known, and a stick of lime held in it glows with dazzling brilliancy, forming the calcium or Drummond light. Mixed with air or oxygen, it explodes violently on contact with a spark.[‡]

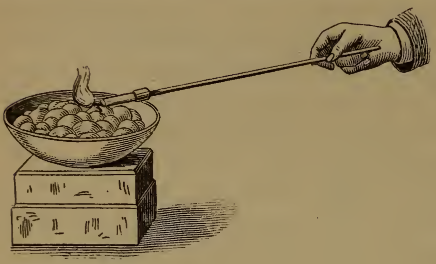

Fig. 4.

Oxygen (O—16). *Sources.* Most abundant of the elements, comprising one fifth of the air, eight ninths of water, one half of the crust of the earth, and three fourths of all organized bodies. *Prepared* most easily by heating potassium chlorate (Fig. 5):

$$KClO_3 = KCl + O_3.$$

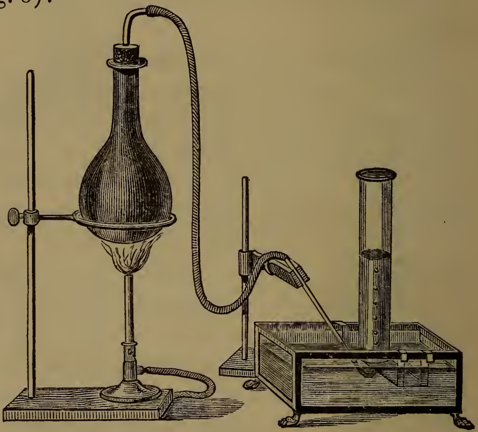

Fig. 5.

If manganese dioxide (MnO_2) be mixed with the chlorate, the gas is liberated more quietly and at a lower temperature. The manganese dioxide is unaltered in the reaction. It seems to act by its mere presence, an influence called *catalysis.*

Physical Properties. Gas; liquefied (Pictet, 1877) by great cold and intense pressure; colorless, odorless, tasteless; 1.10 times as heavy

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[*] **Experiment.** If an inverted jar of the gas is suddenly turned up, and a flame held a foot or two above, the gas escaping from the jar rises rapidly, and in coming in contact with the flame burns with a slight explosion.

[†] **Experiment.** If a jar of the gas be held mouth down and a candle be passed up into it (Fig. 3), the gas ignites and burns quietly at the open end, while the candle passed up into the gas is extinguished, but may be relighted again by the burning gas as it is withdrawn.

[‡] **Experiment.** Fill a bladder or rubber bag with two parts of hydrogen and one of oxygen or five of air. Attach a tube and blow up soap bubbles in a basin. (Fig. 4.) Touched with a flame, they explode.

as air. Water dissolves only three volumes to the hundred, but enough to sustain aquatic life.

Chemical Properties. Intense affinities; combines with every element except fluorine. The product of its action is called an *oxide*, and the process *oxidation*. Oxidation so rapid as to produce heat and light is called *combustion;* if no light, *slow combustion.* Substances that burn in air burn more brilliantly in oxygen,[*] and many substances that do not burn in air will burn in this gas.[†] By this property oxygen is usually recognized and distinguished from most other gases. Oxygen, especially in its diluted form (air), is the great supporter of combustion, for which its abundance and universal presence eminently fit it. Combustible and supporter of combustion are only relative terms. When a combustible substance burns in a supporter of combustion the union is mutual, one being as much a party to the action as the other. A jet of air[‡] or oxygen burns as readily in coal gas as a jet of coal gas burns in air or oxygen. The one in greatest abundance is usually called the supporter of combustion.


Fig. 6.


Fig. 7.


Fig. 8.

Oxidizing agents are compounds in which oxygen is held so feebly it is readily given up to substances having greater affinity for it.

Uses. The process of respiration is a species of combustion, and, as oxygen is the best supporter of combustion, it is the best

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[*] **Experiment.** A bit of phosphorus, dried by pressing between folds of blotting paper, is placed in a combustion spoon (Fig. 6), ignited, and lowered into a jar of oxygen. (Fig. 7.) The combustion is so intense that the phosphorus volatilizes, and its vapor burns throughout the jar with a brilliancy so dazzling that it is called the "phosphorus sun."

[†] **Experiment.** A watch-spring is wound into a spiral, tipped with a bit of tinder or a piece of yarn dipped in sulphur. This is lighted and lowered into a jar of oxygen. (Fig. 8.) The iron catches fire and burns with brilliant scintillations, globules of melted iron falling and melting into the glass, unless the bottom be covered with sand or water.

[‡] **Experiment.** Secure an ordinary lamp chimney (Fig. 9) and a wide cork

(and only) supporter of animal respiration. Administered in capillary bronchitis, œdema glottidis, etc., when the patient can not take in a volume of air sufficient to supply the requisite amount of oxygen.

Fig. 9.

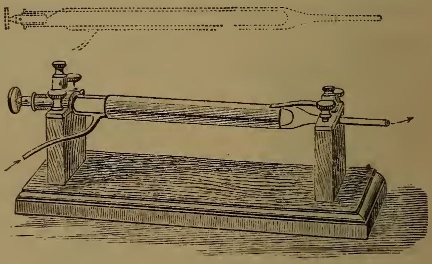

Fig. 10.

OZONE. If through a portion of air or oxygen electric sparks be passed,* a part of the oxygen will acquire a pungent odor and peculiar properties. The same change may be induced by various chemical processes, *e. g.* by mixing permanganate of potassium and sulphuric acid, or when phosphorus partially covered with water is exposed to the air. This modified oxygen is called *ozone*. It is one and a half times as heavy as ordinary oxygen, for its molecule contains three instead of two atoms. Very energetic, oxidizing

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to fit its lower end. Pass through the cork a narrow tube (*c*) connected by rubber hose with the house gas, and a wider one opening into the air. Turn on the coal gas and light it as it issues from the tube (*c*). The cork with the flame (not too large) is then inserted into the chimney, where it continues to burn, sufficient air entering through the wide tube (*a*). Upon turning on more gas the air is crowded out and the chimney filled with coal gas. The gas flame disappears from the tube (*c*), and an air flame appears upon the tube (*a*) as the entering air burns in the atmosphere of coal gas. The excess of coal gas may also be lighted as it escapes, showing a gas flame above and an air flame within the chimney. On lessening the flow of gas the air will again be in excess, and the flame again appear on the narrow tube (*c*.) In the gas flame above the lamp chimney (Fig. 9) heat some potassium chlorate in a combustion spoon until it melts and oxygen begins to bubble up. Then lower it into the atmosphere of coal gas within the chimney. The escaping oxygen burns brilliantly, the coal gas being the supporter of the combustion.

*Siemens's apparatus for ozoning oxygen (Fig. 10) consists of two tubes, the inner surface of the inner and the outer surface of the outer tube being coated with tin-foil, and each connected with the poles of an induction coil or Toepler-Holtz machine. A current of oxygen passing between these tubes may be ozonized to the extent of fifteen or twenty per cent.

substances unaffected by ordinary oxygen. Oxidizes potassium iodide with liberation of iodine, hence its *test:* paper dipped in a solution of potassium iodide and starch is colored blue in the presence of ozone.* Ozone is found in the air, especially after thunder-storms, and when present in considerable amount (as much as .00005 per cent) is apt to irritate the respiratory tract; but by oxidizing infecting germs, etc., it prevents the spread of infectious diseases.

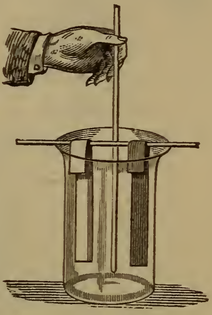

Fig. 11.

COMPOUNDS OF HYDROGEN WITH OXYGEN. Two are known—hydrogen oxide, or water, H_2O; hydrogen peroxide, or oxygenated water, H_2O_2.

WATER (H$_2$O) *occurs* widely distributed in nature; an important constituent of all organized tissues; forms seven eighths of the human body.

Physical Properties. Transparent, colorless, odorless, tasteless—liquid. Below 32° F. (0° C.) it is a solid (ice), and above 212° F. (100° C.) a vapor (steam or water gas). In solidifying, water expands; so ice floats. The boiling point is higher than 212° F. under increased pressure or when it contains solid matter in solution.

Water is the greatest of all solvents. The watery solution of a fixed substance is called a "*liquor*," and of a volatile substance an "*aqua*."

One body is said to dissolve in another when they coalesce and their particles intimately mingle. This is possible only in the liquid and gaseous states. When a substance dissolves it takes on the physical state of the solvent, *e. g.* a solid or gas dissolving in water becomes a liquid and then mixes with the water, the gas elevating the temperature and the solid lowering it. Heat assisting the liquefaction of a solid, and opposing that of a gas, hastens the solution of the one and retards that of the other. Most solid substances when separating from a solution take with

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***Experiment.** Pour a little ether into a beaker, across the top of which is a glass rod supporting a strip of blue litmus paper and one of paper dipped in potassium iodide and starch-water. Hold a hot glass rod in the jar (Fig. 11); the ether will undergo slow combustion, producing acid fumes which redden the litmus, and ozone which blues the other paper.

them, as a necessary part of the crystal, a certain definite amount of water—*water of crystallization.* This water does not modify the chemical nature of the substance, but is necessary for maintaining the crystalline form. If the crystal loses its water of crystallization by heat or exposure, it *effloresces* into an amorphous powder. Some substances when exposed absorb water from the air and *deliquesce* (melt down).

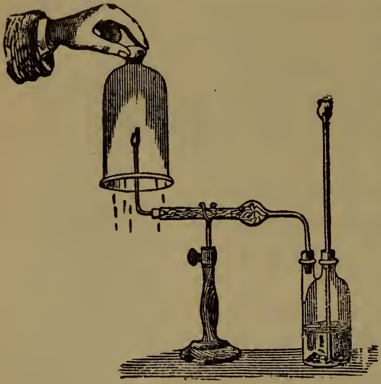

Fig. 12. Water from burning H.

Chemical Properties. The chemical composition of water may be proved by (*synthesis*) combining its constituents ($H_2+O=H_2O$) (Fig. 12), or by (*analysis*) passing the galvanic current through water until it is decomposed into its component gases ($H_2O=H_2+O$). (Fig. 13.) Neutral in reaction; combines with the oxides of the metals to form hydrates (bases), and with the oxides of the non-metals to form acids.

Natural waters are never pure. The nature of the impurities in water depends on the condition of the atmosphere through which it has fallen as rain, and the nature of the geological strata through or over which it has passed, for water dissolves something from almost every thing it touches. Good, potable (drinkable) water should be cool, clear, and odorless. It should contain just enough dissolved gases and solids to give it an agreeable taste, neither flat, salty, nor sweetish; and should dissolve soap without forming a curd. Water impregnated with inorganic matters, especially salts of calcium, is called hard. A much more serious contamination is with organic (animal and vegetable) matters. Such water is a prolific source of disease. It is probable, in fact almost

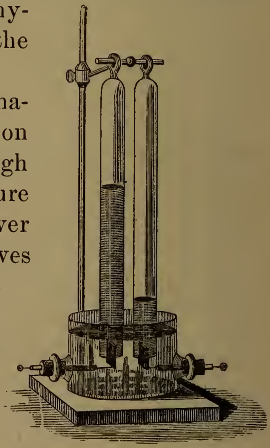

Fig. 13.

proven, that most infectious diseases are due to micro-organisms, many of which find the most favorable conditions for their life and growth in water contaminated with organic, especially animal, matter. Though chemical analysis can not detect the disease-producing elements, it can detect organic impurity, without which they can not exist. This is easily done thus: (1) Half fill a clean bottle with the water, warm, agitate, and critically smell it. A foul odor indicates organic impurity. (2) Fill a clean pint bottle three fourths full, add a teaspoonful of the purest white sugar; set aside in a warm place for two days, when, if it becomes cloudy, it is unfit to use. These rough-and-ready tests are those best suited to the practitioner, the more exact methods being practicable only to the chemist.

To *purify* natural waters, they may be boiled to kill living organisms, and filtered to remove suspended matters; but for chemical purposes, where great purity is desired, they are distilled* (*aqua destillata*, U. S. P.).

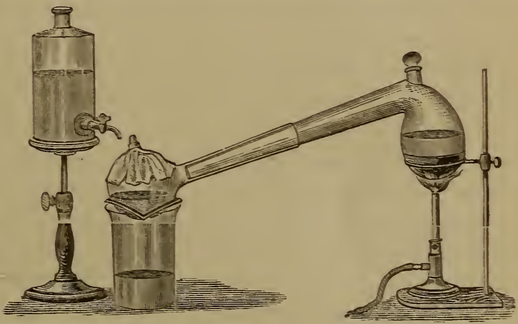

Fig. 14.

Mineral waters are those possessing special therapeutic value. They may be classed as follows:

1. *Carbonated*, those charged with carbonic acid.
2. *Sulphur*, containing H_2S or some soluble sulphide.
3. *Alkaline*, containing alkaline salts of potassium, sodium, or lithium.
4. *Saline*, containing neutral salts.
5. *Chalybeate*, containing iron.
6. *Thermal*, or hot waters.

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*When a liquid is rapidly vaporized, and the vapor, passing over into a colder vessel, is recondensed, the process is called *distillation*. (Fig. 14.) If a solid be similarly treated it is called *sublimation*. When water containing solid matter in solution is distilled, the solids remain in the vessel, while the water passes over, enabling us to obtain perfectly pure water. When a mixture of two or more liquids is heated, the one having the lowest boiling point distills first, leaving the others behind. This is called *fractional distillation*.

HYDROGEN DIOXIDE—*Oxygenated Water* (H_2O_2). *Prepared* most easily diluted by passing CO_2 through water holding barium dioxide in suspension.

$$BaO_2+CO_2+H_2O=BaCO_3+H_2O_2.$$

The $BaCO_3$ may be allowed to subside and the clear solution poured off.

Properties. When concentrated it is a colorless, syrupy liquid, with a pungent odor and taste—prone to decompose into H_2O+O.

Used to bleach* the hair and skin, converting brunettes into blondes; as a disinfectant to ulcers, ozœna, and in diphtheria, especially when the membrane has invaded the nose; also as a test for pus in urine, with which it causes an effervescence.

RADICALS. Every molecule is composed of two parts, called *radicals,* held together by chemical affinity. Both radicals may be elements, as in (H)—(Cl), or one may be elementary and the other compound, as (H)—(NO_3), or both compound, as (NH_4)—(NO_3). Some compound radicals can be isolated, *e. g.* by heat, (Hg)—(CN) $=Hg+CN$. Others decompose whenever set free. Whenever a galvanic current is passed through a compound, the chemical affinity is overcome by the electricity, and the molecule separates into its two radicals, one of which goes to the positive and the other to the negative pole. Unlike electrical conditions attract, so the radical going to the negative pole must be *electro-positive,* and the one going to the positive pole *electro-negative.* The metallic radicals are usually electro-positive and the non-metallic electro-negative.

Some radicals are more intensely electro-negative or electro-positive than others. In the following list the more common elements are so arranged that each is usually positive to those following it and negative to those preceding:

Positive end: + K, Na, Mg, Zn, Fe, Al, Pb, Sn, Bi, Cu, Ag, Hg, Pt, Au, H, Sb, As, C, P, S, N, I, Br, Cl, F, O—Negative end.

A radical is electro-positive or electro-negative only in its relation to other radicals; for, while C is positive to O, it is negative to K.

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* **Experiment.** Secure an old oil painting darkened with age, or take paper dipped in lead acetate and blackened by hydrogen sulphide: wash it with hydrogen dioxide, and the dark stain will be made white by the lead sulphide being oxidized into sulphate.

In formulæ the electro-positive radical is written first and the electro-negative next.

The greater the difference between the electrical condition of two radicals, the greater the energy with which they unite and the more stable the product, and, *vice versa; e. g.* O has a strong affinity for K, a weak one for Cl, and will not unite with F under any circumstances. An idea once prevailed that the relations of affinities were fixed and constant between the same substances, and great pains were taken to construct tables similar to the above to show what was called the "precedence of chemical affinities." These tables showed the order of affinities for the circumstances under which the experiments were made, and nothing else.

The circumstances attending chemical reactions are so complicated that in the greater number of cases it is impossible to predict the precedence of affinities and the result of an untried experiment.

Among these modifying causes may be mentioned:

1. *Temperature*, changes of which often reverse chemical affinities. Moderately heated, mercury and oxygen will readily combine, but when highly heated their affinity is destroyed, and they will refuse to unite, or, if already combined, will separate.

Ordinarily free carbon has no affinity for oxygen, but at high temperatures it surpasses all other elements in its greediness for that substance, even taking it from a metal so extremely electropositive as potassium.

2. *Volatility. Whenever in a mixture of two or more substances it is possible, by a rearrangement of the radicals, to form a compound volatile at the temperature of the experiment, such rearrangement will occur and the volatile compound be formed.* For example:

$$(Fe)\,(S)+(H_2)\,(SO_4)=(Fe)\,(SO_4)+(H_2)\,(S);\text{ or,}$$
$$2\,(NH_4)\,(Cl)+(Ca)\,(CO_3)=(NH_4)_2CO_3+(Ca)\,(Cl_2);\text{ or,}$$
$$H_3BO_3+3\,Na\,Cl=3\,H\,Cl+Na_3\,BO_3.$$

3. *Insolubility. Whenever, on mixing two or more substances in solution, it is possible, by rearrangement of the radicals, to form an*

Experiment. Into the vessel shown in Fig. 13 put some water; add solutions of red litmus, potassium iodide, and boiled starch; connect with the galvanic battery. The electric current decomposes the potassium iodide into *iodine*, which gathers at the positive pole, producing a blue color, with the starch, and *potassium* at the negative, where it produces alkali, turning the red litmus blue.

insoluble compound, that rearrangement will occur and the insoluble compound be formed as a precipitate. For example:

$$(Ca)(Cl_2)+(NH_4)_2(CO_3)=(Ca)(CO_3)+2(NH_4)(Cl).$$

It is especially important to remember this law, for its application in tests, incompatibilities, and antidotes.

4. *Nascent State.* Ordinarily the atoms of an element are grouped in pairs, and hence somewhat indifferent to the attractions of other atoms; but just as they are being liberated (born) from a compound they are alone. Each atom, having no fellow, readily enters into combination with any atom it meets. This state is called *nascent* (*nasci*, to be born).

5. *Catalysis.* This is the name given to the unexplained influence exerted by some substances of inducing chemical reactions between other substances without itself undergoing any change.

The valence of a radical is its combining value, or its value in exchange for other radicals.* Here again hydrogen is taken as the standard. A radical that combines with or takes the place of one atom of hydrogen is said to be *univalent* (one valued); of two atoms, *bivalent;* three, *trivalent;* four, *quadrivalent;* five, *quinquivalent;* six, *sexivalent.* The valence is indicated by a Roman numeral just above and after the radical, thus: $(NH_4{}^{I})$, Ca^{II}, $(PO_4)^{III}$, Si^{IV}, As^{V}, S^{VI}. The two radicals of every saturated compound must possess an equal number of unsatisfied valences. Hence,

In HCl the radical Cl is equivalent to 1 atom of hydrogen;
In H_2O the radical O is equivalent to 2 atoms of hydrogen;
In NH_3 the radical N is equivalent to 3 atoms of hydrogen;
In CH_4 the radical C is equivalent to 4 atoms of hydrogen.

Therefore Cl is univalent, O bivalent, N trivalent, and C quadrivalent.

The same regard for valence is observed when radicals are made to displace each other, thus: $H^{I}{}_2(SO_4)^{II}$ requires two atoms of K^{I} or one of Zn^{II} to form $K^{I}{}_2(SO_4)^{II}$ or $Zn^{II}(SO_4)^{II}$.

Some elements exercise more than one valence: *e. g.* mercury may be univalent, as in HgI, or bivalent, as in HgI_2; or iron may be bivalent, as in $FeCl_2$, or the double atom (Fe_2) sexivalent, as in Fe_2Cl_6. The termination "*ous*" is given to those compounds

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* The student should bear in mind that valence has nothing to do with the combining weight or the chemical activity of an element.

in which the positive element exercises its lower valence, and "*ic*" to those in which the higher valence is exercised, as, $FeCl_2$, ferrous chloride; and Fe_2Cl_6, ferric chloride.

In the following table the most commonly occurring simple or elementary radicals are arranged according to their valence:

TABLE OF VALENCE.

I	II.	III	IV	V	VI
F, Cl	Ba, Sr	Al	C, Si		
Br, I	Ca, Mg	Au	Pt		
H, Ag	Cd, Zn	Bo			
	O				
K, Na	Pb, Sn		Pb, Sn		
(NH_4), Li	S, Se				S, Se
	Fe, Cr				Fe_2, Cr_2
	Mn, Co				Mn_2, Co_2
	Ni				Ni_2
		N, P		N, P	
		Bi, Sb, As.		Bi, Sb, As	
Cu, Hg	Cu, Hg				

The next table shows the valences, together with the symbols and formulæ, of the most common electro-negative (acidulous) radicals:

Univalent Radicals.
Cl is the negative radical of all chlorides.
Br is the negative radical of all bromides.
I is the negative radical of all iodides.
CN is the negative radical of all cyanides.
HO is the negative radical of all hydrates.
NO_3 is the negative radical of all nitrates.
ClO_3 is the negative radical of all chlorates.
$C_2H_3O_2$ is the negative radical of all acetates.

Bivalent Radicals.
O is the negative radical of all oxides.
S is the negative radical of all sulphides.
SO_3 is the negative radical of all sulphites.
SO_4 is the negative radical of all sulphates.
CO_3 is the negative radical of all carbonates.
C_2O_4 is the negative radical of all oxalates.
$C_4H_4O_6$ is the negative radical of all tartrates.

Trivalent Radicals.
$C_6H_5O_7$ is the negative radical of all citrates.
PO_4 is the negative radical of all phosphates.
BO_3 is the negative radical of all borates.

The student should learn these tables thoroughly, for with them he can easily know the formulæ of all the principal inorganic and organic compounds.

II. The Chlorine Group.

Name.	Derivation of Name.	Symbol.	At. Wt.
FLUORINE	Fluor spar	F	19
CHLORINE	χλωρός, green	Cl	35.5
BROMINE	Βρῶμος, stink	Br	80
IODINE	Ιώδης, violet	I	127

The members of this group are all univalent and much alike in their sources and physical and chemical properties. They differ in degree rather than in kind, forming a graded series. Hence we will consider them all together.

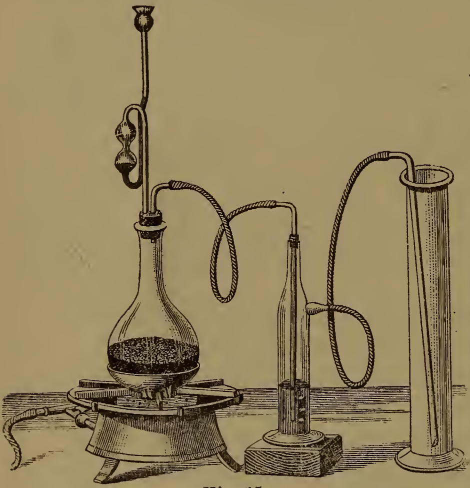

Fig. 15.

Sources. Never free in nature. The principal source of fluorine is fluor spar (CaF_2), while compounds of chlorine, bromine, and iodine are abundant in sea and other salt waters.

Preparation. Fluorine has probably never been isolated; the others may be prepared by removing the hydrogen from their hydrogen salts (hydracids) by means of oxygen derived from manganese dioxide, thus:

$$4HCl+MnO_2=MnCl_2+2H_2O+Cl_2.^*$$
$$4HBr+MnO_2=MnBr_2+2H_2O+Br_2.$$
$$4HI+MnO_2=MnI_2+2H_2O+I_2.$$

Physical Properties. Fluorine is not known in the free state,

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*Experiment. Into a flask standing in a cup of sand over a heater (Fig. 15) pour several ounces of hydrochloric acid and half as much black oxide of manganese, and agitate. The gas passes out through the wash bottle, and, being heavier than air, collects in the tall cylinder, where its yellowish-green color makes it visible.

but probably a colorless gas. Chlorine is a very irritating yellowish-green gas, two and a half times as heavy as air, slightly soluble in water (three volumes), forming "*Aqua chlori*, U. S." Bromine is a red liquid, giving off red vapors of a disagreeable, irritating odor; very slightly soluble in water.

Fig. 16.

Fig. 17.

Iodine is a solid, in bluish-gray scales, which, when warmed, give off violet vapors; insoluble in water except by the intervention of an alkaline iodide;* soluble in alcohol; irritating, even caustic.

Chemical Properties. Intensely electro-negative; great affinity for the metals,† especially hydrogen.‡ In negativeness, and consequently in affinity for the metals, F is greatest, Cl next, Br next, and I least. Therefore, in compounds with the metals, F will displace Cl, and Cl will displace Br, and either F, Cl, or Br will displace I.§ These elements destroy coloring matters and noxious effluvia in two ways: (1) by abstracting their hydrogen; (2) by abstracting the hydrogen of water, setting free nascent oxygen, which oxidizes the matters in question.‖

** **Experiment.** To some water in a test-tube add a few scales of iodine; it does not dissolve. Now add a crystal of potassium iodide; it dissolves easily.

† **Experiment.** Into a jar of chlorine introduce some copper or bronze foil, or sprinkle some powdered antimony. They inflame spontaneously.

‡ **Experiment.** (*a*) Into a jar of chlorine lower a lighted candle. (Fig. 16.) The hydrogen of the tallow burns in the chlorine to form hydrochloric acid, and all the carbon is liberated as smoke. (*b*) Into a similar jar thrust a piece of paper dipped in warm turpentine. It inflames spontaneously and burns, evolving dense clouds of smoke.

§ **Experiment.** Take two large test-tubes half full of water. Into one put a grain of potassium bromide, into the other potassium iodide; add chlorine-water to each. The chlorine will liberate the bromine in one and the iodine in the other. This may be shown (*a*) by their color; (*b*) by adding a few drops of carbon bisulphide or chloroform, which on agitation will gather all the free bromine and iodine, and be colored brown with the one and violet with the other; (*c*) add a few drops of starch-water, which will give brown with bromine and a deep blue with iodine.

‖ **Experiment.** (*a*) Into one bottle of chlorine gas introduce a piece of dry calico, into another a moist piece. The moist calico is rapidly bleached, while the dry is but slowly affected. (*b*) To a solution of indigo, cochineal, or some aniline color add chlorine-water. It is immediately decolorized.

Medical. Chlorine gas and bromine vapor are used for disinfection. Inhaled they cause severe coryza and bronchitis. Taken into the stomach, bromine and iodine cause gastro-enteritis. The antidote is boiled starch. Locally bromine is used as an escharotic and iodine as a counter-irritant.

Pharmaceutical. The following preparations are officinal: *Tinctura Iodi* (℥j–Oj); and *Liquor Iodi Compositus* (Lugol's Solution) (Iodine ʒvi, potassium iodide ℥iss, and water Oj). The so-called *colorless tincture of iodine* is made by adding ammonia-water to the tincture until it is decolorized by converting the iodine into ammonium iodide.

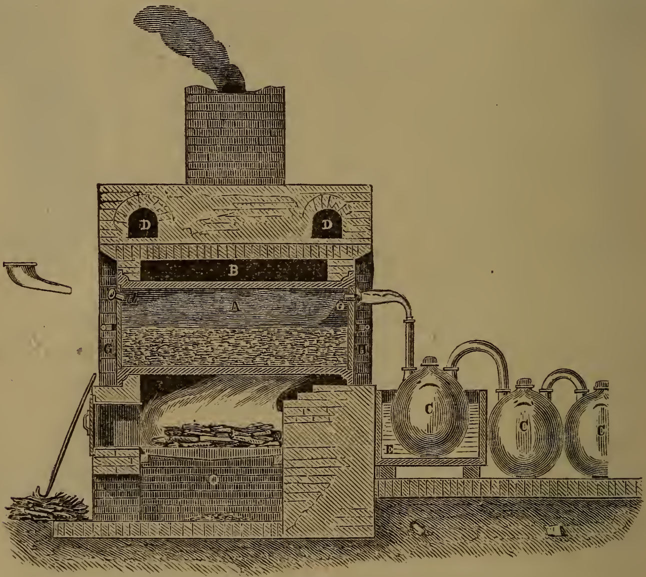

Fig. 18. Commercial Preparation of HCl.

Tests. In the free state chlorine and bromine may be known by their bleaching, color, odor, etc. Iodine is recognized by the blue color it strikes with starch.

Acids. All acids have, as their (positive) basylous radical, hydrogen, which may be replaced by metals to form salts. They may generally be recognized by a sour taste and the property of turning vegetable blues (*e. g.* litmus or purple cabbage) to reds. Acids whose acidulous (negative) radicals contain oxygen are called *oxacids;* those containing no oxygen, *hydracids.* The members of the chlorine group form both classes of acids.

The hydracids of the chlorine group are as follows:

$H+F=HF$—Hydrogen Fluoride—Hydrofluoric acid.

$H+Cl=HCl$—Hydrogen Chloride—Hydrochloric (muriatic) acid.

$H+Br=HBr$—Hydrogen Bromide—Hydrobromic acid.

$H+I=HI$—Hydrogen Iodide—Hydriodic acid.

Binary compounds—*i. e.* those of only two elements—are named by calling first the name of the positive and then that of the negative radical, affixing to the latter the termination "*ide.*"

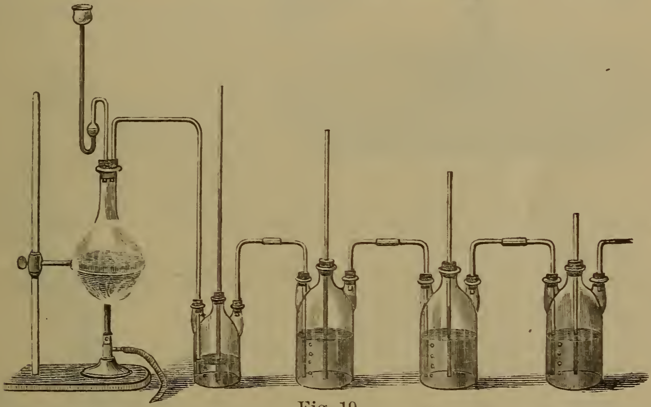

Fig. 19.

Prepared by treating the appropriate salt with H_2SO_4, thus:

$$CaF_2+H_2SO_4=CaSO_4+2HF.$$
$$2NaCl+H_2SO_4=Na_2SO_4+2HCl.^*$$
$$2KBr+H_2SO_4=K_2SO_4+2HBr.$$
$$2KI+H_2SO_4=K_2SO_4+2HI.$$

Physical Properties. Colorless, irritating gases; sharp, sour taste; very soluble, water dissolving several hundred times its own volume, forming *aquæ* known by the simple name of the acid

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* **Experiment.** To prepare hydrochloric acid in the laboratory, put several ounces of common salt and about twice as much sulphuric acid into a flask fitted as shown in Fig. 19. Hydrochloric acid gas is disengaged and may be collected by displacement (like chlorine, Fig. 18), or over mercury. The solution of the gas (the ordinary form) is obtained by passing the gas through a series of Wolff bottles (as shown in the figure) containing cold water.

itself, thus: The so-called hydrochloric acid is a solution of the hydrochloric acid gas in water.*

Chemical Properties. Strong acids; true acids even without water.

Uses. *HF* attacks silicon energetically, hence is used to etch glass; very poisonous, and burns made by it heal with difficulty.

HCl is very useful in the arts. *Aqua regia,* or nitro-muriatic acid, is a mixture of nitric and hydrochloric acids. It is the only solvent of gold and platinum. The metals are attacked by the nascent chlorine evolved by the oxidation of the H of the HCl by the O of the HNO_3. In medicine HCl is often prescribed as a tonic.

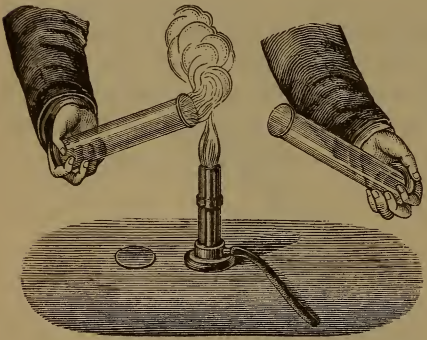

Fig. 20.

HBr, like all bromides, is a sedative. *HI,* like all iodides, is an alterative.

Tests. Fluoride+H_2SO_4—etches glass.†

Chloride+$AgNO_3$—white precipitate, soluble in ammonia.

Bromide+$AgNO_3$—yellowish-white precipitate, slightly soluble in ammonia.

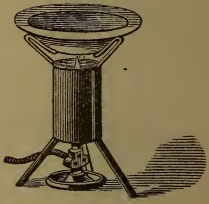

Fig. 21.

Iodide+$AgNO_3$—yellow precipitate, insoluble in ammonia.

If to a bromide or iodide some chlorine-water and starch paste

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***Experiment.** Fill two cylinders, one with hydrogen, the other with chlorine. Apply them mouth to mouth, and agitate until the gases are mixed. Hold their mouths to a flame (Fig. 20), or expose to direct sunlight. They explode with a shrill report, and are filled with hydrochloric acid fumes. Invert one cylinder over water; the gas is absorbed and the water rises to take its place. Into the other thrust a piece of moist blue litmus; it is reddened.

†**Experiment.** On a plate of glass coated with wax or copper-plate varnish (six parts of mastic, one of asphalt, and one of wax dissolved in turpentine) draw a design with a pointed instrument. Invert over a lead dish supported as in Fig. 21, and warmed gently. Hydrofluoric acid gas is evolved and attacks the glass wherever the wax has been scratched off. Upon removing the wax the design is found permanently etched on the glass.

be added, the bromine and iodine will be liberated, the bromine striking a brown and the latter a blue color with the starch.

Oxacids are formed by oxides of non-metals combining with water. The elements of the chlorine group, being very negative, have but little affinity for oxygen. Iodine has most, bromine less, chlorine still less, and fluorine will not unite with oxygen at all.

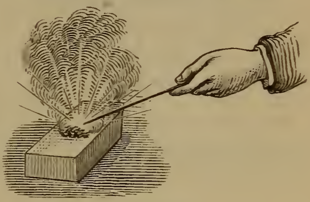
Fig. 22.

Chlorine, bromine, and iodine each forms a series of oxides perfectly analogous, so we will notice only those of one—chlorine.

The several oxides are distinguished by prefixes derived from the Greek numerals indicating the number of oxygen atoms in the formula, thus:

Cl_2O—Chlorine Monoxide.
Cl_2O_2 (?)—Chlorine Dioxide.
Cl_2O_3—Chlorine Trioxide.
Cl_2O_4—Chlorine Tetroxide.
Cl_2O_5—Chlorine Pentoxide.
Cl_2O_7—Chlorine Heptoxide.


Fig, 23.

These oxides combining with water form the corresponding acids, thus:

$Cl_2O+H_2O=2HClO$—Hydrogen Hypochlorite—Hypochlorous acid.

$Cl_2O_3+H_2O=2HClO_2$—Hydrogen Chlorite—Chlorous acid.

$Cl_2O_5+H_2O=2HClO_3$—Hydrogen Chlorate—Chloric acid.

$Cl_2O_7+H_2O=2HClO_4$—Hydrogen Perchlorate—Perchloric acid.

NOTE.—*The names of oxacids are derived from the negative element other than oxygen, and to this certain affixes and prefixes are added to indicate the degree of oxidation. The one containing more oxygen has the affix "-ic," less oxygen, "-ous." If there is in the same series another acid with more oxygen than the "-ic," it is given the prefix "per-;" if less than the "-ous," the prefix "hypo-" (under). Acids ending in "-ic" form salts ending in "-ate;" those ending in "-ous" form salts ending in "-ite." The foregoing chlorine acids illustrate this.*

All these oxides, as well as their corresponding acids, are easily decomposed, sometimes with explosion;* hence much used

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*__Experiment.__ Their oxidizing action on combustibles may be shown by: (*a*) Mix together a dram each of powdered potassium chlorate and sugar; place on a brick (Fig. 22) and touch off with a glass rod dipped in sulphuric acid. A as oxidizing agents and as explosive mixtures.* The most important of these salts is potassium chlorate, used in medicine and in the laboratory for the sake of its oxygen.

III. Sulphur Group.

Sulphur	S	32
Selenium	Se	79.4
Tellurium	Te	128

The elements comprising this group are solid at ordinary temperatures; bivalent and sexivalent; possess electro-negative affinities which, as in other groups, decrease as the atomic weights increase; form hydracids as well as oxacids.

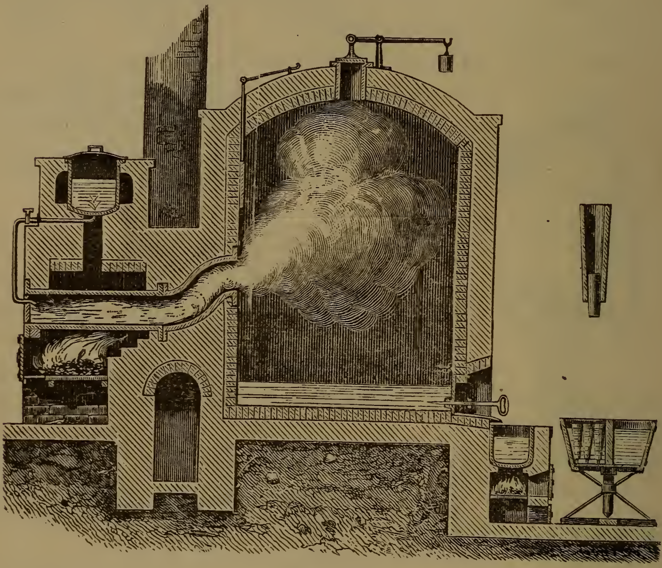

Fig. 24.

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vigorous combustion occurs. (*b*) Drop some crystals of potassium chlorate into a conical glass of water (Fig. 23); add several bits of phosphorus; then by means of a pipette introduce sulphuric acid at the bottom of the glass. The phosphorus takes fire and burns at the expense of the oxygen of the potassium chlorate.

* **Experiment.** Mix on a sheet of paper a scruple of powdered potassium chlorate and five grains of some combustible powder, as sulphur, antimony sulphide, or tannin. Wrap it up in the paper, place upon an anvil, and strike with a hammer. It explodes violently.

The analogy between their compounds is shown in the following table:

Hydro-ic acid.	Dioxide.	Trioxide.	Hypo-ous acid.	-ous acid.	-ic acid.
H_2S	SO_2	SO_3	H_2SO_2	H_2SO_3	H_2SO_4.
H_2Se	SeO_2	SeO_3		H_2SeO_3	$HSeO_4$.
H_2Te	TeO_2	TeO_3		H_2TeO_3	H_2TeO_4.

Selenium and Tellurium are of no medical interest, and will will not be further noticed.

SULPHUR *occurs* free, especially in the neighborhood of volcanoes; occurs combined as sulphides and sulphates in many valuable ores, and in small quantity in the animal and vegetable kingdoms.

Preparation. The native sulphur freed from stones is refined by distillation, as shown in Fig. 24. The crude sulphur is melted in the tank by the hot draft from the fire below, and then runs down through a pipe into the retort, where it is vaporized. This vapor entering a large brick chamber, is condensed in fine, feathery crystals, called *flowers of sulphur* or *sublimed sulphur.** If the chamber be hot, it condenses into a liquid, which is drawn off and molded into rolls, called *roll brimstone.* Sublimed sulphur is apt to contain more or less acid, and is washed (*sulphur lotum*). Boiled with lime and precipitated with HCl, it forms *sulphur precipitatum,* U. S. P. This mixed with water is *milk of sulphur* (*lac sulphuris,* U. S. P.)

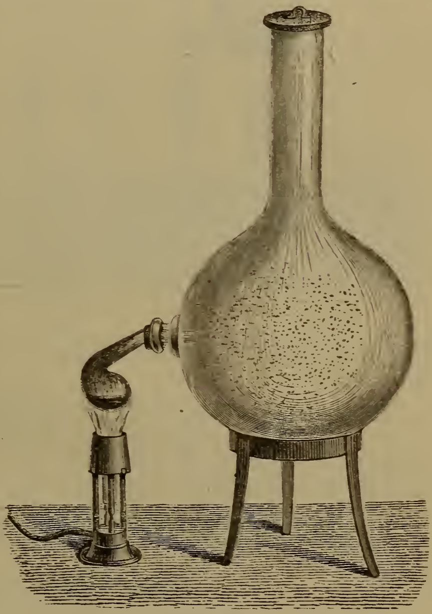

Fig. 25.

***Experiment.** Some sulphur is heated in a small retort or flask connected with the side tubulure of a large receiver. (Fig. 25.) The vapor passes into the receiver, where it condenses in flowers of sulphur.

Physical Properties. A brittle yellow solid; insoluble in water, hence tasteless, etc.

Chemical Properties. Inflammable, hence called "brimstone" (burn-stone). Combines with metals, forming sulphides.* Sulphur forms compounds remarkably analogous to those of oxygen, *e. g.*:

H_2O. KHO. CO_2. H_2CO_3. HCNO.
H_2S. KHS. CS_2. H_2CS_3. HCNS.

Uses. In the arts, to make gunpowder, matches, etc.; in medicine, as a laxative, parasiticide, and alterative. We have only theoretical explanations of the method of its absorption; but that it is absorbed is certain, for persons taking it excrete enough to blacken silver carried on the person.

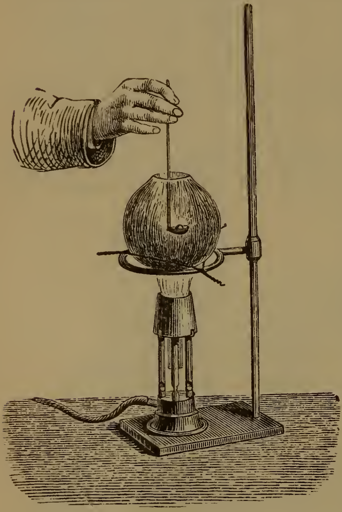

Fig. 26.

HYDROGEN SULPHIDE—H_2S—*Hydrosulphuric Acid or Sulphureted Hydrogen—occurs* in sewer gas and other effluvia from decomposing organic (especially animal) matters, and in the water of sulphur springs.

Prepared in laboratory by decomposing a sulphide, thus:

$$FeS+H_2SO_4=FeSO_4+H_2S.$$

Physical Properties. Colorless gas, having the odor of rotten eggs or intestinal flatus; slightly soluble in water.

Chemical Properties. Very feeble acid; burns with pale blue flame:

$$H_2S+O_3=SO_2+H_2O.†$$

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[*] **Experiment.** In a small glass flask, the neck broken off, a little sulphur is heated to boiling. If now a bundle of fine copper wire or a piece of sodium, in a combustion spoon, be previously heated and then lowered into the vapor (Fig. 26), it burns brilliantly.

Experiment. Mix in a dish equal parts of iron filings and flowers of sulphur; moisten with water and set aside. Within a half hour it gets hot, vaporizes the water, and is converted into a black mass of FeS.

[†] **Experiment.** Burn the gas from a jet; (*a*) Hold near the flame a glass rod dipped in ammonia; white crystals of ammonium sulphide are formed. (*b*) Hold a cold, dry bell glass over the flame; it is bedewed with water.

Forms characteristic precipitates with most metallic salts, hence a valuable test reagent.*

Tests. The presence of H_2S even in minute quantities may be detected by its odor, and by its blackening paper moistened with a solution of lead acetate.

Physiological. When inhaled H_2S is an active poison, combining with the hemaglobulin and destroying its oxygen-carrying power. Even when highly diluted, as in the atmosphere of city dwellings, clumsily "fitted with the modern conveniences," it produces a low febrile condition. When concentrated, or even moderately diluted, the gas proves rapidly fatal.

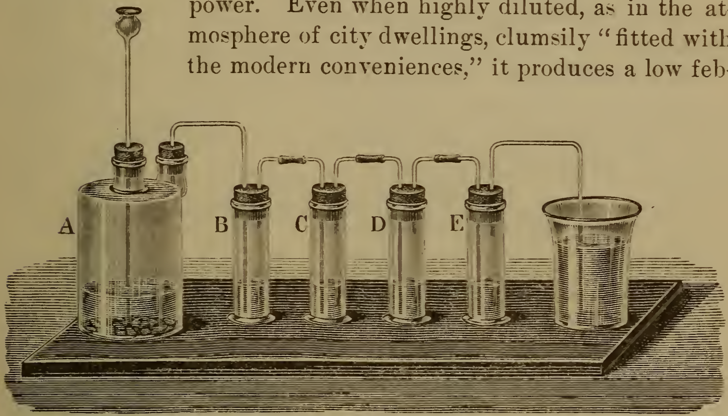

Fig. 27.

Treatment. Fresh air, artificial respiration, and stimulation.

Carbon Disulphide—CS_2. Obtained by bringing S into contact with heated charcoal. A colorless, volatile liquid of a fetid odor, unless it is very pure. A valuable solvent for S, P, india-rubber, etc.

Sulphur Oxides and Acids.

Dioxide—$SO_2+H_2O=H_2SO_3$—Sulphurous acid.
Trioxide—$SO_3+H_2O=H_2SO_4$—Sulphuric acid.

***Experiment.** To show the action of H_2S on metallic salts, connect several wash bottles with the generator A, as shown in Fig. 27, the last tube ending in a beaker of ammonia-water. A dilute solution of lead acetate is put in B, of tartar emetic (antimony) in C, of arsenic in D, of zinc sulphate in E. The gas in passing precipitates lead sulphide (black) in B, antimonious sulphide (orange) in C, arsenious sulphide (yellow) in D, zinc sulphide (white) in E, and is finally taken up by the ammonia to form ammonium sulphide.

SULPHUR DIOXIDE—SO_2 *occurs* whenever sulphur or any of its compounds are burned in air or oxygen.

Prepared in laboratory by decomposing sulphuric acid by copper or charcoal, thus:

$$2H_2SO_4+Cu=CuSO_4+2H_2O+SO_2.$$
$$2H_2SO_4+C=2SO_2+CO_2+2H_2O.$$

Physical Properties. A colorless gas, with a suffocating odor (of burning matches); dissolves in water to form sulphurous acid (H_2SO_3).

Chemical Properties. Neither burns nor supports combustion; a strong deoxidizer; by removing O from coloring matters and infecting germs it bleaches* and disinfects.

Uses. Sulphur dioxide, sulphurous acid, and the sulphites possess the property of destroying micro-organisms and arresting fermentations. A sulphite digested with sulphur forms a hyposulphite, thus:

$$Na_2SO_3+S=Na_2S_2O_3.$$

Sodium hyposulphite has the same uses as the sulphites, and is also a valuable solvent of the silver salts in photography.

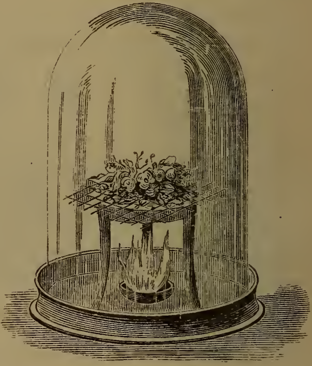

Fig. 28.

SULPHUR TRIOXIDE—SO_3. Made by oxidizing SO_2 in the manufacture of sulphuric acid. This is done upon a large scale by passing SO_2 from burning sulphur into a chamber kept filled with vapor of nitric acid, steam, and air.† The nitric acid gives up a part of its oxygen to oxidize a portion of the SO_2 to SO_3.

$$2HNO_3+3SO_2=3SO_3+H_2O+N_2O_2.$$

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* **Experiment.** Some sulphur is ignited beneath a tripod on which fresh flowers are placed, and the whole covered by a bell glass. (Fig. 28.) The flowers are bleached. The color may be restored by washing with some dilute alkali or acid that will combine with or displace the SO_2, or with very dilute nitric acid, which will restore the oxygen removed by the SO_2.

† The manufacture of sulphuric acid may be illustrated on the lecture-table by the apparatus shown in Fig. 29. The lead chamber is represented by a large flask. Into this are led (*a*) N_2O_2 from the two-necked bottle on the right; (*b*) SO_2 from a mixture of sulphur and manganese dioxide in the flask to the left; (*c*) steam from the other flask, and (*d*) air or oxygen through the rubber tube.

The SO_3 then combines with the water thus produced ($SO^3+H_2O=H_2SO_4$), and more water is supplied by a jet of steam thrown constantly into the chamber.

The N_2O_2 has the power of taking up oxygen from the air and becoming N_2O_4,

$$N_2O_2+O_2=N_2O_4,$$

which in turn parts with this oxygen to oxidize a new quantity of SO_2.

$$N_2O_4+2SO_2=N_2O_2+2SO_3.$$

Thus the process is kept up as long as the SO_2, air, steam, and N_2O_2 are supplied. The acid condenses with the water upon the floor of the chamber, and is drawn off, concentrated, and sold as

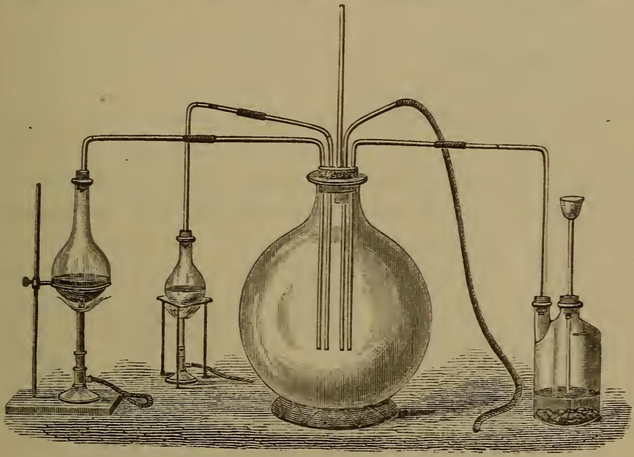

Fig. 29.

SULPHURIC ACID—H_2SO_4—"*Oil of Vitriol.*"

Physical Properties. A dense, colorless, oily-looking liquid, without odor.

Chemical Properties. Strong acid; very avid of water, not only dissolving in it, but combining with it, the act evolving considerable heat; chars organic matters by abstracting H and O to form water (H_2O).*

Uses. So important in the arts that the commercial prosperity of a country may be measured by the amount of H_2SO_4 consumed.

***Experiment.** Pour strong sulphuric acid on an equal quantity of sugar or strong syrup: the sugar is dehydrated and a mass of carbon left.

Properly diluted, it is a refrigerant tonic, but concentrated it is a severe caustic.

Tests. (1) The concentrated acid, if placed on a piece of paper or other organic material, will char it. If dilute, it will char the paper only after being warmed and concentrated by the evaporation of its water. (2) Sulphuric acid, or any other sulphate, will form with a solution of a barium salt a white precipitate ($BaSO_4$) insoluble in nitric or hydrochloric acid.

IV. **Nitrogen Group.**

NITROGEN	N	14
PHOSPHORUS	P	31
ARSENIC	As	75
ANTIMONY (Stibium)	Sb	122
BISMUTH	Bi	210

TRIVALENT AND QUINQUIVALENT. This group, as shown below, forms a graded series from nitrogen and the negative to bismuth at the positive end.

N	P	As	Sb	Bi
14	31	75	122	210
	Sp. gr. 1.83.	Sp. gr. 5.67.	Sp. gr. 6.7.	Sp. gr. 9.8.
Gas, with full negative tendencies.	A soft solid.	Solid.	Hard solid.	Very hard solid.
	Easily volatilizable.	Volatilizable.	Difficultly volatilizable.	Non-volatizable.
	Destitute of metallic luster.	Some metallic luster.	Great metallic luster.	Full metallic luster.
	Negative tendencies.	Both negative and positive tendencies.	More positive tendencies.	Full positive tendencies.

The following will exhibit the relations of some of the most important compounds:

Hydrides.	Chlorides.		Oxides.		Sulphides.	
	-ous,	*-ic.*	*-ous,*	*-ic.*	*-ous,*	*-ic.*
NH_3	NCl_3	……..	N_3O_3,	N_2O_5	……..	……..
PH_3	PCl_3,	PCl_5	P_2O_3,	P_2O_5	P_2S_3,	P_2S_5
AsH_3	$AsCl_3$,	$AsCl_5$	As_2O_3,	As_2O_5	As_2S_3,	As_2S_5
SbH_3	$SbCl_3$,	$SbCl_5$	Sb_2O_3,	Sb_2O_5	Sb_2S_3,	Sb_2S_5
……	$BiCl_3$,	……..	Bi_2O_3,	Bi_2O_5	……..	……..

Nitrogen *occurs* uncombined in the atmosphere; combined in some mineral and in all vegetable and animal bodies, especially in the more highly organized tissues.

Prepared most easily by burning phosphorus in a confined space until the oxygen is removed from the air.* Prepared in this way it contains small quantities of other gases found in air. To prepare it pure, heat ammonium nitrite ($NH_4NO_2=2H_2O+N_2$).

Fig. 30.

Physical Properties. A colorless, tasteless, odorless gas, a little lighter than air.

Chemical Properties. Little tendency to combine with other elements, and its compounds, once formed, are very prone to decompose, either with violent decomposition † or gradual putrefaction; neither combustible nor a supporter of combustion; negatively poisonous.

The Atmosphere. Air, considered by the ancients one of the four elements, is neither an element nor a compound. It is a mixture ‡ mainly of nitrogen and oxygen, the function of the former being to dilute the latter. Miller gives the average composition of air as follows:

	Volumes.
Nitrogen	77.95
Oxygen	20.61
Carbon dioxide	.04
Aqueous vapor	1.40

Also traces of nitric acid, ammonia, sodium chloride, ozone, dust, bacteria, germs, etc. In the neighborhood of large cities various

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* **Experiment.** A flat piece of cork floating on water supports a capsule containing a bit of phosphorus carefully dried. This is ignited and immediately covered with a bell jar. The jar is filled with a dense white cloud from the combustion, which ceases only when the oxygen is all consumed. At first the air expands and some may be forced out. Upon cooling the water rises to take the place of the oxygen, and the white fumes gradually dissolve in the water, and the nitrogen is left clear and comparatively pure.

† **Experiment.** To tincture of iodine add excess of ammonia-water. Filter to separate the precipitated iodide of nitrogen. Put portions of this on separate bits of paper and set aside. When dry they explode on the slightest touch.

‡ Proofs that air is a mixture: (1) Its constituents are not in atomic proportions; (2) air can be made by mechanically mixing the gases; (3) solvents may remove one gas without affecting the others, each dissolving according to its own solubility.

other substances are poured into the air from manufactories. Yet, owing to the rapid diffusion of gases, the composition of the air is almost the same every where.

Watery Vapor. The higher the temperature the more water air will hold. A warm, dry air, when cooled, will appear damp, and the temperature at which it begins to deposit its water is its *dew-point.* A cold, damp air, when heated, becomes capable of holding more water, and appears dry; hence the necessity of supplying water to the heated air of our rooms in winter, especially in cases of bronchitis or catarrhal croup. Even in health a very dry air irritates the air-passages, produces dryness of the skin and *malaise;* while a very moist atmosphere retards evaporation from the skin and lungs, raises the body temperature, and becomes oppressive.

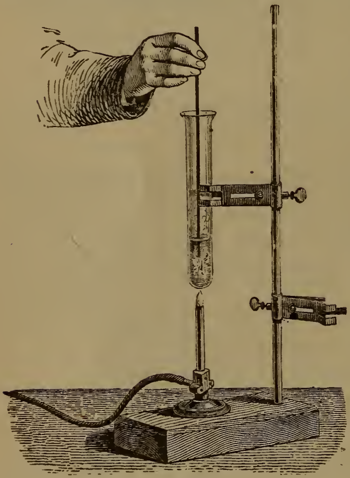

Fig. 31.

Suspended matters in air are of a great variety of substances. The irritation of dust incident to certain trades may cause chronic bronchitis, emphysema, and phthisis. Germs floating in the air are believed to be the cause of many contagious, infectious, and malarial diseases. The best disinfectants* are (*a*) free ventilation and consequent dilution; (*b*) chlorine, bromine, iodine, and sulphur dioxide.

NITROGEN HYDRIDE—AMMONIA, NH_3—*Occurs* in the effluvia from decomposing nitrogenized organic bodies; for nitrogen and hydrogen unite only in the nascent state. (See page 20.) First obtained by distilling camel's dung near the temple of Jupiter Ammon in Libya; hence called "ammonia." Obtained by heating clippings of hides, hoofs, and horns,† especially of deer, in closed retorts (destructive distilla-

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* *Disinfectants* destroy the power to infect, whether it be due to germs or other agent. *Germicides* destroy germs. *Antiseptics* prevent putrefaction. *Antizymotics* prevent fermentation. *Deodorizers* destroy offensive odors.

† **Experiment.** Mix calcium, potassium, or sodium hydrate with some nitrogenous substance, as albumen or clippings of horn, hoofs, flannel, or lean

tion), it was called *spirit of hartshorn.* Coal contains about two per cent of nitrogen, which in the manufacture of coal gas comes off as ammonia. In washing the coal gas the ammonia dissolves in the water. This *aqua* is its commercial source.

Prepared in laboratory by driving the ammonia off from the *aqua* by means of heat. (Fig. 32.)

Physical Properties. Transparent, colorless gas, of an irritating odor; very soluble in water, which dissolves from 500 to 1,000 times its own volume.* Administered by inhalation as a stimulant in fainting fits, etc., but care must be taken, for its too liberal use may cause spasm of the glottis or induce a fatal bronchitis.

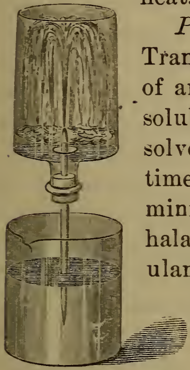
Fig. 33.

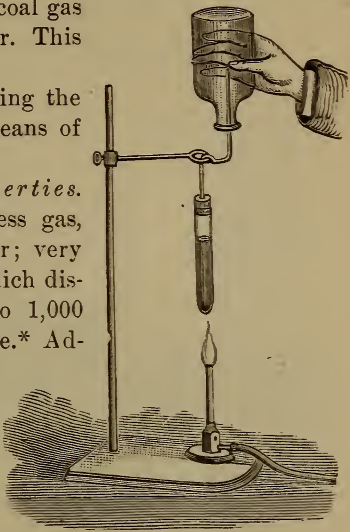
Fig. 32.

Tests. (1) Smell; (2) white fumes with HCl; (3) turns moistened red litmus paper blue.

NITROGEN OXIDES.

Monoxide—$N_2O+H_2O=2HNO$=Hyponitrous acid.
Dioxide—N_2O_2. No corresponding acid.
Trioxide—$N_2O_3+H_2O=2HNO_2$=Nitrous acid.
Tetroxide—N_2O_4 No corresponding acid.
Pentoxide—$N_2O_5+H_2O=2HNO_3$=Nitric acid.

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meat. Heat in a test-tube. (Fig. 31.) Ammonia gas is evolved, recognized by its odor, alkalinity, or by white fumes forming when a glass rod moistened with HCl is thrust into the tube.

***Experiment.** The absorption of ammonia gas by water may be illustrated by filling a large bottle with the gas by upward displacement (Fig. 32), and closing the mouth with a rubber cork through which passes a glass tube sealed at the outer end. If this sealed end be plunged under water and then broken off, the water rushes in, forming a fountain. (Fig. 33.) If the water be colored with red litmus solution it will become blue as it enters the bottle, showing the alkalinity of the solution.

NITROGEN MONOXIDE—N_2O (*Nitrous Oxide—Laughing Gas*). *Prepared* by heating ammonium nitrate, as shown in Fig. 34.

$$NH_4NO_3=N_2O+2H_2O.$$

Physical Properties. Colorless, odorless gas, of sweetish taste. Dentists keep it liquefied under pressure in iron cylinders.

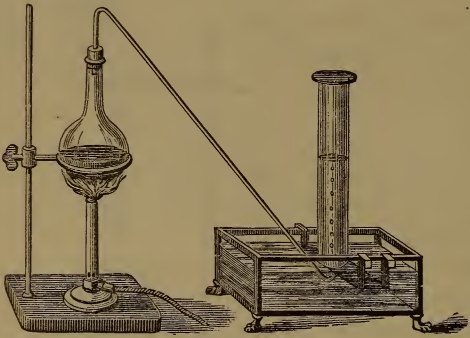

Fig. 34.

Chemical Properties. By the ease with which it gives up its O it is a supporter of combustion and life, next to O itself.

Medical. Inhaled, diluted with air, it produces exhilaration of spirits, muscular activity, and then complete anæsthesia. Used in dental and other brief minor operations.

NITROGEN DIOXIDE—N_2O_2 (*Nitric Oxide*). *Prepared* by action of nitric acid on copper:

$$3Cu+8HNO_3=3Cu(NO_3)_2+4H_2O+N_2O_2.$$

A colorless gas, which, when coming in contact with free O, forms red vapors of N_2O_3 and N_2O_4; hence a test for free O. (Fig. 35.)

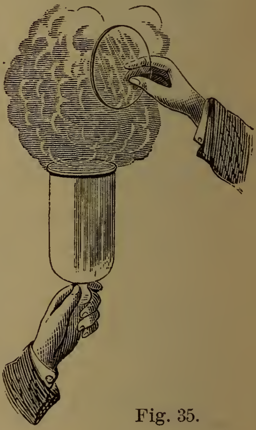

Fig. 35.

NITROGEN TRIOXIDE—N_2O_3 (*Nitrous Acid*—HNO_2). Nitrous acid is known only in its salts, the nitrites. These are produced in nature by the oxidation of nitrogenous organic matter in the presence of certain forms of microscopic life.

This nitrification occurs in waters polluted with organic matter, and normally in the soil, where the acid so formed combines with bases. Hence nitrites in water is evidence of previous contamination with nitrogenous matter. Further oxidation forms nitrates.

NITROGEN TETROXIDE—N_2O_4—*occurs* in company with N_2O

in the brown fumes given off whenever nitric acid is decomposed, as in certain laboratory and manufacturing processes. *The effect* of breathing air thus contaminated is to produce chronic inflammation of the respiratory tract. If the vapor be more concentrated the effects are more acute and serious. At first there is only a cough, in two or three hours a difficulty of breathing, and in about twelve hours death. The remedy is ventilation.

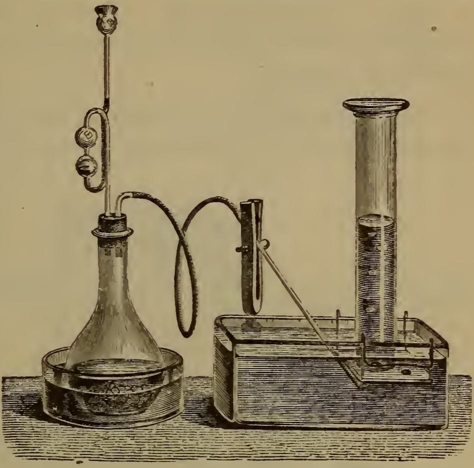

Fig. 36.*

NITROGEN PENTOXIDE—N_2O_5—is of no medical interest.

NITRIC ACID—HNO_3 (*Aqua Fortis*)—*occurs* in traces in the atmosphere and as nitrates in the soil. (See Nitrites.)

Prepared by distilling a nitrate with sulphuric acid.

$$2KNO_3+H_2SO_4=K_2SO_4+2HNO_3.\dagger$$

Physical Properties. Heavy liquid, colorless, but if old and exposed to light it may be yellow or orange from presence of N_2O_3 and N_2O_4. Like all other nitrates, it is soluble in water.

Chemical Properties. Energetic oxidizer;‡ corrosive; stains skin indelibly yellow.

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*Experiment. Copper turnings, clippings, or wires are placed in a flask, and nitric acid diluted with half its volume of water is poured in, and the flask set in cold water. (Fig. 36.) Red fumes soon fill the flask, but when these have escaped the gas appears colorless, turning red, however, on reaching the air. The colorless gas is collected over water.

†**Experiment.** In the laboratory nitric acid may be prepared with the apparatus shown in Fig. 14, page 17. Equal parts of sodium nitrate and sulphuric acid are heated in the retort. The nitric acid produced is vaporized by the heat and recondensed in a receiver supported over a beaker and kept cool by a wet cloth, over which flows a stream of water from an elevated vessel.

‡**Experiment.** Into a mixture of strong sulphuric and nitric acids pour from a beaker tied to a long stick some warm turpentine. The oxidation is so rapid that the turpentine is inflamed.

Medical Uses. The strong acid is an escharotic, coagulating the albumen of the tissues; the dilute, a refrigerant tonic.

Tests. (1) Yellow stain. (2) Add H_2SO_4, and then a crystal of $FeSO_4$ dropped in will be colored brown if nitric acid or any nitrate be present.

PHOSPHORUS (Light-bearer) *occurs* combined with O in the ancient unstratified rocks. These disintegrate and form soil, from which the P passes into the organisms of plants, and thence into the bodies of animals. First isolated by Brandt (1669) from urine; now obtained from bones.

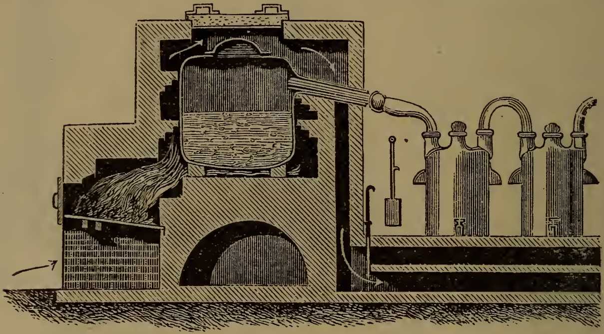

Fig. 37. Commercial Preparation of Nitric Acid.

Physical Properties. A soft, yellowish solid, resembling unbleached wax. Insoluble in water, but soluble in carbon disulphide, ether, chloroform, oils, etc.

Chemical Properties. Very inflammable,* so kept under water; exposed to the air, it undergoes a slow combustion, emits the odor of ozone, and is luminous in the dark.

Physiological. Liable to inflame from careless handling, and burns by it are difficult to heal. In medicinal doses, a nerve tonic and aphrodisiac; in larger quantities a virulent poison and gastro-irritant. Sometimes given with homicidal intent, but more frequently taken accidentally as rat poison, tips of matches, etc. Workmen in match factories suffer from irritation of stomach

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* **Experiment.** Dissolve some phosphorus in carbon disulphide. Pour this on a sheet of filter paper hung on a retort stand. Soon the solvent evaporates and leaves the phosphorus in such a fine state of division that it inflames spontaneously.

and bowels, caries of teeth, necrosis of bones, especially of lower jaw, and from fatty degeneration of various organs. This may be prevented by using a red allotropic variety, which is harmless.

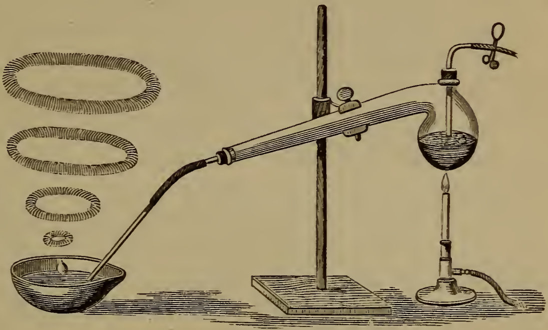

Fig. 38.

No good antidote. Evacuate the stomach; give copper sulphate* as emetic and antidote; old turpentine, the ozone of which oxidizes the P. Avoid fats, for they dissolve it.

Tests. (1) Shines in the dark; (2) emits garlicky odor.

Fig. 39.

Phosphorus Hydride—PH_3 (*Phosphoretted Hydrogen—Phosphine*)—*occurs* mixed with other hydrides of P in the gases arising from decomposing animal or vegetable matters, especially under water; hence seen as the *ignis fatuus* or "Will o' the wisp" over marshes and graveyards.

Prepared by boiling phosphorus in a solution of caustic potash.†

Properties. Colorless gas, of a garlicky odor; inflames spontaneously upon coming in contact with the air; very poisonous.

Phosphorus Oxides. These are analogous to the oxides of nitrogen, and form, on the addition of water, analogous acids.

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* **Experiment.** Place a clean bit of phosphorus for a minute in a solution of copper sulphate. Remove, and note the coating of metallic copper.

† **Experiment.** Into a retort, whose delivery-tube dips under water in a dish, add liquor potassæ and a few bits of phosphorus. Expel the air by passing hydrogen or illuminating gas through the retort, or by adding a few drops of ether, the vapor of which does the same thing. On applying heat the hydrogen or illuminating gas or ether vapor first escapes, then come bubbles of PH_3, each of which, as it bursts into the air, ignites spontaneously, forming beautiful rings of white smoke rotating on their circular axes. These may ascend to the ceiling if the air be still.

Phosphorus Pentoxide (P_2O_5) is produced whenever P burns in air* or O; and forms three different phosphoric acids by combining with one, two, or three molecules of water, thus:

$P_2O_5+3H_2O=H_6P_2O_8=2H_3PO_4$=Orthophosphoric acid.
$P_2O_5+2H_2O=H_4P_2O_7$=Pyrophosphoric acid.
$P_2O_5+H_2O=H_2P_2O_6=2HPO_3$=Metaphosphoric acid.

Orthophosphoric Acid. Never found free, but is widely disseminated in the three kingdoms of nature in its salts, the phosphates. Being the phosphoric acid most used in medicine (the other two are poisonous), it is usually called simply "phosphoric acid." Transparent, odorless, colorless, syrupy liquid. Being tribasic, it forms three classes of phosphates by displacement of one, two, or three atoms of the basic hydrogen, thus:

KH_2PO_4, K_2HPO_4, and K_3PO_4.

In the diluted form (*acidum phosphoricum dilutum*) it is prescribed as a tonic, especially in dyspepsia.

Tests. Add a few drops of the magnesian fluid ($MgSO_4$,NH_4Cl, and NH_4HO, each one part, water eight parts); a white precipitate indicates phosphoric acid or other phosphate.

ARSENIC *occurs* mostly as sulphide, usually associated with other metals. The ore is roasted, and the resulting oxide heated with carbon (charcoal) gives the metal. This is a brittle, steel-gray, crystalline solid, possessing a metallic luster. Heated out of contact with air it sublimes; in air it burns with a bluish-white flame, emitting the odor of garlic and white clouds of As_2O_3. It combines with many elements; the metallic arsenides resemble alloys. *Used* in pyrotechny and in the manufacture of shot, pigments, and fly-poison. All its compounds are poisonous.

Arsenious Hydride—AsH_3—*Arseniuretted Hydrogen—Arsine*—is of great practical interest to the toxicologist, as its formation constitutes one of the most delicate tests for arsenic. Forms whenever hydrogen is generated in presence of an arsenical compound.

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*__Experiment.__ A little stand in the middle of a dinner-plate supports a capsule into which is put a bit of phosphorus freed from adhering water. This is ignited and covered with a bell jar. (Fig. 39.) This jar is filled with clouds of P_2O_5, which, aggregating, fall into the plate like a miniature snow-storm.

ARSENIOUS IODIDE—AsI_3. *Prepared* by fusing together atomic proportions of its constituent elements. It enters into *Donovan's Solution, liq. arsenii et hydrargyri iodidi,* U. S. P.

ARSENIOUS SULPHIDE—As_2S_3—*occurs* native as *orpiment. Prepared* by precipitating an arsenious compound with H_2S. Bright yellow powder, insoluble in water or acid solutions, but soluble in alkaline. Another sulphide is *realgar*, AsS_2. Both are *used* as pigments—the orpiment as a yellow, the realgar as a red.

OXIDES AND ACIDS.

$$As_2O_3+3H_2O=2H_3AsO_3, \text{ (ortho) Arsenious acid.}$$
$$As_2O_5+3H_2O=2H_3AsO_4, \text{ (ortho) Arsenic acid.}$$

ARSENIOUS OXIDE—As_2O_3. *Arsenic, White Arsenic, Ratsbane, Arsenious Acid.* This is not only the most important compound of arsenic, but the most important of toxic agents, whether we consider the deadliness of its effects or the fatal frequency of its administration. When recently made it is in glassy lumps, which on exposure become crystalline and opaque. When sublimed it is deposited again in brilliant octahedral crystals. It is odorless, almost tasteless—slightly sweetish. When powdered arsenic is thrown upon water it does not all sink, notwithstanding its heaviness, but floats, showing a repulsion of the water. Very slightly soluble in water, even boiling water dissolving less than two per cent. If the water be made acid or alkaline, it dissolves more readily. When arsenic dissolves in water it forms *arsenious acid*, H_3AsO_3.

There are two officinal solutions, each containing one per cent of arsenic: (1) *Liq. acidi arseniosi,* in which the water is acidulated with HCl; (2) *Fowler's Solution, liq. potassii arsenitis,* in which the water is made alkaline by K_2CO_3.

ARSENIC OXIDE. *Arsenic pentoxide* is made when arsenious oxide (As_2O_3) is treated with an oxidizing agent, as nitric acid. It is quite soluble in water, with which it forms a series of *arsenic acids* (ortho-, pyro-, and meta-) analogous to the phosphoric acids.

Toxicology of Arsenic. The deadly effect of arsenical compounds has been known from remote antiquity, and they have probably been more used for homicidal purposes than all other toxic agents combined. Although chemistry has made its detection easy and certain, arsenic is so cheap, so readily administered to the unsuspecting victim, and so deadly, that it is still a favorite with the murderer. Owing to the extensive use of arsenical compounds as insect-powders (paris green, etc.), and as pigments for wall-paper, toys, confectionery, etc., cases of accidental poisoning are quite common.

Few physicians have the training and facilities to undertake an extended analysis, but they should all know the simpler tests, so as to promptly recognize the nature of the poison and combat it more intelligently and successfully. Besides, the physician, being early in the case, can by wise precautions prevent breaks in the chain of evidence; protecting the prisoner if innocent, and closing loop-holes of escape if guilty. If foul play is suspected, he should commit all his observations to writing, for notes to be admitted as evidence must be the original ones taken at the time. Having collected the urine, fæces, vomit, and the suspected vehicle of the poison, and having tested some or all of them to verify his suspicion, he should place them under seal or lock and key. He should carefully reserve his opinion, lest he do injustice to the innocent or warn the guilty. In case of death, the coroner should be notified and an autopsy held, in presence of the chemist if possible. The stomach and entire intestinal canal, ligated at both ends, half of the liver, the whole brain, spleen, one kidney, and any urine remaining in the bladder should be saved. These, if possible, should be preserved in separate jars, to which a little pure chloroform may be added to prevent decomposition. These jars must be new and clean, closed with new corks or glass—not zinc caps. They are then to be labeled and also sealed and stamped, so they can not be opened without detection, and as soon as possible turned over to the chemist or prosecuting officer.

The symptoms of arsenical poisoning are those common to all intense irritants, viz., nausea, vomiting, burning pain in the epigastrium, purging, cramps, thirst, fever, rapid pulse, etc., ending in collapse. Smallest fatal dose is two grains, and death usually occurs in twenty-four hours.

Treatment. Remove any unabsorbed poison from the stomach by emetics or stomach-pump. The best *antidote* * is freshly pre-

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* An antidote is something harmless and capable of rendering the poison harmless. Since poisons are inert when insoluble, antidotes are usually such substances as are capable of combining with the poison to form an insoluble and therefore inert compound.

cipitated ferric hydrate, made by adding aqua ammoniæ to a solution of a ferric salt. "Dialysed iron," being a solution of ferric hydrate, may be used. It should be given at frequent intervals and in tablespoonful doses.

TESTS FOR ARSENIC. *In the solid state:* 1. Heated on a knife-blade over a lamp, it *volatilizes* with a white smoke, and leaves no residue.

2. Heated in a test-tube it *sublimes*, and is recondensed in the cooler portion of the tube (Fig. 40) as *octahedral crystals* (Fig. 41).

Fig. 40.

3. Heated in a small tube with powdered charcoal, the arsenic is reduced as it sublimes, and recondenses on the cooler portion of the tube in the metallic state.

In the liquid state: 1. Through the solution, acidulated or rendered neutral, pass H_2S; a yellow precipitate (As_2S_3) falls.

Fig 41.

2. To an aqueous solution add a few drops of nitrate of silver, and then cautiously add ammonia, drop by drop, till a yellow precipitate, *silver arsenite* (Ag_3AsO_3), is obtained, showing the presence of arsenic.

3. Repeat the preceding, adding copper sulphate instead of silver nitrate, and the presence of arsenic is indicated by a green precipitate of copper arsenite (Scheele's green or paris green).

The last two tests may be performed with greater ease and delicacy if the silver nitrate and copper sulphate each be previously treated with ammonia until the precipitate first formed is barely dissolved, forming solutions of *ammonio-nitrate of silver* and *ammonio-sulphate of copper*, which are filtered and set aside as test reagents.

The Plating (*Reinsch's*) *Test.* Place a thin piece of pure copper in the solution acidulated with HCl, and boil. If arsenic be present, it will be deposited as a metallic film on the copper. If the solution be then poured off, and the piece of copper, carefully dried, be heated in a dry test-tube, the film will sublime and condense on the sides of the tube, and the preceding tests may be applied.

The Hydrogen (*Marsh's*) *Test* depends on the fact that AsH_3 is always formed whenever hydrogen is generated in the presence of any arsenical compound. Generate hydrogen in the usual way ($Zn+H_2SO_4$), and if the chemicals are pure (free from arsenic), the gas burns with a pale yellowish flame, without odor, and does not stain a porcelain dish held in the flame. Then pour into the generator some of the suspected solution. If arsenic be present, there is an odor of garlic; the flame becomes bluish-white, and a cold porcelain dish held in the jet so chills the flame that only the H burns, and the As is deposited on the porcelain as a brilliant metallic film. If the delivery-tube be heated, the passing AsH_3 is decomposed, and metallic arsenic deposits in a film of the same character as that on the porcelain.


Fig. 42.

This may be distinguished from the film formed by antimony under similar circumstances by (1) its greater metallic luster, and (2) by its dissolving on the addition of chlorinated soda (Labarraque's solution); (3) moisten the spot with nitric acid; evaporate the acid; a white stain is left, which is colored a red by $AgNO_3$ and yellow by H_2S. The flame should now be extinguished and the delivery-tube made to dip into a solution of $AgNO_3$. This will be blackened, and, if overlaid with aqua ammoniæ, a yellow precipitate will appear at the junction of the two fluids.

ANTIMONY (stibium) *occurs* native, but usually as a sulphide. *Prepared* by roasting the sulphide, and heating the oxide thus obtained with charcoal.

Properties. A bluish-white, brittle, crystalline solid, with a brilliant metallic luster. Resembles metals and forms alloys. In chemical properties it plays the *rôle* of positive and negative radical with equal facility.

Used in alloys, as type metal, Babbitt's metal, britannia, etc., to which it gives hardness and causes them to expand and fill the molds on solidifying. The metal is not used in medicine, most of the compounds being obtained from the sulphide.

Antimonious Hydride — SbH_3 (*Antimoniuretted Hydrogen—Stibine*), corresponding to AsH_3. This gas is formed wherever hydrogen is generated (nascent) in presence of a reducible antimony compound.

Antimonious Chloride — $SbCl_3$. At ordinary temperatures a yellow semi-solid; hence called *butter of antimony*. On addition of considerable water it decomposes, precipitating a white powder, the *oxychloride* ($SbO.Cl$),[*] formerly called *powder of algaroth.*

Oxides and Acids of Antimony.

$Sb_2O_3+H_2O=2HSbO_2$—(meta) Antimonious acid.

$Sb_2O_5+H_2O=2HSbO_3$—(meta) Antimonic acid.

Antimonious Oxide—Sb_2O_3. *Prepared* by treating the oxychloride with sodium carbonate to remove the chlorine. A whitish, insoluble, volatilizable powder.

Antimony and Potassium Tartrate — $(SbO)K\bar{T}$ [†] (*Tartar Emetic*). *Prepared* by treating Sb_2O_3 with the bitartrate of potassium, thus:

$$2KH\bar{T}+Sb_2O_3=2(SbO)K\bar{T}+H_2O.$$

Sweetish, metallic taste; soluble in water and slightly so in alcohol. Dissolved in Sherry wine it forms *vinum antimonii*, U. S. P. It enters also into *unguentum antimonii* and *syrupus scillœ compositus*, U. S. P.

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[*] SbO and BiO, called respectively *antimonyl* and *bismuthyl*, are univalent radicals, because two valences of the trivalent element being satisfied by the bivalent O, only one free valence is left.

[†] ($\bar{T}$) is used to represent the tartaric acidulous radical ($C_4H_4O_6$).

Antimonious Sulphide — Sb_2S_3, the principal ore of antimony; *occurs* native in black, lustrous masses. It may be precipitated from any antimonial solution by H_2S as an orange powder, which is black when thoroughly dried.

Poisoning by antimony occurs oftenest with tartar emetic, for that salt is used more than all the other compounds of antimony. The symptoms are those referable to gastro-enteric irritation. Fortunately the salts of antimony are emetic, and cause spontaneous evacuation of the stomach. Encourage this, and give tannic acid or ferric hydrate, which will form an insoluble compound.

The presence of antimony may be detected by (1) orange precipitate with H_2S; (2) by Marsh's test (see page 46).

BISMUTH *occurs* native and as sulphide. *Prepared* by roasting the sulphide in air, and reducing the resulting oxide with charcoal.

Properties. A brittle, white metal, with a bronze tint; volatilizes at a white heat. Forms compounds closely analogous to those of Sb, but is more positive, and plays the negative *rôle* with less facility.

Used in alloys; *e. g.* pewter and stereotyping metal; the latter melts in boiling water.

Bismuth Nitrate — $Bi\,3NO_3$. Formed by treating bismuth with nitric acid. Dissolves in a little water, but if much water be added it decomposes, with precipitation of

Bismuth Subnitrate — $BiO\,NO_3$ (*Bismuth Oxynitrate*) — a white, tasteless powder, much used in medicine and as a cosmetic (pearl white).

Bismuth Subcarbonate — $(BiO)_2CO_3$. Similar to the preceding in constitution, properties, and uses.

Bismuth and Ammonium Citrate. Obtained in pearly scales by dissolving the citrate in dilute ammonia-water, evaporating to a syrupy consistence and spreading on glass to dry. Being very soluble it is the preparation used in making the popular elixirs of bismuth.

Physiological. The bismuth salts are tonic, sedative, mildly astringent and antifermentative. Used to allay gastro-intestinal irritation. Occasionally the irritation is increased from presence of arsenic which unscrupulous manufacturers often fail to remove as the Pharmacopœia directs. When preparations of bismuth are taken, the stools are blackened by the sulphide formed with the H_2S in the intestines. In severe cases of diarrhœ, with acid fermentation, this blackening does not occur, and its reappearance is a sign of improvement.

Tests. (1) H_2S or NH_4HS gives brownish-black precipitate; (2) the concentrated solution poured into water forms a white precipitate.

V. Boron Group.

BORON	B	11

Occurs as a constituent of boracic acid and borax (sodium borate). Has two allotropic forms, amorphous and crystalline, corresponding to coal and diamond. Forms only one oxide (B_2O_3), which, combining with water, forms an acid:

$$B_2O_3+3H_2O=H_6B_2O_6=2H_3BO_3\text{—Boric acid.}$$

BORIC OR BORACIC ACID *occurs* as pearly scales, soluble in water; feeble acid; an unirritating antiseptic. Boiled with glycerine it forms *boroglyceride*, a soluble, neutral, tasteless liquid used to preserve foods.

Test. Compounds of boron, especially when moistened with sulphuric acid, color the flame green.

VI. Carbon Group.

CARBON (*carbo*, a coal)	C	12
SILICON (*silex*, a flint)	Si	28

Each element is quadrivalent, and occurs in three allatropic forms. The saturated oxide of each forms with water a dibasic acid:

$$CO_2+H_2O=H_2CO_3\text{, Carbonic acid.}$$
$$SiO_2+H_2O=H_2SiO_3\text{, Silicic acid.}$$

CARBON *occurs free* in its three allotropic forms, diamond, graphite, and coal; *combined* in carbonates and in all animal and vegetable substances. All its forms are probably traceable to organized life.

Diamond. Geological history unknown; transparent crystalline body of great brilliancy; hardest substance known. Used as a gem and for cutting glass, etc.

Graphite (to write). Owing to its resemblance to lead it has been called *black lead* or *plumbago;* almost pure carbon. Used for pencils, crucibles, stove polish, etc.

COAL. *Mineral coal* is a black substance, compact in texture, the remains of vegetable life of past ages. *Charcoal* is obtained by burning heaps of wood (Fig. 43) with a limited supply of air. The volatile constituents pass off, leaving the carbon as a light, porous substance, retaining the form and structure of the wood. *Animal charcoal* is made by heating animal matters in closed iron retorts. Charcoal, especially animal, is a valuable absorbent of odorous gases* and coloring matters.†

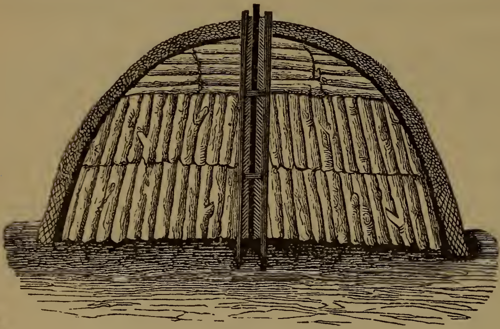
Fig. 43.

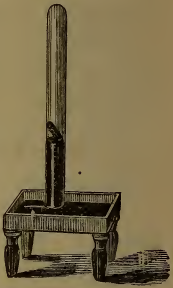
Fig. 44.

Properties. Free carbon is solid at all temperatures, and insoluble in all menstrua. Ordinarily, free carbon is unaffected by chemical agents, but at high temperatures it surpasses all other elements in its avidity for O. Hence it is used to separate the metals from their oxides.‡


Fig. 45.

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* **Experiment.** Fill a test-tube with ammonia gas over mercury. (Fig. 44.) Introduce a piece of charcoal recently heated. The gas is absorbed as is shown by the rapid rise of the mercury.

† **Experiment.** To a solution of indigo, cochineal, or potassium permanganate or beer in a flask, add some animal charcoal, shake up and filter. The filtrate is colorless, and in case beer is used it has also lost its bitter taste. (Fig. 45.)

‡ **Experiment.** Into a slight depression in a piece of charcoal lay some metallic oxide—*e. g.* lead oxide—heat with a blow-pipe. The oxide is reduced by the heated charcoal around it, and globules of the metal appear which coalesce into a bright button.

Carbon Monoxide—CO—*occurs* whenever carbon is burned with an insufficient supply of air, as in anthracite stoves and furnaces, and in coal gas, but never occurs in nature.

Prepared in the laboratory by heating oxalic acid, or potassium ferrocyanide, with sulphuric acid.

Properties. Colorless, odorless, tasteless gas; burns with a pale blue flame; very poisonous, combining with the coloring matter of the blood corpuscles, and destroying their oxygen-carrying power. Artificial respiration is of little use. Transfusion of blood is the most promising treatment. After death the blood remains scarlet. The sources of danger are open charcoal fires, defective draught in stoves and chimneys, and illuminating gas escaping into bed-rooms.

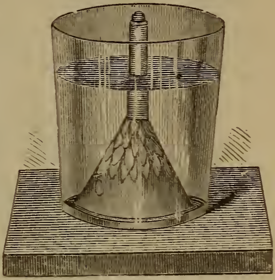

Fig. 46.

Carbon Dioxide—CO_2.

$$CO_2+H_2O+H_2CO_3\text{—Carbonic acid.}$$

Occurs sparingly (.0004) in the atmosphere, as a result of animal respiration, vegetable decay, and combustion. Plants absorb it, appropriating the carbon and returning the oxygen to the air.*

It often accumulates in cellars, beer-vats, wells, etc., where it is called choke-damp.

Prepared by burning carbon, but most conveniently in the laboratory by decomposing a carbonate with an acid.

$$CaCO_3+2HCl=CaCl_2+H_2O+CO_2.$$

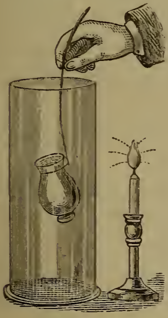

Fig. 47.

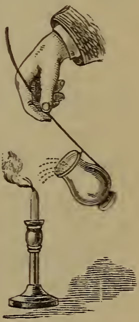

Fig. 48.

Physical Properties. Transparent, colorless gas, of a pungent

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* **Experiment.** Stick several fresh sprigs of mint into a funnel sealed up at the little end. Fill with carbonated water, invert in a glass of the same (Fig. 46), and place in sunlight. Soon minute bubbles will appear on the under surfaces of the leaves and, coalescing into larger bubbles, will rise and collect in the neck of the funnel. When in sufficient quantity, a lighted taper will show it to be oxygen.

odor and sour taste. One and a half times as heavy as air.* Water dissolves its own volume. Soda-water is only a solution of this gas under pressure.†

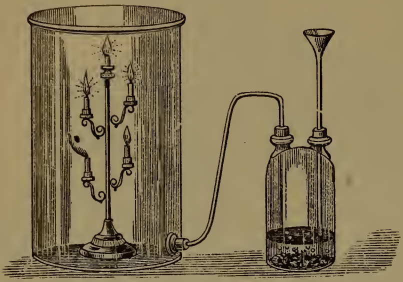

Fig. 49.

Chemical Properties. Neither burns nor supports combustion.‡ In water it exists as carbonic acid—H_2CO_3. On attempting to concentrate this dilute solution the acid decomposes again into water and CO_2; hence wet litmus reddened by it becomes blue again on drying.

Tests. (1) The gas (15 per cent and over) extinguishes a flame (Fig. 49); (2) precipitates lime-water (Fig. 51); (3) carbonates effervesce on adding a strong acid. (Fig. 50.)

Fig. 50.

Physiological. If the gas be undiluted death is immediate from spasm of the glottis. If somewhat dilute (15 to 30 per cent) there is loss of muscular power, anæsthesia, and death without a struggle. If quite dilute (5 to 10 per cent) headache, giddiness, muscular weakness, and sometimes vomiting and convulsions occur.

The effects are more serious if the CO_2 comes from combustion or respiration, because of the removal of oxygen and the admixture of the deadly CO and animal exhalations.

Treatment. Fresh air, artificial respiration, and stimulation. The preventive is ventilation.

VENTILATION. More than 7 parts of CO_2 in 10,000 of air is

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[*] **Experiments.** To show the weight of carbon dioxide: (1) Pour it from one vessel to another. (2) Blow soap bubbles and allow them to fall into a wide vessel containing this gas. As soon as they reach the surface of the gas they stop and float upon it. (3) Pour a large beakerful of the gas into a light pasteboard box that has been balanced on a pair of scales. The box will at once descend. This gas accumulating in wells may be bailed in buckets, and tested by being poured upon a lighted candle. This is illustrated in Figs. 47 and 48.

[†] That water will dissolve a greater quantity of carbon dioxide under pressure is shown by the rapid evolution of the gas whenever a bottle of soda or other carbonated water is opened and the pressure thereby removed.

[‡] Set a candlestick, holding several lighted tapers at different heights, in a large jar. (Fig. 49.) Carbon dioxide is introduced at the bottom, and extinguishes the tapers one by one as the vessel fills up to their levels.

oppressive. Taking this as the maximum impurity allowable, 3,000 cubic feet of fresh air per hour is needed by each person, and more in case of disease or when lamps are burning. To secure this in a room containing 1,000 cubic feet (10x10x10) the air must be changed three times an hour. This would give a draught not uncomfortable or injurious. If the draught be properly distributed, a breathing space of 500 cubic feet changing six times an hour would be unobjectionable. Ventilation may be secured in two ways, by diffusion and by draught.

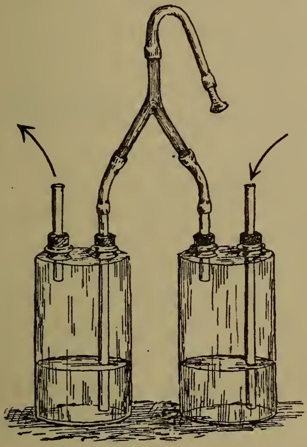

Fig. 51.*

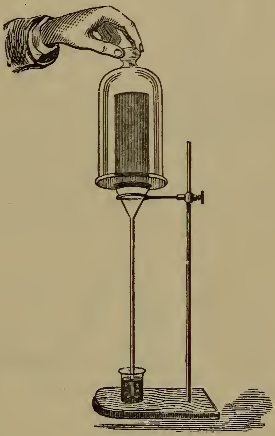

Fig. 52.

Diffusion. Gases mingle more rapidly, liquids more slowly, to make a mixture of uniform density.

When two gases of different densities are separated by a porous partition, they mingle, the lighter passing through more rapidly than the heavier, the rapidity being in inverse ratio to the square roots of their densities.†

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* **Experiment.** Two Wolff bottles are half filled with lime-water and arranged as in Fig. 51. Placing the rubber tube in his mouth, the operator can inspire through one bottle and expire through the other. The small amount of carbon dioxide in the inspired and the larger amount in the expired air is shown by a white precipitate, slight in the one and dense in the other bottle.

† **Experiment.** Cement a porous earthenware battery cup at its open end to the top of a funnel tube, the end of which dips into a beaker of colored water. Support on a stand, as in Fig. 52. Bring down over the cup an inverted bell jar of hydrogen. The light H diffuses so much faster into the cup than

This diffusion is more active in winter than in summer, because of the greater difference in density of the warm air within the house and the cold air without. Damp walls are unhealthful mainly because being no longer porous they prevent this diffusion.

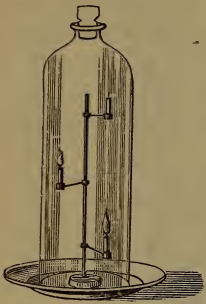
Fig. 53.

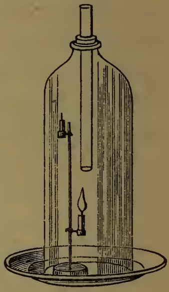
Fig. 54.

Draught is caused by the warm air expanding and rising and cold air rushing in to take its place. The gaseous products of combustion and respiration reach the ceiling before there is time for perfect diffusion, and accumulate there, making this the foulest part of the room. (Fig. 53.) This accumulation extends down to the level of the draught exit. (Fig. 54.) Hence the exit should be near the top of the room.

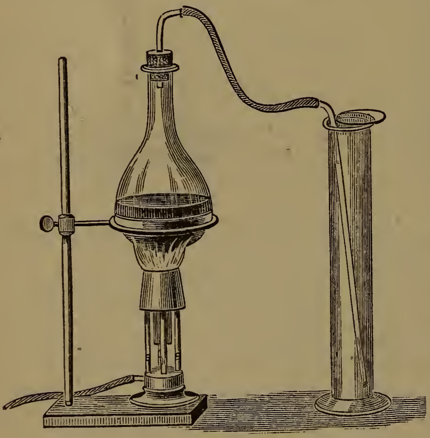
Fig. 55. Preparation of Cyanogen.

CYANOGEN—CN or Cy. Univalent because N^{III} can satisfy only three valences of C^{IV}. A compound negative radical resembling in its chemical behavior the elements of the chlorine group.

Prepared by strongly heating mercuric cyanide.*

$$Hg(CN)_2 = Hg + 2CN.$$

A colorless gas, smelling like peach kernels. Burns with a peach-blossom flame; unites with metals to form cyanides, the most important being,

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the air diffuses out of it, that bubbles of gas escape rapidly through the water. Remove the bell jar and the conditions are reversed. The H now diffuses so rapidly out of the cup that the water is sucked up the tube.

*If mercuric cyanide can not be obtained, a mixture of two parts of thoroughly dried potassium ferrocyanide and three parts mercuric chloride may be used. (Fig. 55.)

Hydrocyanic Acid—H(CN), or HCy—(*Prussic Acid, Hydrogen Cyanide*). *Occurs* in bitter almonds, cherry-laurel water, etc.

Properties. Colorless liquid, having an odor like peach kernels. For medical purposes only a dilute (2 per cent) solution is used, and the dose is from two to five drops.

Toxicology. All the cyanides are very poisonous. One drop of the pure acid produces immediate death, and three grains of potassium cyanide kills in a few minutes. The respiratory centers are paralyzed, and the victim falls and dies in convulsions. Poisoning is liable to occur from handling the acid or the cyanides which are largely used in the arts, or from eating vegetable products, *e. g.* peach and cherry seeds containing *amygdalin*, a substance easily decomposing into prussic acid and other products. Owing to the rapid action of the poison, antidotes are usually impracticable. Use artificial respiration and stimulate. If the patient survive an hour, the prognosis is good.

Tests. (1) Its odor; (2) silver nitrate—white precipitate soluble in boiling HNO_3; (3) add ammonium sulphydrate, evaporate to dryness, and then add ferric chloride—a blood red color.

Cyanates. Cyanic acid (HCyO) and ammonium cyanate (NH_4CyO) are the most interesting. The latter on being heated forms urea.

Sulphocyanates are sulpho-salts corresponding to to the cyanates (oxy-salts), and are good illustrations of the facility with which S forms series of compounds analogous to those of O. They, especially the potassium and sodium salts, are used as test reagents.

Compound Cyanides. Cyanogen shows a great tendency to form complex radicals, especially with iron; as, *ferrocyanogen* $[Fe^{II}(CN)_6{}^{I}]^{IV}$ or $(FeCy_6)^{IV}$, and *ferricyanogen* $[Fe_2{}^{VI}(CN)_{12}{}^{I}]^{VI}$ or $(Fe_2Cy_{12})^{VI}$. These two radicals contain ferrous and ferric iron respectively, and with hydrogen form acids known as *hydroferrocyanic acid*, H_4FeCy_6 (tetrabasic), and *hydroferricyanic acid*, $H_6Fe_2CN_{12}$ or $H_6Fe_2Cy_{12}$ (hexabasic); the salts of these acids are termed ferrocyanides and ferricyanides.

Potassium Ferrocyanide—K_4FeCy_6—commonly called *yellow prussiate of potash*, and *potassium ferricyanide*—K_6FeCy_{12}—*red prussiate of potash*, are important test reagents.

The carbon compounds will be further considered under the head of Organic Chemistry.

SILICON (also called silicium) resembles carbon, and occurs in three allotropic forms corresponding to coal, graphite, and diamond; most abundant element after oxygen. It exists in only a few compounds, but they constitute the larger part of the earth's crust. Its principal compound is—

SILICIC OXIDE—SiO_2—occurring as flint, sand, rock-crystal, etc.; with water it forms silicic acid. Clay, soapstone, asbestos, etc. are native silicates. Melted with the carbonates of the alkalies and alkaline earths it forms glass. If the alkali be in excess the product is soluble glass, which is much used in making surgical dressings.

THE METALS.

Occurrence. Some, as gold and copper, occur free, but most of them are found combined with non-metallic elements, especially sulphur and oxygen.

Preparation. If combined with sulphur the ore is roasted until the sulphur is burned out, leaving the metal as an oxide, which is then heated with carbon to remove the oxygen, thus:

$$ZnS+O_3=ZnO+SO_2;\text{ then, }ZnO+C=CO+Zn.$$

Physical Properties. Very opaque, with a "metallic luster" (in fine powder, a dull black); bluish gray, varying between the pure white of silver and the dull blue of lead. Yellow gold and red copper are exceptions. In weight, varying between lithium, specific gravity 0.58, and platinum, specific gravity 21.50. Most are solid, except mercury (liquid) and hydrogen (gaseous). All are absolutely insoluble.

Chemical Properties. Electro-positive, possessing great affinity for the non-metals and other electro-negative radicals. When two metals are fused together the product is an *alloy.* If one of the metals be mercury, it is called an *amalgam.* Alloys are not chemical compounds, but mixtures, for the metals do not unite in definite proportions, and the alloy is not a new substance, but one with properties intermediate between those of its constituent metals.

Used mostly in the arts. Of the fifty-five metals only about twenty-six, or rather compounds of these, enter the materia medica, and merit our notice.

I. Metals of the Alkalies.

CÆSIUM	Cs	133.
RUBIDIUM	Rb	85.4
POTASSIUM (Kalium)	K	39.1
SODIUM (Natrium)	Na	23
AMMONIUM	(NH_4)	18
LITHIUM	Li	7

Univalent; very electro-positive, so that when combined, unless it be with a strongly electro-negative (acidulous) radical, they form very alkaline compounds (hence the name). The positive affinities, as in the other groups, increase with the atomic weights. All their compounds are soluble.

CÆSIUM AND RUBIDIUM. Rare metals, occurring in small quantities with potassium. Discovered in 1860 by means of the spectroscope, and named from the colors of their lines in the spectrum (*cæsius*, sky blue, and *rubidus*, dark red). Of no medical interest as yet.

POTASSIUM *occurs* only in compounds. *Prepared* by heating one of its oxygen compounds with charcoal in an iron retort (Fig. 56) ($K_2CO_3+2C=3CO+K_2$). The metallic K distills over and condenses in the flat receiver shown in the figure.

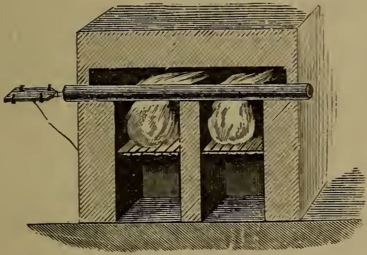

Fig. 56.

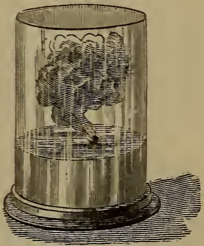

Fig. 57.

Physical Properties. Soft as wax; lighter than water; silvery luster when freshly cut, but quickly tarnishes.

Chemical Properties. Intensely electro-positive; hence it possesses great affinity for the non-metals;* takes O from H_2O, even

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* **Experiment.** Potassium inflames spontaneously when lowered into a jar of chlorine. Warmed with iodine or dropped into bromine it explodes violently. This should be done under a tubulated bell jar, because the potassium is scattered in every direction

as ice,* setting fire to the escaping H, giving the flame the violet color characteristic of K. (Fig. 57.)

POTASSIUM CARBONATE—K_2CO_3. Obtained as an impure solution ("lye") by lixiviating the ashes of plants, especially forest trees. This, evaporated to dryness, forms "concentrated lye." This in turn, when purified, forms "pearl-ash," which is further purified for medicinal use. Sometimes made by burning cream of tartar and lixiviating the residue, hence called *salts of tartar.* A white semi-crystalline or granular powder. CO_3 being a weakly acidulous radical K_2CO_3 is very alkaline, even caustic.

ACID SALTS. Salts are formed by a metallic radical displacing the basic H of an acid. If all the H be displaced, the result is a *normal salt,* as, $H_2SO_4+K_2=K_2SO_4+H_2$. But if part of the basic H of the acid remains, it is called an *acid salt,* as, $H_2SO_4+K=KHSO_4+H$. Sometimes acid salts are called "bi" salts, because the proportion of the acidulous radical to the basylous is twice as great as in the normal; *e. g.* $KHSO_4$ is called potassium bisulphate, because the proportion of the acidulous radical SO_4 to the basylous radical K is twice as great as in the normal sulphate, K_2SO_4.

POTASSIUM BICARBONATE—$KHCO_3$. Although an acid salt in constitution, it is alkaline in reaction, on account of the weakness of its acidulous radical. Made by passing CO_2 into a solution of K_2CO_3. The reaction is as follows:

$$K_2CO_3+H_2O+CO_2=2KHCO_3.$$

POTASSIUM BITARTRATE—$KH\ (C_4H_4O_6)$ or $KH\bar{T}$—*Cream of Tartar.* Prepared similarly to the above, by adding tartaric acid to a solution of the normal tartrate, thus: $K_2\bar{T}+H_2\bar{T}=2KH\bar{T}$. It exists naturally in grape juice, and, being insoluble in an alcoholic menstruum, is precipitated on the sides of the wine casks whenever vinous fermentation sets in. This is its commercial source.

OTHER SALTS. Most salts of K are made by treating the carbonate with the appropriate acid, *e. g.*:

The chloride—$K_2CO_3+2HCl=2KCl+H_2O+CO_2$.
The sulphate—$K_2CO_3+H_2SO_4=K_2SO_4+H_2O+CO_2$, etc.

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* **Experiment.** Load a strong toy cannon with gunpowder. On the fuse lay a small bit of potassium. Touching it with a piece of ice fires the cannon.

The decomposition is attended with an effervescence of CO_2. It is the formation of this volatile compound that determines the reaction. (See Volatility, page 19.)

But the following salts are not made in this way:

Potassium Hydrate—KHO—*Caustic Potash*—may be made experimentally by the reaction of metallic K on water, thus:

$$H_2O+K=KHO+H.$$

But made in the shops by boiling K_2CO_3 with slaked lime, thus:

$$K_2CO_3+Ca2HO=CaCO_3+2KHO.$$

The insoluble $CaCO_3$ (chalk) sinks to the bottom, and the KHO dissolves in the supernatant liquid, which when clear is poured off (decanted). This watery solution, if of proper strength (℥j–Oj), forms "*Liquor potassæ*, U. S. P." If this solution be evaporated to a syrupy consistence and poured into molds, it forms the stick caustic potash. KHO is very alkaline, and a powerful cautery.

Exposed to the air it absorbs CO_2 and forms carbonate:

$$2KHO+CO_2=K_2CO_3+H_2O.$$

Potassium Iodide—KI:

$$6KHO+6I=5KI+KIO_3+3H_2O.$$

The color disappears because the I goes to form colorless salts. Prepared thus, the KI is contaminated with KIO_3 (K-Iodate).* But if the mixture be strongly heated the O_3 is driven off and the KI alone remains. The addition of charcoal facilitates the removal of the O_3.

Potassium Bromide—KBr—may be made similarly to the above.

Sodio-Potassium Tartrate—$NaK\bar{T}$. *Rochelle Salt.* A neutral salt made by boiling acid potassium tartrate with sodium bicarbonate.

$$KH\bar{T}+NaHCO_3=NaK\bar{T}+H_2O+CO_2.$$

This is the reaction that occurs in baking with cream of tartar baking powders.

Potassium Hypochlorite—KClO. Made by passing chlorine into a cold solution of KHO. Yields free chlorine. The ordi-

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***Experiment.** The presence of KIO_3 in a commercial specimen of KI may be recognized by boiling a little starch in a test-tube, dissolving a crystal of the suspected salt, and then adding a few drops of a strong solution of tartaric acid; if KIO_3 be present, I will be liberated, and a blue color struck with the starch.

nary bleaching solutions (Labarraque's Solution or Javelle water) are solutions of impure sodium or potassium hypochlorite.

Tests for Potassium. (1) If the suspected solution be concentrated, add $H_2\bar{T}$ and get a precipitate of $KH\bar{T}$.* (2) Platinic chloride ($PtCl_4$) gives a yellowish precipitate. But the $PtCl_4$ is very costly, and all the potassium compounds so soluble that the above tests are but little used. The most convenient is the (3) flame test: dip the end of a clean platinum wire in the suspected solution, and hold in the colorless Bunsen or alcohol flame and notice the violet color.

SODIUM *occurs* very abundantly as sodium chloride, or common salt, from which almost all the sodium compounds are now obtained instead of from ashes of seaweeds, as formerly. The *preparation* and *properties* of sodium and its compounds are so similar to those of potassium that we will omit their separate consideration. So much alike are the salts of the two metals that the choice between them is usually governed by considerations of economy, convenience, solubility, fashion, etc. On exposure to the atmosphere the sodium salts usually have a tendency to *effloresce*, while the potassium salts tend to *deliquesce.*

Tests. No good precipitant; for all the compounds of sodium are soluble. However, the strong yellow color it gives a flame is a very delicate test; in fact, too delicate, for it shows traces of sodium in almost every thing.

AMMONIUM. When ammonia gas (NH_3) combines with an acid, it appropriates the basic hydrogen and forms a salt in which NH_4 is the positive radical; *e. g.:*

$NH_3+HCl=NH_4Cl$, corresponding to KCl or $NaCl$;
$NH_3+HHO=NH_4HO$, corresponding to KHO or $NaHO$;
$NH_3+HNO_3=NH_4NO_3$, corresponding to KNO_3 or $NaNO_3$;
$2NH_3+H_2SO_4=(NH_4)_2SO_4$, corresponding to K_2SO_4 or Na_2SO_4.

This radical plays the *rôle* of a metal, like K and Na, and is called *Ammonium.* Does not exist uncombined, although Weyl claims to isolate it as a dark-blue liquid metal.† We can obtain

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* The addition of alcohol makes the test much more delicate.

† **Experiment.** *The supposed free ammonium.* Sodio-ammonium is prepared by heating sodium in a sealed tube with ammonia gas. This is in turn heated with ammonium chloride in a sealed tube. A dark blue liquid, with metallic luster, is obtained, but soon decomposes into ammonia gas and hydrogen.

it as an amalgam by the reaction between sodium amalgam and ammonium chloride.*

AMMONIUM HYDRATE—NH_4HO—*Caustic Ammonia*—is formed in solution whenever ammonia gas (NH_3) dissolves in water, thus: $NH_3+H_2O=NH_4HO$. It has been already stated that the watery solution of a fixed substance is called a *liquor;* of a volatile substance, an *aqua.* In like manner alcoholic solutions of fixed substances are called *tinctures,* and of volatile, *spirits.* There are four U. S. P. solutions of ammonia:

Aqua ammoniæ	10 *per cent.*
Aqua ammoniæ fortior	26 "
Spiritus ammoniæ	10 "
Spiritus ammoniæ aromaticus.	

In all these solutions NH_4HO exists, but has never been isolated, because whenever we attempt to evaporate the water or alcohol the NH_4HO decomposes into NH_3+H_2O. Ammonium hydrate is very alkaline.

AMMONIUM HYDROSULPHIDE—NH_4HS—*occurs* in decomposing nitrogenous, sulphurized organic bodies. *Made* by saturating a solution of NH_4HO with H_2S. A yellowish solution; used as a test reagent.

AMMONIUM CARBONATE—$(NH_4)_2CO_3$. *Ammonii Carbonas,* U. S. P.—*Sal volatile*—is prepared by heating a mixture of NH_4Cl and chalk ($CaCO_3$) up to the temperature at which $(NH_4)_2CO_3$ would be volatilized, when the following reaction will occur:

$$2NH_4Cl+CaCO_3=CaCl_2+(NH_4)_2CO_3.$$

(See Volatility, page 19.) Very prone to absorb CO_2 from the atmosphere and become bicarbonate unless NH_4HO be added.

OTHER SALTS may be made by adding the appropriate acid to the carbonate or hydrate of ammonium. If we use the carbonate, we can tell when acid enough has been added by the cessation of effervescence. If the hydrate be used there is no effervescence, and our only guide is the point at which the solution becomes

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***Experiment.** To some mercury in a test-tube add sodium, small bits at a time. On this sodium amalgam pour a strong solution of ammonium chloride. Sodium chloride and ammonium amalgam are formed.

$$(Na+Hg)+NH_4Cl=NaCl+(NH_4+Hg).$$

The ammonium amalgam swells up and soon decomposes—$(NH_4+Hg)=NH_3+H+Hg$—the gaseous NH_3 and hydrogen escape, and only the mercury remains.

neutral in reaction. This is determined by the use of test papers. These are made of white, unsized paper, steeped in a blue vegetable pigment called *litmus*, which is *reddened* by *acids* and *restored* to its *blue* by *alkalies.*

Physiological. The hydrate and carbonate are alkaline irritants, like the corresponding K and Na compounds, though in less degree. They, moreover, give off NH_3, which, though irritating to the respiratory tract, is a valuable stimulant in fainting fits, etc. Two drams of aqua ammoniæ have killed. The treatment, as *for all alkalies,* is to give a dilute acid or some oil.

Tests. (1) An ammonium salt warmed with liq. potassæ gives off NH_3, recognized (*a*) by its odor, (*b*) its forming a white cloud of NH_4Cl when a glass rod dipped in the HCl is held over the vessel, and (*c*) its changing moist red litmus to blue. (2) Heat the dry ammonium salt and it volatilizes.

LITHIUM. Sparingly but widely distributed in nature, especially in the waters of certain springs. Lightest of the solid elements. Its salts closely resemble those of sodium.

Physiological. Lithium urate being by far the most soluble compound of uric acid, salts of lithium, especially the carbonate, are given to gouty persons to promote the elimination of uric acid, which accumulates in that disease.

Test. It colors the flame a beautiful carmine red.

Analytical. To determine whether a salt be a compound of K, Na, NH_4, or Li, heat samples of each; the one that volatilizes is the salt of NH_4. Confirm this by boiling with KHO and getting the odor of ammonia. To the other three salts apply the flame tests, getting the violet for K, yellow for Na, and carmine for Li.*

II. **Metals of the Alkaline Earths.**

BARIUM	Ba	137
STRONTIUM	Sr	87.6
CALCIUM	Ca	40
MAGNESIUM	Mg	24

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*The delicate violet of K may be masked by the intense yellow of Na, but can be seen if observed through a piece of blue glass, a medium that absorbs the yellow light.

Bivalent; their oxides and hydrates are very *alkaline*, but of an *earthy* character. Their positiveness or basicity, as in other groups, is in the order of the atomic weights. Their carbonates are decomposable by heat and insoluble in water, unless it contains H_2CO_3 in solution. Their sulphates increase in solubility from the insoluble barium salt to the very soluble magnesium sulphate.

BARIUM.—Of little interest to the medical student, except that its compounds are poisonous. Barium sulphate is very insoluble; hence (1) the antidote of barium is some soluble sulphate, and (2) barium solutions (nitrate and chloride) are delicate tests for sulphates, and *vice versa*. (See Insolubility, page 19.) Barium gives the flame a green color; hence used (nitrate) in pyrotechny to make the green or Bengal light.*

STRONTIUM. Of little importance, except that its nitrate is used in pyrotechny to make the red light.†

CALCIUM. Never free, but its compounds are very abundant, as limestone, gypsum, etc. Calcium salts are necessary to animal life, the teeth and bones consisting mainly of calcium phosphate.

CALCIUM CHLORIDE—$CaCl_2$.

Made: $CaCO_3+2HCl=CaCl_2+H_2O+CO_2$.

A white salt; very avid of water and deliquescent; used to dry gases.

CALCIUM CARBONATE—$CaCO_3$. Abundant as limestone, marble, corals, chalk, and shells of the crustacea, mollusks, etc. Chalk consists of microscopic shells. *Precipitated chalk* is made by adding a soluble carbonate to a soluble calcium salt, as:

$$Na_2CO_3+CaCl_2=2NaCl+CaCO_3.$$

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* *Green Fire:* Barium nitrate, 450 grains; sulphur, 150 grains; potassium chlorate, 100 grains, and lampblack, 25 grains.

† *Red Fire:* Strontium nitrate, 800 grains; sulphur, 225 grains; potassium chlorate, 200 grains, and lampblack, 50 grains.

For lecture-room experiments the following, without sulphur, are preferable:

Green Fire: Two parts barium nitrate, two parts potassium chlorate, and one part ground shellac.

Red Fire: Two parts strontium nitrate, two parts potassium chlorate, and one part ground shellac.

The ingredients should be *dry*, powdered separately, and mixed with little friction as possible.

The precipitate ($CaCO_3$) may be separated from the $CaCl_2$ in solution, by—

(*a*) *Filtration*—Pouring the mixture into a cone of filter paper placed in a funnel, when the water with the dissolved salt will pass through, leaving the insoluble portion (the precipitate) on the filter. (*b*) *Decantation*—Allowing the precipitate to settle to the bottom, and pouring off the clear fluid. In either case the precipitate may be washed from any remaining $CaCl_2$ by adding pure water and repeating the process.

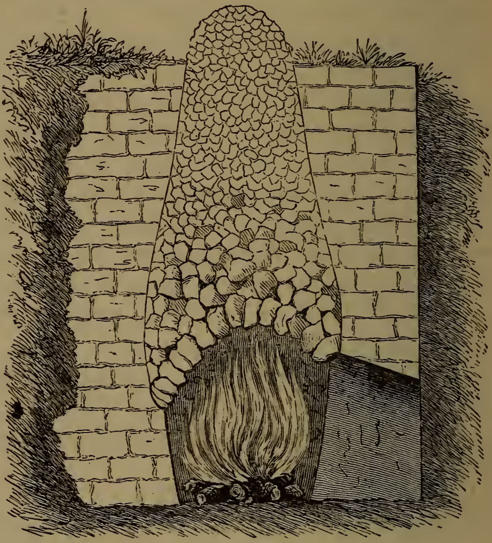

Fig. 58.

CALCIUM OXIDE—CaO—*Lime, quick lime; calx*, U. S. P. A white solid; made by heating limestone in rude furnaces called kilns. (Fig. 58.)

$$CaCO_3 = CaO + CO_2.$$

When water is added to CaO there is a violent chemical union, great heat is evolved, and a hydrate is formed, thus:

$$CaO + H_2O = Ca2HO.$$

CALCIUM HYDRATE—Ca2HO—*Slaked lime.* A white, odorless powder; very slightly soluble in water, less than one grain to the ounce, but enough to give "lime-water" (*liquor calcis*, U. S. P) a decidedly alkaline taste and reaction. The presence of sugar greatly increases its solubility (*liq. calcis saccharatus*, Br.).

CHLORINATED LIME—*Chloride of lime, bleaching powder, calx chlorata*, U. S. P.—is a mixture of chloride of calcium ($CaCl_2$) and calcium hypochlorite (Ca2ClO). It is made by passing chlorine gas over slaked lime until it ceases to be absorbed. It is white, moistens on exposure to the air, absorbing CO_2 and giving off Cl.

It is employed as a source from which to get a gradual supply of chlorine for disinfecting and bleaching purposes.

Calcium Sulphate — $CaSO_4$ — *occurs* native, as gypsum, which, when heated, loses its water of crystallization and forms a white amorphous powder called *plaster of Paris.* If this plaster be mixed with water enough to form a creamy liquid, it will recrystallize or "set" into a hard compact mass. Much used in surgery to make casts to hold broken limbs in position. Very slightly soluble in water.

Calcium Phosphate—$Ca_3(PO_4)_2$. It is the most abundant mineral ingredient of the body; in every tissue and fluid, especially the teeth and bones, to which it gives hardness and rigidity. A white tasteless powder, soluble in dilute acids. Dissolved by lactic acid, it is given as *syrupus calcii lactophosphatis,* U. S. P., in scrofula, rickets, and other diseases of defective nutrition.

Calcium Oxalate—CaC_2O_4, or $CaO\bar{x}$ — *occurs* in the juices of some plants and in the urine. Obtained as a fine white crystalline powder when a soluble oxalate is added to a calcium solution. Insoluble in water or acetic acid, but soluble in the mineral acids.

Hard Waters are such as contain mineral matters, especially calcium (lime) compounds. Often water, in passing through the soil, becomes highly charged with carbonic acid, and dissolves considerable amounts of $CaCO_3$, and is hard. This is called *temporary hardness,* because on exposure or boiling, the carbonic acid is driven off, the $CaCO_3$ is precipitated, and the water becomes *soft.* The solubility of $CaSO_4$ does not depend on the presence of carbonic acid, and boiling will not precipitate it. So water impregnated with $CaSO_4$ is said to be *permanently hard.* Drinking-water should contain a small quantity of lime; but very hard water impairs digestion. Hard water is unfit for washing, because the soluble alkali soap reacts with the lime salt to form an insoluble lime-soap.

MAGNESIUM. Never free; abundant in magnesian limestone ($CaCO_3MgCO_3$). Asbestos, meerschaum, and soapstone are native silicates. Most natural waters contain its salts. Silvery white metal; burns with a brilliant white light, rich in chemical rays, and used in photographing caves and other dark places.

MAGNESIUM SULPHATE—$MgSO_4$—*occurs* in the waters of various springs, as those at Epsom; hence often called *Epsom salts.* Made artificially from the native carbonate, thus:

$$MgCO_3+H_2SO_4=MgSO_4+(H_2O+CO_2).$$

White, crystalline, soluble salt of a nauseous bitter taste. It is a popular purgative. The nauseous taste and griping may be obviated by adding aromatics, acid, sulphate of iron (as in Crab Orchard salts), or by free dilution.

MAGNESIUM CITRATE is the most pleasant of the saline purgatives. Usually given as the *liquor magnesii citratis,* which is prepared by adding a solution of citric acid to $MgCO_3$, and bottling immediately to retain the CO_2.

MAGNESIUM CARBONATE—$MgCO_3$—*occurs* native. For medicinal purposes it is prepared by precipitation, thus:

$$MgSO_4+Na_2CO_3=Na_2SO_4+MgCO_3.$$

Similar to chalk in its physical and chemical properties.

MAGNESIUM OXIDE — MgO — *Magnesia.* Made like CaO, by heating the carbonate:

$$MgCO_3=MgO+CO_2.$$

Insoluble and tasteless (earthy), but its alkalinity is shown by its turning moist red litmus paper blue when the solid MgO is dropped upon it.

MAGNESIUM HYDRATE—$Mg(HO)_2$. Formed by precipitating a magnesium solution with potassium or sodium hydrate. Insoluble in water, but, like other salts of magnesium, soluble in the presence of ammonium compounds with which they form double salts. Suspended in water, it is called *milk of magnesia.*

MAGNESIUM PHOSPHATES. These resemble the calcium phosphates and are associated with them in the body, though in small quantity. The *ammonio-magnesium phosphate* ($MgNH_4PO_4$) is precipitated whenever a soluble phosphate in neutral or alkaline solution finds itself in presence of an ammonium salt, as occurs in the alkaline fermentation of urine.

Physiological. Magnesium oxide and hydrate being alkaline and tasteless, are popular antidotes for acids. These and the carbonate are given to correct acid conditions of the digestive tract, and combining with the acids they form soluble salts that are laxative.

Analytical. To determine whether a solution be one of barium, calcium or magnesium: Add potassium chromate; a precipitate indicates *barium.* If no precipitate, add ammonium chloride and then ammonium carbonate, a precipitate indicates *calcium.* If no precipitate, add sodium phosphate; a precipitate indicates *magnesium.*

III. Metals of the Earths.

[Boron	B	11]
Aluminium	Al	27
Scandium	Sc	44
Yttrium	Y	92
Lanthanum	La	139
Cerium	Ce	141
Didymium	D	145
Samarium	Sm	150
Erbium	E	168
Ytterbium	Yb	173

Trivalent, though in compounds two atoms go together, forming a sexivalent radical, as Al_2Cl_6. Boron is so weakly positive that it is a non-metal. (See page 49.) Aluminium is the most important member of this group, the others being rare metals associated with it in various minerals. Their oxides and hydrates are of a neutral or earthy character.

ALUMINIUM. Never found free, but in the abundance and distribution of its compounds it ranks next to oxygen and silicon—third among the elements and first among the metals. Isolated with difficulty, and therefore costly. Bluish-white metal, ductile and very light; does not tarnish in the air. With copper it forms a golden yellow alloy, known as *aluminium bronze.*

ALUMINIUM CHLORIDE—Al_2Cl_6. *Prepared* industrially in the manufacture of aluminium. A soluble, astringent salt. It absorbs and combines with H_2S, PH_3, and NH_3. An impure solution is sold as a disinfectant under the name *chloralum.*

ALUMINIUM SULPHATE—Al_23SO_4. Made by treating white clay with H_2SO_4. It has properties similar to the above.

ALUM—*Alumen.* An alum is a double sulphate of a trivalent and a univalent radical. Its constitution may be expressed thus:

$$R_2^{III}3SO_4.R_2^{I}SO_4, \text{ or } 2R^{III}R^{I}2SO_4.$$

The trivalent radical (R^{III}) may be Al, Fe, Cr, or Mn. The univalent radical (R^{I}) may be K, Na, NH_4, etc. So, by different combinations of these radicals a variety of alums may be formed. The old *potash alum* ($Al_23SO_4.K_2SO_4$) is giving place in the arts to the cheaper *ammonium alum* ($Al_23SO_4.(NH_4)_2SO_4$). The *ammonio-ferric alum* ($Fe_23SO_4.(NH_4)_2SO_4$) is also much used in medicine. Burnt alum, *alumen exsiccatum*, is a white amorphous powder obtained by heating alum until its water of crystallization is driven off. Alum, like other salts in which the acidulous radical predominates, is astringent; burnt alum, on account of its avidity for water, is a mild escharotic.

ALUMINIUM SILICATES. Very abundant, as granite, clay, sand, etc. Clay is usually of a reddish or brown color from admixture of oxides of Fe, etc. Pure white clay (kaolin) is used in the arts to make porcelain, and in medicine as a vehicle for the external application of acids, etc.

CERIUM is a rare metal. One of its salts, the oxalate, is used as a sedative to irritable stomachs, especially in the vomiting of pregnancy. When pure it is a very efficient remedy; but the commercial article is liable to contain salts of lanthanum, didymium, and other allied metals.

IV. The Zinc Group.

ZINC, Zn, 65.2, and CADMIUM, Cd, 112.

Bivalent; bluish-white metals, closely allied in sources and properties.

ZINC. When heated in air, zinc burns with an intense bluish-white light, forming clouds of oxide. It tarnishes quickly in air or water, but becomes coated with a film of oxide that protects it from further corrosion. Iron coated with zinc ("galvanized iron") will withstand exposure to the weather an indefinite time. Alloyed with copper, zinc forms brass. Pure H_2SO_4 is unaffected by pure zinc or zinc coated with mercury (amalgamated), unless it form a galvanic circuit.* Commercial zinc is rapidly attacked by most acids.

***Experiment.** Into a large test-tube containing bits of zinc pour dilute sulphuric acid; there is a prompt effervescence of hydrogen. Add a little mercury, and agitate; the action ceases. Drop in a piece of copper; it begins again.

ZINC SULPHATE—$ZnSO_4$—*White Vitriol*—is made thus:

$$Zn+H_2SO_4=ZnSO_4+H_2.$$

White, soluble salt, resembling $MgSO_4$ in appearance; astringent and emetic.

ZINC CHLORIDE—$ZnCl_2$. *Made:* $Zn+2HCl=ZnCl_2+H_2$. A white deliquescent salt; strongly astringent; severe caustic. Used as an injection to preserve anatomical subjects.

ZINC CARBONATE—$ZnCO_3$—is a white, insoluble powder made by precipitation:

$$ZnSO_4+Na_2CO_3=Na_2SO_4+ZnCO_3.$$

Used in the arts (*zinc white*) in place of lead carbonate, for it is not blackened by H_2S; in medicine as a dusting powder for excoriated surfaces, and in ointment.

ZINC OXIDE—ZnO—is prepared either by burning metallic zinc* or heating the carbonate, $ZnCO_3=ZnO+CO_2$.

It is a yellowish-white powder, used *externally* in ointment, *internally* as a tonic and astringent, especially in the night-sweats of phthisis and diarrhea of children.

ZINC SULPHIDE—ZnS—is precipitated whenever a solution of a zinc salt is added to the solution of a soluble sulphide, unless the solution be acid in reaction. It is the only white sulphide, therefore a test for zinc.

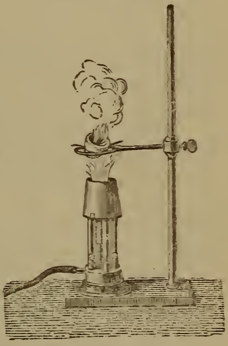

Fig. 59.

Poisoning. All the salts of zinc that are soluble in the digestive fluids act as irritant poisons. Sodium chloride and organic acids dissolve metallic zinc, therefore food kept in galvanized iron vessels is more or less poisonous, especially since commercial zinc usually contains traces of arsenic. For this reason articles intended for toxicological analysis should never be kept in jars with zinc caps.

CADMIUM resembles zinc in its properties and uses, except that its sulphide is yellow and insoluble in acid solutions.

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[*] **Experiment.** Place bits of zinc in a hessian crucible and heat strongly over a triple burner. The metal is volatilized, and the vapor igniting burns with an intense bluish-white flame, yielding white flakes of zinc oxide, the *lana philosophica* (philosopher's wool) of the older chemists. (Fig. 59.)

V. The Iron Group.

CHROMIUM	Cr	52.2
MANGANESE	Mn	55.0
IRON	Fe	56.0
COBALT	Co	58.8
NICKEL	Ni	58.8

These are hard metals and all more or less magnetic.

By a variation in valence they form two classes of compounds: One in which the atom is *bivalent*, as in ferrous chloride ($FeCl_2$); the other in which the atom is *trivalent*, as in ferric chloride (Fe_2Cl_6). With oxygen they form acidulous radicals, which form the chromates, manganates, and ferrates, with the stronger bases.

CHROMIUM. So named because all its compounds are colored. The metal is of but little use. Its compounds are of great importance to the chemist and of considerable utility in the arts, but few are used in medicine.

CHROMIUM TRIOXIDE—CrO_3—is made by treating a strong solution of potassium bichromate with sulphuric acid, thus:

$$K_2Cr_2O_7+H_2SO_4=K_2SO_4+H_2CrO_4+CrO_3.$$

The CrO_3 separates in crimson prisms. It is a powerful oxidant and a caustic. Sometimes improperly called chromic acid.

CHROMATES. The principal ones are potassium chromate, K_2CrO_4, a valuable test reagent, and lead chromate, $PbCrO_4$, a yellow pigment.

BICHROMATES are not regular acid or bi-salts, but compounds of a chromate and chromium trioxide. The most important of these is potassium bichromate, $K_2Cr_2O_7$, or $K_2CrO_4.CrO_3$. It forms large red soluble crystals. It is added to the sulphuric acid in batteries to oxidize* the nascent hydrogen.

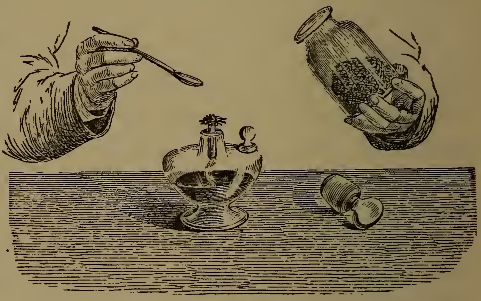

Fig. 60.

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* **Experiments.** The oxidizing action of the chromic salts can be shown in a number of reactions. (*a*) Any organic substance, as sugar, oxalic acid, or

MANGANESE resembles iron in its properties. Used to alloy iron in the preparation of certain kinds of steel. Its most abundant ore is the

MANGANESE DIOXIDE—MnO_2—*Black Oxide of Manganese*—an insoluble steel-gray powder that readily gives up its extra atom of O. Used in large quantities in the preparation of chlorine and oxygen gas.

MANGANOUS SULPHATE—$MnSO_4$.

$$MnO_2+H_2SO_4=MnSO_4+H_2O+O.$$

A soluble, rose-colored salt.

MANGANOUS SULPHIDE—MnS—is precipitated whenever a solution of a salt of manganese is treated with NH_4HS. It is the only flesh-colored sulphide, hence its formation is a test of manganese.

MANGANATES. If a mixture of KHO, $KClO_3$, and MnO_2 be heated together, there results a green mass of *potassium manganate*, K_2MnO_4. If this be dissolved in distilled water, it forms a green solution, which, on boiling, or even standing awhile, is changed to a purple, due to the formation of *potassium permanganate*, $K_2Mn_2O_8$.

The permanganate* gives up its oxygen so readily to organic matter, at the same time losing its purple color, that it is used as a test for organic impurity in water and as a disinfectant.

Physiological. Associated with iron (1 to 20), manganese is a normal constituent of the blood corpuscles; hence its preparations, like those of iron, are blood tonics. Valuable in amenorrhœa.

IRON *occurs* abundantly as oxide, carbonate, and sulphide; occasionally free.

Preparation. The carbonate or sulphide is first roasted until

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a chip of wood, boiled in the sulphuric acid and bichromate mixture, is oxidized, disappearing completely, with evolution of carbon dioxide. (*b*) Rinse out a beaker with strong alcohol, and then drop in a few crystals of chromic acid. The thin layer of alcohol is ignited with the odor of aldehyde. (*c*) Pour a few drops of absolute alcohol on the wick of a spirit lamp (Fig. 60), and lay on several crystals of chromic acid. It ignites.

[*] **Experiment.** Powdered potassium permanganate treated with sulphuric acid gives off ozone. (See page 14.) So powerful an oxidizer is this mixture that alcohol, ether, benzol, carbon disulphide, flowers of sulphur, tannin, etc., are ignited on contact with it.

converted into oxide. The oxide is heated in a blast furnace (Fig. 61) with coal and fluxes (limestone and silicates). The carbon of the coal removes the oxygen from the iron, which melts and sinks beneath the melted fluxes. The fused metal is then drawn off into furrows in the sand called *pigs*. This is *cast iron*, containing 4 or 5 per cent of carbon. *Wrought iron* contains little or no carbon, and *steel* an intermediate amount.

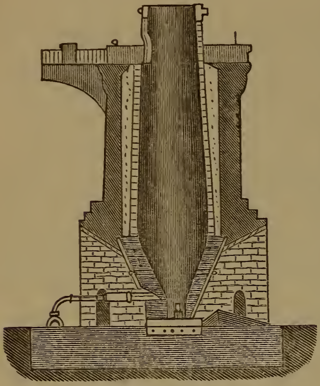

Fig. 61.

Properties. A bluish-gray metal, sp. gr. 7.5; rusts (oxidizes) when exposed to moist air, or water containing air.

Reduced Iron—*Ferrum redactum, iron by hydrogen,* Quevenne's iron. It is prepared by heating ferric oxide nearly to redness in a tube through which hydrogen is passed:

$$Fe_2O_3+H_6=Fe_2+3H_2O.$$

It is a very fine, dark gray powder;* prescribed in pill form.

CHLORIDES.

Ferrous Chloride—$FeCl_2$. Made by adding iron to hydrochloric acid until effervescence ceases, thus:

$$Fe+2HCl=FeCl_2+H_2.$$

Like most ferrous salts, it is green and prone to oxidize with the formation of ferric compounds.

Ferric Chloride—Fe_2Cl_6—is made by first forming the ferrous chloride as above, and then adding nitric and hydrochloric acids. The nascent chlorine evolved by the nitro-hydrochloric acid converts the ferrous into ferric chloride, thus:

$$6FeCl_2+6HCl+2HNO_3=3Fe_2Cl_6+N_2O_2+4H_2O.$$

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*__Experiments.__ (*a*) Reduced iron, if good and fresh, will ignite on contact with a lighted taper and burn with a red glow. (*b*) Faraday used to show that it is more inflammable than gunpowder, by pouring it mixed with gunpowder upon an alcohol flame burning on a white dinner-plate. The iron burns with bright scintillations, while the gunpowder falls through the flame and is only ignited when the flame dies down and reaches the surface of the plate. (*c*) One part of sulphur, two of reduced iron, and three of nitre make an iron gunpowder that burns as quickly and more brilliantly than ordinary gunpowder.

The *liq. ferri chloridi*, U. S. P, is the aqueous solution. This, when diluted with alcohol, forms the *tinct. ferri chloridi*, U. S. P. If citrate of potassium or sodium be added to this tincture, the solution loses its styptic taste, does not affect the teeth, and is not incompatible with solutions containing tannin.

SULPHATES.

FERROUS SULPHATE — $FeSO_4$ — *Copperas, Green Vitriol. Prepared:* $Fe+H_2SO_4=FeSO_4+H_2$. Soluble, green crystals efflorescing upon exposure. A cheap and excellent disinfectant, destroying organic matters by abstracting their oxygen. When given in pill form it is first *exsiccated.*

FERRIC SULPHATE—Fe_23SO_4—*Tersulphate* is made by adding nitro-sulphuric acid ($HNO_3+H_2SO_4$) to a solution of the ferrous sulphate, thus.

$$6FeSO_4+3H_2SO_4+2HNO_3=3Fe_23SO_4+N_2O_2+4H_2O.$$

Its officinal solution is the *liq. ferri tersulphatis. Liq. ferri subsulphatis*, U. S. P., Monsel's Solution, is prepared similarly to the above, using only half the quantity of sulphuric acid.

HYDRATES.

FERROUS HYDRATE—Fe2HO—is precipitated on mixing solutions of a hydrate and a ferrous salt, as—

$$FeSO_4+2NaHO=Na_2So_4+Fe2HO.$$

A green precipitate, which soon oxidizes and becomes brown.

FERRIC HYDRATE—Fe_26HO. A brownish red, gelatinous mass, precipitated by soluble hydrates from ferric solutions, *e. g.:*

$$Fe_2Cl_6+6NH_4HO=6NH_4Cl+Fe_26HO.$$

This is the favorite antidote for arsenic, for which purpose it must be freshly prepared and given in large doses. Ferric hydrate dissolves freely in a solution of ferric chloride, forming a dark red liquid of a styptic taste.

If this liquid be put in a *dialyser* (Fig. 62), a vessel with a bottom of parchment or animal membrane, and suspended in water, the chloride passes out through the membrane into the water. When barely enough ferric chloride remains within the dialyser to hold the ferric hydrate in solution and the styptic taste has disappeared, the liquid is removed and sold under the name of "Dialysed Iron."

FERRIC NITRATE—$Fe_2 6NO_3$.

Made: $Fe_2 6HO+6HNO_3=6H_2O+Fe_2 6NO_3$.

Liq. ferri nitratis, U. S. P., is a reddish acid liquid. Used as an astringent, especially in dysentery.

FERROUS IODIDE—FeI_2. *Prepared:* $Fe+I_2=FeI_2$.

Sometimes given in pill, but better with syrup, which acts as a preservative as well as a vehicle.

FERROUS CARBONATE—$FeCO_3$—is obtained by adding a soluble (alkaline) carbonate to a ferrous salt, thus:

$$FeSO_4+K_2CO_3=K_2SO_4+FeCO_3.$$

It is insoluble in pure water, but slightly soluble in water containing carbonic acid, as in chalybeate springs. On exposure to the air it turns red from formation of ferric hydrate; so it is preserved by mixing with sugar and honey, as in the *ferri carbonas saccharatus*, U. S. P.

Fig. 62.

FERROUS SULPHIDE—FeS—*occurs* native, but may be made by heating together iron filings and flowers of sulphur. Used in the preparation of H_2S.

SCALE COMPOUNDS OF IRON. These are ferric salts, mostly with organic acids. They do not crystallize readily, but are sold as thin scales. Made by evaporating their solutions to a syrupy consistence, poured upon plates, and when dry peeled off in scales. Often other bases, as potassium or ammonium, together with alkaloids, as quinine and strychnine, are incorporated in the compound.

The following are officinal: *Ferri citras, ferri et ammonii citras, ferri et quiniæ citras, ferri et strychniæ citras, ferri et ammonii tartras, ferri et potassii tartras*, and *ferri pyrophosphas*.

Physiological. Iron is a normal constituent of the body, especially the blood corpuscles, where it performs an important function, as is shown by the great increase of blood corpuscles and of bodily vigor attending its administration. Many of its salts, especially the ferric salts of the mineral acids, are astringent and hemostatic. Iron is eliminated by various organs, but is mainly discharged by the bowels as sulphide blackening the fæces.

Tests for Iron. *Ferrous salts* are usually *green;* with NH_4HS they give a black precipitate of FeS.

Ferric Salts are usually *red;* they give a black precipitate with NH_4HS; a black precipitate with tannic acid; and a blood-red with sulphocyanate of potassium.

COBALT. Its chief ore is a compound with arsenic, sold under the name of *cobalt* or *flystone*, for poisoning flies. Its salts are used in preparing sympathetic ink, for when free from moisture they are deep blue, but almost colorless when moist. Writing done with a dilute solution of chloride of cobalt is invisible until warmed, when it becomes blue, the color disappearing when the paper is cooled or moistened.

Test for Cobalt. It imparts a deep blue color to a bead of glass or borax melted in the blow-pipe flame.

NICKEL. This is a hard, grayish-white metal that does not tarnish in the air. Used to electro-plate instruments made of metals more prone to corrode, and to make cheap coin. Mixed with brass, it forms German silver.

VI. The Lead Group.*

TIN (*Stannum*)	Sn	118
LEAD (*Plumbum*)	Pb	207

The members of this group are *bivalent* and *quadrivalent.*

TIN. A bluish-white malleable metal, not corroded by air or water; hence used to form a protective coating for iron and copper. *Tin-ware* is usually sheet-iron coated by being dipped

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* This is really a continuation of the Carbon Group, the metallic character increasing with the atomic weights: C=12; Si=28; —— =73; Sn=118; Pb=207. The third member, corresponding to arsenic of the Nitrogen Group, is yet undiscovered.

into molten tin. Tin alloyed with lead is easily dissolved, and may cause lead-poisoning.

Tin-foil (thin laminæ of tin) is used in wrapping to exclude air and moisture. Tin enters into the composition of a great many alloys. Powdered tin is sometimes used as an anthelmintic.

Tin forms two classes of compounds; the *stannous,* in which the atom is bivalent, and *stannic,* in which the atom is quadrivalent. These are of importance to the chemist, but of little interest to the physician.

LEAD. Its principal ore is its sulphide (PbS), called *galena.* It is a soft, heavy, blue metal, very slowly acted upon by most substances; hence used to make water-pipes and vessels that are exposed to corrosive liquids.

Water containing nitrates or nitrites (from organic matter) dissolves lead slightly; but if it contains carbonates or sulphates, the lead is protected by an insoluble coating of lead carbonate or sulphate.

Lead enters into the composition of many alloys; as pewter, solder, shot, type-metal, etc. The quadrivalent compounds of lead are of so little importance that the term *plumbic* is applied to the bivalent compounds.

LEAD OXIDE—PbO—*Litharge.* A yellow substance, found native; made artificially by heating lead in the air. It is by treating this with the appropriate acid that most of the lead salts are prepared. When rubbed with oils it decomposes the glyceryllic ethers and combines with the fatty acids to form lead soaps, one of which, the oleate, is *lead plaster, emplastrum plumbi,* U. S. P.

LEAD DIOXIDE, or *puce lead,* is a dark brown powder, forming one of the constituents of *red lead* (Pb_3O_4 or $2PbO\ PbO_2$).

Prepared by treating *red lead* with nitric acid to dissolve out the PbO.

LEAD NITRATE—$Pb2NO_3$.

Made: $PbO+2HNO_3=Pb2NO_3+H_2O$.

Ledoyen's disinfectant fluid is a solution of $Pb2NO_3$ (one dram to the ounce). It corrects fetid odors by neutralizing H_2S and NH_4HS.

LEAD ACETATE—$Pb(C_2H_3O_2)_2$, or $PbAc_2$—*Sugar of Lead.*

Made: $PbO+2HAc=PbAc_2+H_2O$.

Used in medicine more than any other lead salt. Its solution will dissolve considerable quantities of PbO, forming the *solution of the subacetate of lead, the liquor plumbi subacetatis,* U. S. P., *Goulard's extract.* It is astringent, and, like all the lead salts, sedative. Much used as a topical application in erysipelas, acute eczema, and other skin affections; and diluted (*lead-water*), it is used in conjunctivitis and other mucous inflammations.

The following insoluble salts may be made by precipitation from solutions of the preceding soluble ones:

LEAD CHLORIDE—$PbCl_2$. *Made:* Soluble lead salt added to a soluble chloride; *e. g.* $PbAc_2+2HCl=PbCl_2+2HAc$. Slightly soluble in warm water, but in cold it is always precipitated from solutions of moderate strength; hence classed with HgCl and AgCl as one of the three insoluble chlorides.

LEAD SULPHATE—$PbSO_4$. Forms as a white precipitate whenever a solution of a lead salt is added to a sulphate solution, thus:

$$PbAc_2+ZnSO_4=PbSO_4+ZnAc_2.$$

LEAD CARBONATE—$PbCO_3$—*White Lead.*

Made: $PbAc_2+Na_2CO_3=PbCO_3+2NaAc$.

Commercially, it is made by some modification of the old Dutch method, which consists in covering bars of lead with the refuse of the wine-press and barn manure. The acetic fumes from the grape husks attack the lead, forming lead acetate, which is decomposed by the carbonic acid from the manure. The acetic acid thus liberated combines with another portion of lead, which is again precipitated by the carbonic acid, and thus the process continues until all the lead is consumed.

Used for painting, but blackens when air contains H_2S.

LEAD SULPHIDE—PbS—is formed as a black precipitate whenever a lead solution is treated with a soluble sulphide, as H_2S or NH_4HS.

LEAD IODIDE—PbI_2. A bright yellow precipitate on adding a soluble iodide to a lead solution; as,

$$PbAc_2+2KI=2KAc+PbI_2.$$

LEAD CHROMATE—$PbCrO_4$.

Made: $PbAc_2+K_2CrO_4=PbCrO_4+2KAc$.

Under the name of *chrome yellow* it is used in painting. Of late it has been used to color food products.

Tests for lead consist in forming precipitates of the foregoing insoluble compounds.

Physiological. All the lead compounds are poisonous. *Acute* poisoning sometimes occurs from the ingestion of a single large dose of a soluble lead salt. The symptoms are those of gastric irritation. *Treatment.* Give $MgSO_4$ to form the insoluble $PbSO_4$.

The *chronic* form of lead intoxication, *painter's colic,* is purely poisonous, and is produced by the continued absorption of minute quantities of lead by the skin of those handling it, and by the lungs and stomachs of those living in painted apartments, or using food and drink from leaden vessels. There is impairment of digestion, constipation, blue line along the edge of the gums, colic, and paralysis, especially of the extensor muscles. Lead once absorbed is eliminated very slowly, having combined with the albuminoids, a combination which is rendered soluble by the administration of iodide of potassium.

The *treatment* for chronic lead-poisoning is to give $MgSO_4$, for the double purpose of overcoming the constipation and precipitating any lead remaining unabsorbed in the alimentary canal; also KI to promote the elimination of that which is combined with the albuminoids. Alum is a favorite treatment, seeming to perform all accomplished by both the $MgSO_4$ and KI. The paralyzed muscles must be treated with electricity, so that when the lead is eliminated and the nerve influence returns, it may not find them degenerated past redemption.

VII. The Copper Group.

COPPER (*Cuprum*)	Cu	63.4
MERCURY (*Hydrargyrum*)	Hg	200

Each of these elements is univalent and bivalent, forming two classes of compounds, "*ous*" and "*ic.*" At ordinary temperatures they are acted upon but slowly by the non-oxidizing acids, as H_2SO_4 and HCl; but HNO_3 attacks them vigorously.

COPPER is usually found combined with sulphur, etc., but often in the metallic state, especially on the southern shores of Lake Superior. Being found free, it was among the first metals wrought by man, so the bronze preceded the iron age. Copper is a red malleable metal; an excellent conductor of electricity.

CUPRIC SULPHATE—$CuSO_4$—*Blue Vitriol, Blue Stone.* Obtained as an incidental product from silver refineries, copper mines, etc.; made experimentally by heating copper with strong H_2SO_4. Forms beautiful blue crystals, soluble in water, but insoluble in alcohol. If the crystals be heated they lose their water of crystallization and form a white powder, which becomes blue again upon the addition of water. Hence, used as a test for water in alcohol. Like other salts in which the acidulous radical predominates, cupric sulphate is astringent and coagulates albumen. A prompt emetic, but not used as much as $ZnSO_4$, because if, by chance, it be not all ejected from the stomach, a gastro-enteritis is liable to be set up.

CUPRIC HYDRATE—$Cu2HO$—is formed as a bluish-white precipitate whenever a soluble copper salt is treated with a soluble hydrate, thus: $CuSO_4+2KHO=K_2SO_4+Cu2HO$.

When heated, even under water, it decomposes—

$$Cu2HO=CuO+H_2O.$$

CUPRIC OXIDE—CuO—*Black Oxide. Prepared* by heating copper turnings in air. It gives up its oxygen easily, hence used as an oxidizer in organic analysis.

CUPROUS OXIDE — Cu_2O — *Suboxide. Made* by boiling the cupric oxide with an oxidizable substance, as glucose (copper tests for glucose), which is oxidized at the expense of the oxygen of the cupric oxide. The precipitate is first yellow (hydrate), but soon becomes a bright red (oxide).

CUPRIC SUBACETATE OR OXYACETATE—sometimes called *verdigris* (green-gray)—is made industrially by exposing plates of copper to the acetic fumes of grape husks, etc. It is apt to be formed whenever fruits containing acetic acid are placed in copper vessels.

TESTS. 1. *Plating Test.* Dip into the suspected solution a more electro-positive metal, as iron, and a plating of metallic copper will be deposited on the iron, an equivalent proportion of which takes the place of the copper in the solution.

2. *Sulphur Test.* Add H_2S or NH_4HS, and if copper be present a black precipitate (CuS) will be formed.

3. *Ammonia Test.* Add ammonia, and if copper be present a deep blue ammonio-salt of copper will be formed.

4. *Arsenic Test.* To the ammonio-salt, described above, add an aqueous solution of As_2O_3, and a green precipitate of *arsenite of copper* (paris green) will be thrown down.

5. *Glucose Test.* Add KHO, ($CuSO_4+2KHO=K_2SO_4+Cu2HO$) and boil ($Cu2HO=CuO+H_2O$), with a little glucose, and a yellowish-red precipitate (Cu_2O) indicates copper.

It will be seen from the two last reactions, above described, that a substance acted upon characteristically by a reagent is as good a test for the reagent as the reagent is for it—*i. e.* arsenic and glucose, being acted upon characteristically by copper, are as good tests for copper as copper is for them.

Physiological. Canned fruits, pickles, etc., that have been colored green with copper, and food, especially if acid, that has been cooked or kept in copper vessels, are apt to give an acute gastro-enteritis. Chronic copper poisoning, so called, is perhaps always due to other substances, as lead or arsenic, and should be treated accordingly.

Antidotes for acute copper poisoning: Encourage vomiting and give albumen (white of egg), which combines with the copper salt to form an insoluble albuminate; or iron filings, which will precipitate the copper in metallic state.

MERCURY is the only metal liquid at ordinary temperatures, and resembles silver in appearance, hence the names hydrargyrum (water silver) and quicksilver (fluid silver). It is so heavy that iron and stone float upon it as corks on water. (Fig. 63 represents a marble and a ball of iron floating on mercury.) It does not tarnish in the air unless contaminated with baser metals; dissolves all metals, except iron, to form amalgams.

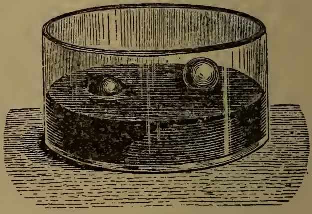

Fig. 63.

Uses. Metallic mercury is used extensively in the refining of silver and gold, in thermometers and other instruments, with tin in silvering mirrors, and in many other branches of the arts. Metallic mercury, rubbed up with various excipients until globules cease to be visible, forms several officinal preparations. Rubbed with chalk it forms "gray powder," *hydrargyrum cum creta;* with confection of roses and licorice powder it forms "blue pill," *pilula hydrargyri;* and with lard and suet it forms "mercurial ointment," *unguentum hydrargyri.* The therapeutic activity of these preparations is not due to the metallic mercury they contain, but to small quantities of mercurous oxide formed by the oxidation of the finely divided metal. So their strength varies with the thoroughness of the rubbing, the extent of the exposure, and the age of the preparation.

MERCUROUS IODIDE—HgI—*Proto-iodide, Green Iodide, Hydrargyri Iodidum Viride,* U. S. P Made by rubbing together chemical equivalents of mercury (200) and iodine (127) until they combine and form a green mass.

MERCURIC IODIDE—HgI_2—*Biniodide, Red Iodide, Hydrargyri Iodidum Rubrum.* Made like the above, except that two equivalents of iodine (2x127) are employed.

Both the iodides, being insoluble may be precipitated by adding a solution of KI to a solution of mercurous salt for the one and a mercuric for the other, thus:

$$HgNO_3+KI=HgI+KNO_3 \text{ and}$$
$$Hg2NO_3+2KI=HgI_2+2KNO_3.$$

The mercuric iodide is dissolved by excess of either the $Hg2NO_3$ or the KI. In precipitating, mercuric iodide is first yellow, but rapidly becomes red. If some of the dry red powder be placed on a sheet of paper and warmed over a lamp, it changes back to yellow, but on shaking or rubbing the red is restored. These changes in color are due to changes in crystalline structure.

MERCUROUS NITRATE—$HgNO_3$—is formed when mercury is treated with cold dilute nitric acid.

MERCURIC NITRATE—$Hg2NO_3$.—*Acid nitrate of mercury,* is formed if the mercury be boiled with strong nitric acid. Like all nitrates, both of the above are soluble. It enters into the *liq. hydrargyri nitratis,* U. S. P., and "citrine ointment," *ung. hydrargyri nitratis,* U. S. P.

MERCUROUS SULPHATE—Hg_2SO_4—is made by digesting sulphuric acid with excess of mercury.

MERCURIC SULPHATE—$HgSO_4$—is made by heating mercury with excess of sulphuric acid. A white, crystalline salt, used in some forms of galvanic batteries. When diluted with water it decomposes into an acid salt, which remains in solution, and a yellow precipitate of oxysulphate, $HgSO_4.2HgO$, called "turpeth mineral," *hydrargyri subsulphas flavus*, U. S. P.

MERCUROUS CHLORIDE—$HgCl$—*Calomel, mild chloride, Hydrargyri Chloridum Mite*, U. S. P.—is made by heating mercurous sulphate with sodium chloride ($Hg_2SO_4+2NaCl=Na_2SO_4+2HgCl$), when the mercurous chloride sublimes and is condensed in a cool receiver.

Calomel is a white, insoluble powder. Exposed to light it is slowly decomposed ($2HgCl=Hg+HgCl_2$). With aqua regia, and more slowly with other soluble chlorides, it is converted into mercuric chloride. Calomel probably passes through the stomach unaltered, but is converted into the mercurous oxide by the alkaline fluids in the small intestine.

MERCURIC CHLORIDE—$HgCl_2$—*Bichloride of Mercury, Corrosive Sublimate*—is prepared by sublimation from a mixture of mercuric sulphate and sodium chloride, thus:

$$HgSO_4+2NaCl=Na_2SO_4+HgCl_2.$$

It is crystalline and soluble, with a disagreeable styptic taste, and is very poisonous.

MERCURIC AMMONIUM CHLORIDE—*Ammoniated Mercury, White Precipitate*, U. S. P. Formed by adding ammonia to a solution of mercuric chloride; mostly used in ointment. It is a double salt of mercury and NH_2 a derivative of ammonium. Its composition is that of NH_4Cl, in which two atoms of H are displaced by one of Hg, forming NH_2HgCl. The ammonio-sulphate of copper previously described has an analogous composition.

MERCUROUS OXIDE—Hg_2O—*Black Oxide of Mercury*—is made by treating a mercurous salt with a soluble hydrate, as:

$$2HgCl+2KHO=Hg_2O+2KCl+H_2O.$$

It is seldom used in medicine.

MERCURIC OXIDE—HgO—*Red or Yellow Oxide.* When prepared by decomposing mercuric nitrate by heat, it is crystalline and of a red color (*hydrargyri oxidum rubrum*, U. S. P.); but when made by precipitating a mercuric solution with a hydrate,

$$HgCl_2+2KHO=HgO+2KCl+H_2O,$$

it is an amorphous yellow powder (*hydrargyri oxidum flavum*, U. S. P.). The yellow variety, being amorphous and more finely divided, is less gritty and has greater therapeutic activity.

Oleate of Mercury is made by warming the yellow oxide with oleic acid. A liquid or semi-solid. Applied to the skin it is rapidly absorbed.

Mercurous Sulphide — Hg_2S — is an unstable compound, which falls as a black precipitate when a mercurous solution is treated with a soluble sulphide.

Mercuric Sulphide—HgS—falls as a black precipitate when a mercuric solution is treated with a soluble sulphide. It is found in nature in crystalline masses called *cinnabar*. By certain processes it may be obtained as a deep-red crystalline powder, called *vermilion*.

Tests. These consist in adding to the suspected liquid solutions of salts containing radicals capable of uniting with mercury and forming precipitates of the foregoing insoluble compounds. But the *galvanic test* is perhaps the best for clinical purposes. On a gold or copper coin put a drop of the suspected solution acidulated with HCl, and with a piece of baser metal, as a knife blade, touch the coin through the drop of fluid. Mercury, if present, will be deposited on the coin in a silvery film.

Physiological. Acute poisoning occurs from swallowing a single large dose of some of the mercuric compounds, especially corrosive sublimate. The minimum fatal dose of corrosive sublimate is three grains; of white precipitate and turpeth mineral forty grains. Children tolerate mercury much better in proportion to their age than adults. The symptoms are those of severe gastro-enteric irritation. Give albumen, with which it forms an insoluble compound. Iron filings also act as a chemical antidote by decomposing the salt, taking the acidulous radical and depositing the mercury in the metallic state.

Chronic poisoning is often called from its most prominent symptom, *salivation* or *ptyalism.* It usually occurs from small, but often repeated doses of the mercurous preparations, as blue pill, calomel, etc. One of the first *symptoms* is a delicate red line along the margin of the gums, then comes a metallic taste, abdominal pains, nausea, vomiting, dysenteric diarrhea, profuse flow of saliva, fetid breath, fever, emaciation, and paralysis. Sphacelation of the mouth and lips sometimes occurs. The *treatment* is to stop the ingestion of poison, and give some astringent, as tannin.

VIII. The Silver Group.

Silver (*Argentum*)	Ag	108
Gold (*Aurum*)	Au	197
Platinum	Pt	197.4

These are heavy, bright metals, not easily corroded, rare and very valuable. Silver is univalent; gold, trivalent; and platinum, quadrivalent.

SILVER *occurs* free, but oftener as a sulphide associated with lead in galena. A white, malleable, ductile metal, capable of a high polish; best known conductor of electricity; dissolved readily by nitric, but not by hydrochloric or sulphuric acid, except by the aid of heat; does not tarnish in air unless ozone or H_2S be present.

Used to plate mirrors and articles made of the more corrodible metals; alloyed with copper as coin; for tubes, sutures, etc., in surgery, for it does not corrode and irritate the tissues.

Silver Nitrate — $AgNO_3$ — *Argenti Nitras*, U. S. P., *Lunar Caustic. Made* by the action of nitric acid on silver. If coin silver be used, the solution is blue from the presence of copper. Silver nitrate is a crystalline salt, very soluble. Its taste is acrid, and in large doses it acts as a corrosive poison, destroying the tissues by coagulating their albumen. For use as a cautery it is fused and molded into sticks.

Silver Oxide—Ag_2O—is precipitated as a brown powder on treating a solution of silver nitrate with caustic potash or soda ($2AgNO_3+2KHO=2KNO_3+Ag_2O+H_2O$). Slightly soluble in water. The other salts of silver are insoluble, and made by precipitating a solution of silver nitrate with a solution containing the appropriate radical.

Silver Cyanide—$AgCN$.

$$AgNO_3+KCN=AgCN+KNO_3.$$

White precipitate, soluble in ammonium hydrate.

Silver Chloride—$AgCl$.*

$$AgNO_3+NaCl=AgCl+NaNO_3.$$

White precipitate; insoluble in nitric acid, but freely soluble in ammonium hydrate.

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*There are three insoluble chlorides, viz., $PbCl_2$, $HgCl$, and $AgCl$. They may be distinguished by ammonia, which dissolves $AgCl$; blackens $HgCl$, and has no effect on $PbCl_2$.

SILVER BROMIDE—$AgBr$.

$$AgNO_3+KBr=AgBr+KNO_3.$$

Yellowish-white precipitate; slightly soluble in ammonium hydrate.

SILVER IODIDE—AgI.

$$AgNO_3+KI=AgI+KNO_3.$$

Yellow precipitate; insoluble in ammonium hydrate.

Effects of Light. Light decomposes salts of silver, especially if organic matter be present, depositing metallic silver in a fine, black powder, hence their uses in photography, and in making indelible inks, hair dyes, etc. The black stain of silver on the hands or clothes may be removed by potassium cyanide or by applying tincture of iodine and washing in ammonia-water. When persons have taken silver salts for a long time, it sometimes occurs that the tissues, especially the skin, are permanently darkened. This is due to the decomposition of the silver salt under the influence of organic matter and light.

Poisoning occurs mostly from swallowing the nitrate, which is the only soluble silver salt. It is a severe corrosive poison, destroying the tissues by coagulating their albumen. Its best antidote is a soluble chloride, as common salt, which forms the insoluble silver chloride. Albumen is also a good antidote.

GOLD *occurs* widely, but sparingly distributed; always free, mixed with sand and quartz, from which it is separated by agitation with water or by dissolving it out with mercury. It is a soft, bright, yellow metal; so malleable that it may be beaten into sheets (gold-leaf) less than one two-hundred-thousandth of an inch in thickness. These transmit green light. For coinage and general use gold is usually hardened by the addition of copper or silver, the amount of which is indicated by the term *carat fine.* Thus, pure gold is twenty-four carat, and eighteen, sixteen, and twelve carat signify so many twenty-fourths of pure gold.

Gold does not tarnish in the air; is unaffected by any single acid, but nitro-muriatic acid (aqua regia) easily dissolves it, forming *auric chloride*, $AuCl_3$, a caustic salt, which is sometimes given as a nerve tonic and aphrodisiac. Dose, one twentieth to one tenth of a grain.

PLATINUM *occurs* free, associated with the allied metals, palladium, rhodium, rhuthinium, and iridium. Owing to its scarcity it is almost as costly as gold. Resembles silver in appearance; can be melted only with great difficulty, and very few substances corrode it; hence it is used to make vessels that are to be exposed to high temperatures, or to contain corrosive chemicals. Platinum wire is also used in flame-testing.

Platinum readily dissolves in nitro-muriatic acid, forming *platinic chloride*, $PtCl_4$, a valuable reagent for potassium, ammonium, and alkaloids.

TABLE I.—TO DETERMINE THE METALLIC RADICAL OF A SALT IN AQUEOUS OR SLIGHTLY ACID SOLUTION BY SYSTEMATIC ANALYSIS:

Add hydrochloric acid.

<table>
<tr>
<td rowspan="6">Precipitate
Hg (ous) Pb Ag.
Collect, wash, and add NH_4HO.
Hg ppt., blackened.
Pb ppt., still white.
Ag ppt., dissolved.
Sb and Bi may also be precipitated by HCl, but are dissolved on adding more HCl.</td>
<td colspan="6">If HCl gave no precipitate the metal is still in the liquid; pass H_2S through the solution.</td>
</tr>
<tr>
<td rowspan="5">Precipitate
Cd Cu Hg(ic) Pb Bi As Sb Sn Au Pt.
Collect, wash, add NH_4HS
Insoluble.
Cd, yellow
Cu, Hg(ic), Pb, Bi } black.
Soluble.
As (ous & ic), Sn (ic) } yellow.
Sb, orange.
Sn (ous), Au, Pt } black.
Apply special tests for each to the original solution. For these, see the previous pages.</td>
<td colspan="5">If H_2S gave no precipitate add NH_4Cl, NH_4HO, and NH_4HS.</td>
</tr>
<tr>
<td rowspan="4">Precipitate
Zn Mn CO Ni Al Fe Cr.
Zn, Al } white.
Cr, green.
Mn, skin-tint.
Ni, Co, Fe } black.
Test specially for each in original solution.
See previous pages.</td>
<td colspan="4">If NH_4HS gave no precipitate add $(NH_4)_2CO_3$.</td>
</tr>
<tr>
<td colspan="2">Precipitate
Ba Sr Ca.
Collect, wash, dissolve in $HC_2H_3O_2$, add K_2CrO_4.</td>
<td colspan="2">If $(NH_4)_2CO_3$ gave no precipitate add $(NH_4)_2HAsO_4$.</td>
</tr>
<tr>
<td rowspan="2">Ppt.
Ba.</td>
<td>Sol.
Sr Ca.
Add dil. H_2SO_4.</td>
<td rowspan="2">Ppt.
Mg.</td>
<td rowspan="2">If no precipitate, test original solution in flame on loop of Pt wire.
Li, crimson.
Na, yellow.
K, violet.
If neither, test orig. sol. for NH_4.</td>
</tr>
<tr>
<td>Ppt. Sol.
Sr.—Ca.</td>
</tr>
</table>

TABLE II.—TO DETERMINE THE ACIDULOUS (NEGATIVE) RADICAL OF AN ORDINARY SALT IN AQUEOUS SOLUTION, POUR SMALL PORTIONS INTO FIVE TEST-TUBES, THE SOLUTION RENDERED NEUTRAL, IF NECESSARY, BY AMMONIA. THEN ADD TO EACH RESPECTIVELY A FEW DROPS OF SULPHURIC ACID, BARIUM CHLORIDE, CALCIUM CHLORIDE, SILVER NITRATE, AND FERRIC CHLORIDE. INTERPRET THEIR EFFECTS ACCORDING TO THE FOLLOWING TABLE:

H_2SO_4 decomposes	$BaCl_2$ precipitates	$CaCl_2$ precipitates	$AgNO_3$ precipitates	Fe_2Cl6 precipitates	Not precipitated
Sulphides, Sulphites, Carbonates } With effervescence of H_2S and SO_2 known by smell, and CO_2 having no noticeable odor. Cyanides } with odor of HCN. Acetates } with odor of acetic acid, especially if warmed.	(All white.) Sulphates, insol. in HCl. Sulphites, Carbonates } Effervesce with acids. Citrates, Tartrates } Char when heated on platinum foil. Borates. Oxalates. Phosphates.	(All white.) Sulphates, sol. in much water. Borates, Carbonates, Citrates } soluble in NH_4Cl. Phosphates. Oxalates. Sulphites. Tartrates. All sol. in acetic acid, except oxalate and some sulphate and tartrate. All sol. in HCl, except much sulphate. Citrate and tartrate char when heated on platinum foil. Carbonate and sulphite effervesce with acids, evolving H_2S and SO_2.	Borates, Carbonates, Chlorides, Citrates, Cyanides, Oxalates, Sulphates, Tartrates } white. Bromides } yellowish-white. Iodides, Phosphates } yellow. Sulphides, black. All soluble in dilute nitric acid, except chloride, bromide, iodide, cyanide, and sulphide.	Phosphates } yellowish-white. Borates, yellowish. Oxalates, yellow. Carbonates, reddish. Acetates, red if neutral. Sulphides, black.	Nitrates. Chlorates. Apply special tests. (See previous pages.)

TABLE III.—THE SOLUBILITY OR INSOLUBILITY OF SALTS IN WATER.

Much time may be saved by not applying special tests to an aqueous solution for salts known to be insoluble in water. S, soluble; Ss, slightly soluble; I, insoluble; ?, unknown or does not exist.

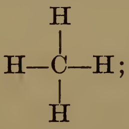
	Acetate.	Arseniate.	Arsenite.	Carbonate.	Chloride.	Citrate.	Chromate.	Cyanide.	Hydrate.	Iodide.	Nitrate.	Oxalate.	Oxide.	Phosphate.	Sulphate.	Sulphide.	Sulphite.	Tartrate.
Aluminium	S	I	I	I	S	Ss	I	?	I	?	S	I	I	I	S	I	I	S
Ammonium	S	S	S	S	S	S	S	S	S	S	S	S	?	S	S	S	S	S
Antimony	S	I	I	?	S	?	I	?	I	I	?	Ss	I	I	I	I	I	I
Barium	S	I	I	I	S	Ss	I	S	S	S	S	I	S	I	I	S	I	I
Bismuth	S	I	?	I	S	Ss	I	?	I	I	S	I	I	I	S	I	I	I
Cadmium	S	?	?	I	S	Ss	?	?	I	S	S	I	I	I	S	I	S	Ss
Calcium	S	I	I	I	S	Ss	S	S	Ss	S	S	I	Ss	I	Ss	Ss	Ss	Ss
Chromium	S	I	?	I	S	S	I	I	I	S	S	S	I	I	S	I	S	S
Cobalt	S	I	I	I	S	S	?	I	I	S	S	I	I	I	S	I	Ss	S
Copper	S	I	I	I	S	S	S	I	I	....	S	I	I	I	S	I	I	Ss
Ferric	S	I	I	?	S	S	S	?	I	S	S	Ss	I	I	S	I	?	S
Ferrous	S	I	I	I	S	S	?	I	I	S	S	Ss	I	I	S	I	Ss	Ss
Gold	?	I	?	?	S	?	?	I	I	I	?	?	I	I	?	I	?	?
Lead	S	I	I	I	Ss	Ss	I	I	Ss	Ss	S	I	I	I	I	I	I	I
Magnesium	S	I	I	I	S	S	S	?	I	S	S	I	I	I	S	?	S	S
Manganese	S	I	?	I	S	I	S	I	I	S	S	I	I	I	S	I	Ss	Ss
Mercuric	S	I	I	I	S	?	Ss	S	I	I	S	I	I	I	S	I	?	I
Mercurous	Ss	I	I	I	I	I	I	?	?	I	S	I	I	I	Ss	I	?	I
Nickel	S	I	I	I	S	S	I	I	I	S	S	I	I	I	S	I	Ss	Ss
Platinum	?	I	?	?	S	?	?	?	I	Ss	S	?	I	?	S	I	S	?
Potassium	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Silver	Ss	I	I	I	I	I	I	I	?	I	S	I	I	I	Ss	I	I	I
Sodium	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S	S
Stannic	S	?	I	?	S	?	I	?	I	S	S	S	I	I	S	I	?	?
Stannous	S	I	I	?	S	?	I	?	I	S	S	I	I	I	S	I	I	S
Strontium	S	I	I	I	S	I	I	S	S	S	S	I	S	I	I	S	I	Ss
Zinc	S	I	?	I	S	Ss	S	I	?	S	S	I	1	I	S	I	Ss	Ss

ORGANIC CHEMISTRY.

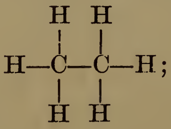
Formerly organic chemistry was defined as the chemistry of the compounds produced only by organized life. Gradually this definition has been abandoned, for with the increase of chemical knowledge many substances identical with the animal and vegetable products have of late years been made in the laboratory without the aid of the vital force, and probably, if their chemical constitution were fully understood, all animal and vegetable products could be duplicated artificially. However, chemistry has not, and probably never will, produce an *organized* body; *i. e.*, one having an anatomical cellular structure.

It is a noticeable fact that every organic compound contains carbon. Hence, *organic chemistry* is now defined to be "the chemistry of the carbon compounds," and the following pages may be considered a resumption of the study of that element.

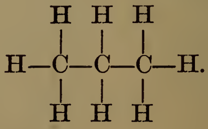
Though carbon forms compounds of infinite number and extreme complexity, it is with the aid of a very few other elements, viz., hydrogen, oxygen, nitrogen, and occasionally sulphur, phosphorus, and iron—sometimes others; but the larger number of even the artificial compounds contain only the above-named elements. This is due to the fact that the carbon atoms possess in the highest degree the power of *combining with each other and interchanging valences*, forming groups or chains around which the other elements are arranged. But for this power carbon could form only one saturated compound with hydrogen, CH_4. Carbon being quadrivalent, the compounds C_2H_6 and C_3H_8 would be unsaturated. Experiment, however, proves that they are saturated compounds. The explanation is that the carbon atoms combine with each other, mutually neutralizing one or more valences, thus:



$$\begin{array}{ccccc} & & H & & \\ & & | & & \\ H & - & C & - & H; \\ & & | & & \\ & & H & & \end{array}$$



$$\begin{array}{ccccccc} & & H & & H & & \\ & & | & & | & & \\ H & - & C & - & C & - & H; \\ & & | & & | & & \\ & & H & & H & & \end{array}$$



$$\begin{array}{ccccccccc} & & H & & H & & H & & \\ & & | & & | & & | & & \\ H & - & C & - & C & - & C & - & H. \\ & & | & & | & & | & & \\ & & H & & H & & H & & \end{array}$$

It will be observed that these formulæ have a common difference of CH_2. They are said to form a *homologous series.* When the carbon remains the same but the hydrogen differs by H_2, the series is said to be *isologous.*

In the following examples each vertical column represents a homologous, each horizontal line an isologous series:

CH_4	CH_2	C
C_2H_6	C_2H_4	C_2H_2
C_3H_8	C_3H_6	C_3H_4
C_4H_{10}	C_4H_8	C_4H_6
C_5H_{12}	C_5H_{10}	C_5H_8
etc.	etc.	etc.

Without this arrangement in series, it would be impossible to remember the composition of organic substances.

Hydrocarbons are compounds of carbon with hydrogen. Of these CH_4 is the type from which all the other members of this class may be regarded as derived in isologous or homologous series. *Petroleum* is a mixture of the homologous derivatives from the first (CH_4) to about the sixteenth ($C_{16}H_{34}$). These are separated by distilling the crude oil. Those having the smallest molecules, being lightest, pass over first, forming *naphtha*, *benzine*, etc. As the heat is increased the medium-weight compounds come over next, forming *kerosene.* The residuum consists of the heaviest carbohydrides, which can be distilled only by high heat, forming *lubricating oil*, *vaseline*, *paraffine*, etc. Kerosene is the one used for ordinary lamps. If it contains too much of the lighter products it is liable to give off vapors which, mixing with air, are explosive. In most States it is illegal to sell kerosene which gives off an inflammable vapor ("flashes") below 100° F.

THE VOLATILE OILS are a class of carbohydrides, all having the same formula, $C_{10}H_{16}$. Though having very different chemical and physical properties, they are composed of the same elements in the same proportion. Such bodies are said to be *isomeric* (ἴσος, equal, μερος, part).

Volatile oils are found in plants, especially the flowers, of which they are usually the odorous essences (hence called also *essential oils*). Obtained by distillation from flowers, etc. Very slightly soluble in water (*aquæ*), but quite soluble in alcohol (*spir-*

itus). A cologne is an alcoholic solution of an assortment of volatile oils. A large number of volatile oils are officinal; as, anise, bergamot, cinnamon, lemon, orange, wintergreen, etc.

Turpentine (*oleum terebinthinæ,* U. S. P.) may be taken as a type of the class. It is a thin, colorless liquid, a valuable solvent of oils and resins; absorbs oxygen and stores it up as ozone, gaining thereby oxidizing, antiseptic, and disinfectant properties. By the action of concentrated sulphuric acid, turpentine is changed into *terebene* ($C_{10}H_{16}$), a valuable remedy for bronchitis and flatulence. On exposure to the air the volatile oils oxidize with production of resins and camphors.

RESINS are formed by the oxidation of volatile oils. Insoluble in water; soluble in alcohol and ether; alkalies dissolve them, forming soapy mixtures. The officinal resin (*resina,* U. S. P.) is formed by the oxidation of turpentine as it exudes from the pine tree.

In the natural state resins are usually mixed with other substances. Mixed with volatile oils they form *oleo-resins* and *balsams, e. g.*, benzoin, tolu, and balsam of Peru, and with gums, *gum-resins, e. g.*, ammoniac myrrh and asafœtida.

CAMPHORS. *Common camphor,* obtained from the camphor laurel, is a white, crystalline, volatile solid of a peculiar pungent odor; slightly soluble in water (*aqua camphoræ,* U. S. P.), and freely soluble in alcohol and ether.

Monobromated camphor, used in medicine as a sedative, is formed by substituting one atom of bromine for one of hydrogen in ordinary camphor.

Menthol is a camphor-like body found in oil of peppermint, and possesses the odor of that plant.

CAOUTCHOUC, or *India-rubber,* and *gutta-percha* are inspissated juices of certain tropical trees. Caoutchouc is elastic; gutta-percha is not. Both are hardened (vulcanized) by combining with sulphur. They are unaffected by most chemicals and solvents. Chloroform is their best solvent.

THE ALCOHOL RADICALS, a homologous series of univalent basylous radicals, so called because they are the bases of the most important alcohols. Their compounds are numerous, and enter largely into the *materia medica.* In the following table a few of these compounds are given:

Radicals.		Alcohols. (Hydrates.)	Oxides, Ethers.	Examples of Compound Ethers.		Aldehydes.	Acids.
				Nitrates.	Sulphates.		
Methyl,	CH_3	CH_3HO	$(CH_3)_2O$	CH_3NO_3	$(CH_3)_2SO_4$	CH_2O	CH_2O_2
Ethyl,	C_2H_5	C_2H_5HO	$(C_2H_5)_2O$	$C_2H_5NO_3$	$(C_2H_5)_2SO_4$	C_2H_4O	$C_2H_4O_2$
Propyl,	C_3H_7	C_3H_7HO	$(C_3H_7)_2O$	$C_3H_7NO_3$	$(C_3H_7)_2SO_4$	C_3H_6O	$C_3H_6O_2$
Butyl,	C_4H_9	C_4H_9HO	$(C_4H_9)_2O$	$C_4H_9NO_3$	$(C_4H_9)_2SO_4$	C_4H_8O	$C_4H_8O_2$
Amyl,	C_5H_{11}	$C_5H_{11}HO$	$(C_5H_{11})_2O$	$C_5H_{11}NO_3$	$(C_5H_{11})_2SO_4$	$C_5H_{10}O$	$C_5H_{10}O_2$
Hexyl,	C_6H_{13}	$C_6H_{13}HO$	$(C_6H_{13})_2O$	$C_6H_{13}NO_3$	$(C_6H_{13})_2SO_4$	$C_6H_{12}O$	$C_6H_{12}O_2$
etc.		etc.	etc.	etc.	etc.	etc.	etc.

In the formation of these compounds the starting point is not the radicals, but their hydrates, the alcohols. When an alcohol is oxidized with a limited supply of oxygen, two atoms of hydrogen are removed and no oxygen is added. This forms the *aldehyde*, thus:

$$\overset{\text{Methyl alcohol.}}{CH_3HO}+O=\overset{\text{Methyl aldehyde.}}{CH_2O}+H_2O.$$

If there is a full oxidation, an atom of oxygen takes the place of the two atoms of hydrogen removed and forms the corresponding acid, as:

$$\overset{\text{Methyl alcohol.}}{CH_3HO}+O_2=\overset{\text{Formic acid.}}{CH_2O_2}+H_2O.$$

In the formation of aldehydes and acids the radical supplies part of the hydrogen removed and loses its identity. As part of the hydrogen in an acid forms the positive radical, it is written first; *e. g.* formic acid is written $HCHO_2$, instead of CH_2O_2. The various other compounds of these radicals are called *ethers;* the oxides being called *simple ethers*, the others *compound ethers.* They are generally formed by treating the appropriate alcohol with the appropriate acid.

Alcohols. The modern chemist accepts as alcohols many substances that bear little resemblance to ordinary alcohol.

METHYLIC ALCOHOL—CH_3HO—*Wood Naphtha, Wood Spirit, Wood Alcohol, Pyroligneous Spirit, Methyl Hydrate*—does not exist in nature. Made by the destructive distillation of wood. The commercial article has a very disagreeable odor and taste from the presence of tarry matters, etc.; but when pure methylic alcohol resembles ordinary alcohol in its properties and physiological action. It is not used in medicine, but is extensively employed in the arts as a solvent, as in the preparation of varnishes, etc.

Methylated spirit is ordinary alcohol to which has been added one tenth part of commercial methylic alcohol to render it unfit for drinking, and thus relieve it of the heavy tax imposed upon alcoholic beverages.

ETHYLIC ALCOHOL—C_2H_5HO—*Ethyl Hydrate, Spirits of Wine, Vinic Alcohol, Alcohol.* Alcohol does not exist in nature, but is produced in a number of reactions. Liquids containing it (wines, etc.) have been known from the remotest antiquity, and are obtained by allowing liquids containing glucose (grape sugar) to ferment.

$$\underset{\text{Glucose.}}{C_6H_{12}O_6} = \underset{\text{Alcohol.}}{2C_2H_5HO} + \underset{\text{Carbon dioxide.}}{2CO_2}.$$

The alcohol is then separated by distillation, for, being more volatile than the water, it passes over first.

Commercial alcohol always contains water, and when pure or *absolute* alcohol is required, the commercial article is mixed with some substance which is very avid of water, as quicklime, and then again distilled.

Alcohol is a light, colorless liquid, of a pleasant, pungent odor and burning taste. Has a great affinity for water, which probably accounts for its preserving animal tissues and coagulating albumen.

It is largely used in the arts and in pharmacy, principally as a solvent; but also in the manufacture of various substances, as vinegar, chloral, chloroform, iodoform,* ether, etc.; and as a fuel when a hot and smokeless flame is needed; and as a menstruum in the preparation of tinctures and spirits.

Alcoholic solutions of fixed medicinal substances are called *tinctures;* those of volatile principles, *spirits.*

Alcohol is employed in various forms and degrees of concentration. *Absolute alcohol* is rarely employed. *Alcohol fortius,* U. S. P., stronger alcohol, contains 92 per cent of alcohol. *Alcohol,* U. S. P., is the ordinary rectified spirit, and contains 85 per cent of alcohol. *Alcohol dilutum,* U. S. P., diluted alcohol, is made by mixing water and alcohol, equal parts.

***Experiment.** To test for alcohol in a solution: Warm; add a few scales of iodine, and then caustic potash until the color is discharged. On cooling yellow scales of iodoform are deposited.

Spiritus frumenti, U. S. P., whisky, and *spiritus vini gallici*, U. S. P., brandy, are obtained by distillation; the former from fermented grain, and the latter from fermented grape juice. They contain about 50 per cent of alcohol. They are colored by the addition of caramel (burnt sugar). Their flavor is due to small quantities of other alcohols produced in the fermentation, and to certain ethers formed from these alcohols, especially as the liquor "*ages*."

A large class of alcoholic beverages are made by fermenting various liquids containing sugar or some substance capable of conversion into sugar.

Beer, *ale*, and *porter* are infusions of malted grain, fermented and flavored with hops. They therefore contain the soluble constituents of the grain. Their alcoholic strength is about 5 per cent. *Wines* are prepared by allowing grape juice to ferment. The alcoholic strength of the different varieties varies from 10 to 25 per cent. Sherry (*vinum Xericum*) and port (*vinum rubrum*) are the only ones officinal. *Cider* is the fermented juice of the apple, prepared very much in the same way as wine is from grape juice, and contains about 5 per cent of alcohol. It is very prone to acetous fermentation and liable to produce colic and diarrhea.

Alcohol, when concentrated, abstracts water from the tissues and coagulates their albuminoid constituents, and is a poison. In full doses (*always best with food*) it produces a sense of warmth in the stomach, general comfort and exhilaration, followed by incoherence of ideas and impairment of muscular co-ordination. Taken habitually, in any of its forms, it impairs the mental and moral force of its victim and produces in the various organs, especially the liver and kidneys, the degenerative changes characteristic of chronic alcoholism. It should never be taken in health, but as a medicine it is the most valuable of stimulants. In cases of acute poisoning by alcohol, the stomach and bladder should be evacuated, and the depression (coma) counteracted by strong coffee, the cold douche, and other stimulants.

AMYLIC ALCOHOL—$C_5H_{11}HO$—*Amyl Hydrate*, *Fusel Oil*. This is a heavy liquid, soluble in alcohol but not in water, hence incorrectly called an oil. It is produced in the fermentation of grain and potatoes, and is the most deleterious impurity in common whisky before it has undergone the refining process.

It has a penetrating, disagreeable odor, resembling that of mean whisky. Although not fragrant itself, its ethers, when dissolved in ethylic alcohol, have the taste and odor of various fruits, and are used in the preparations of artificial fruit essences.*

The other alcohols of this series are of no medical interest.

GLYCERYLIC ALCOHOL—C_3H_53HO—*Glyceryl Hydrate, Glycerine.* This is a sweet, viscid liquid, formed incidentally in the preparation of lead plaster, and in the manufacture of soap and candles. It is colorless and odorless, and neutral in reaction; a solvent of a great many mineral and organic substances. It dissolves freshly precipitated tannate of quinine (five grains to the ounce), and the solution is scarcely bitter. When glycerine is treated with strong nitric acid *nitro-glycerine,* $C_3H_53NO_3$ results, one of the most dangerous of explosives. Dynamite is a mixture of nitro-glycerine and sand. Glycerine is used in medicine mainly as a solvent and as a local application.

MANNITYL ALCOHOL—C_6H_86HO—*Mannite.* This is the principal ingredient in manna, a white, gummy substance exuding from certain trees. Mannite is a crystalline substance closely resembling glucose, except that it does not undergo the vinous fermentation, and does not respond to Trommer's or Fehling's tests.

Ethers. The *simple* ethers (oxides) are the results of dehydrating two molecules of alcohol by means of sulphuric acid; the *compound* ethers are made by treating the appropriate alcohol with the appropriate acid.

ETHYL OXIDE—$(C_2H_5)_2O$—*Sulphuric Ether, Ether.* Ether is made by distilling a mixture of alcohol and sulphuric acid;† hence the misnomer, *sulphuric ether.*

A small quantity of sulphuric acid is capable of converting a large amount of alcohol into ether, for it is unaltered in the reaction; in fact, the process might go on indefinitely but for the

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* **Experiment.** To a half dram of fusel oil in a test-tube add some sodium or potassium acetate and a few drops of sulphuric acid. Warm the mixture, and the acetate of amyl (essence of pear) may be recognized by its odor.

† **Experiment.** Into a large test-tube pour alcohol and half as much sulphuric acid; warm and note the odor of the ether evolved. Next adapt a cork with delivery tube, and slowly distill the ether into a cool test-tube. By adding more alcohol the operation may be repeated again and again.

acid being so diluted with the water derived from the alcohol, as to finally stop the reaction. The sulphuric acid is said to act by its mere presence, by *catalysis;* or, in other words, it acts because it acts, a ready but feminine way of explaining many otherwise inexplicable chemical and physiological phenomena.

The true rationale is as follows:

$$C_2H_5HO+H_2SO_4=C_2H_5HSO_4+H_2O,$$

And then

$$C_2H_5HSO_4+C_2H_5HO=(C_2H_5)_2O+H_2SO_4.$$

Ether is a colorless, very volatile liquid of a peculiar odor, called *ethereal.* It burns easily, and its vapor, mixed with air or oxygen, explodes when ignited;* so ether should never be used near, especially above, a flame. Ether is a valuable solvent, and, as it evaporates very rapidly, it is used to produce cold.† But its chief use in medicine is an anæsthetic. Being less liable to paralyze the nerve centers, it is safer than chloroform.

ETHYL CHLORIDE—C_2H_5Cl. *Hydrochloric ether* must not be confounded with the so-called *chloric ether*, which is an alcoholic solution of chloroform.

ETHYL BROMIDE—C_2H_5Br—*Hydrobromic Ether.* A valuable anæsthetic, but not much used.

ETHYL NITRITE—$C_2H_5NO_2$—*Nitrous Ether.* If nitric acid be treated with copper or starch it loses part of its oxygen, being converted into nitrous acid (HNO_2), which will unite with alcohol, forming nitrous ether and water, thus:

$$C_2H_5HO+HNO_2=C_2H_5NO_2+H_2O.$$

Nitrous ether is a yellowish liquid, of an apple-like odor and sweetish taste. It is used only diluted with alcohol, forming the *spiritus ætheris nitrosi*, U. S. P., commonly called *sweet spirits of niter.*

AMYL NITRITE—$C_5H_{11}NO_2$. Made like ethyl nitrite, except that amylic alcohol is used. Nitrite of amyl is a volatile, oily liquid, of a peculiar odor, resembling that of bananas. It is given by inhalation, especially in epilepsy, for which purpose it is put up in glass bulbs holding about two drops. These are crushed and inhaled during the *aura.*

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*__Experiment.__ Put a dram of ether in a dish and apply a flame. The vapor, mixed with air, explodes; the rest burns rapidly.

†__Experiment.__ Set a test-tube of water in a beaker of ether. Blow air briskly through the ether; the water will freeze.

CHLOROFORM—*Trichlormethane.* Methyl, having only one free valence, must give up two atoms of its hydrogen in order to combine with three atoms of chlorine, making the formula of chloroform $CHCl_3$. This is the most used of all the ethers. Chloroform is made by distilling a mixture of chlorinated lime, water, and ordinary alcohol. It is a colorless, volatile liquid, of a pleasant ethereal odor and sweet taste. It is heavier than water, and does not dissolve in it, but is soluble in alcohol and ether; not easily ignited; a good solvent for phosphorus, iodine, India-rubber, and the alkaloids. Chloroform is sometimes given by the stomach as a sedative, but most frequently administered by inhalation as an anæsthetic, for which purpose it should be of undoubted purity. Pure chloroform is not colored by an equal volume of pure sulphuric acid, nor should its specific gravity be below 1480.

Toxicology. If chloroform be taken by the stomach, it being insoluble, is absorbed very slowly; and its principal action is the local irritation of the mucous surfaces. Recovery has followed a dose of four ounces, and death has been caused by one dram taken into the stomach. The vapor acts more energetically, and seems to owe its potency for evil to its paralyzing influence on the nerve centers, especially those of the heart. So chloroform vapor should never be administered except by a capable physician, and well diluted with atmosphere. However, death has occurred from the inhalation of moderate quantities of chloroform properly diluted, and at the hands of careful physicians, and the autopsy revealed no heart lesion.

There is no chemical antidote for chloroform. When it has been swallowed, evacuate the stomach; when inhaled, lower the head, give fresh air, employ artificial respiration, and apply the induced current.

The poison is usually recognized by its odor.

IODOFORM—CHI_3—Made by the action of iodine and potash on alcohol; yellow scales; insoluble in water, and of a saffron-like odor, which is the chief objection to its use. Light decomposes it, giving it a violet color. Iodoform is used on ulcers, etc., as an anæsthetic, alterative, and antiseptic.

FIXED OILS AND FATS. These are ethers—combinations of glyceryl (C_3H_5) with the fat acids, oleic, stearic, butyric, palmitic, etc. The natural fats are mixtures of these. Those containing mostly oleate of glyceryl (olein) are liquid. Warm-blooded animals yield mostly solid, cold-blooded, liquid fats. Drying oils are such as absorb oxygen from the air and become resinous, *e. g.* linseed. Many fats partially decompose on exposure, producing free acid, and become *rancid*. The fixed oils are insoluble in water, soluble in alcohol, ether, and chloroform.

The compounds of the fatty acids with metallic radicals are called *soaps*. Soaps are made by the *saponification* of a fat with a caustic alkali. For example:

Stearine. Sodium stearate. Glycerine.

$$(C_3H_5)(C_{18}H_{35}O_2)_3+3NaHO=3NaC_{18}H_{35}O_2+C_3H_5(HO)_3$$

All the soaps are insoluble, except those of the alkali metals (K, Na, NH_4, etc.). This explains the curd, soap precipitates in hard waters. Lead soap (lead plaster) is officinal. When soap dissolves in cold water, it decomposes into an acid salt which makes the soapsuds and a small quantity of free alkali which does the cleaning.

Aldehydes. An unimportant class. They constitute the first step in the oxidation of alcohols into acids, viz., the removal of hydrogen (hence the name). Since nothing has taken the place of the hydrogen removed, they are unsaturated and very prone to change, especially to take on oxygen and form the acids.

ETHYL ALDEHYDE, *acetic aldehyde*, or simply *aldehyde** (C_2H_4O), is a colorless, volatile, acid liquid of a pungent odor. One of its modifications, called *paraldehyde*, is used as a hypnotic, which unlike morphine is followed by no unpleasant effects, except a pungent odor to the breath. Dose, ʒss–ʒj.

CHLORAL. If chlorine displace three atoms of hydrogen in ethyl aldehyde, it forms *tri-chlor-aldehyde* or chloral (C_2HCl_3O), a colorless, heavy liquid. With a molecule of water this forms a white crystalline solid, called *chloral hydrate*, having a pungent but agreeable odor and taste. Warmed with an alkali it decomposes, thus:

Chloral. Sod. formate. Chloroform.

$$C_2HCl_3O+NaHO=NaCHO_2 + CHCl_3.$$

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* **Experiment.** To a little bichromate and sulphuric acid mixture in a test-tube add a little alcohol; or hold a hot glass rod in a beaker containing a little ether. The peculiar pungent odor is that of aldehyde.

Liebriech thought this reaction would occur in the warm alkaline blood and the sedative action of chloroform be obtained. Though mistaken in this, he found chloral hydrate a valuable hypnotic. The chloral habit is difficult to cure. In overdoses chloral is a poison, and cases are multiplying as its powers become better known. No chemical antidote. Evacuate the stomach, give stimulants, and maintain the respiration and bodily warmth.

Organic Acids. These are regarded by chemists as the natural results of the oxidation of alcohols. But as most of them were discovered before their relation to the alcohols were known, their names often have no connection with those of the alcohols. We take up only the most important.

ACETIC ACID—$HC_2H_3O_2$. This is the acid of vinegar. Formed in a great many reactions, but made mainly by the destructive distillation of wood, or by the oxidation of ordinary alcohol. If wine, cider, or other alcoholic liquors be exposed to the air, a fungus (*micoderma aceti*) forms on the surface and acts as an oxygen carrier, and the alcohol is converted into acetic acid, thus:

$$C_2H_5HO+O_2=HC_2H_3O_2+H_2O.$$

A more rapid process is to pass the alcohol through barrels filled with beech shavings.

Acetic acid is a colorless liquid, of a pungent, sour taste and smell. When free from water (glacial) it crystallizes at temperatures below 60° F. Acetic acid in dilute solution (vinegar) is much used for domestic purposes. For medicinal use the crude vinegar is purified by distillation, forming *acidum aceticum dilutum*, U. S. P.

As all the acetates are soluble, their best test is to add a strong acid and recognize the acetic acid set free by its odor.

BENZOIC ACID exists in many balsams and gum-resins. When benzoin is heated benzoic acid sublimes in silky needles of a pleasant balsamic odor. Or, if the urine of herbivorous animals be boiled with hydrochloric acid, the hippuric acid is converted into benzoic acid. But the acid obtained from this source may be known by its urinous odor.

CARBAZOTIC OR PICRIC ACID is a yellow substance, of a bitter taste, made by the action of nitric acid on carbolic acid. Used in the arts as a yellow dye.

CARBOLIC ACID—C_6H_5HO—*Phenyl Alcohol, Phenol.* This is an alcohol, the hydrate of phenyl, a radical of the aromatic series. Called an acid because it combines with bases and forms bodies resembling salts called carbolates or phenates.

Carbolic acid is formed in a number of reactions, but the commercial article is obtained exclusively from coal tar. It has a strong, disagreeable odor; occurs as white crystals, which melt on the addition of a small quantity of water; reddens by age; slightly soluble in water, but very soluble in glycerine, the solution being soluble in water; stains skin and mucous membranes white by coagulating their albumen; and is a corrosive poison. Albumen is its best antidote.

Carbolic acid is a powerful antiseptic and disinfectant. Applied locally it is astringent, sedative, and even anæsthetic.

Resorcin—(C_6H_42HO). Closely related to phenol, but a stronger antiseptic and much less poisonous. It occurs in soluble, colorless, odorless crystals of a sweetish taste. It is given as an antizymotic in diseases attended with fermentative changes and in the specific fevers.

Creasote is a complex mixture obtained from wood tar; closely allied to carbolic acid in its properties and uses, but may be readily distinguished from it by being insoluble in glycerine.

CITRIC ACID exists in the juices of many fruits, especially the lemon. Forms colorless crystals which are very soluble, and possess a sour taste. Many of its salts are used in medicine.

FORMIC ACID—$HCHO_2$—is the oxidation product of methylic alcohol. It was formerly obtained from the red ant (*formica rufa*), but now made artificially. It exists in stinging nettle, pine needles, etc., and also in the stings of most insects.

GALLIC ACID. When galls are moistened and exposed to the action of the atmosphere, the tannic acid they contain is converted into gallic acid. It resembles tannic acid, but may be distinguished by its not precipitating a solution of gelatin.

LACTIC ACID (*lactis*, of milk). This is the acid of sour milk, where it is formed by the fermentation of the sugar of milk through the agency of the casein. It is also formed in the body by the decomposition of glucose, thus:

$$C_6H_{12}O_6=2H_2C_3H_4O_3.$$

It is a syrupy liquid, of a very sour taste.

MALIC ACID (*malum*, an apple) exists in many fruits, as apples, cherries, etc., and very abundantly in garden rhubarb.

OXALIC ACID—$H_2C_2O_4$. The acid and its salts are found in many plants, especially the sorrel (oxalis) grasses. In certain pathological conditions it is formed in the body and eliminated in the urine as calcium oxalate. It is made in large quantities by the action of nitric acid on sugar, or of alkalies on saw-dust. Oxalic acid closely resembles Epsom salts, for which it is sometimes taken by mistake. It is a powerful irritant poison. Being cheap and largely used for removing ink stains, cleaning copper, etc., poisoning by oxalic acid is by no means rare. Its best antidote is chalk, or some other compound of calcium, with which it forms a very insoluble compound.

Test. Calcium chloride gives a white precipitate, insoluble in acetic, but soluble in hydrochloric acid.

PYROGALLIC ACID sublimes as white, feathery crystals when gallic acid is heated. Used in gas analysis to absorb oxygen, as a deoxidizer in photography, and as a hair dye.

Test. A blue color with ferrous, and a red with ferric salts.

SALICYLIC ACID. Formerly prepared from salicin, but now made by a patented process from carbolic acid. A very pure acid may be obtained from oil of wintergreen, which consists mainly of methyl salicylate. This, treated with potassium hydrate, forms methyl hydrate (methyl alcohol) and potassium salicylate; and if to this hydrochloric acid be added, potassium chloride will be formed, and salicylic acid will fall in a mass of silky, white crystals. Salicylic acid is scarcely soluble in cold water, hence the salicylate of sodium is usually prescribed, which is not only more soluble, but less irritating to mucous membranes.

Test. Intense violet with a ferric salt.

SUCCINIC ACID was first obtained from amber (*succinum*), but is now made by fermenting malic acid.

TANNIC ACID, OR TANNIN. This is the active principle of the vegetable astringents; usually obtained from oak galls; a greenish or brownish powder, very soluble in water, of a rough, astringent taste. It precipitates solutions of salts of the alkaloids and most metals. It precipitates gelatin and other albuminoid substances, a fact that explains the process of tanning raw hides. With ferric solutions tannin gives a black precipitate (black ink).

TARTARIC ACID—$H_2C_4H_4O_6$, or $H_2\bar{T}$. Tartrates exist in the juices of many fruits. Grape juice contains much acid tartrate of potassium ($KH\bar{T}$), which, being very insoluble in an alcoholic menstruum, is precipitated on the sides of the cask whenever the wine ferments. This forms *argol*, the principal source of tartaric acid. Tartaric acid forms colorless crystals, very soluble, and of a sharp, agreeable, sour taste.

VALERIANIC ACID—$HC_5H_9O_2$. This substance was first obtained from valerian root; but now it is made artificially by oxidizing amylic alcohol by means of sulphuric acid and potassium bichromate. Valerianic acid is a colorless liquid, possessing the disagreeable odor of valerian.

The Carbohydrates. These substances are closely related to the alcohols, and by some classed as such. They are so named because they contain carbon (six or twelve atoms), and the hydrogen and oxygen they contain are in the exact proportion to form water. They constitute the bulk of plants. They are divided into three groups:

1. AMYLOSES ($C_6H_{10}O_5$), which include cellulin, starch, dextrin, glycogen, gums, etc.
2. SACCHAROSES ($C_{12}H_{22}O_{11}$), including cane sugar, milk sugar, etc.
3. GLUCOSES ($C_6H_{12}O_6$), such as grape sugar (glucose), fruit sugar, etc.

Although the members of each of these groups differ widely in their physical and chemical properties, still they consist of the same elements in exactly the same proportions and have the same formula. Such bodies are said to be *isomeric*.

AMYLOSES—($C_6H_{10}O_5$). CELLULIN—*Cellulose, Lignin*—forms the cell-walls and tissues of plants. Woody fiber, cotton, linen, and unsized paper are almost pure cellulin. Dissolves only in a solution of cupric oxide in ammonia. Acids precipitate it as a white mass, which, mixed with camphor and compressed, is *celluloid*. Unsized paper dipped into moderately strong sulphuric acid, washed and dried, has its fibers agglutinated, loses its porosity, becomes very tough, and is sold as *artificial parchment* for dialyzers, diplomas, etc.

Nitro-cellulose, or *Gun Cotton*, a powerful explosive, is cotton that has been dipped into a mixture of nitric and sulphuric acids, and then washed and dried. Its solution in ether is *collodion.* The *flexible* collodion contains a little turpentine and castor oil; the *styptic* collodion contains twenty per cent tannin.

STARCH—*Amylum*—the most important member of the carbohydrates, and a valuable food; found in the roots, stems, and seeds of all plants. Starch is a white powder, consisting of granules, formed of concentric layers, like an onion. These granules have all a similar appearance. Yet those from different kinds of plants differ enough to enable one, by microscopic examination, to determine the source of any starch. (Fig. 64.)

Fig. 64. (*a*) potato, (*b*) corn, (*c*) bean starch.

When starch is boiled the granules swell and burst, casting the starch into the water, forming mucilage of starch, which is used for laundrying and for surgical dressings. Starch is a very valuable food. The best test for starch is iodine, with which it forms a blue. Heat discharges the blue, but it returns on cooling.

DEXTRIN—*British Gum*—is the gum used on postage stamps, etc. It may be made from starch in various ways, one of which is by heating it to 300° F. It is very soluble, and gives no blue with iodine.

GLYCOGEN—(*Generator of Glucose*)—is found in the animal economy, especially in the liver. Like dextrin, it is a derivative of starch, but differs from it in being soluble, and giving only a wine-color with iodine. It seems to be the form in which the carbohydrates are stored up, to be used by the system as necessity arises.

GUMS are a class of substances soluble in water, but insoluble in alcohol. A type of the class is *gum Arabic.*

SACCHAROSES—$C_{12}H_{22}O_{11}$. CANE SUGAR—*Beet Sugar, Sucrose.* Very abundant in the sugar-cane, sugar-maple, beet-root, etc. It is the most soluble, perfectly crystallizable, and sweetest of the sugars, and the one most used for domestic purposes. Its aqueous solution is called *simple syrup* (*syrupus simplex*).

MILK SUGAR, as its name implies, occurs in milk; harder, less soluble, and less sweet than cane-sugar. Used in the trituration of medicines.

GLUCOSES—$C_6H_{12}O_6$. GLUCOSE—*Grape Sugar*, *Diabetic Sugar.* Found associated with other sugars in most plants, especially in the grape; but the source of most interest to the physician is the animal economy. This is the sugar of diabetic urine, and the ability to detect it with ease and certainty in such conditions is a necessity to the practitioner of the present day.

Glucose is not so sweet nor so soluble and crystallizable as cane-sugar. Having great affinity for oxygen it is a valuable reducing agent, and on this property most of its tests depend. Boiled with a dilute mineral acid or allowed to remain under the influence of certain animal and vegetable ferments,* warm and moist, the amyloses and saccharoses are

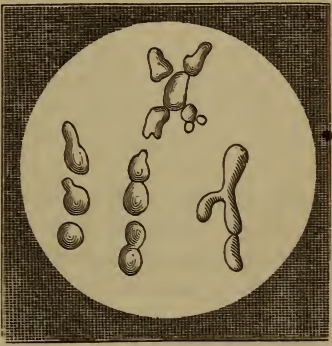

Fig. 65.

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*FERMENTS. These are certain nitrogenous bodies, animal and vegetable, which by some means not clearly understood cause many organic compounds to decompose with the production of other and simpler substances, the ferments themselves being unaffected. Ferments are of two classes:

1. The *unorganized*, or *soluble* ferments. Among these are: (*a*) *Diastase*, or *maltin*, formed from the gluten and serving to convert the starch of the seed into glucose. Malt, which is sprouted barley, contains it in abundance, and is used to convert meal (starch) into glucose for fermentation in the manufacture of alcoholic liquors, and in medicine as a digestive agent. The ptyalin of saliva and a pancreatic ferment act like diastase. (*b*) *Pepsin*, of the gastric juice, and (*c*) *Trypsin*, of the pancreatic fluid, both of which serve to convert the albuminoids into peptones; the one in acid, and the other in alkaline solution.

2. *Organized Ferments.* When their spores are carried by the atmosphere or otherwise into a suitable, fermentable liquid, and kept warm (68° to 105° F.), these ferments grow and proliferate with great rapidity, inducing fermentative changes in a few hours. The most important of these ferments are: (*a*) *Yeast* (*torula cerevisiæ*), shown in Fig. 65. This converts glucose into alcohol and carbon dioxide (vinous fermentation). (*b*) *Acetic acid ferment* (*mycoderma aceti*),

converted into glucose. The reaction consists in the addition of H_2O to the molecule, thus:

$$\underset{\text{Starch.}}{C_6H_{10}O_5}+\underset{\text{Water.}}{H_2O}=\underset{\text{Glucose.}}{C_6H_{12}O_6}. \qquad \underset{\text{Cane Sugar.}}{C_{12}H_{22}O_{11}}+\underset{\text{Water.}}{H_2O}=\underset{\text{Glucose.}}{2C_6H_{12}O_6}.$$

When starch and cane-sugar are eaten the digestive ferments (pancreatin and ptyalin) convert them into glucose. The ferment (diastase), developed in a germinating seed, converts the starch into glucose, which is readily assimilated by the sprouting plant.

Glucose is so easily made by boiling cellulin, but more especially starches, with sulphuric acid, that it has become a common adulterant or substitute for cane sugar, especially syrup. This would be harmless but for the fact that the cheap acid used is apt to be contaminated with lead, arsenic, etc.

Glucosides. This class includes a large number of bodies, mostly of vegetable origin, which, though different in other respects, possess one common property, viz: When acted upon by a ferment or a dilute acid they decompose, producing, among other things, glucose. Their chemical constitution is not definitely known, but probably they are ethers of glucose. In the natural state they are generally associated with an albuminoid body capable of acting as a ferment, and gradually decomposing them with the production of glucose. This explains why some fruits (*e. g.* a persimmon) on ripening lose their rough, astringent taste and become sweet. The tannin (*tannic acid*) is converted into glucose. It also explains why a mustard poultice made with hot water is inert. The ferment is coagulated and the glucoside (*myronic acid*) does not decompose.

The glucosides of greatest medical importance are: Amygdalin (bitter almonds), cathartic acid (senna), colocynthin, digitalin,

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commonly called "mother of vinegar," grows on solutions containing alcohol, which it helps to oxidize into acetic acid. (*c*) *Thrush fungus* (*oidium albicans*) grows within the mouths of ill-kept children. It induces a slight alcoholic fermentation. (*d*) *Lactic* and *Butyric* ferments go together, the one preceding and the other closely following. These fermentations occur in intestinal indigestion, and the gas evolved produces flatulent colic.

Putrefaction (the spontaneous decomposition of nitrogenous organized bodies) is accompanied, if not caused, by micro-organisms, usually bacteria. *Decay*, on the other hand, is the gradual decomposition of organic bodies by the slow action of oxygen, and does not depend on living organisms.

elaterin, glycerrhizin (licorice), indican (source of indigo-blue), jalapin, myronic acid (mustard), santonin, tannin, etc. Their names terminate with the syllable "*-in.*"

Nitrogenous Bodies. Many of these (the proteids, etc.), are described in physiology, and need not be treated here.

Bodies of the Ammonia Type. Taking the molecule of ammonia, NH_3, as a basis, and by substituting for one or more atoms of its hydrogen one or more organic radicals or combinations of radicals, we can obtain a large number of interesting and important substances, known as amines, etc. For example:

$$\begin{array}{ccccc} N\left\{\begin{matrix} H \\ H \\ H \end{matrix}\right. ; & N\left\{\begin{matrix} C_2H_5 \\ H \\ H \end{matrix}\right. ; & N\left\{\begin{matrix} C_6H_5 \\ H \\ H \end{matrix}\right. ; & N\left\{\begin{matrix} CH_3 \\ CH_3 \\ CH_3 \end{matrix}\right. ; & N\left\{\begin{matrix} C_6H_5 \\ H \\ C_2H_3O_2 \end{matrix}\right. \\ \text{Ammonia.} & \text{Ethylamine.} & \text{Phenylamine.} & \text{Trimethylamine.} & \text{Acetanilide.} \end{array}$$

Like ammonia, these bodies are alkaline and combine with acids to form salts, appropriating instead of displacing their hydrogen, *e. g.* $NH_3+HCl=NH_4Cl$, ammonium chloride or ammonia hydrochloride; in like manner $NH_2(C_2H_5)+HCl=NH_2(C_2H_5)HCl$, ethylamine hydrochloride. There is perhaps no other field in chemistry so promising to the discoverer of therapeutic agents. Numerous and important as these substances are, we can mention only a few.

$$N\left\{\begin{matrix} C_6H_5 \\ H \\ H \end{matrix}\right.$$

Aniline (Phenylamine) is a colorless liquid, but its compounds (the aniline dyes) are coloring matters of great brilliancy.* They are sometimes contaminated with arsenic used in their manufacture.

$$N\left\{\begin{matrix} CH_3 \\ CH_3 \\ CH_3 \end{matrix}\right.$$

Trimethylamine is sometimes confounded with propylamine. It is a colorless, volatile alkaloid, with an ammoniacal, fishy odor. It is found in many animal and vegetable substances, but obtained from pickled herring. The hydrochloride is the salt used. Dose, ten to fifteen grains.

$$N\left\{\begin{matrix} C_6H_5 \\ H \\ C_2H_3O_2 \end{matrix}\right.$$

Antifebrin (acetanilide). This is a derivative of aniline in which the acetic radical is made to displace an atom of hydrogen. A crystalline, odorless, solid, slightly soluble in warm water, very soluble in alcohol. In

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*__Experiment.__ Dissolve a few drops of aniline in water in two test-tubes. To one add solution of chlorinated lime—a purple color is produced; to the other add some sulphuric acid and potassium chromate mixture—a blue color appears.

doses of five to ten grains, repeated every two or three hours, it is an antipyretic and sedative. It is said not to affect the healthy temperature, but to rapidly lower a fever.

ANTIPYRINE, a derivative of the artificial alkaloid, chinoline, is a white crystalline powder, soluble in water and alcohol, of a slight tarry taste and odor. The hydrochloride is the salt used. In doses of ten to fifteen grains it is a valuable antipyretic and anodyne.

Urea belongs to this class, but is described further on in connection with the urine.

Alkaloids (alkali-like). These bodies are mostly of vegetable origin and bear a close analogy to the preceding, for they are ammonia substitution compounds, alkaline in reaction, and combine with acids to form salts. Of late years chemists have made substances very similar to, if not identical with, some of the natural alkaloids; and the time seems not far distant when our most costly alkaloids will be made cheaply by artificial means. In plants alkaloids are not found free, but combined with some vegetable acid forming a salt. Their salts (except tannates) are usually soluble and intensely bitter; the free alkaloids being much less soluble, are much less bitter. Those alkaloids (as conine and nicotine) that contain no oxygen are liquid; but the great majority of them are white powders.

Alkaloids are so seldom prescribed in the free state that when the simple name of an alkaloid is written in a prescription the druggist puts up its most common salt. The names of alkaloids end in "*-ine*," and are derived from the names of the plants in which they exist or from some characteristic property.

The intense effect alkaloids exert on the animal organism makes them generally the active principles of the drugs in which they are found.* But the active principle of a drug is

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*PROXIMATE AND ULTIMATE PRINCIPLES. Most organic bodies in their natural state are mixtures of several different substances; *e. g.* common resin is a mixture of two or three hydrocarbons, butter of four or five fats, opium and cinchona of several dozen compounds, while brain matter is of such complex composition that no satisfactory analysis has ever been made. These substances that naturally exist, mixed together to form a body, are called its *proximate principles*. The separation of these unaltered from the body and from each other is called *proximate analysis*. In this, different methods must be devised for different substances. For example: Take a piece of vegetable not always an alkaloid. The alkaloids include the majority of our most potent remedies and powerful poisons. Tannin is a common antidote, but most important is the prompt evacuation of the stomach and the intelligent use of physiological antagonists.

The alkaloids, even those of medical interest, are so numerous that to give each separate consideration would cover a great portion of the materia medica. We can mention but a few of the most important.

OPIUM ALKALOIDS—*Morphine* (*Morpheus, the God of Sleep*), *Codeine*, *Narcotine*, etc.—exist in the plant combined with meconic acid. The red color this acid gives with Fe_2Cl_6 is the best test for opium. Morphine gives a dirty blue with Fe_2Cl_6. *Apomorphine* is made by heating morphine and hydrochloric acid to 300° F. It is an emetic.

CINCHONA ALKALOIDS—*Quinine*, *Quinidine*, *Quinoidine*, *Quinicine*, *Cinchonine*, *Cinchonidine*, *Cinchonicine*, etc. The sulphate and bisulphate are the salts generally used.

Test for Quinine. Add chlorine water, shake, and then add aq. ammoniæ; a green coloration is produced.

NUX VOMICA ALKALOIDS—*Strychnine* and *Brucine*. Violent poisons.

Tests. Strychnine dissolved in sulphuric acid and treated with potassium bichromate forms a purple. Brucine treated with nitric acid gives a deep red.

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tissue containing woody fiber, starch, sugar, resin, and volatile oil. The oil is removed by a gentle heat; the resin is dissolved out by alcohol; the sugar by cold water, and the starch by boiling in water, leaving the woody fiber.

The Ultimate Principles of a body are the elements (carbon, hydrogen, etc.) of which it is composed, and the recognition and measuring of these is *ultimate analysis*. This, while requiring careful manipulation, is simple in principle. The body is burned with a full supply of oxygen, converting the carbon into CO_2 and the hydrogen into H_2O. These are collected and weighed, and the quantities of carbon and hydrogen in them are calculated. The amount of oxygen, if any, is determined by subtracting the sum of the carbon and hydrogen from the weight of the original body. For example: 46 grains of alcohol (C, H and O) burned completely makes:

88 grains of CO_2 equivalent to 24 grains of C.
54 grains of H_2O equivalent to 6 grains of H.
46 grains alcohol, minus 30 grains (24+6), = 16 grains of O.

The less common elements, chlorine, nitrogen, sulphur, phosphorus, etc., are determined by special methods.

Liquid Alkaloids—*Nicotine* from tobacco, and *Conine* from hemlock—contain no oxygen. Virulent poisons.

Cocaine, from coca leaves, is a valuable local anæsthetic; much used in minor surgery. The hydrochloride is the salt used.

Ptomaines (πτῶμα, a corpse). This name is given to certain alkaloids formed in animal and some vegetable bodies during putrefaction, and in some pathological conditions during life. Their chemical properties and physiological actions are similar to those of the vegetable alkaloids. Some of them are extremely poisonous. The severe gastro-intestinal irritation and toxic symptoms often following the eating of spoiled meat are due to ptomaines. Recent investigators have succeeded in separating from pyæmic fluids a definite alkaloid, and called it *septicine.* Our knowledge of the ptomaines is as yet very unsatisfactory.

THE URINE.

The urine is a fluid secreted continuously by the kidneys, and is the chief means by which the nitrogenous waste of the body is discharged. A specimen, to be representative, should be a portion of the whole twenty-four hours' urine, for considerable variation in composition and properties may occur during the day. Especially is this true of traces of albumen and sugar. When this is impracticable, that passed before breakfast is generally preferable, because farthest from a meal. When significant variations during the day are suspected, several specimens may be taken at different hours. For *microscopical examination* a few ounces of the urine in a stoppered vial, or better still, in a covered conical glass (Fig. 66) are set aside for several hours until the sediment has settled to the bottom and can be examined.

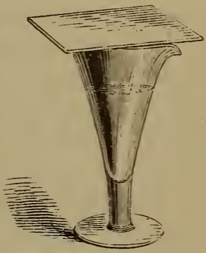


Fig. 66.

Physical Properties. Normal urine is a *transparent, aqueous fluid, of a pale yellow color, characteristic odor, acid reaction, and a specific gravity of* 1020 *when passed in the average quantity of about forty-five fluid ounces in the twenty-four hours.* This description is to be taken with much allowance, for very wide variations occur even in health. With these variations the student must become thoroughly familiar before he is capable of interpreting a specimen. Therefore the physical properties will be considered more particularly.

Quantity. *In health* this depends upon (*a*) the amount of water ingested, and (*b*) its vicarious elimination by the skin, lungs, and bowels. *Pathologically* it is increased in diabetes, also in hysterical conditions associated with convulsions and high arterial pressure, and after the administration of diuretics.

Transparency. Normal urine is not always transparent, nor is transparent always normal. Some degree of opacity may be due to: (*a*) *Mucus*, with some entangled epithelial cells, which may be observed in many specimens of healthy urine, especially of females, because of the larger area of mucous surface in that sex. (*b*) *Urates* (of Na, K, Ca, and Mg), which often form a precipitate in urine, especially when allowed to stand over night in a cold room. The test for this sediment is heat, which quickly dissipates it. (*c*) *Earthy phospates* (of Ca and Mg), which may give an opacity to normal urine, especially if it is alkaline or even weakly acid. The test for this sediment is the addition of a few drops of any acid which promptly clears it up, while heat would only increase it. (*d*) *Fungi* (bacteria, penicillia, sarcinæ, etc.), especially in decomposing urine.

A urine may be abnormally opaque from the above causes, or from the presence of blood or pus. When due to blood or pus the opacity is increased by heat or acids because of the precipitation of albumen always present in *liquor sanguinis* and *liquor puris.*

Fluidity. Healthy urine is never otherwise than an aqueous fluid, flowing and dripping with ease; but in certain diseased conditions, abnormal quantities of mucus, or the presence of pus or fat, especially if the urine be allowed to decompose and become very alkaline, may give rise to viscidity.

Color. *Healthy* variations in color depend mainly upon the amount of water and the consequent degree of concentration or dilution of the solid constituents. Aside from abnormal degrees of the above, *pathological* variations in color may be the result of (*a*) an increase or diminution of the normal coloring matters, as in fevers, etc.; (*b*) the presence of abnormal substances, as biliary and blood coloring matters. Moreover, the urine may be colored after the administration of certain drugs, as senna, santonin, rhubarb, prickly pear, etc.

Odor. When freshly passed, urine has, in addition to its characteristic odor, an aromatic fragrance due to certain volatile ethers. Alkaline urine has an ammoniacal odor, unless the alkalinity be due to fixed alkali, when the smell is fainty and sickening, like that of horses' urine. Diabetic urine exhales a sweetish smell. In certain forms of dyspepsia and liver trouble the odor of the urine is almost pathognomonic. Medicines and certain articles of food often impart a peculiar odor, as turpentine the odor of violets, and asparagus and cauliflower a rank, disgusting smell.

Reaction. Normally the urine of the whole twenty-four hours will average an acid reaction; but great variations occur during the day. Before meals it will have a high degree of acidity, but after eating becomes nearly neutral, or even alkaline. This is due to the ingestion of food which is largely alkaline and to the abstraction of acidulous principles from the blood to form acid gastric juice. It has also been observed that urine passed on rising in the morning is especially acid. This is probably due to the fact that during sleep less carbonic acid is exhaled from the lungs and less perspiration (acid) given off by the skin. The reaction of the urine is important to the physician, as it may favor or prevent the formation of sediments and concretions or irritation of the kidneys and bladder. The acidity of urine is due, not to free acid, but to acid sodium phosphate (NaH_2PO_4) occurring in consequence of carbonic, uric, and hippuric acids, seizing on to a portion of the sodium of the basic phosphate.

Fig. 67. Alkaline fermentation.

An *acid fermentation*, attended with decomposition of mucus and coloring matters and a production of acetic and lactic acids, sometimes occurs in urine that has stood for some time at a moderate temperature. After a while, more quickly in warm weather, the *alkaline fermentation* begins, caused by the development of the micrococcus ureæ. (Pasteur.) The urea is converted into ammonium carbonate, thus:

$$CH_4N_2O+2H_2O=(NH_4)_2CO_3.$$

This gives the urine an ammoniacal odor and alkaline reaction, and it becomes opaque from the precipitation of urate of ammonium and the earthy phosphates and the development of bacteria.

Pus and blood, or vessels tainted with urine previously fermented, greatly hasten this change.

Specific Gravity. Though the average specific gravity is 1020, it exhibits, even in health, great variations, the extremes being 1002 after copious use of water and diuretics, and 1040 after abstinence from fluid and the elimination of water through other means, as profuse perspiration or copious diarrhea. The amount of solids varying but little in health, fluctuations in specific gravity are due mainly to variations in the amount of water, and, as long as the inverse proportion between specific gravity and volume of urine is preserved, variations need cause no alarm.

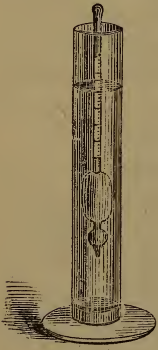
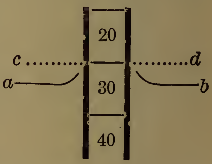


Fig. 68.

Specific gravity is usually measured by an instrument called a hydrometer or *urinometer* (Fig. 68), which is a hollow glass float, weighted with mercury and having a long, graduated neck. The graduation begins above at 1000, because the heavier the urine the less deeply will the instrument sink and the further the neck will protrude from the surface. It is well to test a new urinometer by immersing it in water at 60° F. (15.5° C.), into which it should sink to 0 or 1000 on the scale. Urinometers are usually provided with a cylinder or jar, as shown in the figure, but a large test-tube will answer. This is about three fourths filled; the urinometer is then introduced, and when still, the specific gravity is read off. The cylinder or test-tube should not be too narrow, lest the urinometer be attracted to and catch against the sides, and not rise as high or sink as low as it should. The fluid being attracted up around the stem, the reading should be made not along the line *c d*, as in the diagram, but *a b*, which represents the true level of the liquid. We may approximate the amount of solids in any urine by doubling the last two figures of the specific gravity, which will give the per cent. Thus, if a urine be of specific gravity 1020, doubling the last two figures gives .040, or 4 per cent. If the daily volume be fifty ounces, 4 per cent of this is two ounces, which will represent the quantity of solids passed daily.



Chemical Constituents. The average composition of a thousand parts of urine is about as follows:

Organic.	Water	950.00
	Urea	26.20
	Kreatine and kreatinine	.80
	Urates of sodium and potassium	1.45
	Hippurates of sodium and potassium	.70
	Mucus and coloring matters	.35
Inorganic.	Phosphates of sodium and potassium	3.75
	Phosphates of calcium and magnesium	.90
	Chlorides of sodium and potassium	12.55
	Sulphates of sodium and potassium	3.30
		1000.00

Pathologically there may be present also albumen, glucose, blood, bile, etc., besides various sediments.

Urea—CH_4N_2O. This is the most constant and abundant organic constituent of the urine, and, being the main nitrogenous excretion of the body, it is the index of nitrogenous waste, whether of food or tissue. Its average amount is about one ounce *per diem*.

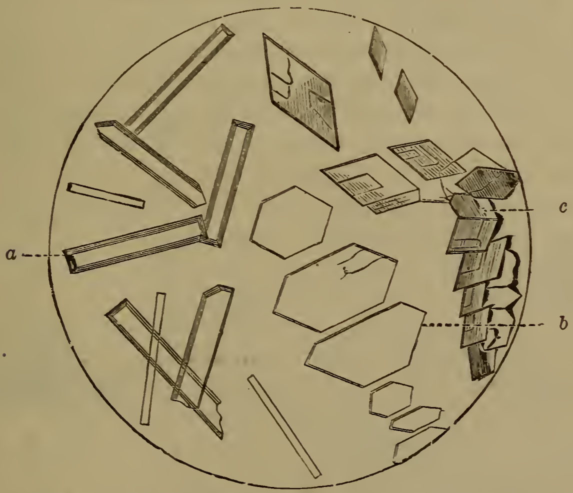


Fig. 69. (*a*) Urea; (*b*) hexagonal plates; and (*c*) smaller scales, or rhombic plates of urea nitrate.

Urea crystallizes in colorless prisms, very soluble in water, and behaves like an alkaloid, combining readily with nitric and oxalic acids to form salts. Both of these salts may be precipitated in colorless plates from concentrated urine by adding nitric or oxalic acid. (Fig. 69, *c.*) In the course of many diseases it is important to estimate the amount of urea excreted day by day. A rough estimate may be based on the specific gravity. For, since urea is the largest solid ingredient in urine, it follows that if sugar be absent, albumen in small amount or removed, and the amount of chlorides normal, variations in specific gravity must be due mainly to variations in amount of urea.

The exact methods most generally employed consist in decomposing the urine, by means of chlorinated soda, into nitrogen and carbon dioxide, and measuring either the volume of gas evolved or the specific gravity lost by the decomposition.

Davy's Method. A graduated tube closed at one end is one third filled with mercury. A measured quantity of the urine (a dram or half dram, according to capacity of tube) is then added, and the tube is next filled to the brim with liquor sod. chloratæ. Closing the opening with the thumb, the tube is inverted over a strong solution of common salt in a dish. (Fig. 70.) The mercury runs out and the salt water rises to take its place, while the urine and soda mixture, being lighter, remain in the upper part of the tube. Here the gas from the decomposing urea collects. The decomposition is complete in three or four hours, when the amount of the gas may be read off by the graduations upon the tube, every cubic inch representing .64 grain (or 1 cubic centimeter representing 2.5 milligrams) of urea.

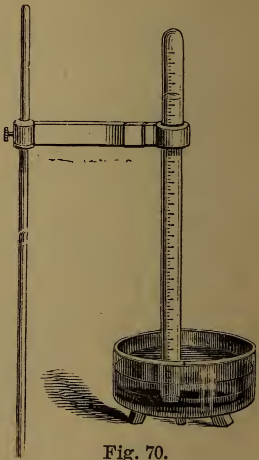


Fig. 70.

Fowler's Method, based on the loss of specific gravity, is easier of application. The specific gravity of the urine is carefully determined as well as that of the liq. sodæ chloratæ (U. S. P.) to be used. One volume of the urine is mixed with exactly seven volumes of the liq. sod. chl. and set aside for two hours, or until effervescence ceases. The specific gravity is again taken. As the reaction begins immediately on mixing the fluids, the specific gravity of the mixture must be calculated. This is done by adding to the specific gravity of the urine seven times that of the liq. sod. chl. and dividing the sum by eight. Each degree of difference in specific gravity of the mixture before and after the decomposition represents three and a half grains of urea to the fluid ounce of the day's urine.

Example:

	Ounces.
Quantity of urine in twenty-four hours..........	46
Sp. gr. of the urine....................................	1020
Sp. gr. liq. sod. chloratæ.............................	1042
(Calculated) Sp. gr. mixture $(\frac{1042\times 7+1020}{8}=)$.	1039.2+
(Actual) Sp. gr. mixture after reaction...........	1036.2

1039.2—1036.2=3; $3\times 3\frac{1}{2}=10\frac{1}{2}$ grs. of urea to the ounce of urine; $10\frac{1}{2}\times 46=483$ grs. of urea passed in twenty-four hours.

Kreatine and *Kreatinine,* substances closely allied to urea, exist in urine in such small amounts as to be of no practical significance, and need only to be mentioned in this connection.

Uric Acid is found in the urine of carnivora: in that of herbivora it is largely replaced by an analogous substance—hippuric acid. Gout is characterized by an increased production of uric acid, and the so-called "chalk-stone" deposit in joints during that disease is sodium urate. Free uric acid is so very insoluble that whenever it exists in urine it is always a precipitate. It ap-

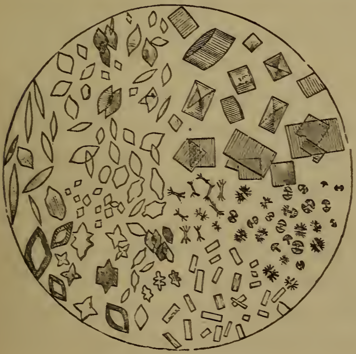


Fig. 71. Uric Acid.


Fig. 72. Uric Acid.

pears as minute reddish grains, which under the microscope are seen to be modifications of rhombic crystals, always stained with the coloring matter of the urine. They often deviate widely from the typical rhomb, as is shown in Figs. 71 and 72, but an experienced eye will readily recognize them. Normally, uric acid, as soon as formed, unites with the alkaline bases to form *urates.* These are very soluble in warm water, but more sparingly so in cold. Therefore a urine, though clear when freshly passed and warm, may exhibit a copious precipitate upon becoming cold, as on a winter night. This precipitate is easily recognized by its dissolving upon warming. Urates of sodium and magnesium generally appear under the microscope as amorphous powders in moss-like aggregations, but occasionally as bundles of small needles, as shown in Fig. 73. The urate of ammonium, a result of the alkaline fermentation, occurs as opaque, brown spherules, smooth or with spiculæ like a thorn-apple. (Fig. 67.) The acid urates are less soluble than the normal, and often precipitate when the urine is very acid or when an acid is added, as in the nitric-acid test for albumen.

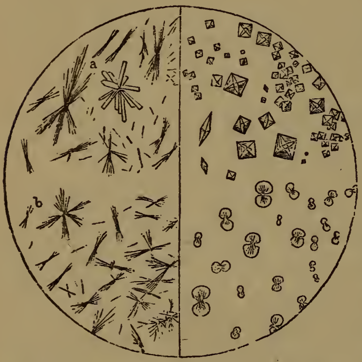


Fig. 73. Urates. Oxalate of Calcium.

The *murexid test* for uric acid and the urates is one of great beauty. Place some of the sediment in a porcelain dish, add a drop or two of nitric acid, and carefully evaporate almost to dryness. If a few drops of ammonia be added, it assumes a beautiful purple color.

Coloring Matters. Our unsatisfactory knowledge of these substances and their clinical significance is to be regretted, since some of them possess an importance next to albumen and sugar. The existence of at least two distinct substances have been demonstrated:

1. *Urobilin* (*urohœmatin*), derived from the coloring matter of the bile, and hence indirectly from the coloring matter of the blood.

2. *Uro-indican* (*uroxanthin*), a substance closely related to, if not identical with, the glucoside *indican*, and, like that substance, it is capable of conversion into indigo-blue.

To estimate the coloring matters, put some urine in a beaker and render it strongly acid with nitric or hydrochloric acid. Let it stand six hours for the pigments to be liberated. Then note the depth of color by transmitted light.

PHOSPHATES. The phosphates are derived mainly from the food, but to some extent also from oxidation of phosphorized tissues:

1. *Earthy Phosphates* (Ca and Mg). Being soluble only in acid solutions, the earthy phosphates are precipitated when the urine is made or becomes alkaline. Furthermore, being less soluble in warm than cold urine, heat often precipitates them, as in the heat test for albumen. Deposits of calcium and magnesium phosphates are generally amorphous, and may be distinguished from the amorphous urates, (*a*) by absence of color and not gathering in mossy forms; (*b*) by a drop of acetic acid added to the sediment on a glass slide under the microscope—phosphates dissolve, while urates gradually lose their base and assume the characteristic forms of uric acid. In ammoniacal urine (alkaline fermentation) the ammonio-magnesian phosphate ($MgNH_4PO_4$), the so-called *triple phosphate*, is formed and deposited in large prismatic, coffin-lid crystals; sometimes, also, in ragged stellate or arborescent crystals, resembling those of snow. (Fig. 74.) In cases of cystitis this may occur within the bladder; hence other calculi often have one or more white layers of the mixed phosphates.

Fig. 74. Triple Phosphate.

2. *Alkaline Phosphates.* These constitute the greater portion of the phosphates, and are made up mainly of *acid sodium phosphate*, with traces of potassium phosphate. Being very soluble, they never form a precipitate.

Magnesian Test. The phosphates are best detected and estimated by precipitation with a solution composed of magnesium sulphate, ammonium chloride, and aq. ammoniæ, each one part, and water eight parts. If the precipitate be thick and creamy, the phosphates are abnormally increased; if it be milky, they are normal, and if translucent, diminished.

Chlorides. These consist almost entirely of sodium chloride, the quantity depending mainly on what is taken in with the food. However, the chlorides are diminished or even disappear from the urine in many fevers, especially in pneumonia, much being eliminated by the sputa. Their reappearance in the urine is often the earliest indication of convalescence. Hence their detection and estimation are important.

Silver-Nitrate Test. First add a few drops of nitric acid to prevent the precipitation of the phosphates. Then, on adding silver nitrate solution, only the chlorides will fall as a white precipitate of chloride of silver. If the precipitate be in curdy masses the chlorides are not diminished; if only a milkiness be produced, they are greatly diminished; and, if no cloudiness, they are entirely absent.

Sulphates. These consist mainly of sodium sulphate, with a little of the potassium salt. They are derived principally from the food, and in small amount from oxidation of albuminoid sulphurized tissues, especially in fevers. They are detected and estimated by precipitation with *barium chloride* or *nitrate*, first adding a little nitric or hydrochloric acid to hold the phosphates in solution. If the precipitate be creamy the sulphates are increased; if milky, normal, and if translucent, diminished.

Albumen. Under this head are included various proteid substances which, not being osmotic, appear in urine only in pathological conditions and functional disturbances. Many of the specific fevers, as pneumonia, typhoid, and diphtheria, produce albuminuria. Albuminous urine is apt to be of diminished transparency from presence of tube casts, fat granules, epithelial cells, etc., and filtering is often necessary before applying the tests.

Heat Test. A test-tube is one third filled with the suspected urine and held in the flame of a spirit lamp, or over the chimney of an ordinary lamp, until it boils. If an opacity occurs it must be either *albumen* or *earthy phosphates.* If earthy phosphates, it clears up on addition of nitric acid, but if albumen, it is slightly increased.

Nitric-Acid Test. This consists in underlaying the urine with nitric acid. Take a test-tube one fourth full, and, holding it aslant, gently pour in an equal volume of the acid, allowing it to trickle down the inside of the tube and pass beneath the urine. Or the acid may be put in first and the urine added afterward. An opacity at the junction of the two liquids is either *albumen* or the *urates.* If urates, it clears up on heating, but if albumen, it is permanent.

Either the heat or nitric-acid test, singly, is unsatisfactory, but both performed together are conclusive. However, the following sources of error should be borne in mind: (*a*) If the urine be very alkaline and the amount of albumen small, heat will cause no opacity; (*b*) if only a drop or two of nitric acid be added, it may hold a small quantity of albumen in solution; (*c*) urea may be precipitated from a concentrated urine by nitric acid, but heat dissolves it; (*d*) decomposed urates containing ammonium carbonate effervesce on addition of an acid; (*e*) often after taking turpentine, copaiba, etc., nitric acid precipitates resin in yellowish flakes, redissolved on addition of alcohol.

Other Tests. Many other substances, as alcohol and certain acids and mineral salts, coagulate albumen and are used as tests for that substance. But they are less used, less convenient, and no more accurate and conclusive than the two already given. Among them may be mentioned (*a*) picric acid, (*b*) potassio-mercuric iodide (KI 50 grs.—$HgCl_2$ 21 grs.), (*c*) sodium tungstate, (*d*) potassium ferrocyanide. These added in saturated solution form white clouds with albumen, provided the urine is first acidulated with citric acid. Strips of filter paper steeped in these chemicals and dried, or pellets, are sometimes carried for use at the bedside.

Quantitative Estimation. It is often very important that we should be able to compare the quantity of albumen in the urine from day to day. The accurate method is by precipitation with acetic acid and boiling, separation by filtration, drying, and weighing by delicate balances, the filter having been previously weighed. But as this involves too much time for the busy practitioner, we must use an approximative method. The easiest way is to precipitate the albumen by heat and nitric acid, set it aside for twelve hours or until next visit, and then note the proportion of volume occupied by the precipitate—one fourth, one eighth, a trace, etc.

Sugar (*Glucose*). It has been proven (Dr. Pavy, 1879) that healthy urine contains traces of glucose, but quantities of clinical significance, and appreciable by the ordinary tests, are present only in glycosuria or diabetes, a pathological condition associated with some disturbance of the glycogenic function of the liver.

High specific gravity in a urine pale and copious, suggests sugar. Before testing, albumen, if present, should be removed by boiling and filtration.

Fermentation Test. Two vials—one for comparison, the other for fermentation—are partly filled with the urine. Into one is put a bit of baker's yeast about the size of a pea. Both vials are loosely plugged with some pervious material, as cotton, and set aside where they will keep warm (60° or 70° F.) until next day or next visit. If sugar be present, fermentation will occur in the vial treated with yeast and CO_2 bubbles up and passes off through the cotton plug, and on taking the specific gravity of each, there will be a difference due to the loss of sugar in the vial fermented.

Alkali Test. Boil the urine with liquor potassæ or sodæ, and if glucose be present it will be oxidized and form a molasses-like coloration, the depth of which indicates the amount of sugar present. On adding nitric acid a molasses-like odor is developed and the coloration discharged.

Alkali-Copper Test. This depends on the power glucose has of reducing the cupric to the cuprous oxide. There are several methods of performing this test:

(1) *Trommer's.* A drop or two of a weak (about 1 to 30) solution of cupric sulphate is added to an inch of urine in a test-tube, and then an equal bulk of liquor potassæ or sodæ. Immediately there falls, in addition to the earthy phosphates, a bluish precipitate. If sugar is present, this precipitate dissolves on agitation, forming a blue solution, which, on boiling, deposits a yellow, orange, or red precipitate of cuprous oxide. (See p. 79.)

(2) *Fehling's.* This differs from Trommer's in the addition of tartaric acid or some tartrate to dissolve the blue precipitate.

Furthermore, the ingredients are in definite proportion, so as to make the solution available for quantitative analysis. Below are given the two formulæ in general use, one in the French and the other in the English measures:

	Fehling's Solution.	Pavy's Solution.
Cupric Sulphate	34.64 grams.	320 grains.
Potassium Tartrate	173.20 grams.	640 grains.
Caustic Potash	80.00 grams.	1280 grains.
Water	1 liter.	20 ounces.

On standing a long time this solution is apt to spoil, the tartaric acid being converted into racemic acid, which, like glucose, will deoxidize the cupric oxide. Hence, it is best to make the solution in two separate parts, the cupric sulphate with one half the water and the tartrate and caustic potash with the other half. For use, mix equal parts, forming Fehling's solution fresh. A convenient amount should be put in a test-tube and boiled alone for a few seconds. If it remains clear it is good, and the urine may then be added gradually. Either immediately, or when the heat is reapplied, if sugar be present the reddish precipitate will appear. Heat should not be applied longer than a minute, for prolonged boiling can cause the reduction of the copper oxide by various other organic substances found in urine.

(3) *Haines*' differs from Fehling's in that glycerine is used instead of the tartrate, and the solution does not spoil.

Alkali-Bismuth Test. (1) To some urine in a test-tube add a pinch of bismuth subnitrate and then an equal volume of liquor potassæ. Boil about two minutes. If sugar be present, the bismuth will be reduced and deposited as a black metallic mirror on the sides and bottom of the tube. (2) A bismuth test solution corresponding to Fehling's is made by warming a scruple each of bismuth subnitrate and tartaric acid in two ounces of water, and adding liquor potassæ till a clear solution is obtained. This boiled with a urine containing glucose gives the black bismuth precipitate.

The elements of the foregoing tests put up in pellets and tablets, while more convenient, are less reliable and spoil sooner than the solution.

Picric-Acid Test. This is an extremely delicate test for glucose, and has the practical advantage of being as good a test for albumen. To the suspected urine add an equal volume of a saturated solution of picric acid. A cloudy precipitate indicates albumen. Next add a few drops of liquor potassæ and warm gently. A *deep red* color indicates sugar, though a lighter coloration may occur in urine free from glucose.

Quantitative—(1) *Fermentation.* Each degree of specific gravity lost in fermenting represents one grain of sugar to the ounce of the twenty-four hours' urine.

(2) *Fehling's. Two hundred minims of the solution is decolorized by one grain of sugar.* Two hundred minims (grains) of the test solution are measured off into a small flask, diluted with twice its bulk of water, and gently boiled. (Fig. 75.) A graduated burette (also shown in figure) is then filled to zero with the urine. To the boiling test solution the urine is added drop by drop till the blue color is discharged. By the graduations on the burette the quantity of urine added is easily read. As that represents one grain of sugar, the amount of sugar in the entire urine is easily calculated.

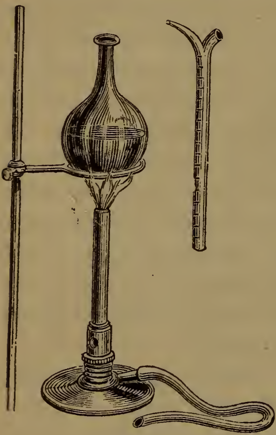


Fig. 75.

Blood gives to urine a smoky hue, or even a dark brown color. Hæmaturia (blood in urine) may occur as the result of (*a*) some disease or injury in the genito-urinary tract, as acute nephritis, calculus, parasites, cancer, wounds, etc.; (*b*) a depraved condition of the blood, as in scurvy, purpura, and eruptive fevers; (*c*) a disturbance of the renal circulation, as in mental emotions, malarial paroxysms, and cardiac obstructions.

If the urine be acid, the blood corpuscles retain their shape for several days and are easily recognized by the microscope. They appear as amber-colored, biconcave disks, either single or laid in rows, like piles of coin. Owing to the biconcavity of the corpuscles their centers and peripheries alternate in brightness and shadow, as the object-glass is made to approach or recede. Their color and smaller size also serve to distinguish them from pus corpuscles. In doubtful cases a minute drop of blood, taken from the finger with a needle, may be used for comparison. After urine containing blood has stood for some time, the corpuscles lose their regular outline and become shriveled and angular. (See *a* in figure.) If the corpuscles be disintegrated and dissolved we must *test* for *blood-coloring matters.*

The spectroscope offers the best means for their detection, but as physicians are seldom provided with that instrument, the following is the *test:* Place the urine in a test-tube and shake up with equal volumes of tincture of guiacum and ozonized ether or old oil of turpentine. If blood-coloring matters are present, the precipitated resin is *blue*, instead of a dirty greenish yellow.

Fig. 76. Blood Corpuscles.

Bile. Urine containing bile is yellow, froths on shaking, and a rag dipped in it and dried is permanently yellow.

1. *Test for Bile-coloring Matters.** Underlay the urine with yellow nitric acid or a mixture of nitric and sulphuric acids; or the urine and acid may be placed adjacent on a white plate. In either method there occurs, at the junction of the liquids, a play of colors, *green* being prominent and characteristic.

2. *Test for Bile Acids.* Add a few grains of cane sugar or glucose to the urine and underlay it with sulphuric acid. At the junction of the liquid a *reddish-purple* color appears. As other substances than the bile acids may produce this reaction, we must, in cases of doubt, evaporate the urine to dryness, extract with alcohol, precipitate with ether, and redissolve in distilled water, and then apply the test as above.

Leucin and Tyrosin occur only in bile urine, for they attend destructive liver disease, especially acute, yellow atrophy and

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*Bilirubin oxidizes so easily that icteric urine often gives only the green coloration, or, if kept long, fails to respond at all. Hence, if fresh icteric urine can not be obtained and bile urine must be prepared for demonstration, fresh bile from a recently killed animal, and not the inspissated, must be used.

phosphorus-poisoning. They form yellowish crystalline deposits (Fig. 77)—leucin as spherules, with concentric striæ, and tyrosin as sheaf-like bundles of fine needles.

Fig. 77. Leucin Spherules and Tyrosin Needles.

Calcium Oxalate *occurs* in extremely small amounts in normal urine, but more abundantly in the so-called oxalic diathesis and in certain forms of dyspepsia, or after eating rhubarb or other things containing it. If persistently present it may form a (mulberry) calculus. It occurs in both acid and alkaline urine, and always as a light delicate precipitate, which under high powers is seen to consist of small, brilliant octahedral crystals, but sometimes dumbbells. (Fig. 73.) In certain aspects the smaller octahedra appear as squares crossed by two bright diagonal lines.

Fig. 78.

Carbonate of Calcium. Hippuric Acid.

Carbonate of Calcium is a very rare deposit in human, but abundant in the urine of cattle. It occurs in small spherules, sometimes coalescing; acetic acid dissolves it with effervescence.

Hippuric Acid (*Horse-uric Acid*) largely replaces uric acid in the urine of herbivorous animals, and to some extent in that of man, especially after a vegetable diet. It occurs in pointed, four-sided prisms and acicular crystals, insoluble in acetic acid but soluble in alcohol. (Fig. 78.)

FAT in such quantities as to float on the urine generally comes from the introduction of a catheter or from foreign admixture.

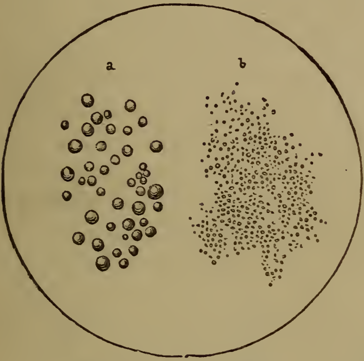

Fig. 79. Fat Globules.

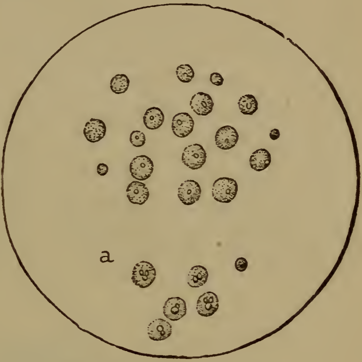

Fig. 80. Pus Corpuscles.

Fatty degeneration of kidney, or leakage of a lymph vessel, or the opening of an abscess into the urinary tract may cause fat in the urine. It occurs as minute, highly refracting globules of various sizes (see *a* in Fig. 79), but sometimes, especially in chylous urine, in more intimate emulsion (as at *b*), the globules appearing under the microscope as mere specks. Fat may be recognized by its dissolving on addition of ether.

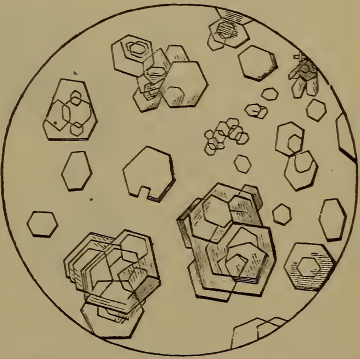

Fig. 81. Cystin.

CYSTIN is a rare urinary sediment, a yellowish deposit of hexagonal plates (Fig. 81), not dissolved by heat or acetic acid but readily by ammonia. It is a highly sulphurized body whose formation in the system is obscure. It sometimes forms calculi.

MUCUS AND PUS. Mucus is a normal constituent of urine. It is a transparent fluid, and would be invisible but for the mucus corpuscles, epithelium, and other sediments entangled in it.

Though closely related to albumen, mucin is coagulated by acetic acid and not by heat. Mucus is increased by irritation of the urinary tract, but as inflammation supervenes albumen appears and the urine is purulent. The mucus and pus corpuscles present the same appearance under the microscope as other leucocytes, viz., rounded, colorless, very granular cells, a little larger than red blood corpuscles. (Fig. 80.) If the urine be greatly diluted, or, better, treated with acetic acid, the cells swell up, lose their granular appearance, become transparent, and show their nuclei (*a* in Fig. 80). The pus cell oftener than the mucus corpuscle has more than one nucleus. Pus may be distinguished from mucus: (1) It is always attended with albumen; (2) (Donne's test) treated with an alkali it forms a gelatinous mass.

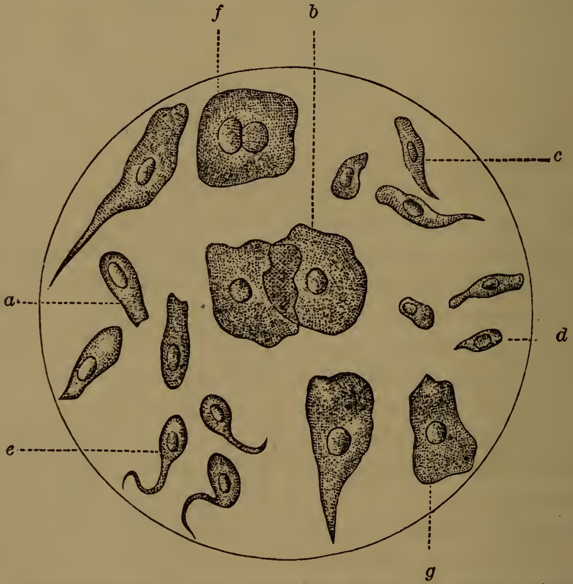

Fig. 82. (*a*) Epithelium from the human urethra; (*b*) vagina; (*c*) prostate; (*d*) Cowper's glands; (*e*) Littre's glands; (*f*) female urethra; (*g*) bladder.

Epithelium in the urine may come from any part of the genito-urinary tract. The accompanying cut shows the typical forms of cells coming from various situations. It is generally impossible to locate the origin of an epithelial cell beyond the vagina and bladder, for their distinctive differences, but slight at best, are rendered still fainter by maceration in the urine. *Renal* epithelium comes from the uriniferous tubules, and are rounded and granular, and, unlike pus cells, they show their nuclei without acetic acid. They are usually associated with albumen and tube casts (Fig. 83), and therefore point to kidney disease.

TUBE CASTS. In hemorrhage from or inflammation of the kidney the urine usually contains microscopic casts or moulds of the uriniferous tubules formed by exudation into the tubule of coagulable material, which afterward contracts, becomes loose, and is washed out with the urine. As they imbed and bring away epithelial cells, granular matter, fat globules, blood discs, etc., they are a valuable index to the condition of the tubules. (1) *Epithelial* casts (see upper portion of figure) are those bearing renal epithelium. They indicate desquamative nephritis.

Fig. 83. Epithelial Cells and Tube Casts.

Fig. 84. Spermatozoa.

(2) *Hyaline* casts (shown in left-hand part of figure) are transparent and comparatively free from entangled material. They come from tubules whose epithelium is sound and adherent or from those bereft of epithelium. In the latter case they are more solid in appearance (*waxy casts*) and indicate serious nephritis. (3) *Granular* casts are opaque from presence of granular *debris*. (4) *Fatty* casts (see larger cast in figure) are such as carry oil globules, either free or contained in epithelial cells. They are proof of fatty degeneration of the kidney. (5) *Blood* casts contain blood corpuscles, and show that the hematuria is of renal origin.

SPERMATOZOA occur in urine as a result of spermatorrhea, nocturnal emissions, or coitus. They are liable to escape observation, for they subside slowly, and are very small and transparent. Under a high power they are seen to consist of a small oval cell with a tail-like prolongation. Their tadpole-like appearance is shown in Fig. 84. They are motionless in urine, and remain for days unaltered.

MICRO-ORGANISMS. Urine being a solution of organic matters becomes as soon as voided a ready medium for the growth of the lower forms of life, the germs of which get in from the air or unclean vessels. Besides various others we may mention: (1) *Yeast fungus* (shown on page 105) is seen during its sporule stage as transparent oval cells, sometimes arranging themselves in branches. It grows only in saccharine urine, though spores closely resembling it are seen in acid urine containing neither sugar or albumen. (2) *Sarcina* is a fungus seldom found in urine but more frequently in matters vomited during certain diseases of the stomach. The cells are arranged in cubes, resembling bales bound with cross-bands. The sarcinæ shown at *a* in figure are from the urine, those at *b* from vomited matters.

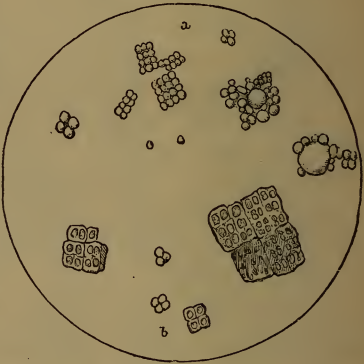

Fig. 85. Sarcina Ventriculi.

1. *Bacteria* (*little rods*). This is the general term given to the minute moving organisms invariably present in putrefying animal and vegetable matter. They consist of simple cells filled with a colorless fluid and presenting several varieties of form: (*a*) *Micrococci* appearing as trembling points, distinguished from other particles by their progressive motion; (*b*) *Rods* about the length of the diameter of blood discs, sometimes at rest, but usually vibrating across the field; (*c*) *Vibriones*, consisting of several rods joined together and moving with greater rapidity; and (*d*) *Zoöglea*, aggregations of bacteria held together by gelatinous material and resembling masses of amorphous urates or phosphates. These various forms are shown in Fig. 67, page 113. Bacteria not only

cause decomposition outside, but may set it up in urine while yet within the bladder, provided they be introduced from without. This may be done by dirty catheters and sounds, or they may work their way down the urethra in the pus of a gleet. The ammoniacal fermentation thus set up soon induces cystitis.

Extraneous bodies, such as hair, wool, or fragments of feathers, are often found in urinary deposits, and ludicrous mistakes have been made by observers not on their guard for such casual admixtures.

SEDIMENTS. *The chemical examination* of unorganized urinary sediments is generally an easy matter, for they usually consist of urates, phosphates, calcium oxalate, or uric acid. Warm the sediment with the supernatant urine, it dissolves—*urates.* If not, warm with acetic acid, it dissolves—*phosphates.* If not, warm with hydrochloric acid, it dissolves—*calcium oxalate.* If not, it is *uric acid*, which may be confirmed by the murexid test.

URINARY CALCULI. Urinary calculi (*calculus*, a pebble) are composed of urinary sediments which have gathered around some nucleus (usually calcium oxalate or uric-acid crystals, or some foreign body) within the bladder, and being slowly deposited, particle upon particle and layer upon layer, the concretion becomes as hard as stone. The concretion often consists of successive layers of different sediments deposited during varying conditions of the urine.

The qualitative analysis of calculi is easy. Saw the stone through the middle and see whether it be composed of the same material throughout or of successive layers of different sediments. If the former, take the sawdust; if the latter, chip off a specimen from a single layer. But this should be pulverized *very fine* (for it is dissolved much less readily than fresh sediments), and then test by means of heat acetic and hydrochloric acids, just as other sediments.

The following method is easier in practice:

I. Heat to redness on a piece of platinum foil. If no residue, see II; if a residue, see III.

II. To a fresh portion apply the murexid test. If it responds it is *ammonium urate* or *uric acid;* if it does not respond it is *cystin* or *xanthin*, see IV.

III. To the residue, when cool, add hydrochloric acid. If it effervesces it is an *oxalate* or *urate*, which may be determined by the murexid test; if it does not effervesce it is a *phosphate.*

IV. Dissolve some of the powder in nitric acid. If the solution is yellow it is *xanthin;* if dark brown it is *cystin.*

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1 gram = 15.434 grains.
1 cubic centimeter = 0.061 cubic inch = 0.27 ʒ.
1 liter = 61 cubic inches = 33.8 ℥.
1 millimeter = $\frac{1}{25}$ inch.

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