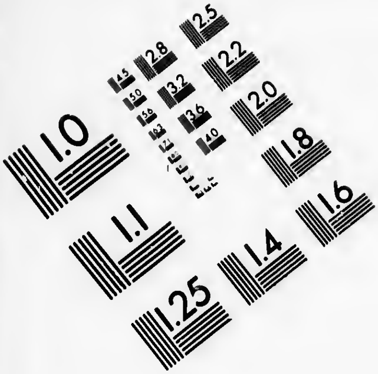


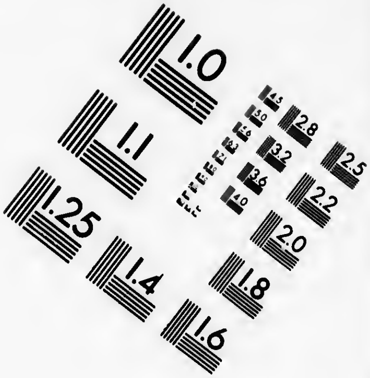
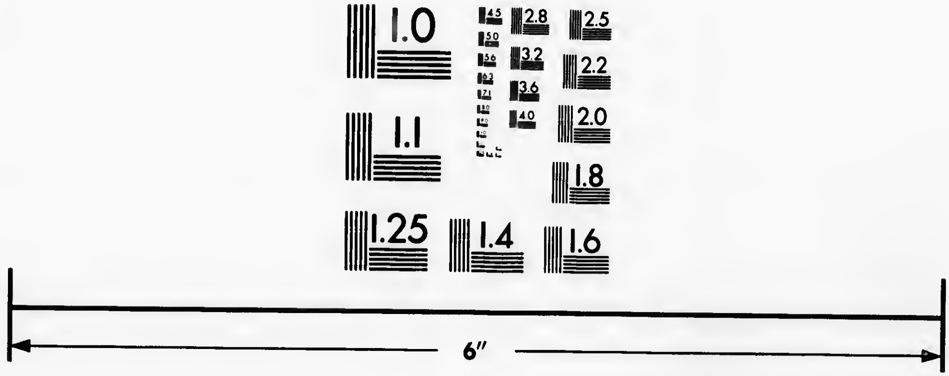
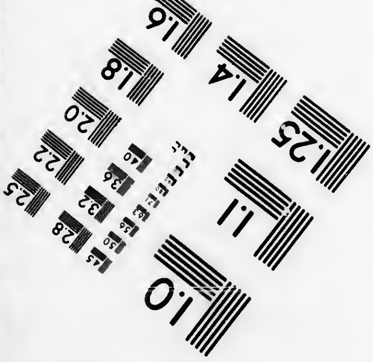


IMAGE EVALUATION
TEST TARGET (MT-3)





Photographic
Sciences
Corporation

23 WEST MAIN STREET
WEBSTER, N.Y. 14580
(716) 872-4503

Technical and Bibliographic Notes / Notes techniques et bibliographiques

The Institute has attempted to obtain the best original copy available for filming. Features of this copy which may be bibliographically unique, which may alter any of the images in the reproduction, or which may significantly change the usual method of filming, are checked below.

L'Institut a microfilmé le meilleur exemplaire qu'il lui a été possible de se procurer. Les détails de cet exemplaire qui sont peut-être uniques du point de vue bibliographique, qui peuvent modifier une image reproduite, ou qui peuvent exiger une modification dans la méthode normale de filmage sont indiqués ci-dessous.

- ☑ Coloured covers/ Couverture de couleur
- ☑ Covers damaged/ Couverture endommagée
- ☐ Covers restored and/or laminated/ Couverture restaurée et/ou pelliculée
- ☐ Cover title missing/ Le titre de couverture manque
- ☐ Coloured maps/ Cartes géographiques en couleur
- ☐ Coloured ink (i.e. other than blue or black)/ Encre de couleur (i.e. autre que bleue ou noire)
- ☐ Coloured plates and/or illustrations/ Planches et/ou illustrations en couleur
- ☐ Bound with other material/ Relié avec d'autres documents
- ☑ Tight binding may cause shadows or distortion along interior margin/ La reliure serrée peut causer de l'ombre ou de la distorsion le long de la marge intérieure
- ☐ Blank leaves added during restoration may appear within the text. Whenever possible, these have been omitted from filming/ Il se peut que certaines pages blanches ajoutées lors d'une restauration apparaissent dans le texte, mais, lorsque cela était possible, ces pages n'ont pas été filmées.

- ☐ Coloured pages/ Pages de couleur
- ☐ Pages damaged/ Pages endommagées
- ☐ Pages restored and/or laminated/ Pages restaurées et/ou pelliculées
- ☑ Pages discoloured, stained or foxed/ Pages décolorées, tachetées ou piquées
- ☐ Pages detached/ Pages détachées
- ☑ Showthrough/ Transparence
- ☑ Quality of print varies/ Qualité inégale de l'impression
- ☐ Continuous pagination/ Pagination continue
- ☑ Includes index(es)/ Comprend un (des) index

Title on header taken from:/
Le titre de l'en-tête provient:

- ☐ Title page of issue/ Page de titre de la livraison
- ☐ Caption of issue/ Titre de départ de la livraison
- ☐ Masthead/ Générique (périodiques) de la livraison

☐ Additional comments:/ Commentaires supplémentaires:

This item is filmed at the reduction ratio checked below/
Ce document est filmé au taux de réduction indiqué ci-dessous.

10X 14X 18X 22X 26X 30X

☐☐☐☐☐☐☐☐☐☐☐☑☐☐☐☐☐☐☐☐☐☐☐

12X 16X 20X 24X 28X 32X

The copy filmed here has been reproduced thanks to the generosity of:

National Library of Canada

The images appearing here are the best quality possible considering the condition and legibility of the original copy and in keeping with the filming contract specifications.

Original copies in printed paper covers are filmed beginning with the front cover and ending on the last page with a printed or illustrated impression, or the back cover when appropriate. All other original copies are filmed beginning on the first page with a printed or illustrated impression, and ending on the last page with a printed or illustrated impression.

The last recorded frame on each microfiche shall contain the symbol → (meaning "CONTINUED"), or the symbol ▼ (meaning "END"), whichever applies.

Maps, plates, charts, etc., may be filmed at different reduction ratios. Those too large to be entirely included in one exposure are filmed beginning in the upper left hand corner, left to right and top to bottom, as many frames as required. The following diagrams illustrate the method:

L'exemplaire filmé fut reproduit grâce à la générosité de:

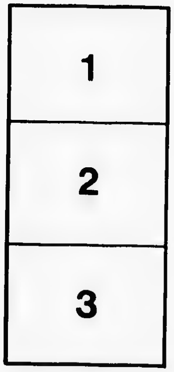
Bibliothèque nationale du Canada

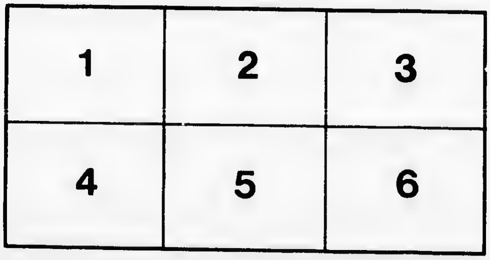
Les images suivantes ont été reproduites avec le plus grand soin, compte tenu de la condition et de la netteté de l'exemplaire filmé, et en conformité avec les conditions du contrat de filmage.

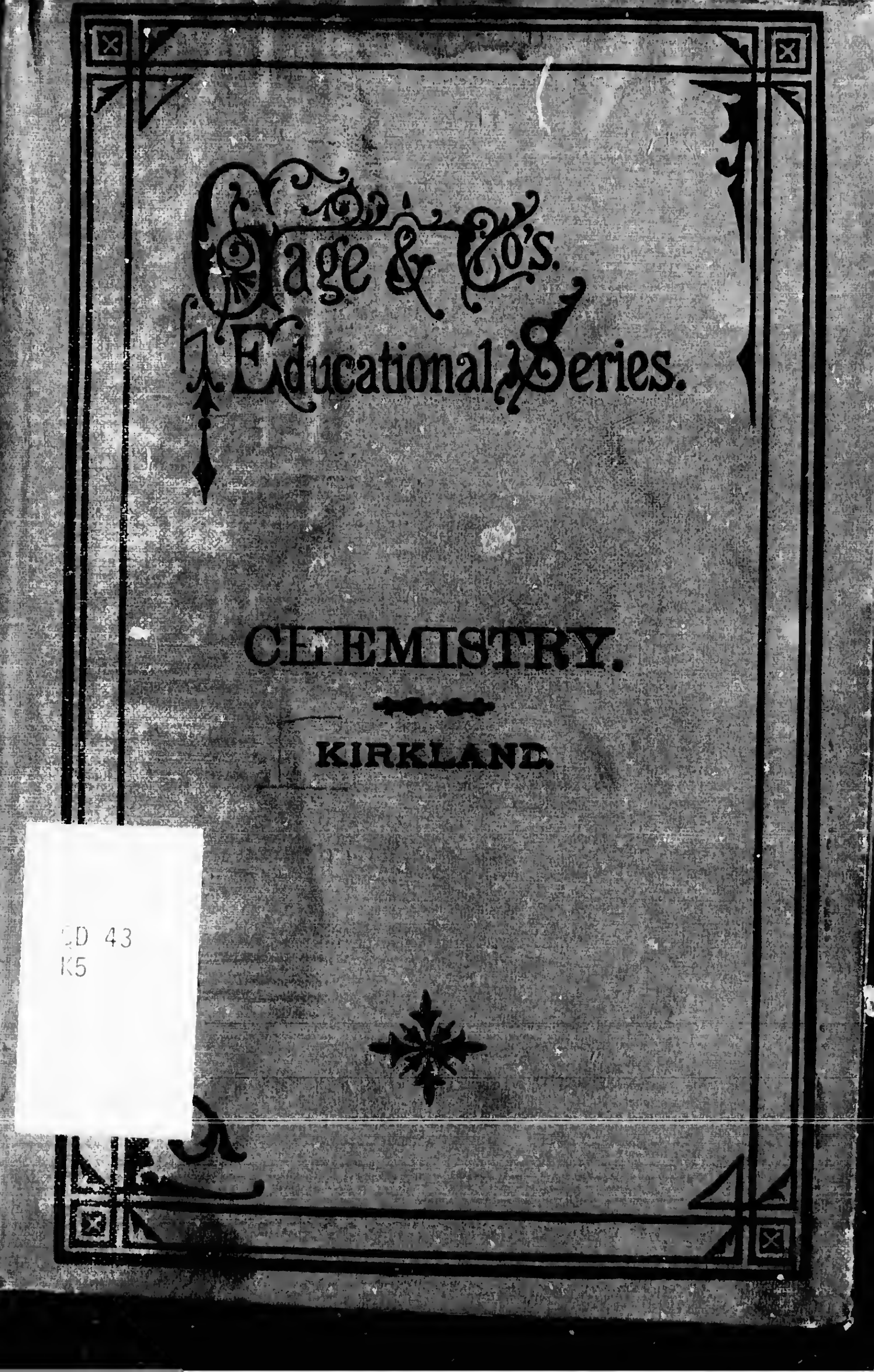
Les exemplaires originaux dont la couverture en papier est imprimée sont filmés en commençant par le premier plat et en terminant soit par la dernière page qui comporte une empreinte d'impression ou d'illustration, soit par le second plat, selon le cas. Tous les autres exemplaires originaux sont filmés en commençant par la première page qui comporte une empreinte d'impression ou d'illustration et en terminant par la dernière page qui comporte une telle empreinte.

Un des symboles suivants apparaîtra sur la dernière image de chaque microfiche, selon le cas: le symbole → signifie "A SUIVRE", le symbole ▼ signifie "FIN".

Les cartes, planches, tableaux, etc., peuvent être filmés à des taux de réduction différents. Lorsque le document est trop grand pour être reproduit en un seul cliché, il est filmé à partir de l'angle supérieur gauche, de gauche à droite, et de haut en bas, en prenant le nombre d'images nécessaire. Les diagrammes suivants illustrent la méthode.







Gage & Co's. Educational Series.

CHEMISTRY.

KIRKLAND.

ELEMENTARY

EXPERIMENTAL CHEMISTRY.

—BY—

THOMAS KIRKLAND, M. A.,

PRINCIPAL, NORMAL SCHOOL, TORONTO.

TORONTO:
W. J. GAGE & CO.,
54 FRONT ST. WEST.

PREFACE.

The following pages contain the substance of lectures in chemistry, delivered to the students in the Toronto Normal School, during the past ten years. The experiments were made in the lecture-room, and afterwards repeated by the students in the laboratory. In 1881-2, portions of the lectures were printed in an educational journal and met with such general acceptance that the author has been induced to publish them in a more extended form.

It has been the author's aim to facilitate the study of chemistry by the experimental and inductive method. It is hoped that the book will enable the careful student to acquaint himself with the main facts and principles of chemistry through the medium of his own perceptive faculties, by a process not unlike that by which these facts and principles were first established. The minute instructions given in the descriptions of the experiments are intended to enable the student to gain a real and adequate knowledge of the things described in the presence of the things themselves. In cases where it is impossible for every student to experiment for himself, the author hopes that this work will enable the teacher to exhibit to his class, in a familiar and inexpensive manner, experiments enough to supply ocular demonstrations of the leading facts and generalizations of the science.

The following are the chief characteristics of this work :—

(1) The experiments, new and old, have been carefully prepared with a view of drawing conclusions from them, showing the intimate connection between these results and the theories based upon them. Full explanations are given as to the best method of performing these experiments, and, when necessary, cautions are given, so as to avoid accidents to the beginner.

(2) The apparatus is of the simplest kind, and much of it may be constructed by the student himself, from inexpensive materials that may be found everywhere.

(3) The chapters and sections are so arranged that each can be discussed with a fair degree of completeness ; thus, the chapters on water and the atmosphere are placed after the chapters on carbon dioxide and ammonia. This arrangement, will also enable the student to become familiar with easy experiments, such as those on carbon dioxide, before undertaking the more difficult.

(4) Among the innumerable facts of chemistry, those only have been established which are either of practical importance, or of such theoretical value as to be necessary for the elucidation of chemical theory. By so doing, space has been gained for information generally confined to more extensive works.

h to supply ocular
and generalizations of

aracteristics of this

have been carefully
clusions from them,
tween these results
Full explanations are
orming these experi-
are given, so as to

st kind, and much
dent himself, from
nd everywhere.

so arranged that
ee of completeness ;
e atmosphere are
xide and ammonia.
student to become
s those on carbon
ifficult.

f chemistry, those
either of practical
e as to be neces-
ory. By so doing,
generally confined

---

(4) At the end of each important section, a series of questions is given to direct the attention of the student to the more important facts which require to be impressed on the mind. Problems, as a test of attainments, can hardly be overestimated, and, therefore, a few carefully selected problems will be found at the end of each section.

(5) For the sake of completeness, facts have been stated in the summaries at the end of the sections which cannot be illustrated without the aid of expensive apparatus.

The book has been specially prepared for those teachers who have not a great abundance of apparatus, and who wish their pupils to become *accurate* observers and *strict* reasoners.

Like all elementary text-books, this work is simply a compilation. It embodies in a somewhat new form previously existing statements of well-recognized facts and principless which have become the common property of science. The author particularly desires to acknowledge his indebtedness to Reynolds' Experimental Chemistry, from which he has introduced several experiments especially in the first chapter.

That errors have crept in, is probable, but it is hoped they will not be found serious, and the author will be glad to have any inaccuracies pointed out for correction in a future edition.

Normal School, Toronto, April, 1886.

CONTENTS.

PAGES.

CHAPTER I.

CHAPTER II.

OXYGEN, HYDROGEN, NITROGEN, AND CARBON.

CHAPTER III.

CHEMICAL CALCULATIONS.

CHAPTER IV.

CARBON DIOXIDE, AND CARBON MONOXIDE.

CHAPTER IV.

NITRIC ACID, AND OXIDES OF NITROGEN.

CHAPTER V.

AMMONIA.

CHAPTER VI.

WATER.

CHAPTER VII.

THE ATMOSPHERE.

CHAPTER VIII.

MARSH GAS, OLEFIANT GAS, AND COMBUSTION.

CHAPTER IX.

CHLORINE, HYDROCHLORIC ACID, BROMINE, IODINE, FLUORINE.

CHAPTER X.

SULPHUR, SULPHUR DIOXIDE, SULPHURIC ACID, SULPHURETTED HYDROGEN.

CHAPTER XI.

PHOSPHORUS, PHOSPHURETTED HYDROGEN ARSENIC.

CHAPTER XII.

BORON, SILICON.

CHAPTER XIII.

ACIDS, BASES, AND SALTS.

CHAPTER XIV.

ATOMICITY OR QUANTIVALENCE.

CHAPTER XV.

CHEMICAL EQUATIONS.

APPENDIX.

ELEMENTARY EXPERIMENTAL CHEMISTRY.

CHAPTER I.

GENERAL PRINCIPLES OF THE SCIENCE.

SECTION I.

Physical and Chemical Changes.

Chemical Affinity. Elements and Compounds.

1. Object of Chemistry. Chemistry has for its object the examination of the properties of all the different substances which occur in nature, so far as they act upon each other, or can be made to act so as to produce something totally different from the substances themselves. It also enables us to ascertain the circumstances under which such changes occur, and to discover the laws upon which they are based.

2. Knowledge of Chemistry, how acquired. Experiment is the fountain-head of all that is known of chemistry; hence, a knowledge of it is only acquired by experiment, by accurate observation of phenomena presented during an experiment, and by carefully reasoning upon the results.

3. Def. of an Experiment. An Experiment is an operation performed under conditions which we ourselves, arrange for the purpose.

4. Physical and Chemical Change.

Exp. 1.—Take a piece of platinum wire and hold it in the flame of a spirit-lamp; observe that it glows as long as held there. Remove it, and it resumes its original state; the action of the heat not leaving any permanent effect upon the wire. The change from cold to red-hot, and from red-hot to cold again, is called a *temporary change*.

Exp. 2.—Take a piece of magnesium wire or ribbon, and observe that it resembles the platinum in many respects such as lustre, tenacity, &c. Hold in the flame of the spirit-lamp till it begins to glow, then remove it, holding it at an angle of 45° so as to keep it burning, and over a piece of blackened paper, to receive the product of combustion; notice that it emits much light, gives out white fumes, and leaves a white substance behind which, though retaining some of the form of the original wire, can easily be powdered when cold, and is seen to be utterly unlike the metal which produced it. In this case the metal has undergone a *permanent change.*

The temporary change in the platinum wire is called a *physical change*, and the force, heat, which produced it is called a *physical force.* The permanent change in the magnesium wire which is accompanied by an entire change in its properties is called a *chemical change*, and that peculiar force which produced it is called a *chemical force*, which is usually termed *Chemical Affinity*, or chemical attraction. And the science of chemistry is almost entirely occupied with the nature and effects of this force.

5. Characteristics of Chemical Affinity.

Exp. 3.—Take two tumblers and hold them mouth downwards for a few seconds over a spirit-lamp until they become slightly warm; into one put a few drops of ammonia, into the other a few drops of hydrochloric acid, shake well and bring the tumblers together mouth to mouth. Observe, that whereas the contents of both were colorless gases, both are now filled with white fumes, which settle on the sides of the tumblers in the form of a white powder. (1). *Hence, Under the influence of Chemical Affinity colorless gases may unite to form a solid.*

Exp. 4. Take a small quantity of water in a tumbler and add as much calcium chloride as the water will dissolve. The solution in the tumbler is now said to be saturated. Now mix the solution with an equal bulk of water, and add to it, all at once, an equal quantity of dilute sulphuric acid (1 to 4), and shake gently; instantly a white solid is formed. (2). Hence, *Under the influence of Chemical Affinity liquids become solids.*

Exp. 5. Rub together in a mortar a small quantity of alum and acetate of lead; the two dry solids become a semi-liquid. (3). Hence, *Under the influence of Chemical Affinity solids may become liquids.*

Exp. 6. Fill a test-tube two-thirds full of water and dissolve a small crystal of copper sulphate in it; the solution will be nearly colorless. Add a few drops of ammonia; a beautiful blue color will be produced. (4). Hence, *Under the influence of Chemical Affinity changes of color frequently take place.*

Exp. 7. Take a thin slice of phosphorus, taking care to cut it under water, dry well with blotting paper and do not handle after drying. Place it on a plate and sprinkle a little iodine upon it and cover it with a wide-mouthed bottle, the two substances combine, much heat being given out. (5). Hence, *Under the influence of Chemical Affinity heat is generally evolved.*

6. Inference from preceding Experiments.—In the second experiment we see that the magnesium wire is utterly transformed; in the third, the white solid formed is altogether unlike the gases which produced it; in the fourth the white solid is an entirely different substance from either of the two liquids used; and in the sixth and seventh the substances formed are totally unlike their constituents. Hence we see that *One of the most characteristic features of Chemical Affinity is the entire change of properties which it occasions in the substances dealt with—a change which no à priori reasoning could possibly predict.*

7. Difference between the Action of Chemical Affinity and other forces.

Exp. 7. Place a small piece of freshly-cut phosphorus on the bottom of a plate, and bring a piece of red-hot iron near it; the phosphorus ignites.

Exp. 8. Rub a glass rod with a piece of silk, or rub a piece of sealing wax with woollen cloth; they will attract a suspended rod of wood or a long thin cylinder of paper. That is, *Heat and Electricity act at appreciable distances.*

Exp. 9. Take a teaspoonful of finely powdered loaf sugar, and two teaspoonfuls of finely powdered potassium chlorate, mix them well together, and form them into a little heap on a piece of card board placed on a tumbler. Dip a glass rod in sulphuric acid and bring it near the powder; as long as there is a measureable distance between the acid and the powder no change will take place. Now bring the acid into actual contact with the mixture and it at once inflames, leaving a black mass of carbon on the plate.

Exp. 10. Into a tumbler put a teaspoonful of baking soda, and the same quantity of finely powdered tartaric acid. However closely the solid particles are brought together by stirring or rubbing no action takes place. Now add water and effervescence immediately ensues, showing that chemical action is taking place; water added to the soda or acid *separately* does not cause any effervescence. The violent action observed on the addition of water to the mixed powders must, therefore, have been due to the mutual attraction of the two solids leading to chemical action; and this could only take place when, by solution, the particles of each body were endued with greater mobility than in the solid state, and were thus enabled to get within the sphere of each other's attraction. (5). Hence we see that *Chemical Affinity acts only at inappreciable distances.*

8. Electricity developed by Chemical Action.

Exp. 11.—Fit a test-tube with a small flat cork, in which pierce two parallel slits with a penknife, and insert in one slit a slip of sheet zinc reaching nearly to the bottom of the tube, and in the other a slip of copper foil of the same length, taking care that they do not touch each other. Insert in each of the slips an end of a piece of fine copper wire about 2 feet in length, so that one wire shall be in contact with each piece of foil. Fill the tube with dilute sulphuric acid (1 part of acid to 12 of water), insert the cork so that the slips shall be in the acid. Such an arrangement is a simple form of a *galvanic cell.* After the zinc has been a few seconds in the acid remove it, put a drop of mercury on a plate and touch it gently with the zinc in three or four places, first on the one side and then on the other, and rub it with the finger till it is all covered with the mercury. The zinc is now said to be *amalgamated.* Place it again in the acid which now does not appear to affect it. Now join the free ends of the wires, taking care that they are quite bright, and observe that a gas rises from the copper. Bring both wires in contact with the tip of the tongue, and notice the sharp metallic taste. Rub a common darning needle with a magnet and suspend it horizontally by a fine silk thread, pass one wire above it and the other below, and join their ends; the needle is immediately deflected. This is the usual test for a galvanic current. (7). Hence we see that *Chemical Action may give rise to a galvanic current.*

Evidences of Chemical Affinity. From the preceding experiments we see that the proof of chemical action taking place between bodies, when they are brought together, may consist in the occurrence of one or more of the following changes :—

(1.) *A change in the state of aggregation, consisting in the conversion of gases into solids, of liquids into solids, or solids into liquids.* These results are not however, to be taken as evidence of chemical action, unless accompanied by other phenomena not produced by simply heating or cooling the substances separately.

(2.) *A change of temperature, usually elevation*

(3.) *The formation of a new substance possessing physical or chemical properties, or both, differing from those of the original substances.*

(4.) *The production of a galvanic current.*

9. Chemical Combination. In several of the preceding experiments we have seen that two substances brought together under certain conditions, unite to form a new substance differing entirely from the original ones. Such a union is called *Chemical Combination, that is, the union of unlike kinds of matter producing a single new substance.*

10. Chemical Decomposition.

Exp. 12.—Accurately weigh enough mercuric oxide to cover the bottom of a clean, dry, not too thin test-tube. Heat the tube gently at first, and then increase the temperature, taking care not to heat intensely any small spot of the tube, and loosely stopping its end with the thumb. Remove the thumb and quickly place inside the tube the glowing end of a splint of wood; it will burst into flame. If the heat is continued till all action is over, there remains in the tube nothing but pure mercury, which is found to weigh less than the mercuric oxide originally taken.

In this Experiment the chemical action brought about by heat resulted in DECOMPOSITION, *which consists in the separation, either partially or completely of the constituents of a compound from one another.* In this case the Mercuric Oxide has been decomposed into mercury, and some other kind of matter driven off as a gas and which is recognised by its property of rekindling a glowing splint of wood. The gas is called OXYGEN.

11. Elements and Compounds. Since chemical action may result in either *combination* or *decomposition*, it follows that all substances in nature may be conveniently divided into two great classes :—

(1) *ELEMENTARY BODIES.* **Elements or simple substances are those out of which no other two or more essentially differing substances have been obtained.**

(2) *COMPOUND BODIES.* **Compounds are those**

bodies out of which two or more essentially differing substances have been obtained.

12. Number of Elements. The researches of Chemists up to the present time have made known to us some sixty-four elements. Of these, or compounds of these with each other, the whole mass of the globe—solid, liquid, gaseous—is composed. And even comparatively few of these are found abundantly. Not more than sixteen are employed in the common changes that take place on the earth, and in the structure of the greater part of the earth's surface only about six are used. Of the others many are of such rare occurrence as to be of little interest except to chemists.

It is important to guard carefully against the idea that the *elements* are *certainly* simple bodies. Chemists, at present, cannot prove them to be compounds, but it is not at all impossible that more powerful analysis may yet decompose them.

13. Subdivisions of Chemistry. Chemistry occupies such an extensive sphere that it has been found desirable for the greater convenience of study to sub-divide the matter on which it treats. It is usually sub-divided into the two branches of *Inorganic* and *Organic Chemistry.* Inorganic Chemistry was formerly defined to be the chemistry of the mineral kingdom, in contradistinction to organic chemistry, which was held to be the chemistry of the animal and vegetable kingdom, since most of the subjects included under this division were of animal or vegetable origin. This definition of organic chemistry is no longer tenable, for by far the greater number of bodies classed by the chemists as organic are neither of vegetable nor of animal origin. At present there is no real difference between organic and inorganic chemistry, but as the substances hitherto comprised under organic chemistry invariably contain carbon, and as the number of these compounds is very large, it is found convenient to study them together under the head of Organic Chemistry.

The bodies treated of under Inorganic Chemistry are divided into two principal divisions,—*Metals*, and *Non-Metals* or *Metalloids*. The Metals, generally speaking, possess a peculiar lustre, called Metallic, a high specific gravity, and are good conductors of heat and electricity; while the *Non-Metals* are noted for the absence of metallic lustre, for their low specific gravity, and for being generally bad conductors of heat and electricity. This division is to a great extent arbitrary. There exists no sharp line of demarcation between the two branches, and it is impossible to say precisely what constitutes a Metal, and what a Non-Metal.

The following pages will treat of what are usually called the Non-metals and their more important Compounds.

SECTION II.

Indestructibility of Matter; Law of Definite Proportions; Chemical Compounds and Mechanical Mixtures; Avogadro's Law; Chemical Notation; Chemical Equations.

14. Indestructibility of Matter.

Exp. 13.—Take a piece of phosphorus about as large as a small pea, put it into a little water in a test-tube, and gently heat till it melts; then place the point of a thin wire in it and let it cool. Bend the lower end of the wire into a spiral so that it will stand upright. Select a thin, light beaker, and a large, thin light flask of about 1500 cubic centimetres (abont 2½ pints). Put about 400 cubic centimetres (¾ of a pint) of water, colored blue with litmus, into the beaker, place the wire in it, and invert the flask over the phosphorus with its mouth reaching almost to the bottom of the colored water. In this condition place the whole apparatus on a pair of scales and exactly balance it by weights in the other pan. In about forty-eight hours it will be found that the water has risen into the flask as if part of the air had been annihilated, and its color has changed from blue to red. The phosphorus has evi-

Chemistry are divided
and *Non-Metals* or
ing, possess a peculiar
vity, and are good con-
Non-Metals are noted
r low specific gravity,
heat and electricity.
ry. There exists no
vo branches, and it is
es a Metal, and what a

re usually called the
pounds.

efinite Proportions;
ixtures; Avogadro's
ons.

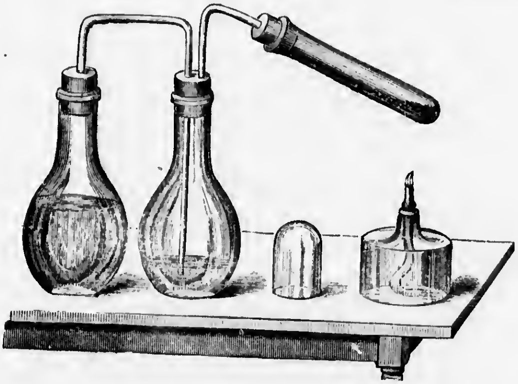
about as large as a
test-tube, and gently
thin wire in it and
vire into a spiral so
, light beaker, and
) cubic centimetres
centimetres ($\frac{3}{4}$ of
s, into the beaker,
ver the phosphorus
tom of the colored
paratus on a pair of
the other pan. In
the water has risen
annihilated, and its
phosphorus has evi-

---

dently wasted away as if part of its substance also had been destroyed. Nevertheless the apparatus weighs exactly as much as at first, showing that no loss can be detected. And yet, if the substance which has disappeared had been destroyed the apparatus would weigh at least, ·43 of a gram less which a good common balance would easily detect. What has taken place is this; the phosphorus has combined with the oxygen which is one of the constituents of the air in the flask. The compound formed has been absorbed by the water, and remains unseen, but indicates its presence by changing the color of the water. This new substance contains all the oxygen and all the phosphorus which seemed to have been lost.

Hence we infer that *No loss of matter occurs in Chemical Combination.*

Exp. 14.—Put 2.16 grams of mercuric oxide into a test-tube, not too thin, which is provided with a bent tube reaching just through the cork of a flask of about 200 cubic centimetres capacity. Let the flask be joined to another of equal size, by means of a bent tube which reaches almost to its bottom, but only just through the cork of the other. The first flask is to be



nearly or quite full of water, and must be quite air-tight at the cork, whilst the second is empty and loosely corked. Place the whole apparatus, which must be perfectly dry outside, upon the scales and accurately balance it by weights. Then heat the test-tube and decompose the mercuric oxide; the red powder will gradually waste away, globules of mercury collect on the side of the tube above the heat, while the oxygen passes into the first flask and drives the water over into the second. Leave the apparatus at rest for a short time till the tube becomes cold, and it will be found that the balance is undisturbed. The whole weight is just the same as at first.

Hence we infer that *No loss of matter takes place in Chemical Decomposition.*

15. Conservation of Matter. Many accurate experiments similar to the foregoing have been made by chemists and have proved beyond doubt that matter is never destroyed. Substances may disappear and seem to be lost, but the loss is only apparent. During all the chemical changes through which a substance may pass, the balance shows that the *weight* remains the same.

The doctrine of the indestructibility of matter or as it is now termed the *Conservation of Matter*, teaches us that we can neither destroy nor create matter, though we can alter its distribution, and its forms and combinations, in innumerable ways. This principle constitutes the basis upon which rests the whole superstructure of the science of Chemistry. *It is the first great principle of Modern Science.*

16. Conservation of Energy. Energy has been defined to be "the capacity or power of one body or system of bodies when in a given condition, to do a measurable amount of work"; and it has been established during the past half-century that we can neither create nor destroy energy, though we can change it from one form to another. *This is the second great principle of modern* *science*, but the consideration of it belongs to Physics rather than to Chemistry.

17. Law of Constant Composition.—The first great law concerning chemical combination discovered by the use of the balance is that of the invariable proportions of the constituents in any chemical compound. In whatever way any given chemical compound may be prepared, or in whatever manner its composition may be accurately ascertained, it is found always to contain a fixed and definite quantity of each of its constituent elements, *and this is a distinguishing characteristic of a chemical compound, as opposed to a mere mechanical mixture, the constituents of which may be present in any varying proportions.* Thus in the last experiment the 2.16 grams of mercuric oxide will yield two grams of mercury and 16 of a gram or 112 cubic centimetres of oxygen, and although the oxide can be prepared in several ways, the weight of mercury and volume of oxygen obtained are always found to be the same from the same weight of the oxide.

A great many experiments have been made in the same direction, and it has been found that every chemical compound which possesses a group of characters serving to define it, and so distinguish it from all other forms of matter, exhibits the remarkable constancy of composition exhibited by mercuric oxide. The inference clearly to be drawn from this is that *Chemical compounds are constant in composition.*

18. Law of Definite Proportions.—The admission of the constancy of composition of chemical compounds leads us to suspect that chemical combination takes place in *definite proportions.* Were it otherwise it would be impossible to give any adequate explanation of the fact that the constituents of mercuric oxide are always found in that body in fixed proportions. This may be put to the test by the following experiment :—

Exp. 15.—Into a small beaker pour about fifty cubic centimetres of hydrochloric acid, and drop in, little by little, powdered sodium carbonate. Effervescence takes place, showing that gas is escaping, and that chemical action is going on. Continue until the last small quantity of the sodium carbonate produces no effervescence. The acid is then all neutralized. Then carefully stir in drops of the acid until with the last drop the last of the small quantity of solid carbonate disappears. The slightest quantity of either, beyond a certain *definite proportion*, remains unchanged. Hence we are led to the following law :—

The Proportion in which bodies unite together chemically is definite and constant.

19. Chemical Compounds and Mechanical Mixtures.—We find a variety of compound bodies in many cases closely resembling chemical compounds. To these various names are applied according to the nature of the substance, such for instance as *mechanical mixture*, solution, alloy, etc. But there is always a marked difference between them and true chemical compounds. The following experiments will illustrate this :—

Exp. 16.—Make a mixture of iron filings and sulphur in the proportion of fifty-six parts by weight of iron to thirty-two parts by weight of sulphur. A greenish gray powder results, but (1) distinct particles of both iron and sulphur can easily be recognized by a good magnifying glass. (2) Gently stir a portion of the powder into a tumbler of water; the heavy particles of iron fall quickly to the bottom of the tumbler, while the lighter sulphur more slowly subsides and collects as a distinct layer. (3) Stir the mixture with a small magnet, and the particles of the iron will firmly adhere to the magnet, while the sulphur can easily be blown away.

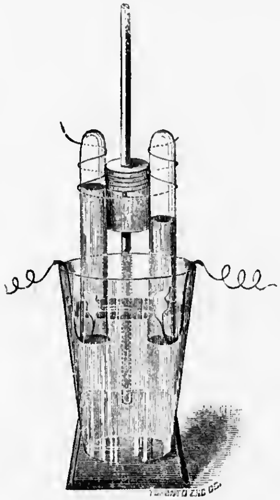
Hence we see that *The constituents of the mixture can easily be separated by mechanical means, and that it partakes of the properties of both iron and sulphur.*

Exp. 17.—Heat a small portion of the mixture of iron and sulphur in a test-tube. The mixture becomes pasty and then glows for a short time, showing that chemical action is taking place. Break the test-tube and grind up its contents in a mortar. (1) When examined by a magnifying glass no particle of iron or sulphur can be detected. (2) It is no longer attracted by the magnet, or at least very little, and therefore contains little or no free iron. (3) The iron and sulphur are no longer separable by mechanical means. (4) If a small quantity is put into a test-tube and dilute sulphuric acid added, a gas possessing a very offensive odor is evolved. Neither iron nor sulphur alone, possesses the property of evolving this gas. The iron and sulphur have chemically combined, forming iron sulphide, which possesses a definite group of characters which not only serve to distinguish it from the free elements, iron and sulphur, or a mixture of them, but from all other bodies.

Hence we see that *the constituents of a chemical compound cannot be separated by mechanical means, that its properties differ entirely from those of its constituents, and that it contains a fixed and definite quantity of its constituents.*

20. Composition of Water.

Exp. 13.—Take two copper wires, each about twenty centimetres in length, flatten an end of each, and to the flattened ends solder a strip of platinum, about two centimetres long by five millimetres broad When these wires are connected with the wires from the battery they are usually spoken of as the *poles* of the battery. Dip the wires in melted paraffine, and wrap round each of them a thread of lamp-wick, previously soaked in paraffine. This will protect the copper from the action of the acid. Take a tumbler three-fourths full of water and add to it a teaspoonful of sulphuric acid for the purpose of increasing the conducting power of the water. Bend the copper wires over the sides of the tumbler so that the tops of the platinum strips may be about two centimetres below the surface of the water Fill two test-tubes with the acidulated water and place them over the platinum strips, keeping the tubes as near together as possible. Connect the wires with the galvanic battery and minute bubbles of gas will immediately be given off. It will soon be seen that twice as much gas is given off from the pole connected with the zinc end of the battery as from the pole connected with the other end; when the former is full the latter is only half full. As soon as the tube connected with the zinc end of the battery is full, close its mouth with the thumb, raise it out of the water, and examine its contents:



(1). Observe that the gas is colorless.

(2). Invert the tube and apply a match to its mouth; the gas takes fire and burns with a pale blue flame.

(3). Refill the tube and turn its mouth upwards and smell it; no odor is perceived. Hold the tube in this position for a few seconds, and then apply a lighted match to its mouth; no combustible gas is found in it. The gas has escaped, and is, therefore, lighter than air.

The gas possessing the above properties is called **Hydrogen** It is considered to be an elementary body, and will be fully treated of in a future chapter.

Examine the gas in the other test-tube in the same way ; it will not take fire. Immerse in it a glowing splint of wood ; the wood instantly takes fire and burns with great brilliancy. Hence the gas has the same characteristics as that obtained by heating mercuric oxide and is therefore *Oxygen.*

The above process is called *electrolysis* and is frequently employed in decomposing chemical compounds.

When the above experiment is carefully made, it is found that the only substance that undergoes permanent change is the water. The weight of the water is diminished in exact proportion to the amount of gas evolved. Besides, if the operation is continued the water will be *completely* resolved into the two gases. The composition of a substance determined as in the preceding experiment, by resolving the compound into its elementary constituents, is called *Analysis.*

Exp. 19.—Take a strong glass tube, about thirty centimetres long and one centimetre in diameter. Choose a good, sound cork, pass two short copper wires through it connecting their extremities within the tube with fine platinum wire. Insert the cork tightly in the tube and cover it with sealing wax. Such a tube is called a *Eudiometer*, of which there are many forms. Now, fill the tube with acidulated water, and place it over the platinum strips, taking care that they do not touch each other. When the tube is about one-half full of the gases, press it firmly against an india rubber or paper pad on the bottom of the tumbler, wrapping a towel loosely round the tube, and connect the wires in the cork with the battery. The fine platinum wire will soon become red-hot and explode the gases. On raising the tube from the pad, the water will rush up and fill the tube, showing that the two gases have united to form water which appears as vapor on the tube before it is raised from the pad.

Water is, therefore, composed of two gases, Oxygen and Hydrogen, in the proportion of one volume of the former to two volumes of the latter.

When the composition of a substance is determined as above by causing its constituents to enter into combination, the process is called *Synthesis.*

21. Combination by Volume.—In the preceding experiment if the apparatus is so arranged that the gases before explosion are heated beyond the temperature of boiling water and kept at the same temperature after explosion, it is found that the two volumes of hydrogen and one volume of oxygen which were mixed together have become chemically united into *two volumes* of steam. It is found in other cases also that **Whatever the number of volumes which enter into combination, the resulting compound is two volumes.**

22. Physical Constitution of Gases.

Ex 20.—Take two glass tubes, about one metre in length and five millimetres in diameter. Close one end of each and bend so that the short limb may be about twenty centimetres long. Fill one of them with acidulated water, colored with litmus or indigo, and place it over the pole from which the hydrogen is escaping, until enough is collected to half fill the short limb of the tube. Turn the short limb uppermost and the gas will pass into it. Half fill the short limb of the other tube with oxygen in the same manner.

(1). Observe that when both limbs are full of water the gases are equally compressed.

(2). By means of a pipette to which a piece of rubber tubing or a piece of fine glass tubing is attached, adjust the water in each tube so that it may stand at various heights, but always at the same height in each tube; the gases expand equally as the water is removed.

(3). If you have mercury pour an equal quantity into each tube, and observe that the gases contract to the same extent.

(4). Plunge the tubes into boiling water; the gases expand equally.

(5). Plunge the tubes into ice-cold water or a freezing mixture; the gases contract to the same extent.

Hence we infer that *Oxygen and Hydrogen gases when compared under the same conditions are affected in the same way and to the same extent by equal alterations of pressure and temperature.*

When the same mode of investigation is applied to other gases, whether elementary or compound, the following characteristics are observed :—

(1). **All true gases obey the same law of compressibility.**

(2). **Equal volumes of all true gases expand equally on the same increase of temperature.**

The conclusion that must necessarily be drawn from the preceding facts is, that all gases, however different chemically must be physically constituted alike. Upon these facts an Italian chemist, named Avogadro, based a most important hypothesis. He assumed that all substances, solid, liquid, and gaseous, are made up of an innumerable number of little particles which he called *molecules*, and was thence led to the enunciation of the following law :—

23. AVOGADRO'S LAW. When in the condition of a perfect gas, all substances under like conditions of temperature and pressure contain in equal volumes the same number of molecules.

The above law is considered the most important in the whole range of chemical science. It is to the chemist what the law of gravitation is to the astronomer. But, like every other part of the molecular theory, this law owes its recognition not to any direct proof that can be adduced from experimental sources in support of it, but to the fact that nearly all observed chemical phenomena not only harmonise with it but also find in it complete and satisfactory explanation. It not only explains the known facts of chemistry, but is also constantly leading to new discoveries. Admit this law and we at once see why gases are equally expanded by heat and why they were equally contracted by cold and pressure and why they combine by volume according to the law in Article 21.

24. Relative Weights of the Molecules of Gases. Avogadro's law gives at once the means of determining directly the relative weights of the molecules of all substances that can exist in the state of gas. For it is obvious that **The ratio of the weights of equal volumes of gases, compared at the same pressure and temperature, must represent the relative weights of the molecules of these gases.** Now it is found by experiment that a given volume of oxygen is sixteen times heavier than an equal volume of hydrogen, under the same conditions; therefore, *the molecule of oxygen is sixteen times heavier than the molecule of hydrogen.*

25. Relative Weights of the Atoms of Elementary Gases. Two volumes of hydrogen and one volume of oxygen combine to form two volumes of steam (Art. 20); and by Avogadro's Law we know that the two volumes of steam contain the same number of molecules as the two volumes of hydrogen, hence we have—

2 vols. hydrogen + 1 vol. oxygen = 2 vols. steam;
therefore, 2 mols. " + 1 mol. " = 2 mols. "

Now, in the two molecules of steam there is but one molecule of oxygen; therefore in one molecule of steam there can only be half a molecule of oxygen, and the weight of the oxygen corresponding to the semi-molecule is the smallest quantity of that gas that can take part in any chemical change, and as it cannot be further divided by any chemical means it is called an *atom*. We shall find that the molecule of hydrogen can also be divided into two parts, and that one of these parts is the least quantity of hydrogen known to take part in any chemical action, and is, therefore, called an *atom* of hydrogen. It has been found convenient to take 1 as the *atomic weight*, or weight of a semi-molecule of hydrogen. The weight of a molecule of hydrogen is, therefore, 2; and since oxygen is sixteen times heavier than hydrogen, the molecule of oxygen is 32, and consequently its atomic weight is 16.

Hence, **To find the atomic weight of an elementary gas, it is only necessary to find its weight compared with hydrogen as the unit.**

26. Specific Gravity of a Compound Gas.—Suppose that in a given volume of steam there is a certain number of molecules, then by Avogadro's Law the same volume of hydrogen will contain the same number of molecules. Therefore, the weight of a given volume of steam has the same ratio to the weight of an equal volume of hydrogen that a molecule of steam has to a molecule of hydrogen. But the molecule of steam is made up of two atoms of hydrogen and one atom of oxygen; its *molecular weight* is, therefore, 18, and the molecular weight of hydrogen is 2, or the ratio is 9 to 1. Therefore, the specific gravity of steam compared with hydrogen as the unit is 9.

Hence, **The specific gravity of a gas simple or compound compared with hydrogen as the unit is found by taking half its molecular weight.**

As a rule the molecule of an element is composed of two atoms; but in *mercury*, *zinc*, and *cadmium* the molecule contains only one atom, and in *phosphorus* and *arsenic* the molecule contains four atoms.

27. Size of Atoms. Sir William Thompson has shown that the size of a molecule of water is about one 500,000,000 th of an inch in diameter. This number is so large that it is difficult to derive any idea from it. A clear conception can be got by supposing, as Thompson does, a drop of water to be magnified until a drop the size of a pea is as large as the earth itself, when, if each constituent molecule were enlarged in the same proportion, the magnified structure would be coarser grained than a heap of small shot, but probably less coarse grained than a heap of cricket-balls.

28. Definitions. — From the preceding paragraphs we collect the following definitions :—

Molecule.—*A molecule is the smallest particle of a compound or element that can exist in a free state.*

Atom.—*An atom is the smallest portion of a chemical element that is known to take part in a chemical change, and is almost invariably the semi-molecule.*

Atomic Weight.—*The atomic weight of an element is the smallest proportion, by weight, in which it enters into or is expelled from a chemical compound, the weight of hydrogen being taken as unity.*

Molecular Weight —*The molecular weight of an element or compound is the sum of the atomic weights of the elements comprising a molecule of it*

29. Chemical Notation.—Instead of writing the names of the elements in full, chemists have agreed to use a set of symbols to represent them. These symbols, however, not only represent the particular element but also *a certain definite quantity of it ;* thus, the letter H always stands for 1 atom, or one part by weight of hydrogen ; the letter O stands for 1 atom, or 16 parts by weight of oxygen. Compounds are in like manner represented by writing the symbols of their constituent elements side by side, and if more than 1 atom of each element is present, the number is indicated by a numeral placed on the right of the symbol of the element below the line. Thus, water is represented by the symbol H_2O, that is, a compound of 2 atoms of hydrogen with 1 atom of oxygen, or 2 parts by weight of hydrogen with 16 parts by weight of oxygen. Again, sulphuric acid is represented by the formula H_2SO_4, which is a statement

rged in the same
coarser grained
ss coarse grained

g paragraphs we

icle of a compound

a chemical element
ge, and is almost

' an element is the
's into or is expelled
ygen being taken as

ght of an element or
its of the elements

riting the names of
ise a set of symbols
, not only represent
nite quantity of it ;
m, or one part by
1 atom, or 16 parts
like manner repre-
tuent elements side
nent is present, the
n the right of the
ius, water is repre-
und of 2 atoms of
by weight of hydro-
gain, sulphuric acid
ich is a statement

---

that it consists of two atoms of hydrogen, one atom of sulphur and four atoms of oxygen, and consequently a certain relative weight of these elements. A figure placed to the right of a symbol only affects the symbol to which it is attached, but when placed to the left all the symbols are affected by it; thus $2H_2O$ means two molecules of water.

30. Chemical Formulæ.—From the preceding it appears that a chemical formula is simply a representation of the composition of a body by means of symbols. *It represents a molecule of either an element or compound*; thus, H_2 represents a molecule of hydrogen, and H_2SO_4 a molecule of sulphuric acid. Such formulæ express the elementary composition of a compound much more tersely than words can; they are written and read more rapidly than the sentences of the same signification could be, and by their brevity clearness and precision, greatly facilitate the study of chemistry. Their great comprehensiveness may be learned from the amount of information condensed into the formula for the molecule of water, H_2O. We learn from it:—

(1) The *number* and *names* of the elements entering into the composition of water.

(2) The *ratio* in which the elements are united in this compound by *weight.*

(3) The ratio in which the elements are united therein by *volume.*

(4) The *ratio* in which the volume of the compound *when formed* stands to the volume of the constituents *before combination.*

(5) The relative *volume-weight* or specific gravity in the state of gas (water-gas), hydrogen being taken as unity.

31. Chemical Equations.—The changes which occur when two or more substances act chemically upon one another are represented by placing on the left the formula of a substance *before* the change, connected with each other by the sign, +, which simply means *and*; on the right are placed the substances as they exist *after* the change connected as before by the sign + The two sides are connected by the sign of equality, =, which means *yields*, or *produces*, and it only retains the meaning of equality in so far as it indicates that the weight *before* the change is equal to the weight *after* the change. A chemical equation is, therefore *a collection of formulæ so arranged as to illustrate the state of things before and after the action of the molecules on each other which results in the formation of new substances.* As these changes almost invariably result from the reaction of one substance on another, they are called *chemical reactions.* In every chemical reaction the substances which are involved in the change are called the *factors*, and the substances produced the *products* of the reaction. As matter is indestructible it follows that *the sum of the products of any reaction must always be equal to the sum of the factors*, and further that *the number of atoms of each element in the products must be the same as the number of atoms of the same kind in the factors.*

It is necessary to bear carefully in mind that a chemical equation differs essentially from an algebraic equation, anything that can be deduced from an algebraic equation according to the rules of the science must be true. This is not the case with respect to a chemical equation. We cannot deduce facts from chemical symbols, *which are merely the language by which we express the results of experiments.* Although important inferences may sometimes be drawn from the mere form of expression, yet these are of no value unless confirmed by experiment.

An exact knowledge of chemical equation is of vital importance. The result of almost every experiment will be expressed by means of these equations, and the student must thoroughly master them.

QUESTION ON CHAPTER I.

1. What is meant (*a*) by chemical change, (*b*) by physical change? Give familiar examples of each.

2. What is chemical affinity? How does it differ from other forms of force? Describe some experiments which illustrate this difference.

3. Explain the terms *chemical combination* and *chemical decomposition*. Give examples of each.

4. Define the terms *element* and *compound* and give examples of each. What is the present number of elements, and what reason have we for believing that additional elements may be discovered? By what means can the compound character of a substance be proved? Why is platinum, magnesium, and mercury considered to be elements?

5. The following substances are heated to redness in air—(*a*) platinum, (*b*) magnesium, (*c*) mercuric oxide. What occurs in each case?

6. Matter is said to be indestructible. Explain this statement and describe experiments in support of it.

7. No weight is lost in chemical combination. How may this be shown?

8. Chemical combination is said to take place in certain definite proportions. Explain this law and describe experiments in support of it.

9. State clearly the difference between a *chemical compound* and a *mechanical mixture*.

Two powders are given you one is a mechanical mixture, the other a chemical compound of sulphur and iron filings. What experiments would you make to distinguish one from the other.

10. One hundred pounds of iron filings are heated with the same weight of sulphur till combination takes place. Which element would be found in excess after the reaction, and how much of it?

11. In what proportion (*a*) by volume, (*b*) by weight do oxygen and hydrogen unite together?

12. Under what circumstances may gases combine together without undergoing a contraction in volume? Under what circumstances will contraction occur? To what is the volume of the resulting gas invariably equal?

13. State Avogadro's Law. Enumerate the circumstances that tend to establish it. Show how it accounts for the simple ratios in which gases combine by volume.

14. Define the terms *molecule, atom, atomic weight*, and *molecular weight.*

15. How may the molecular weight of a gaseous substance be readily ascertained by experiment?

16. What relation exists between the molecular weight of a compound gaseous substance and its density compared with hydrogen? If the specific gravity of a gas be given as compared with hydrogen, how may its specific gravity as compared with air be ascertained?

17. Explain briefly how the atomic weights of oxygen and hydrogen may be ascertained.

18. Define the terms *chemical symbol, chemical formula*, and *chemical equation.* How does a chemical equation differ from an algebraic equation?

CHAPTER II.

SECTION I.

OXYGEN.

Symbol, O. Atomic Weight, 16 (15.96). Molecular Weight, O_2, *32.*

PREPARATION.

32. **From Mercuric Oxide.** — This method possesses special interest, since it is that by which oxygen was first prepared by its discoverer, Priestly. When mercury is heated in contact with air to a temperature of about 450°C. it gradually becomes covered with red scales of Mercuric Oxide, HgO, and when these red scales are exposed to a temperature of about 630°C. they begin to be decomposed into oxygen and metallic mercury. Atoms when liberated form molecules, and in order to satisfy Avogadro's Law, the molecule of mercury must consist of a single atom, and the molecule of oxygen of two atoms; and since a single molecule of mercuric oxide does not yield an even number of atoms, we take two molecules and represent the reaction by the following equation :—

$$\underset{\text{Mercuric Oxide.}}{2HgO} = \underset{\text{Mercury.}}{2Hg} + \underset{\text{Oxygen.}}{O_2}$$

33. **Dissociation.**—Decomposition like that of mercuric oxide under the influence of heat is an example of what is known as *dissociation.* This term is applied to those cases of decomposition in which certain bodies are resolved at an elevated temperature into simpler bodies, which are capable of reuniting and reproducing the original compound when the temperature is again allowed to fall.

34. From Potassium Chlorate.

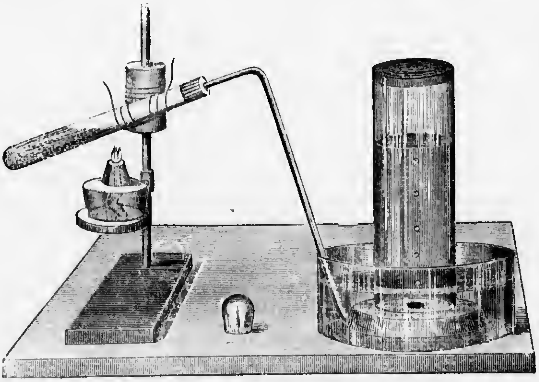
Exp. 1.—A molecule of the white crystalline substance called Potassium Chlorate consists of one atom of Potassium, the symbol of which is *K*, one atom of Chlorine, Cl, and three atoms of Oxygen, O_3; hence its formula is denoted by $KClO_3$. Take a few crystals of this substance, put them into a test-tube, and heat, gently at first, loosely stopping the mouth of the test-tube with the thumb. The salt begins to spirt or "decrepitate," then fuses, and on further heating appears to boil. Plunge a glowing splint of wood into the tube; the splint immediately bursts into flame with a slight explosion, showing that oxygen has been produced. Now, if oxygen has been produced from potassium chlorate, that substance must have undergone a change. To put this to the test, take a few crystals of the chlorate, dissolve them in water in a test-tube, and add a drop or two of a solution of silver nitrate; the liquid remains clear. Dissolve a small quantity of the substance remaining in the test tube, and add a drop or two of silver nitrate solution; at once a white solid forms in the liquid, showing that the potassium chlorate has been changed by the heat. The substance remaining in the test-tube when the oxygen is given off is called Potassium Chloride, KCl. The reaction which has taken place is expressed by the following equation :—

$$\underset{\text{Potassium chlorate}}{2KClO_3} = \underset{\text{Potassium chloride}}{2KCl} + \underset{\text{Oxygen.}}{3O_2}$$

35. From Potassium Chlorate and Manganese Dioxide.

To obtain oxygen from potassium chlorate alone, greater heat is required than glass vessels generally bear, and it is, therefore, seldom adopted in practice, except when very pure oxygen is required. It is found, however, that if the potassium chlorate is mixed with about one-fourth its weight of manganese dioxide, MnO_2, the oxygen is given off at a much lower temperature.

Exp. 2.—Coarsely powder 20 grams (nearly $\frac{3}{4}$ oz.) of potassium chlorate, place it upon a piece of paper, turned up at the edges so as to form a kind of trough, and by means of this transfer the chlorate to a test-tube 6in. by $\frac{3}{4}$in. This quantity will about half fill the tube. Place upon the same paper about 5 grams of manganese dioxide, and transfer it to the test-tube. Place a small piece of paper on the mouth of the tube to keep the thumb clean, and shake vigorously till the two substances are thoroughly mixed. Choose a sound cork, very slightly larger than the mouth of the test-tube, squeeze it well till it becomes soft and elastic, and with a cork-borer or rat-tail file bore a hole through it, so that a glass tube may be inserted tightly into it. Take a piece of glass tubing, about 45 centimeters (18 inches) in length, bend it in two places, one downwards about 5 centimeters (2 inches) from the end of the tube; the other should be upwards, and as near the other end of the tube as possible. The object of the second bend is to facilitate the escape of the bubbles of gas, by giving them a forward and upward direction. The downward bend should now, with gentle pressure and twisting motion, be inserted in the cork.



Take a large cork, bore a hole in it to fit the rod of the retort-stand, twist a piece of copper wire around the cork, leaving about eight inches of the end free. Twist the free ends of the wire around the test-tube, insert the cork in the tube, and arrange as represented in the Figure.

Take two quart and five pint bottles, fill two of them with water, and place mouth downwards, one on the shelf of the pneumatic trough, the other in the water in the trough, ready to take the place of the one on the shelf. On gently heating the test-tube, oxygen will readily be given off. When the bottle is full, slide it off the shelf with one hand, and with the other slide the full bottle into its place. Bring a glass plate under its mouth while still under water, raise it out of the water, and place it mouth upwards on the table. If any of the bottles have to be kept for a considerable time, they should be placed mouth downwards in a saucer containing a little water, which will act as a valve to prevent the gas from escaping.

One ounce or 28.3 grains of potassium chlorate should yield 1¾ gallons, or nearly 7.75 litres.

36. **Precautions.**—In performing this experiment, the following precautions should be observed :—

(1) In the preparation of the mixture, care must be taken to avoid the accidental introduction of small fragments of organic matter. A small quantity of the oxygen mixture should be heated in a test-tube before using it on a large scale, as occasionally the manganese dioxide contains lamp-black and substances of kindred nature, the presence of which would lead to an explosion.

(2) When the delivery-tube is fitted into the cork, and the cork into the test-tube, blow down the open end of the delivery-tube; no air must be heard to escape, or must be seen to bubble out on moistening the cork.

(3) As soon as the oxygen begins to be delivered, the heat

beneath the test-tube should be so regulated that the evolution of the gas should be tranquil and uniform.

(4) The uppermost portions of the mixture should be heated before the lower, which should be heated only after the uppermost parts refuse to yield any more gas.

(5) The test-tube should never be more than three-fourths full, lest particles of solid matter be projected into the delivery-tube.

(6) The test-tube should be inclined at an angle of about 45°, and must never be placed upright.

(7) When the process is to be stopped, the end of the delivery-tube must be immediately removed from the water, otherwise the cold water is liable to go back into the hot test-tube, and break it; and the test-tube must not be allowed to touch cold or wet objects, which would cause the hot glass to crack.

37. Manner in which the Manganese Dioxide acts.

Exp. 3.—Mix one gram of potassium chlorate with half a gram of each of the following substances, viz.:—Manganese dioxide, MnO_2, copper oxide; CuO, ferric oxide, Fe_2O_3, lead oxide, PbO, zinc oxide, ZnO, magnesium oxide, MgO, sand, and powdered glass Put these several mixtures into as many test-tubes, and into another test-tube put one gram of potassium, chlorate alone. Imbed the tubes side by side in sand to about the same depth as they are filled within, apply a gradually increasing heat, and by means of glowing splints frequently plunged into the different tubes, observe the difference in the readiness with which the oxygen is evolved from the several mixtures. It will be seen that there is little difference among the first four tubes, and that the heat will probably not be sufficient to cause the oxygen to be evolved from the other five. After all the oxygen has been given off, about half-fill the first tube with water and gently heat till its contents are dissolved, filter through a piece of blotting paper, and gently dry on hot sand the black substance remaining on the filter. It will be found to be manganese dioxide, and that it is in the same state after the experiment as at the commencement. The same is true of the next three oxides. These oxides may, however, undergo a temporary change. We know that MnO_2 is capable of taking up more oxygen and, combining with a molecule of water to form manganic acid, H_2MnO_4, and it is possible that when heated with potassium chlorate the manganese dioxide may absorb oxygen from this substance, and pass to the state of the higher oxide which is immediately decomposed, the oxygen being evolved and the manganese dioxide returning to its original state. The same applies to the next three oxides, all of which are known to be susceptible of higher oxidation. The zinc oxide and magnesium oxide, on the other hand, which do not form higher oxides, do not facilitate the decomposition of the chlorate.

38. Physical Properties of Oxygen.

Exp. 4.—Take the first bottle filled, which will contain a little air, but will suit well enough for this experiment. Observe the physicial properties of the gas; it has neither color, taste, nor smell. Plunge a glowing splint of wood into it; the splint bursts into flame as in the preceding experiment. Hold the glowing splint at different heights *above* the mouth of the bottle; it will not burst into flame. Leave the bottle for a short time standing on the table with its mouth open; the glowing splint will show that the gas has not escaped. Turn the bottle mouth downwards, and hold the glowing splint *near* its mouth; it will burst into flame. Hold the bottle in this position for a short time; the splint will no longer burst into flame when plunged into the bottle. The gas has escaped. *Oxygen is, therefore, heavier than air.*

39. Combustion of Phosphorus in Oxygen.

Exp. 5.—Take a quart bottle of oxygen, and adjust the deflagrating spoon by holding it against the outside of the bottle, so that the little metallic cup may be about 5 centimetres (2 inches) from the bottom of the bottle, and put a little lime into it to absorb the moisture. Take a piece of phosphorus not larger than a small pea, dry it carefully with soft blotting paper, and place it on the lime in the cup. The phosphorus must not be touched with the fingers after it is dry, and it should be dried only when about to be used. Place the spoon in the bottle, and touch the phosphorus with the end of a heated glass rod. The phosphorus burns brilliantly, and when the heat volatilizes it a flash of light fills the whole vessel, owing to the points of contact between the phosphorus and the oxygen being indefinitely increased, and the bottle is filled with dense white fumes. The phosphorus has combined with the oxygen, forming the white, snow-like substance called *Phosphorus Pentoxide*, P_2O_5. The molecule of phosphorus consists of *four* atoms, and the reaction is expressed as follows:—

FIG. 4.

$$\underset{\text{Phosphorus.}}{P_4} + \underset{\text{Oxygen.}}{5O_2} = \underset{\text{Phosphorus pentoxide.}}{2P_2O_5}$$

40. Acids.—Phosphoric Acid.—Pour a little water into the bottle and shake it up, the phosphorus pentoxide dissolves in the water. Taste a few drops of the solution; it is exceedingly sour. Pour into it some blue litmus solution; it will be immediately reddened. A substance possessing these characteristics is called **an** *acid.* The phosphorus pentoxide has combined with a molecule of water; thus:—

$$\underset{\text{Phosphorus pentoxide.}}{P_2O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Metaphosphoric acid.}}{2HPO_3}$$

41. Combustion of Sulphur in Oxygen.

Exp. 6—Wash the lime out of the cup of the deflagrating spoon, and dry it by holding it in the flame of the spirit-lamp. Place in it a piece of sulphur about as large as a pea. Heat the spoon in the flame until the sulphur melts and begins to burn with a pale-blue almost invisible flame, then place the spoon in a bottle of oxygen; the sulphur burns with a much brighter flame, emitting a beautiful violet-colored light. When the combustion is over, remove the spoon and observe, (1) the pungent suffocating smell of the gaseous combination of sulphur and oxygen which has been produced, (2) the seeming absence of anything in the bottle, the product of combustion being an invisible gas, (3) that a lighted paper plunged into the bottle will be immediately extinguished. The gas is called *Sulphur Dioxide,* SO_2, and the reaction is as follows:—

$$\underset{\text{Sulphur.}}{S_2} + \underset{\text{Oxygen.}}{2O_2} = \underset{\text{Sulphur dioxide.}}{2SO_2}$$

42. Sulphurous Acid. Pour a little water into the bottle and shake it up. Observe that the hand is drawn into the bottte, showing that a vacuum has been produced Taste a few drops of the water; it is sour. Pour a little of it into litmus solution in a test-tube; the solution is immediately reddened. The reddening of the vegetable blue coloring-matter is called the "*acid reaction.*" The gas has combined with a molecule of water to form *Sulphurous Acid,* H_2SO_3; thus:—

$$\underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphurous acid.}}{H_2SO_3}$$

43 Combustion of Carbon in Oxygen.

Exp. 7.—Cleanse the deflagrating spoon from sulphur, by holding it in the flame of the spirit-lamp till no smell of sulphur is perceived. Select a few splints of wood charcoal, place them upon the deflagrating spoon, and adjust it in a bottle of oxygen, as in the first experiment. The charcoal burns energetically in the gas, emitting much light and heat but little or no flame; observe that the product of combustion is an inodorous, invisible gas. Plunge a lighted taper into the bottle; the taper is extinguished. The charcoal (carbon) combines with the oxygen, forming *Carbon Dioxide*, CO_2; thus :—

$$\underset{\text{Carbon}}{C^*} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Carbon dioxide.}}{CO_2}$$

44. Carbonic Acid. Pour about half a tumblerful of water into the bottle and shake it vigorously. The hand will be drawn in, but not so forcibly as in Exp. 6; carbon dioxide is is only moderately soluble in water. Pour some of the water into a glass and taste it; it is slightly sour. Pour some of it into litmus solution; the solution is turned a dark-red color. The carbon dioxide has combined with the water, forming *Carbonic Acid,* H_2CO_3; thus :—

$$\underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Carbonic acid.}}{H_2CO_3}$$

45. Meaning of Test. Pour some clear lime water into the carbonic acid reserved from the last experiment; the clear liquid becomes milky, indicating the presence of carbonic acid, as will be hereafter explained. Lime-water is said to be a "test" for carbon dioxide or carbonic acid. A *test is a material for some experiment intended to bring out a property characteristic of the substance under examination, and by which the presence of that body may be detected.*

46. Origin of name Oxygen.—In the preceding experiments the products of combustion in oxygen when combined

---

[*] The number of atoms in a molecule of carbon is unknown, and in such cases the symbol for the atom is used in equations.

with water formed acids. The name OXYGEN (from the Greek *oxus*, sour, and *gennao*, I produce) was given by Lavoisier, under the mistaken impression that this element contained a principle common to all acids. This is now known to be an error. Later researches have brought to light a number of compounds containing hydrogen possessed of acid properties in which no oxygen is present. Nevertheless the name was not ill-chosen, for of the many hundreds of acids known there are only about six which do not contain oxygen.

47. Combustion of Sodium in Oxygen.

Exp. 8.—Take a small piece of metallic sodium, scrape it clean with a knife, heat it in the deflagrating spoon till it melts and begins to burn, then plunge it into a bottle of oxygen; it will burn with great brilliancy, and with a bright yellow flame. A white solid called *Sodium Oxide*, Na_2O, is formed. The reaction is represented by the equation :—

$$\underset{\text{Sodium.}}{2Na_2} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Sodium oxide.}}{2Na_2O}$$

48. Alkalies—Sodium Hydrate.—Add a little water to the bottle, shake it up and taste a few drops of the solution: it does not taste sour, but has a peculiar nauseous taste, and is soapy to the touch. Add a little to blue litmus solution; it is not reddened, but on the contrary becomes rather darker in color. Dip a glass rod into hydrochloric acid, and with it redden some blue litmus solution. Pour into this some of the solution of the sodium oxide; the red solution at once becomes blue. *The solution of sodium oxide acts upon vegetable colors in just the opposite way from acids, and will, in fact, neutralize their action.* It is called an *alkali*, and substances like this, which will restore the blue color of reddened litmus, are said to have an "*alkaline reaction.*" The sodium oxide has combined with a molecule of water, forming a substance called *Sodium*

Hydroxide or *Sodium Hydrate*, $NaOH$, better known by its common name of *Caustic Soda;* thus :—

$$\underset{\text{Sodium oxide.}}{Na_2O} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sodium hydrate.}}{2NaOH}$$

Similarly, if we burn the metal potassium in oxygen, we obtain *Potassium Oxide*, K_2O, which combines with water to form *Potassium Hydrate*, KOH.

49. Combustion of Magnesium in Oxygen.

Exp. 9.—Burn 10 or 12 centimetres (4 or 5 inches) of Magnesium ribbon in oxygen. A white solid called *Magnesium Oxide*, MgO, is formed; thus :—

$$\underset{\text{Magnesium.}}{2Mg} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Magnesium oxide.}}{2MgO}$$

50. Bases—Magnesium Hydrate. Pour a small quantity of water into the bottle containing the magnesium oxide, and shake it vigorously; it does not seem to dissolve in the water. Add some of it to blue and reddened litmus solutions; it has apparently neither *acid* nor *alkaline* reactions. Dip a piece of white blotting paper in reddened litmus solution, put it into the liquid and leave it for some time; it becomes blue. Half fill a test-tube with water, and add to it *one* drop of nitric acid; the solution will readily redden blue litmus paper when dipped into it. Pour the solution into the bottle in which the magnesium was burnt, and shake it vigorously; it will no longer redden either blue litmus paper or blue litmus solution. The acid has been *neutralized*. *A body possessing the characteristic of neutralizing an acid, either partly or entirely, is called a* BASE. An *alkali* is only a base, which is freely soluble in water. Acids and bases will be fully discussed in a future chapter. It may be stated here that the characters of taste and reaction belong to all well-marked acids and bases which are soluble in water; but they do not belong to *all* the acids and bases. In this case a molecule of water has combined with the magnesium oxide to form Magnesium Hydrate, $Mg(OH)_2$; thus:—

MgO	+	H_2O	=	$Mg(OH)_2$
Magnesium oxide		Water.		Magnesium hydrate.

51. Combustion of Iron or Steel in Oxygen.

Exp. 10.—Take a piece of thin watch-spring, which may be obtained from any watchmaker, heat it in the flame of a spirit-lamp till it is red-hot, and allow it to cool; it will then have lost its elasticity. Coil it into a spiral around a glass tube clean one end with a file, twist it round a bit of charcoal, and fasten the other end to the cap of the deflagrating spoon, and plunge it into a quart bottle of oxygen, on the bottom of which there is at least an inch of water. The burning charcoal heats the steel to redness, which then combines with the oxygen burning brilliantly, forming *Magnetic* or *Black Oxide* of Iron, Fe_3O_4, and throwing out abundance of sparks. The sparks are red-hot carbon contained in the steel, which also combine with the oxygen, forming carbon dioxide. The oxidized iron falls to the bottom in black globules, which are so hot that they are apt to melt into the glass and crack it, unless they have to pass through a considerable depth of water. The reaction is—

$3Fe$	+	$2O_2$	=	Fe_3O_4
Iron.		Oxygen.		Magnetic oxide of iron.

52. Magnetic Oxide.—Pick out the black globules and examine them, as in the preceding experiments. They do not produce an acid with water under any conditions, nor do they exhibit any alkaline or basic characters, and may, therefore, be classified as *indifferent* or neutral oxides.

53. Ferric Oxide.—Observe that the inside of the bottle becomes covered during the experiment with brick-red oxide, Fe_2O_3, which is called *Ferric Oxide.* This ferric oxide is insoluble in water, and, therefore, cannot be tested with litmus. It will, however, neutralize acids, and is, therefore, a base.

54. Oxides.—These experiments show that there are three kinds of oxides :—

I. *Acid producing oxides or Anhydrides.*—These oxides combine with water to form *acids*, or substances usually possessing a sour taste, which turn blue litmus red, and which neutralize basic oxides, forming a class of compounds termed *Salts.* All the non-metallic elements, with the exception of hydrogen and fluorine, form one or more compounds with oxygen, which, when united with water, yield acids, and in many cases intensely powerful acids. The following are of this class :—

Phosphorus pentoxide,	P_2O_5	yields with water	Metaphosphoric acid,	HPO_3.
Sulphur dioxide,	SO_2	"	Sulphurous	" H_2SO_3.
Carbon "	CO_2	"	Carbonic	" H_2CO_3.

II. *Indifferent Oxides.*—These oxides are indisposed to enter into combination with either acids or bases. They usually contain more oxygen than the basic oxides, a portion of it being loosely combined, and given off on heating. The following belong to this class :—

Manganese dioxide, MnO_2.
Black oxide of iron, Fe_3O_4.
Lead dioxide, PbO_2.

III. *Basic Oxides.*—Many of the metals, by their union with oxygen, give rise to oxides, the characteristic property of which is their power of neutralizing acids, thereby forming salts. Many of them combine with water to form a class of compounds called *Hydroxides* or *Hydrates.* The following are of this class :—

Sodium oxide,	Na_2O,	yields with water,	Sodium hydrate,	$NaOH$.
Magnesium oxide,	MgO,	"	Magnesium hydrate,	$Mg(OH)_2$.
Ferric oxide,	Fe_2O_3,	"	Ferric hydrate,	$Fe_2(OH)_6$.

55. Other Methods of Preparing Oxygen.

From Manganese Dioxide.—By heating the manganese dioxide to redness in a piece of gas-pipe or gun-barrel, it gives off the one-third of its oxygen; thus:—

$$\underset{\text{Manganese dioxide.}}{3MnO_2} = \underset{\text{Red oxide of manganese.}}{Mn_3O_4} + \underset{\text{Oxygen.}}{O_2}$$

From Manganese Dioxide and Sulphuric Acid.—By gently heating manganese dioxide and sulphuric acid when manganese sulphate and water are formed, and half the oxygen which the dioxide contained is disengaged; thus:—

$$\underset{\text{Manganese dioxide.}}{MnO_2} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Manganese sulphate.}}{MnSO_4} + \underset{\text{Water.}}{H_2O} + \underset{\text{Oxygen.}}{O_2}$$

By decomposing Water by Electrolysis.—This has already been effected.

56. Tests for Oxygen.

(1) **Exp. 11.**—Fill a test-tube with oxygen, pour into it a small quantity of a strong solution of caustic potash, and add an equal quantity of a strong solution of pyrogallic acid. Cut a piece about the size of a cent from an old rubber over-shoe, place it on the mouth of the tube to protect the thumb, and shake vigorously; *the solution in the tube becomes intensely brown.* This test will indicate the presence of very small quantities of oxygen. Invert the tube in water and withdraw the thumb; the water will rush into the tube and entirely fill it, showing that the solution has absorbed the oxygen.

(2) **Exp. 12.**—Put a few bits of clippings of sheet-copper into a test-tube. Add a small quantity of dilute nitric acid; an effervescence begins, and the tube is filled with red vapors. These soon disappear, and a colorless gas comes over. Pass this gas into a bottle of oxygen standing mouth downwards on the shelf of the pneumatic trough; immediately red vapors are formed which the water soon absorbs. The colorless gas is nitric oxide,

NO, which unites with the oxygen, forming higher nitrogen oxides; thus:—

$$\underset{\text{Nitric oxide.}}{2NO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Nitrogen peroxide.}}{2NO_2}$$

SUMMARY AND ADDITIONAL FACTS.

57. **History.**—Oxygen was discovered by Priestly in 1774, who, heating mercuric oxide with a burning glass, decomposed the oxide into oxygen and metallic mercury. Scheele, a Swedish apothecary, obtained it independently a year later. It was liquefied by Pictet, of Geneva, in December, 1877, under a pressure of 300 atmospheres and the temperature produced by the evaporation of liquid carbon dioxide in a vacuum.

58. **Occurrence.**—It is the most abundant and most important of all the elements. About one-fifth of the atmosphere consists of free oxygen; no less than eight-ninths of water is formed of the same gas. It makes up three-fourths of our own bodies, not less than four-fifths of every plant, and about one-half of the solid rock. Indeed, it seems to be the preponderating element of the globe.

59. **Properties.**—Oxygen is a colorless, odorless, and tasteless gas. It is a little heavier than air, its specific gravity being 1·1056, air being taken as unity, and 16, or more accurately 15·96, when hydrogen is taken as the unit. It is only slightly soluble in water, 100 volumes of which at ordinary atmospheric temperature dissolve about three volumes of water. On the solution of this small quantity of water depends the existence of fishes and other forms of animal life in the waters of rivers, lakes, and seas. Fishes pass water continually through their gills, in which the oxygenation of their blood takes place.

Oxygen is a great supporter of animal life. Pure oxygen, however, differs from dilute oxygen in its effects on life. A small animal placed in a vessel of oxygen soon dies of excessive excitement, produced by a too rapid oxidation of its blood. It is equally necessary to vegetable life, being needful to the development of the buds of maturer plants, and also being absorbed by their roots. Hence, if much earth is piled about a healthy tree, the tree soon sickens; while, on the contrary, if its roots find their way into a drain, the tree grows better than ever.

In all cases in which bodies burn in oxygen, the substance burnt combines with the oxygen to form new substances, and the heat and light developed are a consequence of this union. The weight of the product corresponds exactly with the weight of the body burned, *plus* the weight of the oxygen consumed. *The combination of two or more bodies, when accompanied by the development of heat and light, is termed* COMBUSTION, the body burnt being the *combustible*, and the body in which the burning takes place the *supporter* of combustion. These are, however, merely relative terms, as will be shown when combustion is discussed.

The combination of oxygen with other elements is termed oxidation. In all cases of oxidation heat is developed, but it depends on the rapidity with which the oxidation is effected whether light is also produced. Thus, when iron is burnt in oxygen, the combination of the two elements is effected with great rapidity, a large amount of heat is developed within a short space of time, and the product of combustion becomes incandescent. When, however, iron oxidizes or rusts slowly, no light is produced, although more heat is developed than when the same weight of iron is burnt in oxygen, the oxide, Fe_3O_4, being formed in the latter, and the oxide, Fe_2O_3, in the former case.

Oxygen combines with all the non-metallic elements, except fluorine. All metals combine with oxygen. When metals are easily volatile, as is the case with magnesium and zinc, they burn with a bright flame in oxygen. A few metals, such as gold and platinum, cannot be made to unite directly with oxygen, even at the highest temperatures. The oxides of such metals can, however, be prepared by indirect means.

Oxygen readily combines with oils, woollen refuse, greasy tow, moist hay, etc., the heat developed by their slow oxidation being to a great extent retained, as they are bad conductors of heat, until finally it is sufficient to inflame them. This is termed *Spontaneous Combustion*, and it is one of the most common sources of fire, both in manufactories and on board ship, and in hay-ricks in which the hay has been put up damp, for moisture greatly assists the process of slow combustion.

By oxidation earth, air, and sea are purified from contamination; noxious vapors and pestilential effluvia being destroyed by a process of burning, more slow indeed, but as real as takes place in a furnace. The offensive impurities which flow into rivers and lakes are oxidized by the dissolved gas, and thus the waters are kept pure.

60. Tests.

(1) A glowing splint of wood bursts into flame when plunged into oxygen, and the purity of the gas may be roughly estimated by the degree of sharpness with which the wood bursts into flame.

(2) Pure oxygen turns a solution of pyrogallic acid in strong potash brownish-black, and is completely absorbed by it.

(3) With nitric oxide, oxygen gives red fumes of nitrogen peroxide.

SECTION II.

OZONE.

Symbol, O_3. *Molecular Weight*, 48.

PREPARATION.

61. Formed in Cases of Slow Oxidation or Combustion.

Exp. 13.—Carefully scrape a stick of phosphorus until quite clean, under water, place it in a wide-mouthed bottle, pour in enough water to half cover it, and place a glass plate upon the mouth of the bottle; white fumes of phosphorus trioxide, P_2O_3, arise from the phosphorus, but are soon absorbed by the water, at the same time Ozone is formed. The phosphorus combines gradually with part of the oxygen of the air in the bottle, while some of the remainder is converted into ozone, $3O_2$ becoming $2O_3$. Put a piece of starch about the size of a large shot into a test-tube, and quarter fill with water, shake up and then boil; add a fragment of about the same size of potassium iodide, KI, and allow it to dissolve. Dip some pieces of white paper in the solution, and, after the phosphorus has been in the jar for twenty minutes or half an hour, introduce the paper; it immediately becomes blue. This is the ordinary test for the presence of Ozone.

Ozone can also be formed by the passage of a series of electric sparks through air or pure oxygen; and it may be recognized by its odor whenever an electric machine is worked. The quantity of oxygen thus changed is small, but if a *silent electric discharge* is passed through the gas, care being taken to avoid sparks, a much larger proportion of oxygen undergoes this transformation.

PROPERTIES OF OZONE.

62. Heavier than Air.

Exp. 14.—Lay a piece of test paper in the bottom of a tumbler and gradually invert a bottle of Ozone, prepared as in Exp. 13, over it; the test paper immediately becomes blue.

63. Oxidizing Power.

Exp. 15.—Suspend a bright silver coin in a bottle of ozone; in a few minutes it will be covered with a grey deposit of silver oxide.

64. Bleaching Power.

Exp. 16.—Into a jar of air ozonized by phosphorus pour a little dilute solution of indigo; it is at once decolorized. Moistened litmus-paper is immediately bleached when introduced into a bottle of ozonized air.

The bleaching and disinfecting of bodies by ozone are owing to their oxidation. Strips of test-paper exposed to the air, and shaded from the sun for a few hours, will frequently be found to have turned blue, especially in country places.

When substances are oxidized by ozone no diminution of the volume of the gas takes place. The density of ozone is found to be $24 (H = 1)$, that of oxygen being 16, so that ozone is half as heavy again as oxygen; therefore the molecule of ozone must contain three atoms. At a temperature of about 260° C. it is reconverted into ordinary oxygen, the gas returning to its original volume; thus:—$2O_3 = 3O_2$.

QUESTIONS ON OXYGEN.

1. From what substance was oxygen first prepared, and by whom?

2. Describe the preparation of oxygen from potassium chlorate, and express the reaction by an equation.

3. I heat potassium chlorate in a flask until effervescence commences; what chemical change takes place? I then drop into the flask some manganese dioxide; what happens?

4. Why is manganese dioxide generally mixed with potassium chlorate for the preparation of oxygen? What other substances might be substituted for it? Explain the action of such substances.

5. The following bodies are separately burned in oxygen: (*a*) phosphorus, (*b*) sulphur, (*c*) carbon, (*d*) magnesium, (*e*) sodium, (*f*) iron. Give the names of the oxides formed in each case, and express all the reactions by equations.

6. State the action of water on each of the oxides in the preceding question, and express each reaction by an equation.

7. Describe by equations as many processes as you can for the preparation of oxygen gas.

8. Oxygen is said to be a great supporter of combustion. Illustrate this statement by a description of one or two experiments.

9. Explain the difference between the red rust of iron and the black scales found about a forge.

10. A watch-spring is burned in a closed vessel of oxygen. State (*a*) whether the weight of the bottle and its contents is affected by the combustion; (*b*) what is the nature of the product formed by the combustion? (*c*) whether the grey metallic-looking globules, found at the bottom of the bottle after the experiment, consists of metallic iron; (*d*) whether the whole of the oxygen originally filling the bottle is still present, and, if so, in what form.

11. At the bottom of a glass flask some sand is placed, and on the top of it a fragment of phosphorus; the flask is then filled with dry oxygen, closed air-tight, and the whole weighed; the flask is then warmed until the phosphorus kindles, and when quite cold re-weighed; the stopper is then removed for an instant, and a third weighing is made. How will the three weighings differ?

12. If air were used instead of oxygen in the previous question, what difference would be noticed in the weighings?

13. What are the principal properties of oxygen, and how is it distinguished from other gases?

14. How does ozone occur in nature? How may oxygen be converted into ozone?

15. How is the presence of ozone usually detected?

16. What change in bulk occurs when oxygen is converted into ozone?

17. Compare the properties of ozone with those of oxygen, and show how they differ from each other.

xygen : (*a*)
) sodium,
case, and

preceding

can for the

Illustrate
ts.
on and the

gen. State
affected by
formed by
g globules,
consists of
nally filling

ced, and on
filled with
the flask is
quite cold
and a third
?
us question,

nd how is it

gen be con-

nverted into

oxygen, and

SECTION III.

HYDROGEN.

Symbol, H. *Atomic Weight*, 1. *Molecular Weight*, H_2, 2.

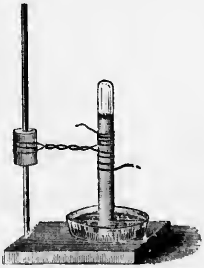
PREPARATION.

65. By Decomposing Water by the Galvanic Current.

We have already seen (Chap. I., Art. 19) that hydrogen may be obtained from water by sending a galvanic current through it, causing it to be resolved into its constituent gases. This method, however, is by far too costly to be employed on a large scale.

66. By the Decomposition of Water by Sodium or Potassium.

Exp. 1.—Boil some water ten or fifteen minutes, that all the air may be expelled from it; let it cool, and fill a saucer and a large and strong test-tube with it; close the mouth of the test-tube with the thumb and insert it under the water in the saucer. Should the mouth of the test-tube be too wide to be closed by the thumb, place a small watch-glass or a piece of thick blotting paper under it and rapidly insert it in the saucer. Support the test-tube with its mouth just under the water by means of a copper wire twisted tightly round a cork fitted on to a retort stand Now place on the end of a wire a piece of sodium, not larger than a small pea, and thrust it rapidly under the mouth of the tube. The metal frees itself from the wire, and, as it is lighter than water, ascends into the tube, floating there with a rotary motion. A gas is evolved from the water, and collects in the upper part of the tube. When the tube is full, place a glass plate under it and raise it from the water; invert it and rapidly apply a



light to its mouth: the gas will burn with a pale blue flame—as in Exp. 19, Art. 19—and is readily recognized as hydrogen. Add reddened litmus to some of the water in the saucer, and it will immediately become blue, showing that the water now contains an *alkali*. On evaporating the water in the saucer, this alkali is found to be sodium hydrate, $NaOH$. The sodium must, therefore, have replaced one-half of the hydrogen in the water, in the manner shown in the following equation:—

$$\underset{\text{Water.}}{2H_2O} + \underset{\text{Sodium.}}{Na_2} = \underset{\text{Sodium hydrate.}}{2NaOH} + \underset{\text{Hydrogen.}}{H_2}$$

Exp. 2.—Lay a piece of blotting paper on the surface of the water in the saucer, and throw upon it a small piece of sodium; an energetic decomposition of the water takes place, and in a few seconds the sodium will apparently burst into flame, and burn with a bright golden color. The apparent combustion of the sodium is really due to the burning of the hydrogen set free by the metal, which is inflamed by the intense heat which accompanies its evolution. This experiment differs from the preceding one only inasmuch as the hydrogen is burnt as it is liberated, while in the former case it is collected. The sodium hydrate may be rendered evident as before by the addition of reddened litmus solution to the water.

If potassium had been used instead of sodium in the preceding experiment, the blotting paper might have been dispensed with. The potassium glides about with a hissing noise, decomposing the water much more violently than sodium, the hydrogen evolved burning with a violet flame, potassium hydrate, KOH, remaining in solution in the water.

In these experiments care must be taken not to hold the face too near when the flame has ceased; for there remains a globule of the metal, which is in a melted state, and when it cools down to such a temperature as to permit the water to come in contact with it, steam is rapidly generated, and the melted metal is blown out of the water.

lue flame—
s hydrogen.
saucer, and
water now
the saucer.
The sodium
rogen in the
on :—

en.

rface of the
of sodium;
ce, and in a
flame, and
mbustion of
ydrogen set
heat which
rs from the
urnt as it is
The sodium
addition of

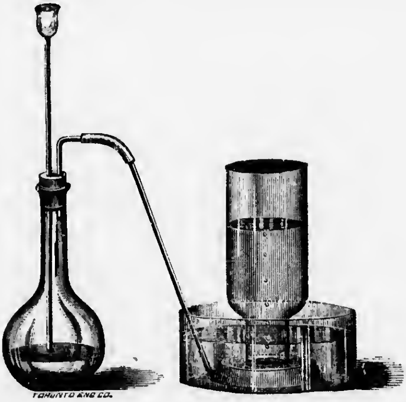
he preceding
pensed with.
decomposing
he hydrogen
drate, KOH,

hold the face
ins a globule
t cools down
ne in contact
ted metal is

---

67. By the action of Zinc on dilute Sulphuric Acid.

Exp. 3.—The most convenient mode of preparing hydrogen gas for ordinary use, where absolute purity is not requisite, is by the action of dilute sulphuric acid on zinc. Take a strong flask, with a flat bottom, of about 250 (10 oz.) cubic centimetres capacity, fit to it a good sound cork which has been previously well soaked in melted paraffine. Take a funnel-tube and a piece of glass tubing bent once at right angles, and bore in the cork two holes of such a size as to fit them, taking care not to make the holes too near the edge of the cork nor too near each other. Fit the funnel-tube into one of these holes so that it may reach nearly to the bottom of the flask, and join to the other tube, by means of a short piece of india-rubber tubing, a bent delivery-tube. Put 30 grams (10 oz.) of granulated zinc*



---

* Zinc may be *granulated* by melting it in an iron ladle and pouring it into a pail of water. If the melted metal is poured from the height of a yard or more above the surface of the water, the granules are spongy and very thin, presenting a large surface compared with their weight; whilst solid heavy granules are obtained if the zinc is poured at a distance of a few inches only above the water. The former kind is most convenient when a rapid current of hydrogen is required.

or zinc clippings into the flask, inclining it to one side, and gently sliding the zinc down the neck, taking care that it does not fall heavily against the bottom. Fit the cork into the neck of the flask, and arrange the apparatus so that the delivery-tube may be under the shelf of the trough. A straight tube, to the upper end of which a small funnel is united by a cork, may be used instead of a funnel-tube. Pour through the funnel enough water to cover the zinc to the depth of about one centimetre, and try whether the joints are tight by blowing through the delivery-tube till the water rises in the funnel, then pressing the connecting india-rubber with the fingers, and observing if the water remains in the funnel or descends very slowly. If it descends rapidly, wet the cork and push it still farther into the flask ; if there is still leakage, it may be detected by the bubbling of the water through the cork, and may be stopped by a little sealing wax, or more conveniently by a mixture of equal parts of bees-wax and turpentine, with a little Venetian red to give it color. Now pour through the funnel-tube sulphuric acid in small quantities at a time. The disengagement of the gas commences immediately, and when it slackens it may be invigorated by a little more acid. Great care must be taken not to add too much acid, or the liquid in the flask will froth over. Should it exhibit a tendency to do so, pour some water down the funnel to dilute and cool the acid. If the zinc happens to be very pure the sulphuric acid will act upon it very slowly. In that case a few drops of copper sulphate will at once cause energetic action. Fill two bottles with the mixture of air and hydrogen which first escapes from the flask and reject it. This precaution is important, as it will be shown that air forms with hydrogen a mixture which explodes upon contact with a light. As soon as a bottle is filled, place a glass plate or small saucer under it, lift it out of the trough, and place it on the table, mouth downwards. It is necessary to collect only one or two bottles of the gas at first, as the collection may go on while the experiments are being performed.

One ounce of zinc is sufficient to liberate from the acid about 2½ gallons of gas, or 30 grams will furnish about 10 litres.

Scraps of iron may be substituted for zinc; but in this case the gas is less pure. It has a disagreeable odor, due to the presence of compounds of carbon and hydrogen, but these may be removed by passing the gas through tubes filled with fragments of wood-charcoal.

68. Whence comes the Hydrogen? The hydrogen must either come from the zinc, the water, or the sulphuric acid. It cannot come from the zinc, for that is an element; nor can it come from the water for that is not decomposed by zinc alone, and besides it remains in the bottle unchanged when the experiment is finished. It must, therefore, come from the sulphuric acid. Now a molecule of sulphuric acid is represented by the formula H_2SO_4. There remains in the flask, when all the hydrogen has come off, water and zinc sulphate, and the formula of the latter is $ZnSO_4$. It appears, therefore, that one atom of zinc has *replaced* two atoms of hydrogen to form a molecule of zinc sulphate. The molecule of zinc, like the molecule of mercury, is supposed to contain only one atom; hence the reaction is expressed by the following equation:—

$$\begin{array}{ccccccc} Zn & + & H_2SO_4 & = & ZnSO_4 & + & H_2 \\ \text{Zinc.} & & \text{Sulphuric acid.} & & \text{Zinc Sulphate.} & & \text{Hydrogen.} \end{array}$$

69. Use of the Water. In the preceding equation no account has been taken of the water which was added. The water remains unchanged in the flask after the experiment is finished. Had no water been present, the zinc sulphate formed would have coated the surface of the metal, and thus have protected it from the further action of the acid.

PROPERTIES OF HYDROGEN.

70. Combustibility.

Exp. 4.—Fill a test-tube with the gas, and observe that it is colorless, and that it has a disagreeable smell. This is almost always the case with hydrogen prepared by this method. The smell is caused by the presence of minute quantities of compounds of hydrogen with sulphur, arsenic, or carbon ; but the gas prepared with pure zinc and pure sulphuric acid is quite free from smell. Take a bottle of hydrogen, hold it mouth downwards, and apply a lighted taper to its mouth ; the gas takes fire, and burns with a pale, almost invisible flame. Pass the taper further up into the jar ; it is extinguished Draw it out slowly ; it is rekindled. Now turn the mouth of the bottle upwards ; the flame will pass quickly down the jar, and the gas will be found to have entirely disappeared.

Hence we see that *Hydrogen is a combustible gas, but does not support combustion in the ordinary sense of the term.*

71. Levity.

Exp. 5—Take two bottles of hydrogen, place one on the table, mouth upwards, and hold the other in the left hand, mouth downwards. After about ten seconds, apply the burning taper to the bottle in the left hand ; the hydrogen takes fire, with a slight explosion, and burns, showing that the gas still remains in the bottle. Lift the other bottle from the table, invert it, and thrust the taper into it, the taper burns as in the outer air, showing that the gas has escaped.

Exp. 6.—Take an empty bottle and hold it mouth downwards in the left hand. Take a bottle of hydrogen in the right hand, and bring its mouth close to the edge of the first bottle, slowly depressing the closed end till its mouth is brought under the empty bottle. Place the bottle in the left hand on the table, mouth downwards. Thrust a burning taper into the mouth of the bottle in the right hand ; the gas will be found to have escaped. Now raise the other bottle and bring the burning taper to its mouth ; the gas will take fire with a slight explosion, showing that the hydrogen has ascended, and displaced the air in the bottle just as it displaces the water in filling a bottle at the pneumatic trough.

Hence we see that *Hydrogen is much lighter than air.*

72. Displacement.

Exp. 7.—Owing to the lightness of hydrogen, it may be collected by *upward displacement.* Hold a bottle mouth downwards, and put the delivery-tube in it so that it may reach nearly to the bottom. In a short time the gas will have displaced the heavier air, and the bottle will be found full of hydrogen. To ascertain when the bottle is full, hold a piece of smoking paper under it; the smoke readily ascends through the air, but not through the lighter hydrogen.

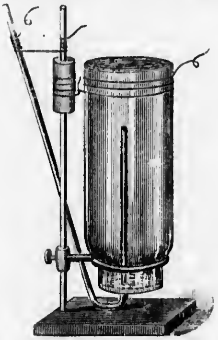
Exp. 8.—HYDROGEN SOAP-BUBBLES. The lightness of hydrogen may also be shown as follows: Cut a little castile-soap into thin shavings, and dissolve it in rain-water, making a saturated solution. To two volumes of this solution add one volume of glycerine. Attach to the delivery-tube of the hydrogen flask, by a piece of rubber tubing, the stem of a tobacco pipe. Pour some of the soap-solution into a saucer, dip the bowl of the pipe into it, and let the gas blow a bubble. While the bubble is small, turn the mouth of the pipe upwards, The bubble, having attained a diameter of several inches, will break away, or else may be easily detached by a sudden movement of the pipe downwards. It will then rapidly rise.

73. Explosive Mixture of Hydrogen and Air.

Exp. 9.—Take a soda-water bottle and fit to it a perforated cork without a tube, or with only a quill, so that, if the cork flies to the ceiling and falls, it will do no harm. Place some granulated zinc in it, pour in some cold dilute sulphuric acid (about 1 of acid to 8 of water), and insert the cork, but not too tightly.

Hold a lighted taper to the orifice, and in a short time a loud explosion will occur. Allow sufficient time to elapse so that the air may be expelled. No explosion occurs, but the gas burns quietly at the orifice.

Exp. 10.—Take a wide-mouthed bottle, and bend a glass tube in the form of a siphon, so that the shorter arm may be a little longer than the bottle. Support the bottle, mouth downwards, on one of the rings of the retort stand. Place the shorter arm in the bottle, the longer arm extending upwards and tied to the retort-stand. Cover the mouth of the bottle with a piece of brown paper gummed to the glass, the siphon-tube passing through it. Connect the longer arm by rubber-tubing with the generating flask, and fill the bottle by upward displacement. Remove the rubber tubing, and the hydrogen being lighter than air, will be siphoned upwards, just as water is siphoned downwards. Apply a light to the end of the longer arm, and hydrogen is seen to burn with its usual non-luminous flame. After a short time, however, this flame flickers, emits a musical note, at first shrill, but gradually deepens to a bass sound, until after a time distinct beats are heard, and at last, when the exact proportions between the hydrogen and the air, which enters through the pores of the paper, have been reached, the flame is seen to pass down the tube, enter the bottle, and the whole mass unites with a sudden and violent explosion, but quite harmless if the mouth of the bottle is sufficiently wide. Explosions are dangerous only when the gases are so confined that they cannot easily escape when expanded by heat. Hence bottles should be wide-mouthed, and thin glass vessels should not be used.



74. Precautions.

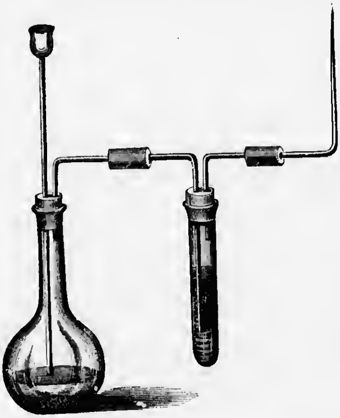
From the preceding experiments we see that no light should ever be brought in contact with the contents of the bottle in which hydrogen is generated, nor with any large quantity of the gas, until its non-explosive character has been demonstrated. This may be done by bringing the delivery-tube within the mouth of the test-tube. We have seen that hydrogen can be poured upwards, therefore, the tube will soon fill with the gas. After about ten seconds, remove it, still holding it mouth downwards, and apply a light to its mouth. If the hydrogen burns tranquilly at the mouth of the tube, the gas may be considered pure, but if it explodes with a whistling pop, further time must be allowed.

75. Cause of Explosions.

The explosion of the mixture of hydrogen and air is due to the sudden expansion caused by the heat generated in the combination of the hydrogen with the oxygen through the mixture. After the explosion of the mixture of hydrogen and air (oxygen and nitrogen), the substances present are steam and nitrogen, which are expanded by the heat developed in the combination to a volume far greater than the vessel can contain, so that a portion of the gas and vapor issues very suddenly into the air around, the collision with which produces the report.

76. Heat of Combustion.

Exp. 11.—Take a test-tube 6 in. by 1 in., fit it with a cork through which pass two tubes bent at right-angles, one of them reaching nearly to the bottom of the tube, the other just passing through the cork. Twist a short piece of copper-wire into a spiral and put it in the bottom of the tube, and nearly fill the tube with pumice stone, moistened with sulphuric acid, which readily absorbs moisture, and will thoroughly dry the gas. Attach the longer tube to the generating flask by



means of a cork, and to the other tube attach a small glass tube, about 20 centimetres long, drawn out so as to form a rather fine jet, and supported in a vertical position. If the hydrogen is not coming off with sufficient rapidity, add a little sulphuric acid, and before applying a light to the jet, ascertain that the hydrogen is not mixed with air. Kindle the gas and note that the flame at first is of a pale bluish color, but soon becomes a bright golden, owing to the sodium in the glass. Twist a small piece of thin sheet tin into a funnel, and place it over the jet; the flame becomes nearly colorless. Hold a fine platinium wire in it; the wire becomes nearly white hot and emits much light. Hold in it a small piece of caustic lime or chalk, with a fine point or edge; it soon produces a brilliant white light.

77. Product of Combustion.

Exp. 12.—Invert over the burning jet a large dry, wide mouthed bottle, the inner surface is quickly bedewed with moisture, and presently drops of liquid trickle down the sides and collect at the shoulder. When some drops of the liquid have been collected, test it with blue and reddened litmus paper; it is neutral, that is, it has neither an acid nor an alkaline reaction. Throw a little bit of potassium upon it; the potassium bursts into flame. Water is the only neutral liquid on the surface of which potassium will burn. Now, since the drying-tube completely removes moisture from the unburnt gas, and the latter does not bedew a cold surface against which it may be allowed to impinge, the liquid we observe to be deposited from the flame must be a product of the combustion of hydrogen in air. It is from this property that hydrogen derives its name (Greek, *hudor*, water, and *gennao*, I give rise to).

78. The Hydrogen Harmonicum.

Exp. 13.—Take a glass tube, open at both ends, about one centimetre wide and 30 centimetres long, and slowly pass the jet up into it, the flame is seen suddenly to elongate and a musical note results. The note emitted depends on the diameter and length of the tube, consequently tubes varying in these particulars may be used to produce different sounds. By raising or depressing the tube, the intensity of the sounds may be greatly varied. Ordinary wide-mouthed bottles may be used instead of tubes. These musical notes are really a succession of detonations due to the periodic combination of the atmospheric oxygen with the issuing jet of hydrogen, and succeeding each other with such swiftness as to prevent the ear observing the intervals between them. They may be produced by any combustible gas burned in the same way.

79. Reducing action of Hydrogen.

Exp. 14.—Take a hard-glass tube, about 20 centimetres long and 1 centimetre in diameter, and draw out one end to a moderately wide jet. Place midway in it a thin layer of copper oxide, CuO, connect it by a cork and glass tube with the drying tube, which is itself joined to the generating flask. Generate hydrogen as usual and keep up a steady current through the apparatus. When all the air has been expelled, apply heat to the tube so as to raise the temperature of the copper oxide to a low red heat. The oxide soon begins to glow, and steam issues from the end of the tube, and may be condensed in a cold flask. The lamp may be removed, and when the glowing ceases, the tube is seen to contain a red body, easily identified as metallic copper. The change is expressed by the following equation:—

$$\underset{\text{Copper oxide.}}{CuO} + \underset{\text{Hydrogen.}}{H_2} = \underset{\text{Metallic copper.}}{Cu} + \underset{\text{Water.}}{H_2O}$$

This reaction has been employed to determine the composition of water by weight.

Iron rust, Fe_2O_3, may be substituted for the copper oxide. Metallic iron will be left in the tube, and in a very fine state of division, in which condition the metal easily takes fire, when scattered out of the tube into the air, so rapidly does it combine with the oxygen again. The reaction is expressed by the following equation:—

$$\underset{\text{Ferric oxide.}}{Fe_2O_3} + \underset{\text{Hydrogen.}}{3H_2} = \underset{\text{Iron.}}{2Fe} + \underset{\text{Water.}}{3H_2O}$$

OTHER METHODS OF PREPARING HYDROGEN.

80. By the action of Zinc on dilute Hydrochloric Acid.

Exp. 15.—Add to a few pieces of granulated zinc, contained in a test-tube, some dilute hydrochloric acid till there is a brisk effervescence. Apply a light to the mouth of the tube, the sharp explosion and the well-known lambent flame show the presence of hydrogen. The reaction is expressed by the following equation:—

$$\underset{\text{Hydrochloric acid.}}{2HCl} + \underset{\text{Zinc.}}{Zn} = \underset{\text{Zinc chloride.}}{ZnCl_2} + \underset{\text{Hydrogen.}}{H_2}$$

81. By the action of Zinc on a strong Solution of Potash.

Exp. 16.—Add a little granulated zinc to a strong aqueous solution of caustic potash in a test-tube, to which adapt a cork and delivery-tube. On boiling, a gas will be slowly given off, which may be collected over water in the usual way. Some steam will pass over, but this will condense. On applying a light to a test-tube full of the gas, it will give the well-known hydrogen flame. The following equation expresses the reaction :—

$$\underset{\text{Zinc.}}{Zn} + \underset{\text{Potassium hydrate.}}{2KOH} = \underset{\text{Zinc potassate.}}{Zn(OK)_2} + \underset{\text{Hydrogen.}}{H_2}$$

This method is interesting from its theoretical bearing rather than from any practical utility. But if iron filings are added with the zinc, hydrogen is given off without the application of heat. The zinc dissolves, as above, but not the iron, which forms a galvanic circuit, and thus hastens the solution of the zinc. By this process very pure hydrogen can be prepared.

82. From the Decomposition of Water by Iron at a red heat.

Clean iron turnings or filings, free from rust, are placed in a piece of clean gas-piping, and are heated to low redness in a furnace. The cheapest furnace for this purpose is an ordinary plumber's furnace with holes pierced through its sides. Steam generated from a flask of boiling water is then conducted through the tube, and the liberated hydrogen is collected over water in the usual way. The reaction is as follows :—

$$\underset{\text{Iron.}}{3Fe} + \underset{\text{Water.}}{4H_2O} = \underset{\text{Magnetic oxide of Iron.}}{Fe_3O_4} + \underset{\text{Hydrogen.}}{4H_2}$$

The magnetic oxide of iron produced in this experiment is adherent, and a protection from further rust. Barff's process for preventing articles from rusting, is an application of this principle.

83. From the decomposition of Water by Magnesium.

The preceding experiment, though interesting, is quite a troublesome one. By substituting magnesium for iron the experiment may be conducted in a glass tube, and will be much more satisfactory.

Exp. 17.—Place about 3 feet of magnesium ribbon, in folds, in a hard-glass tube in such a way that the metal touches the glass in a number of points. Draw out one end of the tube to a pretty wide jet, and attach the other to a flask of water. Boil the water in the flask and allow the steam to flow until the air is expelled. Heat the tube sufficiently to prevent condensation at the mouth of the jet. Then heat the metal strongly at the extreme end. After a few moments it takes fire, burning brilliantly, and the escaping hydrogen may be lighted at the jet. It is best to keep the metal quite hot throughout. The reaction is as follows:—

$$\begin{array}{ccccccc} Mg & + & H_2O & = & MgO & + & H_2 \\ \text{Magnesium.} & & \text{Water.} & & \text{Magnesia.} & & \text{Hydrogen.} \end{array}$$

SUMMARY AND ADDITIONAL FACTS.

84. History.—Hydrogen was probably known as early as the sixteenth century, but its true nature was first ascertained by Cavendish in 1766. It was named hydrogen by Lavoisier. It was liquefied almost simultaneously and independently by two distinguished physicists, M. Cailletet, of Chatillon-sur-Seine, and M. Pictet, of Geneva. On January 10th, 1878, Pictet succeeded in liquefying hydrogen by a pressure of 650 atmospheres and at a temperature of —140°C. Cailletet demonstrated its liquefaction on December 30th, previously. On opening the stopcock, a *steel-blue* colored opaque jet of liquid hydrogen rushed out with a hissing noise, and at the same time a rattling was heard, as if small shot or hail had fallen to the ground. This was caused by the reduction of temperature due to its reassumption of the gaseous state. It is, therefore, only an accident of temperature and pressure that prevents it from possessing the ordinary metallic properties with which we are familiar in lead, silver, or copper. It is simply the vapor of a highly volatile metal.

85. Occurrence.—Hydrogen occurs in nature almost solely in a state of combination, although it has been found in the free state in small quantities, with other gases issuing from volcanoes. Free hydrogen has been found in the sun and other heavenly bodies. As it constitutes one-ninth of water, it is necessarily present in large proportions in all animals and plants. It is a constituent of all acids, and forms an essential portion of nearly all organic substances.

86. Properties.—Hydrogen is a tasteless, colorless, inodorous gas. It is the lightest substance known, being 14·435 times lighter than atmospheric air. It burns, but does not support ordinary combustion or animal life. In burning, it produces a greater heat than an equal bulk of any known substance. On combining with oxygen to form water, one gram of hydrogen yields heat enough to raise 34,462 grams of water from 0° to 1° C., and this is termed its *calorific power.*

Hydrogen is only slightly soluble in water, 100 volumes of the latter dissolving only 1·93 volumes of the former. Some metals absorb hydrogen in large quantities, especially platinum and palladium, the latter taking up no less than 370 volumes of the gas at ordinary temperature. On comparing hydrogen with oxygen, note their remarkable chemical dissimilarity. Oxygen combines with all the elements except fluorine, whereas the hydrogen compound with fluorine is easily formed, and is of great stability. The combining power between oxygen and metals is intense, whilst that between hydrogen and metals is almost nothing.

QUESTIONS ON HYDROGEN.

1. Give three distinct reactions by means of which hydrogen may be obtained from water. Give equations for each reaction.

2. How is hydrogen usually prepared, and what impurities may it contain when prepared by this method? Give an explanation of the process.

3. What occurs when zinc is heated with a strong solution of caustic potash or soda? Give an equation.

4. How could you obtain hydrogen from hydrochloric acid?

5. Name any four metals which, if dropped into hydrochloric or dilute sulphuric acids decompose these acids with the evolution of hydrogen. Give equations.

6. How would you prove that hydrogen is lighter than air?

7. A burning jet of hydrogen is placed successively in air, in oxygen, and in hydrogen. State exactly what occurs in each case.

8. What impurities are generally found in hydrogen prepared by the action of acids on iron, and how may they be removed?

9. I have two bottles of hydrogen; I hold one of them mouth upwards, and the other mouth downwards. At the expiration of half a minute, I plunge a lighted taper into each gas. Describe exactly what you would expect to take place in each case.

10. Describe the chemical change which takes place when steam is passed over red-hot iron filings.

11. Describe exactly how you would show the formation of water from burning hydrogen, and sketch the apparatus employed.

12. How is hydrogen found in the free state in nature? How does it occur in combination? Who first obtained it pure?

13. Hydrogen is said to be a metal. What evidence have we of this?

SECTION IV.

NITROGEN.

Symbol, N. Atomic Weight, 14 (14.01). Molecular Weight, N_2, *28.*

PRRPARATION.

By Abstracting the Oxygen from Air.

87. By burning Hydrogen Gas.

Exp. 1.—Take a quart bottle of hydrogen, hold it mouth downwards, and apply a light to it; the hydrogen burns only at the mouth of the bottle. When the hydrogen ceases to burn, place the bottle mouth downwards in water till it cools; the bottle is filled with a colorless gas, otherwise the water would rise in it. Still keeping the bottle inverted, immerse a lighted taper in it; the taper is immediately extinguished, and the gas does not burn. The gas in the bottle must be one of the constituents of the atmosphere. It is called *Nitrogen.* The other constituent, Oxygen, has united with the hydrogen, forming water, which is seen on the inside of the bottle. The nitrogen, heated by the burning hydrogen, becomes lighter, and therefore ascends in the bottle, and presses the hydrogen downwards.

88. By the active Combustion of Phosphorus.

Exp. 2.—Cover the bottom of a soup-plate, to the depth of half an inch, with water. Take a piece of chalk or crayon, hollow it out into a little cup, and float it on the water by means of a large flat cork. Into the cup put a piece of dry phosphorus, about the size of a large pea, set it on fire, and cover it with a quart bottle. Keep the hand on the bottle till the phosphorus goes out. The phosphorus combines with the oxygen of the atmosphere, forming phosphorus pentoxide, P_2O_5; thus:—

$$\underset{\text{Phosphorus.}}{P_4} + \underset{\text{Oxygen.}}{5O_2} = \underset{\text{Phosphorus pentoxide.}}{2P_2O_5}$$

The combustion goes on until nearly all the oxygen is removed from the included air. The air is at first expanded by the heat of the flame, and a portion of it escapes from the vessel; afterwards it diminishes in volume as its oxygen is removed, so that it is necessary to pour water into the plate to prevent the external air from passing into the bottle. After a time the white fumes will be absorbed by the water, leaving the enclosed nitrogen quite clear.

89. By the slow Combustion of Phosphorus in Moist Air.

Exp. 3.—If, instead of setting fire to the phosphorus, as in the last experiment, the bottle is simply placed over it, the phosphorus will gradually combine with the oxygen of the air, forming phosphorus trioxide, P_2O_3 ; thus:—

$$\underset{\text{Phosphorus.}}{P_4} + \underset{\text{Oxygen.}}{3O_2} = \underset{\text{Phosphorus trioxide.}}{2P_2O_3}$$

In two or three days all the oxygen will be removed, leaving nearly pure nitrogen, amounting to about four-fifths of the original bulk of the air.

Ordinary combustibles are not available for the preparation of nitrogen for two reasons: (1) They do not remove all the oxygen. (2) They introduce a contaminating gas. Phosphorus has neither disadvantage, and is generally employed.

PROPERTIES.

Exp. 3.—Place a glass plate under the mouth of the bottle, in the first experiment, and set it on the table, mouth upwards. Plunge a lighted taper into it; the taper is extinguished, while the gas itself does not take fire, thus showing it to differ from oxygen, in which the taper continued to burn, and from hydrogen, which extinguished the taper, but was itself inflamed.

Hence, *Nitrogen is neither combustible nor a supporter of combustion.*

Exp. 4.—Pour some clear lime-water into the bottle, close its mouth with the palm of the hand, and briskly agitate it; the lime-water is not rendered turbid. *This test serves to distinguish nitrogen from carbon dioxide.*

SUMMARY AND ADDITIONAL FACTS.

90. History.—Nitrogen was discovered by Dr. Rutherford, Professor of Botany in the University of Edinburgh, in 1772.

Scheele and Lavoisier independently proved that air is a mixture of the newly discovered oxygen and another gas which Lavoisier named *azote*. This azote, Chaptal in 1789 recognized as a constituent of nitre or saltpetre, and he, therefore, named it nitrogen. It was liquefied in 1878 by a pressure of 200 atmospheres.

91. Sources.—In nature the great store and source of nitrogen is the atmosphere, of which it forms four-fifths. In the mineral kingdom, especially in soils, it occurs in small quantities as an ingredient of saltpetre and of ammonia. It is a small but constant constituent of all plants, and in the animal it is a never-failing component of the working tissues, the muscles, tendons, and nerves, and is, therefore, an indispensable ingredient of food. Nitrogen also constitutes an essential part of many of the most potent and valuable medicines, as well as some of the most dangerous poisons, such as quinine, morphine, prussic acid, and strychnine.

92. Properties.—Free nitrogen, under ordinary circumstances, has scarcely any active properties, but it is best characterized by its chemical indifference to most other bodies. We have seen that it does not support combustion, neither does it burn. It cannot maintain respiration, so that animals perish if confined in it. Decay does not proceed in an atmosphere of this gas, and, in general, it is difficult to effect its direct union with other bodies.

Its specific gravity is ·9713; it is, therefore, a little lighter than air. It is very slightly soluble in water. Being so feebly adapted for combination, most of its compounds are more or less unstable, and many are explosive.

93. Tests.—In a free state we know it by its negative character; in combination as nitric acid or as ammonia, it may be recognised by the special tests for these bodies.

Uses.—These are confined to Nature as a diluent to air, and as a constituent of tissue.

QUESTIONS ON NITROGEN.

1. How does nitrogen occur in nature? Mention some gaseous and solid bodies which contain it.
2. When and by whom was nitrogen discovered?
3. How can you obtain nitrogen from atmospheric air? Express the reactions by equations, and give drawings of the apparatus.
4. Describe the characters of nitrogen, and show how it may be distinguished from oxygen and hydrogen.
5. What effect would an atmosphere of nitrogen produce on life and combustion?

SECTION V.

CARBON.

Symbol C. Atomic Weight, 12 (11.97). Molecular Weight Unknown.

94. Carbon exists in three allotropic states:—1. Charcoal, of which there are several varieties; 2. Graphite; 3. The Diamond.

Def.—When an elementary substance is capable of making its appearance in two or more forms having different properties, these are said to be *allotropic* modifications of the element.

I. PREPARATION OF CHARCOAL.

95. From Wood.

Exp. 1.—Light a thin splint of wood, and plunge it into a test-tube, held with its mouth downwards. The tube prevents the air from getting freely to the wood, causing a smothered burning, and thus a slender piece of charcoal is produced.

Wood consists substantially of carbon, hydrogen, and oxygen. The heat drives the hydrogen and oxygen off in the form of water, leaving the carbon behind.

96. Preparation on the large scale.—Charcoal is prepared by covering a pile of wood with earth or sods, in which openings are made at the top and bottom to admit a small quantity of air. The wood is then set on fire, and some of it slowly burns. The heat from the burning wood is used to drive off the hydrogen and oxygen from the rest of the logs, care being taken that sufficient heat is not produced to cause the whole mass of wood to burn. Since carbon is, under all circumstances, infusible and non-volatile, the charcoal retains the form of the wood, but the bulk has been much reduced, and its weight does not exceed one-fourth the weight of the wood.

97. From Vegetable Substances.

Exp. 2.—Take a small piece of writing paper, crush it into a lump, and put it into a narrow test-tube; ignite the paper by holding the end of the tube in the lamp. Take a slip of blue and a slip of reddened litmus paper, and as soon as a white smoke appears in the tube, dip into it the blue test-paper. After a moment take out the blue paper, and put in the red one. Observe (1) that the white paper heated in the tube is converted into a black substance, preserving the same size and shape; (2) that a brown oily liquid is deposited on the sides of the tube; (3) that the blue test-paper turns red, and the red test-paper remains unchanged in the volatile matter given off during the ignition; (4) that the black substance in the tube, when placed on platinum foil and strongly heated, burns away without flame, leaving nothing but a very small quantity of incombustible white ash. The black substance is *charcoal.*

From these experiments we infer (1) *That vegetable substances contain charcoal;* (2) *That they contain the elements of a volatile acid, which acid they produce when subjected to a red heat in close vessels.*

98. From Animal Substances.

Exp. 3.—Take a bit of woollen cloth, put it into a small test-tube. Prepare moistened slips of test-papers as in the preceding experiments. Hold the tube in the flame of the spirit-lamp, and put into it first the blue paper and then the red one. Observe (1) that the woollen cloth is converted into a black substance like charcoal; (2) that a brown oily liquid is deposited on the sides of the tube; (3) that the blue paper remains unaltered, and the red paper turns blue; (4) that the black substance taken from the tube, placed on platinum foil, and held in the flame of the spirit-lamp, burns away without flame, leaving only a very small quantity of white ashes.

From these experiments we infer (1) *That animal bodies contain charcoal;* (2) *That they contain the elements of a volatile alkali, which alkali they produce when subjected to a red heat in a close vessel.*

PROPERTIES OF CHARCOAL.

99. Absorptive Power.

Exp. 4.—Take a narrow-necked bottle and fit it with a good sound cork, previously steeped in paraffine, and fill it with ammonia gas. This may be done either by putting a few drops of ammonia into the bottle and shaking it well, or by placing it, mouth downwards, for a few minutes over the mouth of a bottle containing strong ammonia. Place in it a few pieces of newly-burned charcoal, and cork it up. After a few hours withdraw the cork, and it will be found that there is no odor remaining; the whole of the ammonia has been absorbed.

This power of absorbing gases depends upon the fact that all gases condense in greater or less degree on the surface of solid bodies with which they come in contact, and as charcoal is very porous, or possesses a large surface to a given mass, its absorbent power is proportionately great.

100. Decolorizing Power.

Exp. 5.—Take three tumblers, and place in each a tablespoonful of bone-black. Into the first tumbler put some blue or reddened litmus solution, into the second a solution of logwood or any other vegetable coloring matter, into the third a solution of potassium permanganate. Stir the solutions well with a glass rod. Fold three pieces of filtering or white blotting paper, so as to fit into a funnel. Pour the contents of each tumbler on a filter; observe that the filtrate is in each instance colorless or nearly so. In case the first portions of the filtrate happens to come through colored, they may be poured back upon the filter, and allowed to pass again through the charcoal.

In the purification of brown sugar, the coloring matters are removed in a manner similar to the foregoing, the colored syrup being filtered through layers of bone-black.

Beer or ale, thus treated, lose not only their color but their bitter taste. A solution of quinine sulphate, filtered thus through bone-black, is deprived of its bitter taste.

101. Deodorizing Power.

Exp. 6.—Put a few drops of a sulphuretted hydrogen solution into a test-tube, add a little animal charcoal, and shake the mixture thoroughly for half a minute; the liquid has now lost all the offensive smell of the gas, which has been absorbed by the charcoal.

Charcoal not only absorbs unpleasant effluvia evolved in the process of decay and putrefaction, but has the power, especially in contact with air, of oxidizing and destroying them. This property is retained by charcoal for a long time, and when lost it may be renewed by ignition. Hence charcoal filters are largely used for preventing foul sewer gases from polluting the air of houses, and charcoal respirators have been used to prevent the ingress of deleterious gases into the lungs. Trays filled with heated wood charcoal, placed in infected apartments, have proved very effective in absorbing noxious emanations. Charcoal filters are also largely employed for filtering water for drinking purposes. In its passage through the charcoal, the water has not only the organic and soluble coloring matters removed, but also undergoes aeration.

102. Reducing Power.

Exp. 7.—Heat a glass tube in the spirit-lamp, and draw it out to a point. Drop into the point of the tube a very small quantity of arsenious oxide, As_2O_3, and above it place a splint of wood charcoal. Heat the charcoal red-hot in the flame of the lamp, and gently raise the hand so as to bring the oxide into the flame without taking the charcoal out of it. The arsenious oxide will volatilize, giving up its oxygen to the carbon, and depositing metallic arsenic on the cold part of the tube.

The affinity of carbon for oxygen at a high temperature is very great. It deprives most metallic oxides of their oxygen, and thus brings them into the metallic state. It might almost be said that the art of metallurgy, as it now exists, is based upon the affinity of carbon for oxygen at a high temperature.

103. Indestructibility.—Charcoal, and carbon in most of its forms, is extremely indestructible, unless exposed to an elevated temperature. Hence stakes and fence posts, if charred before they are put into the ground, last very much longer than when this treatment is neglected. For the same reason it is a common practice to char the interior of tubs and casks destined to hold liquids.

104. Lampblack.—This form of charcoal is obtained by burning turpentine, resin, or other vegetable matters rich in carbon, with a limited supply of air. It is not pure carbon. It always retains a portion of incompletely burned compounds of carbon and hydrogen. It furnishes the most indestructible of black pigments, and has long been employed on this account as the basis of printing ink.

105. Animal Charcoal.—Charcoal, manufactured from animal substances, is called *animal charcoal*. When bones are strongly heated out of contact with the air, the variety of charcoal thus produced is called **Bone-Black**, and is much used by sugar refiners.

106. Graphite.—This is a crystalline form of carbon occurring in massive or hexagonal plates. It is also called plumbago, and is more familiarly known as black-lead. It is obtained from the earth in large quantities, and is used for the manufacture of lead pencils, and for giving a black polish to iron articles, such as stoves, etc., and for protecting them from rust.

107. The Diamond.—The diamond is another form of crystalline carbon, occurring in well-defined crystals belonging to the regular system. It is the hardest substance known. Besides its extraordinary value as a gem, it is used for cutting glass. Very small diamonds are said to have been lately prepared artificially by a Glasgow chemist. If the diamond is suspended in a cage of platinum wire, heated to bright redness, and then plunged in oxygen gas, it burns with a steady red light, and with the production of pure carbon dioxide.

108. Allotropic Forms.—Charcoal, graphite, and the diamond are but different forms of the element carbon. They differ in hardness, in color, in specific gravity, and in many other physical properties. They are alike infusible, alike able to resist the action of substances which attack most other bodies, alike in being combustible, and alike in the same weight of each yielding the same quantity of carbon dioxide when burned. Such phenomena as these afford strong grounds for believing that our present elementary substances may have a composite structure.

QUESTIONS ON CARBON.

1. Give an account of the different methods employed for preparing charcoal from wood. How would you demonstrate the preparation of charcoal on the small scale?

2. Carbon is said to exist in three allotropic modifications. Describe why diamond, graphite, and charcoal are considered to be modifications of the element carbon.

3. How may the presence of carbon in organic matter be shown?

4. What happens when a piece of perfectly dry charcoal is placed in a jar of ammonia gas? If the jar of ammonia is standing over mercury and a piece of dry charcoal is placed in it, what happens? What is the cause of the change?

5. What happens when charcoal is heated with a solution of indigo or logwood? Which kind of charcoal acts most readily on these bodies, and to what useful purpose is it applied?

6. How does charcoal act as a disinfectant?

7. How is animal charcoal prepared, and what are its properties?

8. Describe the allotropic forms of carbon. How would you prove that these different substances consist of the same element?

CHAPTER III.

CHEMICAL CALCULATIONS.

109. Amount of Material required to produce a given Weight of a Substance.

We have hitherto employed equations to express chemical changes only, but they also express the relative quantities of the elements which form the compounds acting on each other, or which take part in the changes, and hence they furnish the basis for *quantitative* calculations. Thus, the formation of carbon dioxide, by the action of oxygen on carbon, is expressed by the equation—

$$C + O_2 = CO_2,$$

which not only tells us that one atom of carbon unites with two atoms of oxygen to form a molecule of carbon dioxide, but also that 12 parts by weight of carbon (atomic weight of carbon $= 12$) unite with 32 parts by weight of oxygen (atomic weight of oxygen $= 16$) to form 44 parts by weight of carbon dioxide. From these considerations it is easy to calculate the quantities of carbon and oxygen involved in the formation of a given quantity of carbon dioxide obtained from known quantities of carbon and oxygen.

Ex. 1.—When carbon is burnt in oxygen gas, it forms carbon dioxide, CO_2. How much carbon dioxide by weight can be obtained by thus burning 15 grams of carbon?

Weight of CO_2 obtained from 12 grams of carbon $=$ 44 grams;
" " 1 " $= \frac{44}{12}$ "
" " 15 " $= \frac{15 \times 44}{12}$ "
$= 55$ "

Ex. 2.—How much mercury is required to form 20 grams of mercuric oxide?

For the sake of simplicity we shall use atomic instead of molecular equations.

Hence,

$$\underset{200}{\mathrm{Hg}} + \underset{16}{\mathrm{O}} = \underset{216}{\mathrm{HgO}}\ ;$$

∴ mercury required to form 216 grams of mercuric oxide = 200 grams;
" " 1 " " = $\frac{200}{216}$ "
" " 20 " " = $\frac{20\times200}{216}$ "
= 18·52 "

Ex. 3.—What weight of oxygen can be obtained from 100 grams of potassium chlorate ?

$$\mathrm{KClO_3} = \mathrm{KCl} + 3\mathrm{O}.$$

Weight of a molecule of potassium chlorate = 39·1 + 35·5 + 48 = 122·6
Oxygen obtained from 122·6 grams $\mathrm{KClO_3}$ = 48 grams;
" 1 " = $\frac{48}{122\cdot6}$ "
" 100 " = $\frac{100\times48}{122\cdot6}$ "
= 39·15 "

EXERCISE I.

1. I want 100 lbs. of oxygen; how many pounds of potassium chlorate must I take ?

2. I require 2 kilograms of oxygen; how much (1) mercuric oxide, (2) potassium chlorate, (3) manganese dioxide, shall I need ?

3. On heating some potassium chlorate 298 grams of potassium chloride were left; how many grams of chlorate were heated, and how many grams of oxygen were formed ?

4. What weight of hydrogen could I obtain from 70 grams of water by the action of sodium ?

5. How much potassium chlorate will furnish sufficient oxygen to unite with the hydrogen evolved by the action of 200 grams of sodium upon water ?

6. What weight of zinc must be added to dilute sulphuric acid to liberate 5 grams of hydrogen ?

7. What weight of iron is required to prepare 35 grams of hydrogen from hydrochloric acid ?

110. Relation of the Volume of Gases to Pressure. In Art. 22, Exp. 20, we see that when a confined mass of gas is compressed, its volume diminishes with increased pressure, and conversely increases in volume as the pressure diminishes. Thus, if the pressure on a given mass of gas is doubled, the volume is reduced to one-half, if trebled, to one-third, and so on. From experiments such as these, performed with diffierent gases, the following law has been established :—

111. *Boyle's Law.*—The volume which a gas occupies is inversely proportional to the pressure to which it is subjected, when the temperature remains constant.

This law was enunciated independently by Boyle and Mariotte; and Dalton further showed that it was applicable to the case of a mixture of gases.

Ex. 4.—At a constant temperature a quantity of gas occupies 25 cubic feet under a pressure of 10 lbs.; what space will it occupy under a pressure of 24 lbs.?

Space occupied under a pressure of 10 lbs. = 25 cubic feet;
" " 1 " = 10×25 "
" " 24 " = $\frac{10 \times 25}{24}$ "
= $10\frac{5}{12}$. "

112. Standard Pressure.—The average weight of the atmosphere at the level of the sea, in the latitude of Paris, is that of a column of mercury 760 millimetres (30 inches) in height, and this is taken as the *standard barometric pressure.*

Ex. 5.—A balloon containing 1200 cubic metres of hydrogen under a pressure of 770 millimetres of mercury ascends until the barometer stands at 530 millimetres; what volume would the gas in the balloon now occupy, supposing none to have escaped.

Space occupied under pressure of 770 mm. = 1200 cubic metres;
" " 1 " = 770×1200 "
" " 530 " = $\frac{770 \times 1200}{530}$ "
= 1743·4 "

EXERCISE I.

1. The height of the barometer is said to be, on a particular day, 740 mm.; what is meant by this statement?

2. The standard pressure is 760 mm. of mercury; how many inches of mercury corresponds to this number?

3. 1000 c.c. of hydrogen are measured under a barometric pressure of 740 mm.; what will the volume become under the standard pressure of 760 mm.?

4. At a constant temperature a mass of air occupies 18 cubic feet under a pressure of 7·5 lbs.; what space will it occupy under a pressure of 25 lbs.?

5. If, under the pressure of one atmosphere, a certain quantity of hydrogen gas occupies 50 cubic inches, under what pressure would it occupy 30 cubic inches?

113. The Relation of the Volume of Gases to Temperature.

From Experiment 20, Art. 22, we learn that gases increase in equal volume when heated, and decrease in equal volume when cooled. If we begin with a given volume of gas at 0°C. and measure the gas as we raise its temperature at a definite rate, we find that for each increase in the temperature of 1°C. the gas expands $\frac{1}{273}$rd of its volume at 0°C.; for example, 273 c.c. of a gas at 0°C. expands to 274 c.c. when the temperature is raised 1°C., or to 280 c.c. when the temperature is raised 7°C. Conversely, 273 c.c. becomes 272 c.c. when temperature is cooled 1°C.; that is, a gas diminishes by $\frac{1}{273}$rd of its volume for every degree of temperature, travelling down the scale. If the same ratio of volume to temperature were maintained, it follows that if a given mass of gas were cooled down to the temperature of -273°C. it would be reduced to a mathematical point; that is, all the molecular motion would cease, and so the point of no heat would be reached. Such a temperature has never been reached; nevertheless it is found convenient to take 273°C. below zero as the *absolute zero of temperature*, and the temperatures reckoned from this point are called *absolute temperatures.* On this scale all the temperatures are evidently positive.

114. **Coefficient of Expansion.** The fraction $\frac{1}{273}$, by which gases increase their volume at 0°C. for each degree of rise of temperature, is called the *coefficient of expansion of gases.*

From what has been stated respecting the expansion of gases we have the following law :—

115. ***Law of Charles.*** **The volume of a given mass of gas, under a constant pressure, varies directly as the absolute temperature.**

Ex. 6.—A certain mass of air measures 100 c.c. at 0°C. ; find its volume at 10°C.

The absolute temperatures are $10° + 273° = 283°$,
and $0° + 273° = 273°$.
Volume occupied by the gases at $273° = 100$ cubic centimetres ;
“ “ “ “ $1° = \frac{100}{273}$ “
“ “ “ “ $283° = \frac{283 \times 100}{273}$ “
$= 103{\cdot}66.$ “

Ex. 7.—A gas occupies 500 c.c. at 10°C. ; find its volume at − 10°C.

The absolute temperatures are $10° + 273° = 283°$,
and $-10° + 273° = 263°$.
Volume occupied at $283° = 500$ cubic centimetres ;
“ “ $1° = \frac{500}{283}$ “
“ “ $263° = \frac{263 \times 500}{283}$ “
$= 464{\cdot}6$ “

Ex. 8.—A litre of air is measured at 0°C. and 760 mm. pressure ; what volume will it occupy at 740 mm. at 15°.5C.?

The absolute temperatures are $0° + 273° = 273°$,
and $15°.5 + 273° = 288°.5$.
Volume at 273° and 760 mm. $= 1000$ cubic centimetres ;
“ 1° “ 760 “ $= \frac{1000}{273}$ “
“ 1° “ 1 “ $= \frac{760 \times 1000}{273}$ “
“ 288°.5 “ 740 “ $= \frac{288{\cdot}5 \times 760 \times 1000}{740 \times 273}$ “
$= 1085{\cdot}34$ “

EXERCISE III.

1. A certain quantity of gas occupies 67 cubic inches when the temperature is 10°C., and the barometer 28 inches; how many cubic inches will it occupy at 0°C., with the barometer at 30 inches?

2. A certain quantity of oxygen measures 155 c.c. at 10°C., and under a pressure of 530 mm. of mercury; what will the volume become at 18°.7C., and under a pressure of 590 mm. of mercury?

3. A glass globe holds ten litres. It is filled with oxygen at 0°C., under a pressure of 760 mm.; how much gas will escape when the temperature rises to 15°C., and the barometer falls to 752 mm.?

4. A room is calculated to contain 3000 cubic feet of air at 10°C., and under a pressure of 30 cubic inches of mercury; find what would be the volume of the same quantity of air if it were measured at 0°C. and 31 inches pressure.

5. 10 litres of oxygen are measured at 14°F., required the volume of the gas at 15°C.

6. A flask is filled with oxygen at 0°C. and 760°mm. pressure, and the flask is then tightly corked. The flask would burst if exposed to an outward pressure of 1500 mm. At what temperature would the oxygen exert this pressure, assuming the capacity of the flask to remain unaltered?

116. Unit of Volume—The Crith. One litre of hydrogen at 0°C. and under a pressure of 760°mm. of mercury weighs ·089578 grams, or approximately, ·0896 grams. So important is this hydrogen-litre-weight that Dr. Hofman denotes it by the term *Crith* (Gr. *Krithe*, a barley corn, and hence any small weight), so that

1 Crith = ·0896 grams.

It is of great importance to remember this number, since the weight of a litre of any other gas may be at once found by multiplying this number by half the molecular weight of the gas; thus :—

1 litre of oxygen $= \frac{32}{2} \times \cdot 0896 = 1 \cdot 433$ grams.
1 " carbon dioxide $= \frac{44}{2} \times$ " $= 1 \cdot 971$ "

Ex. 9.—What weight of potassium chlorate will be required to fill a gas bag of a capacity of 20 litres with oxygen at 0.C. and 760°mm. pressure, the temperature of the room being 15°C., and the pressure of the air at the same time being 750 mm.?

(1) *Find the relation between the potassium chlorate and the oxygen given off.*

$$\underset{39 \cdot 1 + 35 \cdot 5 + 48 = 122 \cdot 6}{KClO_3} = \underset{39 \cdot 1 + 35 \cdot 5}{KCl} + \underset{48}{3O}$$

(2) *Reduce the given volume to the standard conditions of temperature and pressure.*

The absolute temperatures are $15^\circ + 273^\circ = 283^\circ$,
and $0^\circ + 273^\circ = 273^\circ$.
Volume at 288° and 750 mm. = 20 litres;
" 1° " 1 " $= \frac{750 \times 20}{288}$ litres;
" 273 " 760 " $= \frac{273 \times 750 \times 20}{760 \times 288}$ litres;
$= 18 \cdot 70$ litres.

(3) *Find the weight of potassium chlorate necessary to furnish the above quantity.*

At 0°C. and 760 mm. 1 litre of oxygen = 16 criths $= 16 \times \cdot 0896$ grams;
" " 18·7 " $= 26 \cdot 808$ "
Potassium chlorate which yields 48 grams of oxygen $= 122 \cdot 6$ "
" " " 1 " $= \frac{122 \cdot 6}{48}$
" " " 26·808 " $= \frac{26 \cdot 308 \times 122 \cdot 6}{48}$
$= 685.47$ grams.

EXERCISE IV.

1. Exactly 100 litres of oxygen, at the normal temperature and pressure, were obtained by heating potassium chlorate: how much of this salt was used?

2. How much potassium chlorate must be used to yield 100 litres of oxygen at 30°C. and 380 mm. pressure?

3. How many litres of oxygen at 720 mm. pressure and 15°C. can be obtained by heating 261 grains of manganese dioxide (1 grain = ·0648 grams)?

4. What volume of hydrogen measured at 12°C. and 750 mm. is disengaged when 100 grams of zinc dissolve in dilute sulphuric acid?

5. A balloon requires 5 cubic metres of gas to inflate it; how many kilograms of sulphuric acid must be converted into zinc sulphate in order to evolve sufficient hydrogen to fill it?

6. A rectangular india-rubber bag, 1 metre long, 50 cm. broad and 30 cm. deep, is to be filled with hydrogen at 0°C. and 760 mm. pressure; how much zinc is required for the purpose?

117. Volume and Weight of Gases.

Since 1 litre of hydrogen = ·0896 grams,
22·32 " = 22·32 × ·0896 "
= 1·999 "
Therefore, 22·32 litres hydrogen = 2 grams approximately;
22·32 " oxygen = 32 "
22·32 " carbon dioxide = 44 "

Hence, *22·32 litres of any gas at 0°C. and 760 mm. pressure weigh a number of grams equal to the number expressing the molecular weight of the gas.*

If great accuracy is not required, the more convenient number, 22·4 litres, may be used instead of 22·32 litres.

Ex. 10.—Find the volume of 244 grams of carbon dioxide at 0°C. and 760 mm. pressure.

Vol. of 44 grams carbon dioxide = 22·4 litres;
" 1 " " = $\frac{22 \cdot 4}{44}$ "
" 244 " " = $\frac{244 \times 22 \cdot 4}{44}$ "
= 124·2 "

Ex. 11.—What volume of sulphur dioxide is formed on burning 8 grams of sulphur?

$$\underset{82}{S} + O_2 = \underset{64}{SO_2}$$

Sulphur dioxide formed from 32 grams sulphur = 64 grams;
" 1 " = 2 "
" 8 " = [illegible] "
Volume occupied by 64 grams sulphur dioxide = 22·4 litres;
∴ 16 = 5·6 "

EXERCISE V.

1. 10 grams of carbon are burnt; what volume of carbon dioxide at 39°C. and 380 mm. is formed?

2. What volume of oxygen can be obtained from 100 grams of mercuric oxide?

3. How much potassium chlorate is required to make 70 litres of oxygen?

4. 174 grams of manganese dioxide are heated; what volume of oxygen is given off?

5. 100 grams of zinc are dissolved in dilute sulphuric acid; what volume of hydrogen is given off?

6. 100 grams of steam are passed over red-hot iron; what volume of hydrogen at 10°C. and 742 mm. is formed?

118. Given the percentage Composition of a Compound to find its formula.

Ex. 12.—The percentage composition of a compound is—

Hydrogen	2·04
Oxygen	65·31
Sulphur	32·65

What is its formula?

The composition may be represented thus:—

Hyd.	Sulp.	Ox.
2.04	82·65	65·31

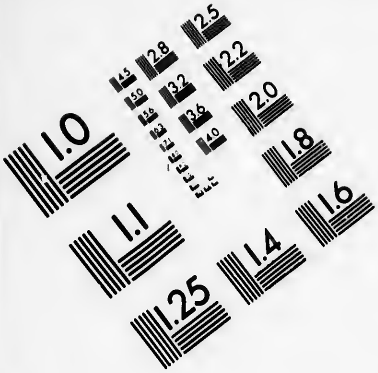
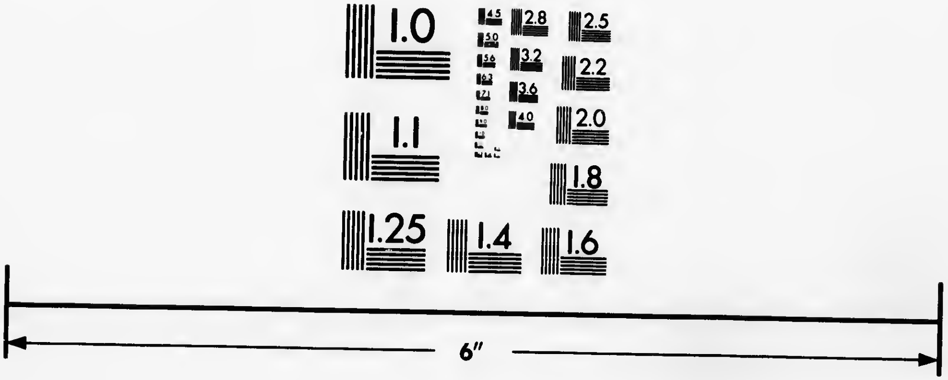
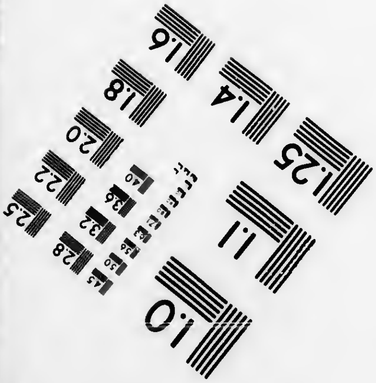
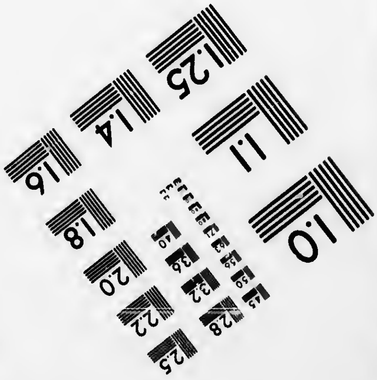


IMAGE EVALUATION
TEST TARGET (MT-3)







Now, since the symbols stand for definite quantities, we have

$$\begin{array}{lccc} & \underset{\frac{2\cdot04}{1}}{H} & \underset{\frac{32\cdot65}{32}}{S} & \underset{\frac{65\cdot31}{16}}{O} \\ = & \underset{2\cdot04}{H} & \underset{1\cdot02}{S} & \underset{4\cdot08}{O} \\ = & H_2SO_4 & & \end{array}$$

EXERCISE VI.

1. The percentage composition of a compound is—

Potassium	31·89
Chlorine	26·95
Oxygen	39·15

What is its formula?

2. A compound is found to have the following percentage composition:—

Oxygen	72·73
Carbon	27·27

What is its formula?

3. A compound of iron and oxygen possesses the following percentage composition. Calculate its formula.

Iron	70·01
Oxygen	29·99

119. Given the formula of a Compound to find its percentage Composition.

Ex. 13.—The formula of water is H_2O; what is the percentage composition?

The composition is—

H_2	Hydrogen	2
O	Oxygen	16

Hydrogen contained in 18 parts of water $= 2$ pts.
" 1 " $= \frac{2}{18}$ "
" 100 " $= \frac{100 \times 2}{18}$ "
$= 11\cdot11$ of hyd.

Oxygen contained in 18 parts of water $= 16$
" 1 " $= \frac{16}{18}$
" 100 " $= \frac{100 \times 16}{18}$
$= 88\cdot88$ of oxy.

Therefore, the percentage composition required is—

Hydrogen	11·11
Oxygen	88·88

EXERCISE VII.

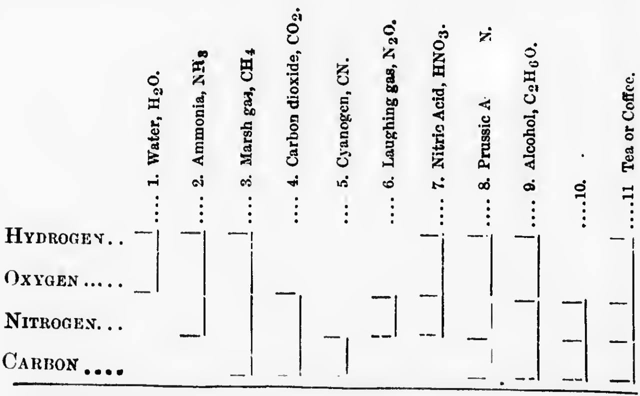
1. The formula of sulphuric acid is H_2SO_4; what is its percentage composition?

2. The formula of potassium chlorate is $KClO_3$; what is its percentage composition?

CHAPTER III.

120. Combination of Elements Already Studied.

Having considered the preparation and properties of the four elementary bodies, oxygen, hydrogen, nitrogen, and carbon, we will proceed to study some of their more important compounds with each other. From the following diagram it will be seen that, leaving out of consideration the *proportions* of each element in the different compounds, ten sets are formed in which occur every possible combination of the four elements, taking two, three, and four together, with one exception. The brackets show which elements are united together in each set, and at the top of each bracket is placed the name of some body which serves as an example of the class to which it belongs in the arrangement:—



Observe that the tenth bracket has no name above it; no compound has yet been discovered which consists of oxygen, nitrogen, and carbon. We shall find that more than one compound belongs to each class. For instance in the fourth class, besides carbon dioxide, there is another oxide of carbon, called carbon monoxide, CO, which contains just half as much oxygen. So in the sixth class, besides N_2O, there are four other oxides of nitrogen.

For convenience we will discuss these subjects in the following order :—

1. Compounds of carbon and oxygen (4).
2. " " " and hydrogen (7).
3. " nitrogen and " (6).
4. " nitrogen and hydrogen (2).
5. " hydrogen and oxygen (1).
6. A mixture of several of these substances, the atmosphere.
7. Compounds of carbon and hydrogen (3).
8. A mixture of these, or coal gas. Combustion.

SECTION I.

CARBON DIOXIDE.

Symbol, CO_2. *Molecular Weight*, 44.

121. Carbonates.—100 parts by weight of calcium carbonate, chalk, or marble consists of—

Calcium	40 Ca.
Carbon	12 C.
Oxygen	48 O_3.

From these and other considerations it is inferred that a molecule of calcium carbonate is denoted by the formula $CaCO_3$. If we compare this formula with the formula of carbonic acid, H_2CO_3 (Art. 44.), we see that the two formulæ are identical, except in the one case we have Ca, the symbol for calcium, and in the other we have H_2. The former may be considered as derived from the latter by the replacement of H_2 by Ca. *Compounds formed by replacing the hydrogen of carbonic acid by* *a metal are called* CARBONATES. Most of the common metals, such as zinc, copper, mercury, act like calcium, one atom of the metal replacing two atoms of hydrogen. *These metals are called dyads.* Others, such as potassium, sodium, and silver, replace the hydrogen, atom for atom. *These are called monads.*

Bicarbonates.—When only half the hydrogen of the acid is replaced by a metal the resulting compound is called a *bicarbonate.* Thus, Na_2CO_3 is sodium carbonate, and $NaHCO_3$ is sodium bicarbonate.

SYSTEMATIC NAME.	COMMON NAME.	FORMULA.
Calcium Carbonate............	Chalk, marble, limestone..	$CaCO_3$.
Sodium "	Barilla, washing soda....	Na_2CO_3.
Hydrogen Sodium Carbonate....	Bicarbonate of soda, baking soda....	$NaHCO_3$.
Potassium Carbonate..........	Potash, pearlash	K_2CO_3.
Hydrogen Potassium Carbonate.	Saleratus...............	$KHCO_3$.
Magnesium Carbonate.........	Magnesia...............	$MgCO_3$.
Lead Carbonate	White lead............	$PbCO_3$.
Copper Carbonate.	Malachite	$CuCO_3$.

122. Properties of Carbonates.

(1) All carbonates are insoluble in water except potassium carbonate, K_2CO; sodium carbonate, Na_2CO_3, and ammonium carbonate $(NH_4)_2CO_3$.

(2) All carbonates evolve carbon dioxide, CO_2, when heated to redness, except the alkaline carbonates, such as potassium carbonate and sodium carbonate.

(3) All carbonates effervesce on the addition of any strong acid; the escaping carbon dioxide is without smell.

(4) All bicarbonates are decomposed by heat.

123. Calcium hydrate.—If calcium carbonate is heated to redness it loses a molecule of carbon dioxide, and there is left a molecule of calcium oxide or quicklime; thus:—

$$\underset{\text{Calcium carbonate.}}{CaCO_3} = \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Calcium oxide, or lime.}}{CaO}$$

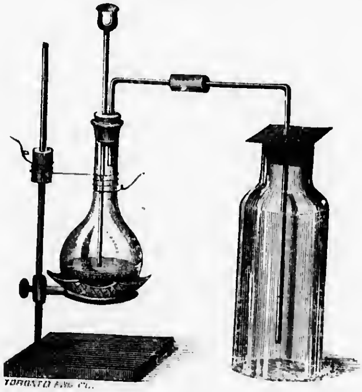
This change is identical with that which takes place in a lime-kiln when lime is made by burning limestone. When water is added to lime it combines with a definite amount of it, 56 parts by weight of lime with 18 parts by weight of water forming *calcium hydrate*, a white powder familiarly called slaked lime; thus:—

$$\underset{\text{Lime.}}{CaO} + \underset{\text{Water.}}{H_2O} = \underset{\text{Calcium hydrate.}}{Ca(OH)_2}$$

124. Lime-water.—If this calcium hydrate is mixed with about 700 times its volume of water it dissolves, forming a clear solution which is familiarly called lime-water. This lime-water is alkaline, turning red litmus blue. It quickly absorbs carbon dioxide from the air, and is used in medicine, and in the laboratory to detect the presence of carbon dioxide and carbonic acid.

125. Preparation of Carbon Dioxide.

Exp. 1.—Take the apparatus used for the preparation of hydrogen, place in it some marble broken into lumps, and pour



in enough of water to cover them. Fit to the flask a delivery-tube, as in the figure, and secure the flask on the retort-stand at such a height that the longer limb of the tube may reach nearly to the table. Place the delivery-tube in a bottle, covering the mouth with a disk of cardboard with a slit in it for the tube. Pour in hydrochloric acid, HCl, in small quantities at a time, until a brisk effervescence is set up. After the lapse of half a minute, pass a lighted taper into the bottle; as it approaches the bottom it goes out. There is evidently some gas collecting there. The reaction is expressed by the following equation:—

$$\underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Calcium chloride.}}{CaCl_2} + \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}$$

This reaction takes place in two stages. The acid exchanges its hydrogen for the calcium, producing calcium chloride on the one hand, and carbonic acid on the other; thus:—

$$CaCO_3 + 2HCl = CaCl_2 + H_2CO_3 \ldots\ldots (1)$$

The carbonic acid, at the moment of its formation, breaks up into water and gaseous carbon dioxide, the latter of which escapes with brisk effervescence; thus:—

$$H_2CO_3 = H_2O + CO_2 \ldots\ldots (2)$$

A cubic inch of marble will yield about four gallons of the gas.

126. Metathesis or Double Decomposition.—The reaction expressed by equation (1) consists merely in an interchange between the hydrogen atoms of the acid and the metallic atom of the carbonate; and of the different modes of chemical action, this is by far the most frequent. It has received the name of *Metathesis or double decomposition, and consists simply in the exchange of the elements or group of elements in one body for the elements or group of elements in another body.*

PROPERTIES.

127. Neither Burns nor Supports Combustion.

Exp. 2.—Plunge a lighted taper into a bottle of carbon dioxide; it is quickly and decidedly extinguished, and the gas does not burn. Its power to extinguish flame is not owing to any chemical action of the gas, but is merely owing to the exclusion of atmospheric oxygen.

This property of carbon dioxide has led to its being used to extinguish fires in mines, and in the "Chemical Fire Extinguisher" it affords a ready means of extinguishing fires in their early stages.

128. Heavier than Air.—The density of carbon dioxide has already been shown by the method of collecting it. It may be further shown as follows:—

Exp. 3.—Take a wide-mouthed bottle and place in it a lighted taper. Bring the mouth of a bottle of the gas close to the edge of the bottle, and pour the gas over the taper, not directly over the centre of the bottle, but at its edge, since the gas receives a forward as well as a downward impulse whilst the bottle is being inverted; the taper will be immediately extinguished, showing that the gas has been poured from one bottle into the other. The molecular weight of carbon dioxide is 44; it is therefore $\frac{44}{2} = 22$ times heavier than hydrogen (Art. 26). But hydrogen is 14·47 times lighter than air; carbon dioxide is, therefore, about $1\frac{1}{2}$ times heavier than air.

129. Solubility in Water.

Exp. 4.—Half fill a bottle with cold water, and fill the other half with carbon dioxide by displacement. Now tightly close the mouth of the bottle with the wetted palm of the hand, and shake it vigorously for a short time; the bottle will adhere to the hand, owing to a partial vacuum being produced by the combination of the gas with the water. The reaction is as follows:—

$$\underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Carbonic acid.}}{H_2CO_3}$$

Invert the bottle in water, and remove the hand; the water will rush in and nearly fill the bottle, showing that the whole of the carbon dioxide has been absorbed. Water, at common temperatures, absorbs its own volume of the gas, acquiring an agreeable acidulous taste, and sparkles when agitated. Its solubility increases if the temperature is diminished or the pressure increased.

130. Acid character.

Exp. 5.—Fill a test-tube to the depth of about two inches with a solution of blue litmus, place the delivery-tube in it, and pass a stream of carbon dioxide through it for a short time; the color is changed to a wine-red, differing entirely from the pure red produced by the action of sulphuric or hydrochloric acid upon the litmus. Boil the reddened solution; it becomes blue again, the carbon dioxide passing off with the steam. Hence, *Carbonic acid is decomposed into carbon dioxide and water, by boiling.*

131. Action on Lime-water.

Exp. 6.—Half-fill a test-tube with clear lime-water, place the delivery-tube in it, and allow the carbon dioxide to bubble through it; the solution becomes milky. The carbon dioxide first combines with the water which holds the calcium hydrate in solution, forming carbonic acid, which then combines with the calcium hydrate, the calcium of the hydrate and the hydrogen of the carbonic acid exchanging places; thus:—

$$(1)\quad \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Carbonic acid.}}{H_2CO_3}$$

$$(2)\quad \underset{\text{Calcium hydrate.}}{Ca(OH)_2} + \underset{\text{Carbonic acid.}}{H_2CO_3} = \underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Water.}}{2H_2O}$$

The calcium carbonate, being insoluble in water, gives the milky appearance.

132. Calcium Bicarbonate.

Exp. 7.—Continue to pass carbon dioxide through the liquid in the test-tube; the turbidity disappears. The calcium carbonate combines with a molecule of carbonic acid, forming *calcium bicarbonate*, which is soluble in water; thus:—

$$\underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Carbonic acid.}}{H_2CO_3} = \underset{\text{Calcium bicarbonate.}}{H_2Ca(CO_3)_2}$$

Calcium Bicarbonate, like Carbonic Acid, has never yet been isolated.

Exp. 8.—Boil the clear liquid in the test-tube, and turbidity again makes its appearance. The calcium bicarbonate is decomposed into the insoluble carbonate, carbon dioxide and water being formed; thus:—

$$\underset{\text{Calcium bicarbonate.}}{H_2Ca(CO_3)_2} = \underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}$$

This experiment explains the origin of the incrustation which is deposited inside kettles and steam-boilers. The calcium bicarbonate is decomposed as above, leaving the insoluble carbonate on the bottom and sides of the vessel. It may be removed from kettles by pouring in a little dilute hydrochloric acid, and may be prevented from forming in boilers by adding ammonium chloride:—

$$\underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Ammonium chloride.}}{2NH_4Cl} = \underset{\text{Calcium chloride.}}{CaCl_2} + \underset{\text{Ammonium carbonate.}}{(HN_4)_2CO_3}$$

The ammonium carbonate volatilizes with the steam, and the very soluble calcium chloride remains in the boiler.

133. Supports Combustion of Substances that have a strong affinity for Oxygen.

Exp. 9.—Attach a piece of magnesium ribbon to the cap of the deflagrating spoon, so that its extremity may reach nearly to the bottom of a large bottle of carbon dioxide. Hold the ribbon in the flame of the spirit-lamp till it begins to burn, and then plunge it slowly into the gas; it will continue to burn brilliantly, forming white flakes of magnesium oxide, interspersed with black particles which consist of carbon. The magnesium combines with the oxygen to form magnesium oxide, setting the carbon free:—

$$\underset{\text{Carbon dioxide.}}{CO} + \underset{\text{Magnesium.}}{2Mg} = \underset{\text{Magnesium oxide.}}{2MgO} + \underset{\text{Carbon.}}{C}$$

Pour a little water into the bottle, and add a small quantity of hydrochloric acid, pour into a test-tube and heat. The magnesium oxide will disappear, while black flakes of carbon will remain floating undissolved in the clear liquid. These may be collected on a filter and shown to be carbon. *This experiment proves that carbon dioxide contains carbon.*

134. Decomposition of Carbon Dioxide by Plants in Sun-light.

Exp. 10.—Fill a large tumbler with water, saturated with carbon dioxide. Fill a glass funnel with fresh green leaves (mint is best). Place the funnel inverted in the tumbler, carefully displacing all the air adhering to the leaves by agitation, and close the neck of the funnel by a cork well saturated with paraffine. Pour off a portion of the water from the tumbler, and place it in direct sunlight. Soon minute bubbles will gather in the leaves and rise into the neck of the funnel. Two or three days in spring or summer, and four or five days in winter will be required. When a sufficient quantity of gas has accumulated, bring the water outside the neck to a level with that inside, remove the cork and insert a glowing splint; the splint will be rekindled, showing the gas to be oxygen. The probable reaction is—

$$\underset{\text{Carbon dioxide.}}{12CO_2} + \underset{\text{Water.}}{11H_2O} = \underset{\text{Sugar.}}{C_{12}H_{22}O_{11}} + \underset{\text{Oxygen.}}{12O_2}$$

This experiment is of great importance as it explains the natural production of oxygen from carbon dioxide and water. The plant leaf is the laboratory in which is constructed the material of which the plant consists, such as woody fibre, sugar, starch, gums, etc. All these consist essentially of carbon and water, and they differ from each other only by a certain number of molecules of water; thus:—

$$\underset{\text{Sugar.}}{C_{12}H_{22}O_{11}} + \underset{\text{Water.}}{H_2O} = \underset{\text{Glucose or fruit sugar.}}{2C_6H_{12}O_6};$$

$$\underset{\text{Fruit sugar.}}{C_6H_{12}O_6} - \underset{\text{Water.}}{H_2O} = \underset{\text{Starch.}}{C_6H_{10}O_5}.$$

Observe also that the volume of oxygen liberated is equal to the volume of carbon dioxide decomposed, so that the volume of the atmosphere remains constant.

OTHER METHODS OF OBTAINING CARBON DIOXIDE.

135. By boiling or heating a solution of Sodium Bicarbonate.

Exp. 11.—Dissolve a tea-spoonful of sodium bicarbonate in water in a test-tube, fitted with a cork and delivery-tube, and boil; carbon dioxide will be given off. The reaction is—

$$\underset{\text{Sodium bicarbonate.}}{2NaHCO_3} = \underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}$$

136. By the action of Hydrochloric Acid on Sodium Bicarbonate.

Exp. 12.—Pour dilute hydrochloric acid on a tea-spoonful of sodium bicarbonate in a test-tube; carbon dioxide will be rapidly given off; thus:—

$$\underset{\text{Sodium bicarbonate.}}{NaHCO_3} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\substack{\text{Sodium chloride}\\\text{or common salt.}}}{NaCl} + \underset{\text{Water.}}{H_2O} + \underset{\text{Carbon dioxide.}}{CO_2}$$

This reaction has been employed as a means of raising dough in the process of bread making. The escaping carbon dioxide puffs up the dough, common salt remaining in the bread. Hydrochloric acid is seldom found sufficiently pure for culinary purposes. Tartaric acid and cream of tartar, however, will answer the same purpose. Indeed, all the baking powders, and yeast powders, and the so-called self-raising flour, depend for their action on the mixture of sodium bicarbonate with some organic acid or other substance that will liberate carbon dioxide from the sodium bicarbonate.

137. By Combustion.

Exp. 13.—Hold a wide-mouthed bottle over the flame of a spirit-lamp for a few moments. Invert the bottle, add a little lime-water, and shake it up; the milky deposit of calcium carbonate will indicate the presence of carbon dioxide. All our ordinary combustibles produce carbon dioxide in this way. The combustion of a bushel of charcoal produces 2500 gallons of the gas.

138. By Respiration.

Exp. 14.—Put a small quantity of lime-water in a test-tube and breathe through it by means of a glass tube. The lime-water will become milky, showing the presence of carbon dioxide. In the same way blue litmus solution may be turned a wine-red color, which becomes blue again on boiling. A man emits by respiration about 1260 cubic inches, or 20·6 litres of carbon dioxide per hour. Two candles in burning will produce the same quantity.

139. By Fermentation.

Exp. 15.—Dissolve a little sugar in eight or ten times its weight of warm (not hot) water in a flask, the delivery-tube of which passes into lime-water. Add to the flask a little dried yeast, previously rubbed down with water; fermentation will begin in the course of an hour or so, and carbon dioxide will pass over into the bottle, and turn the lime-water milky.

Under the action of the yeast, cane sugar, $C_{12}H_{22}O_{11}$, becomes grape sugar, $C_6H_{12}O_6$, and the grape sugar is then changed into alcohol and carbon dioxide; thus:—

$$\underset{\textbf{Grape sugar.}}{C_6H_{12}O_6} \quad = \quad \underset{\textbf{Alcohol.}}{2C_2H_6O} \quad + \quad \underset{\textbf{Carbon dioxide.}}{2CO_2}$$

140. By Germination.

Exp. 16.—Moisten some seeds, put them under a tumbler containing common air, and set them in a moderately warm place; signs of vegetation will soon begin. After the seeds have sprouted, examine the air for carbon dioxide in the usual way. It will be found that a portion of the oxygen has disappeared, and a corresponding volume of carbon dioxide has been produced. The presence of oxygen is as favorable to germination as that of large quantities of carbon dioxide is unfavorable; hence the process is hastened by the introduction into the soil of slaked lime, in order to absorb the carbon dioxide as fast as it is produced by the sprouting seeds.

141. By Decay of Animal or Vegetable Substances.

Exp. 17.—Place some dead leaves in an air-tight bottle; the air will soon cease to have the power of supporting combustion, its oxygen having combined with the carbon of the leaves to form carbon dioxide.

SUMMARY AND ADDITIONAL FACTS.

History.—Carbon dioxide was known as early as the sixteenth century. It was examined by Black in 1757, and called by him fixed air, because it was fixed in the carbonates. In 1775, Lavoisier determined its exact nature, and named it carbonic acid.

Sources.—We have seen that carbon dioxide is a product of respiration in man and animals, that it is a product of combustion, and a product of fermentation. It is a small but constant constituent of the atmosphere; it is likewise invariably contained in the soil, being one of the chief products of decay of all organic substances. From the soil it is taken up by rain and spring water, and it is to this substance that spring water, to a great extent, owes its fresh and pleasant taste. It is evolved from the craters of active volcanoes, from fissures in the earth, and is contained in immense quantities in the carbonates.

Properties.—Carbon dioxide is a colorless gas, possessing a slightly pungent smell and acid taste. It supports neither combustion nor respiration. When pure, carbon dioxide cannot be breathed. When so far diluted as to admit of being received into the lungs, it acts as a narcotic poison, causing drowsiness and insensibility. It is not, however, poisonous in the strict sense of the term. On the contrary, it is always present in the blood in large quantities, and is constantly secreted from the lungs and from other parts of the body. If the atmosphere contains more than a small percentage of this gas, it arrests this secretion, and fatal results necessarily follow. No rule can be laid down as to the precise quantity of carbon dioxide that may be present in the air without injury to respiration. According to Dr. Parks, an eminent authority on this subject, air is unhealthy when the carbon dioxide in it exceeds ·06 per cent., or 6 volumes in 10,000.

Carbon dioxide accumulates in old wells, cellars, etc., being either exhaled from the earth or produced by the decay of organic matter. The ordinary test is to lower a lighted candle before the workman descends. Air containing 4 per cent. of carbon dioxide will extinguish a candle, but will not support respiration for any length of time. The carbon dioxide may be expelled by frequently letting down a bucket into it, and turning it upside down away from the well; or it may be converted into calcium carbonate by pouring in lime-water.

At ordinary pressure, one volume of water absorbs one volume of carbon dioxide; at two pressures, two volumes; at three pressures, three volumes, etc.; but on the removal of the extra pressure all the dissolved gas escapes except the original volume. "Soda water" is simply ordinary water with carbon dioxide in solution. The foaming of soda water, drawn from the fountain, is due to the escape of the carbon dioxide. The sparkling character of champagne, bottled ale, etc., is due to the liberation of carbon dioxide which has been produced by fermentation, and retained in the liquid under pressure.

Carbon dioxide can be liquefied by both cold and pressure. Under a pressure of 36 atmospheres at 0°C. it is converted into a colorless, mobile liquid. When this liquid is suddenly relieved from the pressure under which it alone can exist, part of it at once passes back into the state of gas, and heat is absorbed so rapidly that the remaining portion of the liquid solidifies. By mixing the solid with ether, and evaporating under the exhausted receiver of an air-pump, Faraday obtained a cold of −110°C. Above 32°.5C. carbon dioxide cannot be condensed to a liquid by any pressure. In the same way all other gases show a *critical point in temperature* at which they are no longer able to be condensed to liquids. That the so-called permanent gases, oxygen, hydrogen, etc., could not formerly be condensed was due to the fact that they were compressed at temperatures lying above their critical points.

Tests:

(1) Extinguishes flame.

(2) Lime-water throws down a white precipitate of calcium carbonate.

(3) The gas is soluble in a solution of caustic potash.

(4) With water it forms carbonic acid.

SECTION II.

CARBON MONOXIDE.

Symbol, CO. Molecular Weight, 28.

The element carbon forms with oxygen, besides the compound carbon dioxide, a second compound called carbon monoxide o carbonic oxide, which has the symbol, CO.

PREPARATION.

142. By the decomposition of Oxalic Acid by Sulphuric Acid.

Exp. 1.—Put a few crystals of oxalic acid, $H_2C_2O_4$, into a test-tube, add sufficient sulphuric acid to cover them, and gently heat; effervescence soon takes place. After a few moments bring a lighted taper to the mouth of the tube; a gas takes fire and burns with a pale blue flame. Extinguish the flame, incline the tube and hold a bottle over it for a few minutes, pour some lime-water into the bottle and shake it briskly; the lime-water becomes turbid, showing the presence of carbon dioxide. Since carbon dioxide is not an inflammable gas, two gases must have been produced. The one is carbon dioxide, the other carbon monoxide. The reaction is expressed by the following equation:—

$$\underset{\text{Oxalic acid.}}{H_2C_2O_4} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Dilute sulphuric acid.}}{H_2O, H_2SO_4} + \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Carbon monoxide.}}{CO}$$

The sulphuric acid takes no part in the reaction, except removing the water, and setting free the two gases.

Exp. 2.—Place 10 grams of crystallized oxalic acid in the hydrogen flask, and through the funnel-tube pour about 30 c.c. of strong sulphuric acid. Heat the flask gently, and after allowing the air to escape, collect a large bottle full of the gases over water in the pneumatic trough. Remove the bottle when full, and place it mouth downwards on a piece of glass or in a saucer. The generating flask should be placed in a draught or out of doors, as carbon monoxide is very poisonous.

PROPERTIES.

Exp. 3.—Pour about 20 c.c. of a strong solution of caustic potash or soda into the bottle, close its mouth with the hand and shake briskly; the hand feels pressed into the bottle, showing that some gas has been absorbed. The caustic potash combines with the carbon dioxide, leaving the carbon monoxide untouched. Invert the bottle beneath the water and withdraw the hand; the water rushes in until the bottle is half full. *The gases are, therefore, set free in equal volumes.*

Exp. 4.—Fill a small bottle with water, place it mouth downwards in the trough, and bring the mouth of the bottle containing the carbon monoxide under it, and gently pour the gas from one bottle into the other. Bring a light to the mouth of the bottle; the gas will take fire and burn with the characteristic blue flame noticed in the first experiment. Add a little lime-water to the bottle, and shake it up; the lime-water becomes turbid, showing that carbon dioxide is present. The carbon monoxide has combined with the atmospheric oxygen, forming carbon dioxide:—

$$\underset{\text{Carbon monoxide.}}{2CO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Carbon dioxide.}}{2CO_2}$$

OTHER METHODS OF PREPARING CARBON MONOXIDE.

143. By the decomposition of Potassium Ferrocyanide by strong Sulphuric Acid.

Exp. 5.—Well-dried and finely-powdered potassium ferrocyanide (yellow prussiate of potash), $K_4Fe(CN)_6$, is heated with about nine times its weight of strong sulphuric acid. The reaction is at first slow, and then violently quick as the temperature rises. The gas evolved is carbon monoxide, only very slightly contaminated with carbon dioxide. This is the best method of preparing the gas.

144. By the incomplete Combustion of Carbon.

Exp. 6.—Fill a porcelain or hard glass tube with small lumps of charcoal, and place it in a small furnace, or in some way heat it through its entire length, and pass a stream of air through it. If the coals are glowing strongly, and the stream of air very slow, the gas issuing from the tube will be carbon monoxide:—

$$\underset{\text{Carbon.}}{2C} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Carbon monoxide.}}{2CO}$$

145. By the reduction of Carbon Dioxide by glowing Charcoal.

Exp. 7.—Use the same apparatus as in the last experiment, and pass carbon dioxide in a slow stream over the heated coals. The red-hot charcoal reduces the carbon dioxide; thus:—

$$\underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Carbon.}}{C} = \underset{\text{Carbon monoxide.}}{2CO}$$

146. By heating a Carbonate with Carbon.

Exp. 8.—Mix together finely powdered chalk and charcoal, place the mixture in an iron tube, and heat in a small furnace; the calcium carbonate is reduced to an oxide, and carbon monoxide set free. This is the change which takes place in lime-kilns:—

$$\underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Carbon.}}{C} = \underset{\text{Lime.}}{CaO} + \underset{\text{Carbon monoxide.}}{2CO}$$

SUMMARY AND ADDITIONAL PROPERTIES.

147. History.—Carbon monoxide was discovered by Priestley when igniting chalk in a gun barrel.

148. Occurrence.—It is never found except as an artificial product, as in the neighborhood of brick or lime-kilns.

149. Properties.—Carbon monoxide is a colorless, tasteless gas, possessing a peculiar though slight smell. It is very slightly soluble in water. It is a very poisonous gas, and much of the ill repute which attaches to carbon dioxide really belongs to this gas. Small animals when placed in it die almost instantly. It is the presence of this gas which occasions the peculiar sensation of oppression and headache which is experienced in rooms into which the products of combustion have escaped from fires of charcoal and anthracite. One per cent. is a sufficient quantity to prove fatal. The characteristic blue flame of carbon monoxide may often be observed playing over the surface of clear fires. In stoves the air enters at the lower surface where the oxygen is abundant, forming carbon dioxide as the first product of combustion; this carbon dioxide ascends through a mass of ignited carbon, where it is exposed to a great excess of red-hot charcoal which reduces it to carbon monoxide; when this reaches the surface it combines with oxygen, if present, reproducing carbon dioxide. It is also formed when steam is passed over ignited coal, and is, therefore, a chief ingredient in the so-called water-gas. Common coal gas contains from 4 to 7 per cent. of this gas.

150. **Tests.**—Carbon monoxide is recognised by burning with a pale-blue flame, producing carbon dioxide, which renders lime-water milky.

QUESTIONS AND EXERCISES ON CARBON DIOXIDE AND CARBON MONOXIDE.

1. What weight of quicklime may be obtained from a ton of limestone?

2. Quicklime is slaked with water, diffused through more water and filtered. When carbon dioxide is passed through the filtered liquid it becomes turbid; what is the composition of the substance which causes the turbidity? Give an equation.

3. Express in symbols the action which takes place between hydrochloric acid and chalk, and describe in words the changes which occur, and the proportions by weight of each constituent.

4. Describe the physical and chemical properties of carbon dioxide.

5. What is lime-water, and how is it made? What happens when carbon dioxide is led into it (*a*) in small quantity, (*b*) in excess? Express all reactions by equations.

6. How would you prove that carbon dioxide consists of carbon and oxygen?

7. By what natural agency is carbon dioxide constantly entering the atmosphere withdrawn from it?

8. Describe the composition of ordinary baking-powder, and explain its action in raising bread.

9. What volume of carbon dioxide measured at 745 mm. and 15°C. can be obtained from 150 grams of marble containing 3 per cent. of silica?

10. What weight of carbon is needed for the formation of 88 grams of carbon dioxide? What is the volume of this weight of gas at 0°C. and 30 inches barometric pressure?

11. A substance has the following percentage composition:—Carbon 27·27, oxygen 72·73; what is its formula?

12. How is carbon monoxide usually prepared? How would you prepare it from carbon dioxide? Give an account of its properties.

13. A litre of carbon dioxide is completely reduced to carbon monoxide by hot carbon; what volume of carbon monoxide is obtained? What volume of oxygen is needed for the complete combustion of this carbon monoxide?

14. Explain the production of carbon monoxide. When coals burn in an ordinary stove.

CHAPTER IV.

SECTION I.

NITRIC ACID.

Symbol, HNO_3. *Molecular Weight, 63.*

151. Nitrates.—It has been already stated (Art. 90) that nitrogen is so called because it is a characteristic constituent of *nitre* or *saltpetre* This nitre is a white crystalline substance, occurring as an incrustation on the soil in various localities, especially in India. On analysis it is found to contain potassium, nitrogen and oxygen in the proportion indicated by the formula KNO_3. Comparing a molecule of this substance with a molecule of nitric acid, HNO_3, we see that it may be considered as derived from the acid by the replacement of the hydrogen by potassium. *The compounds formed by replacing the hydrogen of nitric acid by a metal are called Nitrates, and are particularized by the metal which replaces the hydrogen.* KNO_3 is, therefore, called potassium nitrate.

EXAMPLES OF COMMON NITRATES.

SYSTEMATIC NAME.	COMMON NAME.	FORMULA.
Potassium Nitrate	Saltpetre, Nitre	KNO_3.
Sodium Nitrate..............	Cubic Nitre, Chili Saltpetre.	$NaNO_3$.
Ammonium Nitrate..........	....	NH_4NO_3.
Silver Nitrate, Argentic Nitrate	Lunar Caustic	$AgNO_3$.
Barium Nitrate.............	....	$Ba(NO_3)_2$.

152. Properties of Nitrates.

(1) All nitrates are soluble in water.

(2) When a nitrate is heated on charcoal it deflagrates.

PREPARATION OF NITRIC ACID.

153. By the action of Sulphuric Acid on a Nitrate.

Exp. 1.—Take 15 grams (about ½ oz.) of potassium nitrate and transfer it to a small retort. To this add an equal weight of sulphuric acid. The nitrate, when coarsely powdered, will occupy a little over two inches in a test-tube ¾ in. wide, and the sulphuric acid nearly the same. Place the retort on a piece of wire-gauze on a ring of the retort-stand. Insert the beak of the retort in a flask, floating on water, in the pneumatic trough. Apply a gentle heat; ruddy vapors come off, which condense in the neck of the retort and trickle down into the receiving flask, which must be kept cool by occasionally turning it in the water or covering it with a piece of wet blotting paper. As the distillation advances the red fumes which first arise disappear, but towards the end of the process they again appear. When this happens the operation may be stopped, as very little liquid is passing over. The nitric acid is formed by the interchange of the potassium of the nitrate with an atom of hydrogen of the sulphuric acid; thus:—

$$\underset{\text{Potassium nitrate.}}{KNO_3} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Hydrogen potassium sulphate.}}{KHSO_4} + \underset{\text{Nitric acid.}}{HNO_3}$$

The red fumes are due to the decomposition of a little of the nitric acid, which at 260°C. is resolved into nitrogen tetroxide, oxygen, and water; thus:—

$$\underset{\text{Nitric acid.}}{4HNO_3} = \underset{\text{Nitrogen tetroxide.}}{4NO_2} + \underset{\text{Oxygen.}}{O_2} + \underset{\text{Water.}}{2H_2O}$$

PROPERTIES.

154. Acid Properties.

Exp. 2.—Half-fill a test-tube with a dilute solution of litmus, and add one drop of nitric acid. The blue color will be at once changed to red, showing the powerful acid properties of the liquid.

155. Oxidizing Action of Nitric Acid.

Exp. 3.—Half-fill a test-tube with water, and add a solution of sulphate of indigo till it is of a blue color. Now add a few drops of nitric acid, and gently heat; the blue color is changed to a dirty brown. The nitric acid gives up part of its oxygen to the indigo, changing it to *isatin;* thus:—

$$\underset{\text{Indigo.}}{C_{16}H_{10}N_2O_2} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Isatin.}}{2C_8H_5NO_2}$$

Exp. 4.—Put a little nitric acid into a test-tube, and loosely close the mouth of the tube with fine dry shavings or dry hay. Boil the nitric acid, and the shavings take fire; the acid is decomposed into nitrogen tetroxide, water, and oxygen, the latter uniting with the shavings. In the same way sulphur and carbon when boiled in nitric acid are oxidized to sulphuric and carbonic acids respectively.

156. Action on Organic Substance.

Exp. 5.—Dip a piece of white wool or a piece of white flannel into dilute nitric acid, and gently dry it; the wool is stained yellow. In the same manner horn, skin, silk, are stained of a yellow color by dilute nitric acid. By concentrated nitric acid many organic substances are changed into compounds of great practical importance, such as sawdust into oxalic acid, cotton into gun-cotton, benzol into nitro-benzol, and glycerine into nitro-glycerine.

Exp. 6.—Put a drop of nitric acid, a drop of hydrochloric acid, and a drop of sulphuric acid on a piece of black cloth; the cloth is colored red by each of the acids. Now put a drop of ammonia on each of the spots; the red color produced by the hydrocloric and sulphuric acids disappears, while that due to the nitric acid remains. Although the red spots produced by nitric acid cannot be obliterated by the application of ammonia, they should always be wetted with it, or a hole will probably result.

157. Action on Metals.—The action of nitric acid on metals varies with the strength and concentration of the acid. With highly concentrated acid, nitrogen tetroxide is disengaged.

Exp. 7.—Pour a little strong nitric acid on a few scraps of copper, or any other metal except gold or platinum; deep yellow or ruddy fumes of nitrogen tetroxide are produced; thus:—

$$\underset{\text{Copper.}}{Cu} + \underset{\text{Nitric Acid.}}{4HNO_3} = \underset{\text{Copper Nitrate}}{Cu(NO_3)_2} + \underset{\text{Nitrogen Tetroxide.}}{2NO_2} + \underset{\text{Water.}}{2HO_2}$$

This reaction takes place in two stages; thus:—

$$(1) \quad Cu + 2HNO_3 = Cu(NO)_3 + H_2$$

The hydrogen immediately attacks another portion of the nitric acid and nitrogen tetroxide is evolved; thus:—

$$(2) \quad H_2 + 2HNO_3 = 2H_2O + 2NO_2$$

For the action of dilute nitric acid on metals see nitrogen dioxide.

158. Action of Metals on Acids.—There is little doubt that in all cases the metal simply displaces the hydrogen of the acid forming a salt. If under the condition under which the experiment is made, the acid has no tendency to enter into reaction with the hydrogen which is displaced, the hydrogen is evolved; but if the acid can enter into reaction with the hydrogen, the products of this secondary reaction are obtained.

159. Commercial preparation of Nitric Acid.

Sodium nitrate is preferred to potassium nitrate on the manufacturing scale, because it is cheaper, and also that a given weight of it can be made to yield a larger quantity of acid than can be obtained from the same weight of nitre. Only half the quantity of sulphuric acid is used, but a much higher temperature is required. The reaction is expressed by the following equation:—

$$\underset{\text{Sodium nitrate.}}{2NaNO_3} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Sodium sulphate.}}{Na_2SO_4} + \underset{\text{Nitric acid.}}{2HNO_3}$$

160. Tests for free Nitric Acid.

(1) The production of red fumes, when brought in contact with a metal is characteristic of nitric acid.

(2) When a drop of nitric acid is added to water, tinged blue with sulphate of indigo and the mixture boiled, it is bleached.

161. Tests for Nitrates.

(1) The preceding tests apply to nitrates if a few drops of sulphuric acid are added to liberate the nitric acid.

(2) Make a solution of potassium nitrate in a little water in a test-tube, add some copper clippings and a little sulphuric acid, and gently heat; brown vapors are given off. The reaction is—

$$\underset{\text{Potassium nitrate.}}{2KNO_3} + \underset{\text{Sulphuric acid.}}{4H_2SO_4} + \underset{\text{Copper.}}{3Cu} = \underset{\text{Potassium sulphate.}}{K_2SO_4} + \underset{\text{Copper sulphate}}{3CuSO_4} + \underset{\text{Water.}}{4H_2O} + \underset{\text{Nitric oxide.}}{2NO^*}$$

$$2NO + O_2 = 2NO_2$$

If six molecules of *Ferrous Sulphate*, $6Fe_2SO_4$, are added instead of the three atoms of copper, the equation will be the same except that instead of $3CuSO_4$ we will have three molecules of *Ferric Sulphate*, $3Fe_2(SO_4)_3$ If the ferrous sulphate is in excess the nitric oxide combines with two molecules of it forming a *black-brown liquid*, $(FeSO_4)_2NO$, readily decomposed by heat and otherwise unstable. This affords an exceedingly delicate test for nitric acid and nitrates in solution. It may be conducted as follows:—

Exp. 9.—Dissolve a small crystal of nitre in a strong solution of ferrous sulphate in a test-tube; incline the tube and pour a small quantity of sulphuric acid down its side, so that the heavy acid falls to the bottom of the tube without sensibly raising the temperature of the liquid. A "*brown ring*" forms between the two layers of liquids—violet, red, brown, or black, according to proportions and conditions.

Free Nitric Acid may be distinguished from nitrates by giving the *brown ring without the addition of sulphuric acid.*

---

* For method of obtaining coefficients see chapter.

SUMMARY AND ADDITIONAL FACTS.

162. History.—Liquid nitric acid was known as far back as the seventh century. In the time of the alchemists it was known as *aqua fortis*—a name still retained. The nature of its constituents was shown by Cavendish in 1785.

163. Sources.—Its principal source is potassium nitrate, found in large quantities in certain districts in India, and sodium nitrate which occurs in immense quantities in Peru. It is found in the atmosphere after thunder storms, and in the dried leaves of certain plants. Nitrates form whenever organic matters are decomposed at a temperature of from 20°C. to 30°C. in presence of water and alkaline carbonates.

164. Properties.—Pure nitric acid is a colorless liquid, but in commerce is usually found of a yellow color, from the presence of the lower oxides of nitrogen. It possesses a peculiar but not a very powerful smell. It absorbs moisture from the air with avidity. It is an extremely corrosive liquid, an intensely irritant poison, and acts very destructively on organic bodies, communicating a yellow color to such as contain nitrogen. Strong nitric acid has a specific gravity of about 1·5. With very few exceptions, it oxidizes all elementary substances, converting them into oxides, acids, or nitrates, as the case may be. It acts on all metals except gold, platinum, and a few of the rarer metals.

Nitrates greatly promote vegetable growth, intensifying the color and increasing the quantity of nitrogen of the plant to which they are supplied. Their effect when in excess is to favor the development of foliage at the expense of fruit. Sodium nitrate is extensively used as a fertilizer.

165. Uses.—Nitric acid is extensively used for etching on copper, as a solvent for tin in the preparation of valuable mordants used by dyers and calico printers. In medicine it is prescribed as a tonic, and is used in surgery as an energetic caustic.

SECTION II.

NITROGEN MONOXIDE, OR NITROUS OXIDE.

Symbol, N_2O. Molecular Weight, 44.

PREPARATION.

166. By the decomposition of Ammonium Nitrate by heat.

Exp. 1.—Put into the retort used for the preparation of nitric acid 15 grams of ammonium nitrate. This quantity will about two-thirds fill a test-tube six inches by three-fourths of an inch. Support the retort on the retort stand so that its beak may dip under the shelf of the pneumatic trough. using *warm* water in the trough instead of cold. Heat the retort gently until the salt fuses into a clear liquid, then gradually increase the heat; the ammonium nitrate is decomposed with effervescence into nitrogen monoxide and water; thus:—

$$\underset{\text{Ammonium Nitrate.}}{NH_4NO_3} = \underset{\text{Nitrogen monoxide.}}{N_2O} + \underset{\text{Water.}}{2H_2O}$$

Fill five bottles, and as each bottle is filled cover its mouth with a glass plate and place it mouth upwards on the table.

One oz. or 30 grams of the salt furnishes 500 cubic inches, or rather more than 8 litres of the gas.

167. Precautions.

(1) Do not use more heat than is necessary to keep the fluid steadily effervescing. If white fumes increase to fill the retort, especially towards the close of the operation, the heat must be moderated.

(2) Do not dip the neck of the retort into the trough till the gas passes out rapidly, which may be ascertained by holding the mouth of the retort in a capsule containing water.

(3) Be careful not to withdraw the lamp while the beak of the retort is under water, or owing to the sudden condensation

of the steam, the water will be drawn into the hot retort, and crack it.

PROPERTIES.

168. Supports Combustion.

Exp. 2.—Plunge a glowing splint into a bottle of the gas; the splint bursts into flame and burns nearly as brilliantly as in oxygen, but the gas itself does not take fire.

Exp. 3.—Pour the gas from one bottle into another, and test both bottles with a glowing splint of wood. The taper is rekindled in the bottle into which the gas was poured, showing the gas is so much heavier than air that it can displace the lighter air just like liquid.

Exp. 4.—Burn a small piece of phosphorus in nitrous oxide, taking the same precautions as when burning phosphorus in oxygen. When the white fumes have somewhat subsided plunge a lighted taper into the bottle; the taper is extinguished. In this respect the gas in the bottle resembles nitrogen. Now add a little blue litmus solution, and shake it up; the litmus is reddened, as when phosphorus is burnt in pure oxygen. The burning phosphorus has decomposed the nitrous oxide, uniting with its oxygen to form phosphorus pentoxide, and setting the nitrogen free; thus:—

$$\underset{\text{Nitrous oxide.}}{10N_2O} + \underset{\text{Phosphorus.}}{P_4} = \underset{\text{Phosphorus pentoxide.}}{2P_2O_5} + \underset{\text{Nitrogen.}}{10N_2}$$

Exp. 5.—Place a piece of sulphur in the deflagrating spoon, just light it, and plunge it into a bottle of the gas; the sulphur is extinguished. Had this gas been pure oxygen, the sulphur would have burned with increased brilliancy. Hold the spoon in the flame of the spirit-lamp till the sulphur is well kindled, and then plunge it into the gas; it burns with considerably increased degree of brilliancy, and the product of combustion is found to be sulphur dioxide. Probably brown vapors may make their appearance in the bottle; these consist of one of the higher oxides of nitrogen.

The extinction of the feebly burning sulphur, in a gas containing one-half of its volume of oxygen, while it continues to burn in an atmosphere which contains only one-fifth of its volume of oxygen, is due to the fact that the oxygen and nitrogen of the air are in a state of *simple mixture*, while in nitrous oxide they are *chemically combined*. The combustible body must, therefore, break down this chemical union before it can itself combine with the oxygen; to do this a moderately high temperature is required.

169. Solubility in Water.

Exp. 6.—Pour some cold water into a bottle of nitrous oxide, close the mouth with the hand and agitate briskly. Now bring the mouth of the bottle under the water and remove the hand; the water rises in the bottle. This serves to distinguish nitrous oxide from oxygen. Pour the water into a glass and taste it; it has acquired a sweetish taste and a faint agreeable smell.

SUMMARY AND ADDITIONAL FACTS.

170. History.—Nitrous oxide was discovered by Priestley in 1776. Its physiological action was investigated by Davy in 1809.

171. Properties.—Nitrous oxide is a colorless gas possessing a slightly sweetish taste and faint agreeable smell. Its density is the same as carbon dioxide. 100 volumes of water at 0°C. dissolve 130 volumes of the gas, but its solubility diminishes rapidly as the temperature rises. It may be condensed by a pressure of about 50 atmospheres at 0°C. to a colorless liquid, and even frozen by the cold produced by its own evaporation. When liquid nitrous oxide is mixed with carbon disulphide and evaporated in vacuo a reduction of temperature, estimated at −140°C., is produced which is lower than that which has been obtained by any other means. It readily supports combustion, but all combustion in this gas is simply combustion in oxygen, the burning body not uniting with the nitrous oxide but with its oxygen.

The most remarkable property which nitrous oxide possesses is that of causing loss of sensibility when inhaled. When respired for a short time it produces a singular species of transient intoxication attended in some instances with an irresistible tendency to muscular exertion, and often uncontrollable laughter; hence it has received the name of "*laughing gas.*" It is now frequently used as an anæsthetic for producing insensibility to pain, more especially in the extraction of teeth. Great care must be taken that the gas used for inhalation is free from chlorine and nitric oxide.

Nitrous oxide may be distinguished from oxygen as follows:—

(1) It is much more soluble in water.
(2) It is much heavier.
(3) It does not form red fumes with nitric oxide.
(4) When phosphorus is burnt in it the residual nitrogen is identical in volume with that of the original gas.

SECTION III.

NITRIC OXIDE, NITRIC TRIOXIDE, NITROGEN TETROXIDE, AND NITROGEN PENTOXIDE, SECOND LAW OF CHEMICAL COMBINATION.

Nitric Oxide. Symbol, NO. Molecular Weight, 30.

PREPARATION.

172. By the reduction of Nitric Acid.

Exp. 1.—Take the hydrogen bottle and place it in about 15 grams of copper clippings, cover them with warm water, and pour down the funnel-tube a little strong nitric acid. Effervescence almost immediately begins, the flask is filled with red fumes which soon disappear, and then a colorless gas is produced which may be collected over water in the usual way. The reaction is expressed by the following equation :—

$$\underset{\text{Copper.}}{3Cu} + \underset{\text{Nitric Acid.}}{8HNO_3} = \underset{\text{Copper Nitrate,}}{3Cu(NO_3)_2} + \underset{\text{Water.}}{4H_2O} + \underset{\text{Nitric Oxide.}}{2NO}$$

This reaction takes place in two stages, thus :—

$$(1)_3Cu + 6HNO_3 = 3Cu(NO_3)_2 + 3H_2\,;$$
$$(2)\ 3H_2 + 2HNO_3 = 4H_2O + 2NO.$$

PROPERTIES.

173. Supports Combustion.

Exp. 2.—Introduce a lighted taper into a bottle of the gas; it is extinguished. Place a small piece of phosphorus in the deflagrating spoon, touch it with the end of a warm wire, and just as it begins to burn plunge it into the gas; it is immediately extinguished. Now let the phosphorus burn briskly for a short time, and again plunge it into the gas; the phosphorus burns brilliantly. In the first case the temperature was not sufficiently high to decompose the gas and render its oxygen available. In the second case the gas was decomposed, phosphorus pentoxide was produced, and free nitrogen left. Similar experiments may be tried with magnesium wire, charcoal, and sulphur.

174. Affinity for Oxygen.

Exp. 3.—Withdraw the glass plate from a bottle of the gas and observe that red fumes appear when the gas meets the air. The oxygen of the air combines with the nitric oxide to form nitrogen tetroxide; thus :—

$$\underset{\text{Nitric Oxide.}}{2NO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Nitrogen Tetroxide.}}{2NO_2}$$

Pour a little water into the bottle and shake it up with the gas; the water combines with the nitrogen tetroxide to form nitric and nitrous acids; thus:—

$2NO_2$	+	H_2O	=	HNO_3	+	HNO_2
Nitric oxide.		Water.		Nitric acid.		Nitrous acid.

If the quantity of oxygen present is only equal to one volume for four volumes of nitric oxide, nitrogen tetroxide is no longer produced, but nitrogen trioxide is formed; thus:—

$4NO$	+	O_2	=	$2N_2O_3$
Nitric oxide.		Oxygen.		Nitrogen trioxide.

175. Oxidation of Nitric Oxide by pure Oxygen.

Exp. 4.—Half-fill a bottle with nitric oxide, and leave it mouth downwards in the pneumatic trough. Half-fill a similar bottle with pure oxygen. Bring the mouth of the bottle containing the oxygen under the mouth of the bottle containing the nitric oxide, and pass the oxygen very slowly into the nitric oxide. As each bubble of oxygen comes in contact with the gas, brown ruddy fumes are produced. Shake the bottle from time to time, taking care to keep its mouth under the water; the water rises in the bottle, showing that the gas is being absorbed. Continue the addition of the oxygen, and ultimately the bottle is entirely filled with water. The two gases here combine to form nitrogen tetroxide, which combines with the water to form nitric and nitrous acids. As nitric oxide entirely removes the oxygen from a gaseous mixture, it was much used by Priestly, Dalton, and other chemist in the analysis of atmospheric air; but the results are not found to be trustworthy, in consequence of the formation in uncertain proportions of other oxides of iron.

176. Absorption by Ferrous Substance.

Exp. 5.—Pour a solution of ferrous sulphate into a bottle of the gas and shake it up; the hand is strongly drawn into the bottle showing that the gas is absorbed, and the ferrous sulphate becomes black. A definite compound is formed, the formula of which is $(FeSO_4)_2NO$. This is the substance of the dark ring which occurs in testing for nitrates. Boil the dark liquid and nitric oxide is driven off, and a somewhat brownish solution is left.

NITROGEN TRIOXIDE OR NITROUS ANHYDRIDE.

Symbol, N_2O_3. *Molecular Weight, 76.*

PREPARATION.

Exp. 1.—Put a gram or two of arsenious oxide, AS_2O_3, into a hard glass tube and add some nitric acid diluted with half its bulk of water. Apply a gentle heat and brown fumes are given off which, when cooled below 0°C., afford a blue liquid. The reaction is—

$$\underset{\text{Nitric Acid}}{2HNO_3} + \underset{\text{Arsenious oxide.}}{As_2O_3} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Nitrogen trioxide.}}{N_2O_3} + \underset{\text{Arsenic acid.}}{2H_3AsO_4}$$

Exp. 2.—Collect some of the brown fumes in a bottle, add litmus solution and shake it up; it is colored red. The water combines with the oxide to form *nitrous acid:*

$$\underset{\text{Nitrous anhydride.}}{N_2O_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitrous acid.}}{2HNO_2}$$

177. Nitrous Acid, HNO_2.

Nitrous acid is an ill-defined, unstable compound. When its hydrogen is replaced by a metal a class of salts called *nitrites* is formed. Observe that the acid ending in *—ous* forms a salt ending in *ite;* and that an acid with the termination *—ic* gives a salt terminating in *—ate.*

NITROGEN TETROXIDE OR NITROGEN PEROXIDE.

Symbol, NO_2. *Molecular Weight, 46.*

This substance is obtained when perfectly dry lead nitrate is decomposed by heat:—

$$\underset{\text{Lead nitrate.}}{2Pb(NO_3)_2} = \underset{\text{Nitrogen tetroxide.}}{4NO_2} + \underset{\text{Lead oxide.}}{2PbO} + \underset{\text{Oxygen.}}{O_2}$$

Nitrogen peroxide is a liquid at the ordinary atmospheric temperature, but at −9°C. it solidifies.

NITROGEN PENTOXIDE OR NITRIC ANHYDRIDE.

Symbol, N_2O_5. Molecular Weight, 108.

This substance is obtained when dry chlorine gas is conducted over dry silver nitrate :—

$$\underset{\text{Silver nitrate.}}{4AgNO_3} + \underset{\text{Chlorine.}}{2Cl_2} = \underset{\text{Silver chloride.}}{4AgCl} + \underset{\text{Nitrogen pentoxide.}}{2N_2O_5} + \underset{\text{Oxygen.}}{O_2}$$

It is a solid body, very unstable, combining with water to form nitric acid, and is, therefore, an anhydride :—

$$\underset{\text{Nitric anhydride.}}{N_2O_5} + \underset{\text{Water.}}{H_2O} = \underset{\text{Nitric acid.}}{2HNO_3}$$

We have seen that there are five oxides of nitrogen, which are represented by the following formulæ :—

$$N_2O,\quad NO,\quad N_2O_3,\quad NO_2,\quad N_2O_5.$$

Two of these, N_2O_3 and N_2O_5, are anhydrides of distinct acids. Analogy and their place in the series would seem to indicate that the formula of nitric oxide should be N_2O_2, and that of nitrogen peroxide N_2O_4; but this is contrary to the molecular weight, according to Avogadro's Law (Art. 23). Possibly when the gases are condensed it may be found that their molecules may possess twice their ordinary weight. With respect to NO_2 this seems to be the case, for when the gas is cooled down its density increases from 23 ($H = 1$) to nearly 46 when close to its condensing point; the latter number gives the molecular weight 92, and the formula N_2O_4.

178. Second Law of Chemical Combination.—From the study of the preceding compounds we see that :—

(1.) Nitrogen monoxide contains 28 parts by weight of N to 16 parts of O.

(2.) Nitrogen dioxide contains 28 parts by weight of N to 32 parts of O.

(3.) Nitrogen trioxide contains 28 parts by weight of N to 48 parts of O.

(4.) Nitrogen tetroxide contains 28 parts by weight of N to 64 parts of O.

(5.) Nitrogen pentoxide contains 28 parts by weight of N to 80 parts of O.

The proportion of nitrogen in these compounds is the same, whilst that of the oxygen is increased by successive additions of 16 parts by weight. No oxide of nitrogen is found composed of 28 parts of nitrogen to 17 of oxygen. If the amount of oxygen is greater than 16 parts it cannot be less than twice 16, or 32 parts; if it is greater than 32 parts it cannot be less than three times 16, or 48 parts; and so on. The same law holds good in every series ot chemical compounds, and hence we are led to the following law, which is known as Dalton's *Second Law of Chemical Combination*:—

179. *Law of Multiple Proportion.*—When an element unites with another in different proportions, the higher proportions are invariably multiples of the lowest.

180. Atomic Theory.—To explain the facts which are expressed by the laws of combining proportions, Dalton revived what is known as the Atomic Theory. Ancient philosophers were divided in opinion on the question of the finite or infinite divisibility of matter. The Epicureans held that matter was incapable of infinite division, and even if we possessed the mechanical appliances to make the minutest possible subdivision, we should at length reach a point at which the particle would resist all attempts to effect a further subdivision. These ultimate particles Dalton called atoms, and he assumed (1) That all matter consists of ultimate and unchangeable particles or atoms;

(2) that atoms of the same element have a uniform weight, but that in different elements they have different weights ; (3) that the combining numbers of chemistry represent these relative weights ; and (4) that between these different atoms there are attractions, which unite them by juxtaposition in the formation of chemical compounds. Dalton maintained that if these ideas are accepted, the laws of definite proportions and multiple proportions follow as necessary consequences. This theory has been of great service in the modern development of the science ; but it has been gradually extended and altered, until now it assumes a quite different form from that which it had at first.

QUESTIONS AND EXERCISES ON NITRIC ACID.

1. How is nitric acid usually prepared, (*a*) on the small scale, (*b*) on the large scale?

2. Describe in detail the composition and properties of nitric acid. Describe the experiments which illustrate its properties.

3. Point out in what respect nitric and sulphuric acids have similar properties, and on the other hand, what differences there are which enable you readily to distinguish the one from the other.

4. Describe fully what takes place when strong nitric acid is added to copper. Give equations.

5. What happens when nitric acid is added to caustic potash, to sodium carbonate, and to gold respectively? Give equations.

6. You prepare hydrogen by the action of zinc on diluted sulphuric acid. Why cannot you prepare hydrogen by the action of zinc on diluted nitric acid ?

7. What are the distinguishing tests of nitric acid, and how may these tests be applied to the detection of nitrates?

8. How many pounds of nitric acid are obtained on distilling 400 lbs. of sodium nitrate with sulphuric acid?

9. How much nitre and sulphuric acid shall I need to prepare nitric acid enough to neutralize exactly 5 lbs. of chalk?

10. Calculate the percentage composition of nitric acid.

11. Express in symbols what takes place when ammonium nitrate is subjected to heat.

eight, but
; (3) that
se relative
there are
formation
hese ideas
ltiple pro-
heory has
ie science;
til now it
at first.

ll scale, (*b*)

s of nitric
perties.
acids have
ences there
a the other.
tric acid is

c potash, to
ations.
diluted sul-
he action of

nd how may

listilling 400

l to prepare
?
id.
nium nitrate

---

12. Point out the resemblances between nitrous oxide and oxygen, and state how they may be distinguished (*a*) by physical tests, (*b*) by chemical tests.

13. If 100 litres of nitrous oxide at 0°C. and 760 mm. be decomposed into free oxygen and nitrogen, what increase in bulk will occur?

14. Give the name and chemical formula of the substance formed by the action of dilute nitric acid on copper. What happens when the gas produced in this reaction comes in contact with air?

15. How is nitrogen trioxide obtained? What occurs when it is added to water?

CHAPTER V.

AMMONIA.

Symbol, NH_3. *Molecular Weight,* 17

While there are five compounds of nitrogen and oxygen, there is but one known compound of nitrogen and hydrogen. Analysis has shown that this compound is made up of one volume of nitrogen combined with three volumes of hydrogen, the four volumes of the elementary gases being condensed to two volumes; the formula of its molecule is, therefore, NH_3.

PREPARATION.

181. By the reduction of Nitric Acid or a Nitrate by Nascent Hydrogen.

Def.—*An element is said to be in the nascent state at the instant it is set free from combination.*

Exp. 1.—Add a little fine granulated zinc or zinc filings to a strong solution of caustic potash in a test-tube. On boiling, hydrogen is given off (Art. 81). Now add a little potassium nitrate or free nitric acid. If the latter is used it must not be in sufficient quantity to neutralize the alkali. On heating the test-tube, the pungent smell of ammonia is readily recognized, and a piece of red litmus paper, held in the mouth of the tube, is rendered blue, showing the presence of alkaline vapor. One part of the nascent hydrogen, developed by the action of the zinc and the caustic potash. deoxidizes the nitre forming water, whilst another portion of it unites with the nitrogen of the nitrate to form ammonia; thus:—

$$\underset{\text{Potassium nitrate.}}{KNO_3} + \underset{\text{Hydrogen.}}{4H_2} = \underset{\text{Ammonia.}}{NH_3} + \underset{\text{Water.}}{2H_2O} + \underset{\text{Potassium hydrate.}}{KOH}$$

Exp. 2.—Add a little nitre to a strong solution of caustic potash in a test-tube, and when boiling pass a current of hydrogen through it; not a trace of ammonia can be detected.

182. Difference between Elements in the Free and in the Nascent State.

The two preceding experiments illustrate the difference in the chemical activity of hydrogen in the nascent state and in the free state. This is explained by supposing that the moment the hydrogen is liberated from the caustic potash two of its atoms unite together to form a molecule and thus it makes its appearance in the free state. If, however, substances are present on which it can act chemically, they are decomposed by the liberated atoms which are more active in that state and before they have partly expended their energies by uniting to form molecules. Experience proves that all bodies, when in the nascent state, have a greater tendency to combine than when in the free state.

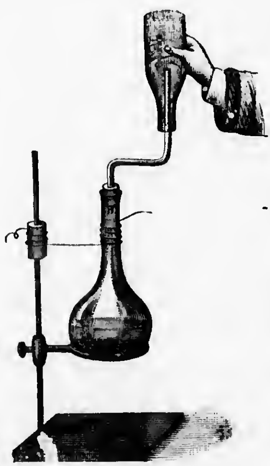
183. By the decomposition of Organic Nitrogenous Substances

Exp 3.—Put a few small pieces of woollen cloth, worsted, feathers, cheese, or indeed of any animal matter, into a test-tube and apply heat; the substance is decomposed, various strongly smelling substances being produced, amongst which ammonia is readily recognized by its alkalinity. All the substances mentioned contain nitrogen and hydrogen, which, when heated out of contact with air, unite and form ammonia. This process is called *destructive distillation*, because the whole of the original matter is destroyed, and new compounds, not originally contained in it, are obtained in the portion distilled.

184. By the action of Slaked Lime on Ammonium Chloride.—Coal contains about 2 per cent. of nitrogen, which during the process of destructive distillation, carried on during the manufacture of coal gas, unites with the hydrogen present in the coal to form ammonia, which is dissolved in water along with other products formed at the same time. This liquor is boiled with milk of lime to liberate the ammonia which distils over and is received into hydrochloric acid, forming ammonium chloride or sal-ammoniac, NH_4Cl; thus:—

$$\underset{\text{Ammonia.}}{NH_3} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Sal-ammoniac.}}{NH_4Cl}$$

Exp. 4.—Take 15 grams (about ½ oz.) of powdered sal-ammoniac and about twice this quantity of slaked lime, mix well together, and introduce into a dry flask, fitted with a cork and glass tube bent twice at right angles, the longer limb being sufficiently long to reach to the bottom of the bottle in which the gas is to be collected. Apply a gentle heat, taking care to move the spirit-lamp to and fro under the flask, the flame never being allowed to rest in one place, otherwise the flask is apt to crack, the substance to be heated being a solid of low conducting power, and not a liquid which would distribute the heat by convection. The gas comes off at a comparatively low temperature. It cannot be collected over water, as it is very soluble in that liquid, but its molecular weight being 17, its specific gravity compared with hydrogen is 8½ (Art. 26); it is, therefore, only about half as heavy as air, and may be collected by *upward displacement*. To ascertain when the bottle is full, hold a piece of red litmus-paper near its neck and slightly above its mouth; if the gas is overflowing, it will quickly restore the blue color; or hold a glass rod moistened with hydrochloric acid in the same place and white fumes are produced. Fill six



bottles, placing them mouth downwards on glass plates. The reaction is expressed by the following equation :—

$$\underset{\text{Ammonium chloride.}}{2NH_4Cl} + \underset{\text{Calcium hydrate.}}{Ca(OH)_2} = \underset{\text{Ammonia.}}{2NH_3} + \underset{\text{Calcium chloride.}}{CaCl_2} + \underset{\text{Water.}}{2H_2O}$$

In order to ensure the decomposition of all the ammonium chloride, a large excess of lime is employed.

One ounce of ammonium chloride, if fully decomposed, will yield about 750 cubic inches; or 30 grams will yield 12½ litres of the gas.

Any other ammonium salt may be used instead of the chloride, and potassium or sodium hydrate instead of lime.

PROPERTIES.

185. Is lighter than Air.

Exp. 5.—Take a dry bottle, and by means of a little gum attach to the inside, near the bottom, a piece of moistened red litmus-paper. Take a bottle of ammonia, bring its mouth under that of the empty bottle, and slowly invert it; the red paper quickly changes to blue, showing the ascent of the ammonia.

186. Solubility in Water.

Exp. 6.—Fill a soup-plate with water, add a little litmus till the water is of a blue color, and then change it red by one or two drops of nitric acid. Bring the mouth of a bottle of ammonia under this water, and gradually withdraw the glass plate; the water rises in the bottle, becoming blue again, and would gradually fill it if the gas were unmixed with air. Shaking the bottle gently hastens the rise of the gas.

The following is a more elegant method of showing the solubility of ammonia in water :—

Exp. 7.—Take a large bottle, fit it with a good sound cork well saturated with paraffine, and fit into the cork a glass tube about 20 cm. long, ending in a jet projecting about 6 cm. into the bottle. Twist a piece of moistened lamp-wick round the tube just below the point. Fill a bottle with red litmus-solution. Now fill the large bottle with ammonia, wet the lamp-wick, insert the cork, and place the tube in the bottle of colored water The ammonia will be absorbed, a fountain will be formed, the red solution becoming blue under the action of the ammonia.

187. Ammonium Hydrate.

Exp. 8.—Take a little water in a narrow test-tube, and place a small rubber band around the tube to mark the height of the water. Place the tube in cold water, and pass ammonia into it; the gas is absorbed by the water, and when saturated it will be seen that the water has increased to the extent of about one-half its volume. This solution of ammonia is colorless, has a very characteristic and pungent smell, and strong alkaline reaction.

The great solubility of ammonia gas in water, accompanied as it is by a considerable evolution of heat, is commonly regarded as due to true chemical combination, a new body being formed which closely resembles potassium and sodium hydrates in its highly alkaline character, and in its power of neutralizing acids and forming salts analogous to potassium and sodium salts. Hence the liquid may be regarded as a hydrate of a monad radicle, NH_4, which acts like a monad metal, and is called *ammonium*. The reaction is—

$$\underset{\text{Ammonia.}}{NH_3} + \underset{\text{Water.}}{H_2O} = \underset{\text{Ammonium hydrate.}}{NH_4OH}$$

If ammonium hydrate, the *liquor ammoniæ* of the druggist, is heated in a test-tube, almost every trace of the gas is removed; thus:—

$$\underset{\text{Ammonium hydrate.}}{NH_4OH} = \underset{\text{Ammonia.}}{NH_3} + \underset{\text{Water.}}{H_2O}$$

This is the readiest method of obtaining gaseous ammonia for the study of its properties. A slight elevation of temperature suffices to decompose it, and nearly all the gas is expelled before the liquid reaches 100°C.

188. Combustion of Ammonia.

Exp. 9.—Pass a taper into a bottle of ammonia held mouth downwards; the flame is first slightly enlarged and then extinguished. The enlargement is due to a very slight and evanescent combustion of the gas. Ammonia, therefore, does not burn when a light is applied to the cold gas. It burns, however, when raised to a high temperature.

Exp. 10.—Put a small quantity of liquor ammoniæ into a flask fitted with a cork, through which passes a small glass tube bent near to the point, so that it can be readily heated by means of a spirit-lamp. Gently heat the flask, and heat the tube to redness near the point; the ammonia burns with a greenish-yellow flame. The reaction is—

$$\underset{\text{Ammonia.}}{4NH_3} + \underset{\text{Oxygen.}}{3O_2} = \underset{\text{Nitrogen.}}{2N_2} + \underset{\text{Water.}}{6H_2O}$$

Ammonia gas may also be burned in an atmosphere of oxygen by the following arrangement:—

Exp. 11.—Take a short tube, about 3 cm ($1\frac{1}{4}$ inches) in diameter, corked at one end. Through the cork pass two tubes, one nearly reaching to the open end of the wide tube, from which to burn ammonia, and the other just passing through the cork to supply oxygen. Place a loose packing of cotton-wool in the tube, in order to diffuse the oxygen pretty equally. On passing currents of the gases through the tubes and lighting the ammonia, it burns as in the last experiment. A similar experiment may be made with a smaller supply of oxygen, by drawing out the tube delivering the ammonia to the form of a jet, bending it, and lowering it into a bottle of oxygen, and applying a light to it just as it enters the mouth of the bottle.

189. Combination with Acids.

As might be expected from its strong alkaline character, ammonia combines with acids, neutralizing them perfectly.

Exp. 12.—Gently heat a wide-mouth bottle, pour into it a small quantity of hydrochloric acid, cover with a glass plate and shake it well. Place this bottle mouth downwards on a bottle of ammonia; on withdrawing the glass plate, the gases combine, disengaging much heat and forming a white solid, ammonium chloride, NH_4Cl, in which the acid and alkali have neutralized each other; thus:—

$$\underset{\text{Ammonia.}}{NH_3} + \underset{\text{Hydrochloric acid.}}{HCl} = \underset{\text{Ammonium chloride.}}{NH_4Cl}$$

Exp. 13—Fill a dry bottle with carbon dioxide, and place it mouth downwards over a bottle of ammonia; a white cloud appears. The gases combine, forming carbonate of ammonia, a salt which occurs in commerce under the name of "salts of hartshorn," and which, with the addition of some perfume, forms the contents of the so-called smelling bottles. It rapidly vaporizes, exhaling the odor of ammonia very strongly, and hence called *sal-volatile.*

190. Ammonium.

It was stated (Exp. 8) that liquor ammoniæ may be considered a hydrate of the monad radicle, NH_4. This radicle has never been obtained in a free state, but it can be obtained combined with mercury as an amalgam.

Exp. 14.—Put a small quantity of mercury into a test-tube, gentle heat, and add a few small pieces of sodium; the metals combine, forming a semi-fluid mass called *sodium amalgam.* Pour out this amalgam before it gets cold into a tumbler one-third full of a cold saturated solution of sal-ammoniac. The amalgam immediately begins to swell, increases to about 15 times its original volume, becoming light enough to float on the liquid. This mass presents a brilliant metallic appearance, and since mercury is not known to unite with any non-metallic substance without losing its metallic appearance, it seems fair to conclude that the soft solid is really an amalgam of mercury and ammonium, the latter taking the place of the sodium; thus:—

$$\underset{\text{Sodium amalgam.}}{Hg_xNa} + \underset{\text{Ammonium chloride.}}{NH_4Cl} = \underset{\text{Ammonium amalgam.}}{Hg_xNH_4} + \underset{\text{Sodium chloride.}}{NaCl}$$

There is, therefore, some experimental evidence as to the existence of the compound metal ammonium, and the close analogy traceable between its salts and those of potassium and sodium tend to confirm this view.

191. Isomorphism.

The resemblance of the salts of ammonium to those of potassium is rendered more striking from the fact that in many cases the crystalline form of the corresponding salts, as well as their texture, color and lustre, is identical. If solutions of these two salts are mixed, neither of them can be crystallized out by itself. When the solution is evaporated the crystals obtained are composed of the two salts mixed in the most varied proportions. Bodies which are thus capable of crystallizing together in all proportions, without alteration of the crystalline form, are called *isomorphous.*

SUMMARY AND ADDITIONAL FACTS.

192. History.—Ammonia was known to the alchemists in its aqueous solution. Priestly first obtained it in the gaseous state in 1774. Scheele, in 1777, discovered that it contained nitrogen, and its true composition was ascertained by Berthollet about 1785. The origin of the name is uncertain. Some consider that it is derived from Jupiter Ammon, near whose temple in the Lybian desert sal-ammoniac, one of its most important compounds, was first prepared. The name hartshorn, or spirits of hartshorn, was adopted from the circumstances of its preparation by distilling the horns of the stag or hart.

193. Sources.—Ammonia is found in the atmosphere, principally in the form of carbonate, in very small quantities. It is present in manure, of which it forms one of the most essential constituents, in all fertile soils, and its salts are found in the juices of plants and in most animal fluids. It can be synthetically prepared by the direct combination of its elements by the silent electric discharge passing through a mixture of nitrogen and hydrogen. The source from which ammonium salts are now obtained is the watery liquid which distils over in the manufacture of coal gas.

194. Properties.—Ammonia is a colorless gas, of a pungent smell and alkaline taste and reaction. It does not support combustion or respiration, and is only feebly combustible. It is dissolved by water with great avidity, much heat being developed and great increase of volume taking place, one volume of water dissolving about 700 volumes of the gas. It is easily liquefied by conducting it into a tube plunged in a mixture of crystallized calcium chloride and ice, which gives a temperature of $-40°$; at a temperature of $-75°$C. it becomes solid.

195. Uses.—In *medicine* its applications are numerous; its pungent odor is reviving and stimulating. In the *arts* it is used for producing artificial cold, for the production of carbonate of soda, for the production of aniline colors, and in the manufacture of indigo. On account of its alkalinity it is used for removing grease from cloth, and in restoring colors which have been changed by acids. In *nature* it furnishes nitrogen to plants. Its salts applied as manures have produced most beneficial effects. Plants watered with a very dilute solution of ammonia have far surpassed in luxuriance similar plants supplied with ordinary water only.

196. Tests.—If salts of ammonia are rubbed together with slaked lime or caustic potash, and the mixture then gently warmed in a test tube, with the addition of a few drops of water, the ammonia is liberated in the gaseous state, and is recognized (1) by its characteristic odor; (2) by turning moistened red litmus-paper to blue; (3) by changing turmeric paper to brown; (4) by changing red logwood paper blue; (5) by rendering paper wet with a solution of copper sulphate blue; (6) by giving rise to the formation of *white fumes* when a glass rod moistened with dilute hydrochloric or acetic acid is brought in contact with it.

Nessler's Test.—For the detection of minute traces of ammonia a reaction, first pointed out by Nessler, is employed. This reagent may be prepared as follows:—Dissolve 5 grams of potassium iodide in a very small quantity of hot water; add to this liquid a saturated solution of mercuric chloride until the red iodide just ceases to redissolve. Now add 12 grams of caustic potash, previously dissolved in a little water; mix and make up the total volume to 100 cubic centimetres with distilled water; finally add a few drops more of mercuric chloride solution, allow to stand and draw off the clear liquid for use. It must not be filtered through paper. This very delicate test is applicable to ammonium hydrate, or ammonium salts. The liquid to be tested should be in a clear glass vessel standing on a white plate. If only a trace of ammonia or ammoniacal salts is present, a yellow tint is produced, whilst with larger quantities a brown precipitate is formed.

QUESTIONS AND EXERCISES ON AMMONIA.

1. How is ammonia prepared? Give the reaction in symbols (*a*) when lime is used, (*b*) when caustic potash is used.

2. What is the chief commercial source of ammonia? Describe the preparation of sal-ammoniac.

3. Enumerate the principal properties of ammonia.

4. What is meant by the *basic* properties of ammonia?

5. Illustrate by symbols the compounds formed when ammonia is passed into sulphuric acid, nitric acid, hydrochloric acid, and water?

6. A gaseous mixture contains ammonia and hydrogen; how can the hydrogen be obtained entirely free from ammonia?

7. How is ammonia usually tested for? What is Nessler's test and how is it used?

8. An alkaline solution (which you are told is either potash or ammonia) is saturated with hydrochloric acid, and evaporated to dryness; how would you ascertain (*a*) by physical tests, (*b*) by chemical tests, the nature of the product?

9. How would you distinguish a mixture of three volumes of hydrogen and one volume of nitrogen from ammonia, which is a chemical combination of the same elements.

10. The salts of ammonia are conceived to contain the radical ammonium. On what grounds is this assumption based?

11. What volume is occupied by 100 grams of ammonia gas?

12. 29 cc. of ammonia are decomposed by a series of electric sparks, and the mixed gases are exploded with 30 c.c. of oxygen. What volume of what gases remains?

13 How much ammonium chloride must be used in order to get 100 grams of ammonia? How many cubic centimetres of hydrogen will the ammonia contain?

14. 100 cubic centimetres of ammonia are completely decomposed by a series of electric sparks, yielding 200 c.c. of mixed hydrogen and nitrogen; an excess of oxygen is next added, when the volume of mixed gases is found to amount to 290 c.c. The mixture is now exploded, when 65 c.c. of gas remain. Show from these data that the symbol for ammonia is NH_3.

15. A body yields, by analysis, 43·75 per cent. of nitrogen, 6·25 per cent of hydrogen, and 50 per cent. of oxygen. What is its formula and its name?

CHAPTER VI.

WATER.

Symbol, H_2O. Molecular Weight, 18.

COMPOSITION—(A) SYNTHETICALLY.

197. By uniting two volumes of Hydrogen and one of Oxygen.

It has already been shown (Art. 20) that water is composed of two volumes of hydrogen and one volume of oxygen.

Exp. 1.—The above experiment can be varied by admitting hydrogen into the eudiometer and noting the volume; taking care not to more than about half-fill the tube with the gases. Press down the mouth of the tube on a piece of india-rubber or pad of blotting paper, and explode the gases. When the tube has cooled admit the water, and note the volume of the residual gas. If 100 volumes of hydrogen and 75 volumes of oxygen be admitted, the gas which remains will be found to occupy 25 volumes. Hence 100 volumes of hydrogen have combined with 50 volumes of oxygen.

198. By the reduction of Copper Oxide by Hydrogen.

The method of performing the experiment has already been described (Art. 79).

(B) ANALYTICALLY.

199. By Electrolysis.

This has been fully described (Art. 19).

200. By decomposing Water by red-hot Iron.

That water is made up of hydrogen and oxygen in the proportion of 2 parts by weight of hydrogen to 16 parts by weight of oxygen, may be shown by passing steam over a weighed quantity of red-hot iron. The water is decomposed, the hydrogen passing over in a free state, and an oxide of iron is formed. The hydrogen may be measured and its weight estimated, whilst the weight of the oxygen in combination with the iron may also be determined (Art. 82).

201. Pure Water.

The preceding experiments have shown that pure water is composed of hydrogen and oxygen. But the water which occurs in nature is never absolutely pure. It may be obtained in a pure state by the following methods:—

202. Purification of Water.

Exp. 2.—(1) *Distillation.*—Half-fill a moderately large sized retort with water, place its neck in a large flask, and place the flask in a pan of cold water. Cover the flask with a cloth or with blotting paper, and pour cold water from time to time upon it. Place the retort upon wire-gauze on a ring of the retort-stand, and boil the water. As fast as the water in the retort is converted into steam, the vapor will pass over into the cold receiver, and will there be condensed again to the liquid state. Reject the first portion of the water that comes over, as it contains carbon dioxide, ammonia, and many volatile substances which the water may have held in solution.

Observe that the *pure water* obtained by distillation in the preceding experiment is colorless, inodorless, and insipid. As regards color, however, when a layer of it, about six feet in thickness, is examined by transmitted light, it appears of a bluish green tint.

(2) *Filtration.*—The separation of suspended matter is effected on the small scale for laboratory purposes by filtration through porous paper placed on a glass tunnel; and on the large scale by employing filtering beds of sand and gravel. In order to separate suspended matter from water used for drinking purposes filter through a layer of charcoal; animal charcoal in coarse grains is the best, but if that cannot be obtained common wood charcoal will do very well. A common garden flower-pot, having the hole covered with a bit of clean-washed flannel, which should be changed from time to time, will make a good filter. Into this put some small gravel, then some white sand, and press down the charcoal on the top of this, and the filter is ready for use. When the charcoal gets clogged take off the top, boil well, dry and it will be as good as ever.

(3) *Freezing.* — Under ordinary circumstances, pure water freezes at 0°C. or 32°F. It becomes more difficult to freeze when certain substances are dissolved in it; thus, sea water freezes at about 3½ degrees lower than fresh water. Dissolved matters are to a great extent, although not completely, separated from ice, and retained in the unfrozen water. Water obtained by melting the ice of sea water is used for drinking in the Arctic regions.

203. Relation of Water to Heat.

(1) *Expansion and Contraction.*

Exp. 3.—Take a small thin flask, fit it with a cork, through which passes a small glass tube about 2ft. long. Fill the flask with cold boiled water, insert the cork, taking care that there are no air-bubbles entangled beneath it, and press it into the bottle until the small tube is about half full. Now heat the flask, and the liquid slowly rises in the tube. *Water, therefore, expands by heat.* This property water shares with other substances. It may be said to be a law of nature that inorganic substances expand when heated.

Next immerse the flask in a mixture of snow or pounded ice and half its weight of common salt; or a mixture of sulphate of soda (Glauber's salt) and hydrochloric acid will do equally well, the salt being just covered with the acid. As the water cools the liquid sinks in the tube, becomes stationary, and then begins to rise. If a thermometer could be plunged into the water in the flask it would be found to mark 4°C. or 39·2°F. when the water began to rise. The expansion of the liquid goes on till a sudden check is observed; if the flask is then examined it will probably be found to be cracked and to contain ice. *Thus water when cooling contracts till the temperature of 4°C. or 39·2°F. is reached, and when further cooled it expands.* If similar experiments are made with other liquids, such as alcohol, oils, etc., they will be found to contract, but not to expand again as the temperature is reduced. Thus water is the great exception to the general law, and in this respect stands alone amongst liquids hitherto examined.

But water not only expands in this unique way before it freezes, but in the act of freezing, it undergoes a large and further expansion. This expansion, exerting an almost irresistible force, plays an important part in the disintegration and splitting of rocks during the winter. The same cause leads to the bursting of water-pipes. Water shares this property with a few other substances, such as cast-iron, bismuth, and antimony. This property of water is of the greatest moment to mankind. If it obeyed the ordinary law, our lakes and rivers would become masses of solid ice, and all animal life in them would perish. The heat of summer would be unable to undo the effects of the winter's cold, and the climate would be so altered as to render any but equatorial regions almost uninhabitable.

(2) *Boiling.*

At ordinary temperatures water is continually giving off invisible vapor, which diffuses into the surrounding atmosphere. *When the temperature is reached at which the pressure (tension) of its vapor is equal to that of the atmosphere at the time, it is said to boil.* This temperature is 100°C. at 760° mm. pressure, or 212°F. at 30 inches barometric pressure. Since the boiling point depends on atmospheric pressure, it will be lowered at high elevations. It has been found that in ascending mountains the boiling is lowered 1°F. for every 590 feet.

(3) *Latent Heat of Water.*

When ice at 0°C. melts it absorbs *without elevation of temperature* as much heat as would raise the temperature of an equal weight of water from 0°C. to 79°C This quantity of heat is required to change the state from solid to liquid water, and is spoken of as its *latent heat.* When water freezes, or becomes solid, this amount of heat, which is necessary to keep the water in the liquid form, and is, therefore, termed the *heat of liquidity,* is evolved or rendered sensible.

(4) *Latent Heat of Steam.*

When water is converted into steam a large quantity of heat becomes latent, since, notwithstanding the continuous action of the fire the temperature remains constant. It has been found that it will take as much heat to change one gram of water at 100°C. into steam of the same temperature as it would to raise 537 grams of water one degree in temperature. *The latent heat of steam is, therefore, 536 thermal units*—a thermal unit being the *amount of heat required to raise a unit weight of water through 1°C.* When water evaporates or passes into the gaseous state, heat is absorbed, and so much heat may thus be abstracted from water that it may be made to freeze by its own evaporation.

204. **Water as a Solvent.**—Water is the most valuable known solvent, there being few substances, solid, liquid or gaseous, which are absolutely insoluble in it. As a rule, the

power of water to dissolve solids increases with the temperature while the solubility of gases is greater at low than at high temperatures.

205. Impurities in Water.—Natural waters are never free from dissolved impurities. They contain gaseous, liquid, or solid impurities, varying according to the source from which they are derived, and the nature of the soil or rocks over which they have flowed. If water containing carbonic acid percolates through cretaceous rocks, the carbonic acid combines with the insoluble calcium carbonate, forming soluble calcium bicarbonate (Art. 132). Of the solid matter dissolved in drinkable waters, the greater part is usually made up of calcium salts, generally accompanied by small quantities of magnesium salts.

206. Lime in Water.

Exp. 4.—Half-fill a test-tube with water, add a little lime-water, and pass carbon dioxide through it; it first becomes milky and then clear again. Now add a solution of ammonium oxalate, $(NH_4)_2C_2O_4$, or oxalic acid, $H_2C_2O_4$; the water becomes turbid. Again, half-fill the test-tube with water, and add a little plaster of Paris, $CaSO_4$, shake well and filter. Now add ammonium oxalate or oxalic acid, and the water becomes turbid. *Hence, ammonium oxalate or oxalic acid is a test for lime.*

Six grains of lime per gallon will yield a slight turbidity, 16 grains a distinct precipitate, and 30 grains a large precipitate soluble in nitric acid.

If the water contains calcium bicarbonate the reaction is—

$$\underset{\text{Calcium bicarbonate.}}{H_2Ca(CO_3)_2} + \underset{\text{Ammonium oxalate.}}{(NH_4)_2C_2O_4} = \underset{\text{Calcium oxalate}}{CaC_2O_4} + \underset{\text{Ammonium bicarbonate.}}{2NH_4HCO_3}$$

207. Hard Water.

Water that contains an excess of calcium or magnesium salts is said to be *hard*, while one not so charged is said to be *soft*.

208. Action of Hard Water on Soap.

Exp. 5.—Dissolve a small quantity of soap in hot alcohol (strong whiskey), and add a little of it to rain water; a "lather" is produced. Now add a little of the soap solution to water containing calcium bicarbonate or calcium sulphate; the soap immediately curdles. There are two organic acids: *stearic acid*, $C_{18}H_{36}O_2$, and *oleic acid*, $C_{18}H_{34}O_2$, which are formed in most natural fats, the former being about three times the latter. These acids combine with soda and potash to form soap, the combination with soda forming *hard soap*, $NaC_{18}H_{35}O_2$. The lime in the water decomposes the soap, forming *calcium stearate*. This substance constitutes the thin scum seen on the surface of hard water when treated with soap. Before it is possible to obtain a *lather* with a hard water, it is necessary to convert the whole of the calcium and magnesium salts present into insoluble stearates. Since stearates have no detergent action, the presence of large quantities of the salts of lime in water seriously impair its economic value.

Water which readily curdles soap is called Hard Water.

209. Temporary Hardness.—We have seen (Art. 132) that when water containing calcium bicarbonate is boiled, the lime is precipitated in the form of calcium carbonate.

Exp. 6.—Half-fill a test-tube with water containing calcium bicarbonate in solution, and carefully add lime-water to it; the solution becomes milky. The lime-water converts the soluble bicarbonate into the insoluble carbonate. The reaction is expressed by the following equation:—

$$\underset{\text{Calcium bicarbonate.}}{H_2Ca(CO_3)_2} + \underset{\text{Calcium hydrate.}}{Ca(OH)_2} = \underset{\text{Calcium carbonate.}}{2CaCO_3} + \underset{\text{Water.}}{2H_2O}$$

Hardness which may be got rid of by boiling or by the addition of lime-water, is called "Temporary Hardness."

210. Permanent Hardness.

Exp. 7.—Shake up a little calcium sulphate (Plaster of Paris) with water and filter. The water now contains calcium sulphate, which is soluble to the extent of one part by weight in about 460 parts of water, while calcium carbonate is only soluble to the extent of 2 grains in a gallon (70,000 grains). Boil a portion of the water; no change is observed. Add lime-water to another portion; it still remains clear, the lime remaining in the water.

Water which cannot be made soft by boiling, or by the addition of lime-water, is said to be "Permanently Hard."

211. Total Hardness—Degree of Hardness.

The term "total hardness" implies both temporary and permanent hardness. In expressing hardness in degrees, it is to be understood that every degree of hardness represents 1 grain of calcium carbonate or its equivalent in soap-destroying power, in a gallon of water. Thus, water that has a hardness of 10 degrees contains 10 grains of carbonate of lime in a gallon. Most chemists now, however, estimate the hardness in 100,000 grains of water.

A soft water is one below six degrees of hardness; a water of 8 or 10 degrees is moderately soft, while a water of 20 degrees and upwards is hard A good drinking water should not have a hardness of more than 15 degrees. A very hard water is injurious for drinking purposes, because its power as a solvent for food is impaired, and because it is absorbed by the stomach with greater difficulty than a soft water, thus giving rise to indigestion and dyspepsia.

212. The Softening of Hard Water.

(1) *By Distillation.* This method is chiefly, if not solely, used for the preparation of fresh from salt water, for the supply of ships at sea. When freshly prepared it has a vapid and un-

pleasant flavor, due partly to deficient aëration and partly to the presence of traces of soluble organic substances; both of these objections may be removed by filtration through charcoal.

(2) *By the Action of Sodium Carbonate.*

Exp. 8.—Add sodium carbonate, Na_2CO_3 (washing soda), to water containing calcium bicarbonate; a white precipitate is formed. The sodium carbonate forming soluble sodium bicarbonate, and insoluble calcium carbonate:—

$$\underset{\text{Calcium bicarbonate.}}{H_2Ca(CO_3)} + \underset{\text{Sodium carbonate}}{Na_2CO_3} = \underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Sodium bicarbonate.}}{2HNaCO_3.}$$

This action takes place slowly in the cold, and rapidly at a higher temperature. As the sodium salts have an unpleasant taste, water thus softened cannot be used for drinking purposes, but the method is commonly in use by laundresses, who thus avoid the waste of soap.

Exp. 9.—To water containing calcium sulphate in solution add washing soda; soluble sulphate of soda is formed and insoluble calcium carbonate precipitated:—

$$\underset{\text{Calcium sulphate.}}{CaSO_4} + \underset{\text{Sodium carbonate.}}{Na_2CO_3} = \underset{\text{Calcium carbonate.}}{CaCO_3} + \underset{\text{Sodium sulphate.}}{NO_2SO_4}$$

Water containing salts of magnesium is softened in the same manner. Instead of sodium carbonate, wood-ashes, which contain potassium carbonate, may be used.

(3) *By Boiling.* This method only applies to that part of the hardness of water which is due to the presence of the bicarbonates of lime, magnesia, and iron, but is not applicable to that arising from sulphates. When water containing bicarbonates is boiled, the bicarbonates are gradually decomposed, the insoluble carbonates being precipitated and carbon dioxide expelled with the steam. Ebullition for half an hour is sufficient.

(4) *By the addition of Lime.* This process is applicable to the same kind of water as the preceding, and to the same extent.

The reaction is given in Experiment 12. The process is now in use on a large scale in numerous places, being applied to the entire water supply of several towns.

213. Lead in Water.—As lead acts as a cumulative poison its salts produce serious results if taken into the system, even in very minute quantities, for a length of time. It is, therefore, of great importance that suspected water should be carefully examined. Pure recently-boiled water has no action on lead, provided air is excluded; but if lead is exposed to the united action of air and water, it is oxidized to the hydrate which dissolves. Potable water always contains a certain amount of salts in solution, and the corrosive action on lead depends upon the nature and quantity of the salts thus present. Salts of ammonia act most prejudicially in this respect, while certain salts of lime, such as the carbonate and sulphate, by forming an insoluble film on its surface, either retard or altogether prevent the action of the water on the lead. This does not apply, however, to water containing a large excess of carbon dioxide, in which case a soluble bicarbonate is formed. Also, some kinds of vegetable matter form insoluble compounds with lead.

Exp. 10.—Boil 100 c. c. of water, acidulated with a few drops of hydrochloric acid, and add sulphuretted hydrogen water (Art. 19). If a brown or black coloration is produced, the presence of lead may be inferred. This black coloration is lead sulphide, PbS. Few waters kept in lead cisterns can be met with which do not yield some trace of lead.

ORGANIC IMPURITIES.

The organic matter may be either of vegetable or animal origin, the latter being by far the most injurious; but water containing any considerable quantity of vegetable matter, partly in suspension and partly in solution, is decidedly unwholesome. Water may be tested for organic matter by the following methods :—

(1) *Potassium Permanganate.*

Exp. 11.—Fill a test-tube nearly full with the water to be examined, and add, by means of a glass tube, as much potassium permanganate as will impart a distinct pink tinge after stirring with the tube. Then fill another test-tube of the same size with distilled or rain water, and add the same quantity of permanganate solution. Place the test-tubes side by side on a sheet of white paper, and note any difference between the two tubes which may speedily or subsequently take place. If decoloration takes place rapidly, there is a strong probability that organic matter of animal origin is present, whereas slower changes indicate that vegetable matter is present. There are other substances, such as nitrates, iron, and sulphuretted hydrogen, which would produce the same effect as the organic matter. The presence of nitrates indicates that the water is unfit for use; the other two are not likely to be present in ordinary water. The decoloration is owing to the oxidation of the organic matter by the oxygen of the permanganate. A sort of rough estimate of the amount of organic matter may be made by observing the relative quantities of permanganate which different waters decolorize.

(2) *Chlorides.*—The presence of chlorides in water is always very suspicious, not that chlorides are in themselves of importance, but because their presence serves as an indication of sewage contamination, for pure waters are almost free from sodium chloride, whilst sewage is highly contaminated with it.

Exp. 12.—Half-fill a test-tube with water, acidulate with a few drops of nitric acid, and add silver nitrate solution. Four grains per gallon of sodium chloride give a turbidity; ten grains a slight precipitate; twenty grains a considerable precipitate soluble in ammonia. Good water should only yield a slight haziness.

(3) *Ammonia.*—The presence of free ammonia in considerable

quantities in water points to sewage contamination. Good potable water should contain very little free ammonia. Its presence may be detected as follows :—

Exp. 13.—To a test-tube half full of well-water add five or six drops of Nessler's Test (Art. 196). A yellow or brown color indicates sewage contamination.

214. Naturally Occurring Waters.

(1) *Rain Water.*—Although this is the purest form of natural water, still it contains certain impurities which are washed out by it from the atmosphere. It invariably contains ammoniacal salts, sodium chloride, and organic matter of various kinds.

(2) *Spring Water.*—The nature and amount of material in spring water depends on the nature of the strata through which it passes. The salts which most commonly occur are (1) the bicarbonates and carbonates of calcium and magnesium, (2) the sulphates of calcium and magnesium, (3) the alkaline carbonates, chlorides, sulphates, nitrates, or silicates. The gaseous constituents consist of oxygen, nitrogen, and carbon dioxide.

(3) *River Water.*—Although river waters contain a smaller amount of salts, it is usually less fitted for drinking purposes than ordinary spring water, as it usually holds in solution a larger proportion of organic matter of vegetable origin, derived from the extensive surface of the country which has been drained by the stream.

(4) *Sea Water.*—This usually contains about 3½ per cent. by weight of substances in solution, the one which is present in by far the largest quantity being common salt, NaCl.

QUESTIONS AND EXERCISES ON WATER.

1. Water is said to be a compound of oxygen and hydrogen; describe experiments in proof of this view.
2. How is the composition of water ascertained by the eudiometer?

After exploding a mixture of oxygen and hydrogen, ·017 gram of water was obtained, and 11·5 c. c. of hydrogen remained. The temperature was 15°C. and the pressure 760 mm.; find the volume of the mixed gases.

3. Explain how the action of hydrogen on copper oxide may be used as a means of determining the composition of water.

Berzelius and Dulong heated 53·821 grams of copper oxide in contact with hydrogen. The residual copper weighed 42·989 grams, and 12·197 grams of water were obtained. Calculate from this data the percentage composition of water.

4. What are the characteristics of pure water, and how may water be obtained in a pure state? How may drinkable water be obtained from sea water?

5. At what temperature is water at its point of greatest density? How may this be shown experimentally?

What effect would continuous frosty weather have on lakes and rivers if water expanded and contracted according to the same rule as a piece of solid iron?

6. What is meant by the boiling point of water? How does the pressure of the atmosphere affect the temperature at which water boils?

Two thin flasks are filled with water and sealed up. One is placed in boiling water and the other in a freezing mixture. What occurs in each case?

7. What is meant by the term *hardness* as applied to water? Give the earthy impurities which are the frequent cause of hardness in water and of deposits in boilers, and state how they are removed. What is meant by saying that a given specimen of water is 10 degrees of hardness?

8. Some hard waters can be softened by boiling, whilst others cannot. Explain the cause of this, and describe any other methods for softening waters.

9. A sample of water contains in 100,000 parts, 16 parts of calcium carbonate, 5 parts of calcium sulphate, and 7 parts of common salt; what is its hardness, and how much will this hardness be reduced by boiling the water for half an hour?

10. Supposing a certain water contains 20 grams of calcium carbonate in the gallon, and that the following equation represents the

action of a solution of soap on calcium carbonate, $2NaC_{18}H_{35}O_2 + CaCO_3 = Ca(C_{18}H_{35}O_2)_2 + Na_2CO_3$. What weight of soap must be added to the water before a permanent lather can be produced in one gallon of water?

11. How may the presence of lead in waters be accounted for, and how may its presence be detected?

12. How may the presence of organic matter in water be detected, and how would you distinguish organic matter of vegetable origin from that of animal origin?

CHAPTER VII.

ATMOSPHERIC AIR.

CONSTITUENTS.

215. Oxygen and Nitrogen.—The proportion of these gases respectively present in the air may be estimated in various ways :—

(1) By withdrawing the Oxygen from a given volume of Air by Phosphorus.

This can be done, as already shown in the preparation of Nitrogen, either by the active combustion of phosphorus (Art. 88) or more accurately by the slow combustion of the same element (Art. 89). In the latter case the phosphorus may be fixed on the end of a copper wire. In each case when the water is adjusted to the same level within and without the bottle, by depressing it to the requisite extent, it will be seen that about one-fifth of the air in the bottle has disappeared.

(2) By absorbing the Oxygen with Pyrogallate of Potash.

Exp. 1.—Take a glass tube, about 1 metre long and 15 mm. in diameter, and divide it into six equal parts by means of small india-rubber bands. Pour through a small glass tube a strong solution of pyrogallic acid till the first division is about one-third full. Wash the tube, place it so that it may pass through the pyrogallic acid, and again pour through it a strong solution of caustic potash till on withdrawing it the liquid may stand a little above the first ring. Place a small piece of india-rubber on the mouth of the tube, shake it well, and invert in a tumbler of water. On withdrawing the thumb, the water rises in the tube, and on adjusting the water to the same level within and without the tube, the water should stand a little above the second ring, showing that about one-fifth of the air has been absorbed. This one-fifth is oxygen, which has been absorbed by the pyrogallate of potash.

(3) By the Eudiometer.

Exp. 2.—To 100 volumes of air in the Eudiometer add 50 volumes of hydrogen and explode. The 150 volumes will shrink (steam being condensed) to 87 volumes; therefore,

$$150 - 87 = 63 \text{ vols. loss.}$$

Of these 63 vols. one-third is oxygen; there are, therefore, 21 volumes of oxygen in 100 volumes of air.

(4) By drawing a measured volume of Air over red-hot Copper, and then weighing the Copper Oxide produced.

Exp. 3.—Take the hydrogen bottle, press the funnel-tube nearly to the bottom of the flask, and add a little water to cover the end of it. Attach a drying-tube, filled with calcium chloride, and connect this with a tube of hard glass filled with bright copper turnings. Bring the copper turnings to a red heat, and then pour water through the funnel-tube to expel the air and cause it to pass over the red-hot copper, which combines with the oxygen to form copper oxide. The nitrogen may be collected in the usual manner.

This experiment illustrates the principle of the method adopted by Dumas and Boussingault in their precise determination of the composition of air by weight. They passed a given volume of air (1) over *calcium chloride*, then (2) over *caustic potash*, and finally (3) over *ignited copper* reduced from its oxide, severally contained in glass tubes accurately weighed before the experiment was commenced. The increase in the weight of the calcium chloride indicated the moisture, of the caustic potash tube, the carbon dioxide, and of the copper tube the oxygen, severally contained in the volume of air operated upon. The residual gas, which is nitrogen, was collected in an exhausted and weighed globe, the increase in the weight of which gave the nitrogen. The mean of a large number of experiments of this kind, in which every possible precaution against error was taken, gave the following results, with which we give the results of the volumetric analyses :—

		PER CENT.	IN ROUND NUMBERS.
By weight	Nitrogen	76·995	77
	Oxygen	23·005	23
By volume	Nitrogen	79·04	79
	Oxygen	20·96	21

216. Constancy of Composition.

Atmospheric air is nearly constant in composition. The results of numerous analyses at various points of the earth's surface and at considerable heights above the level of the sea, show but little variation. Angus Smith has found that the percentage of oxygen in air from the sea-shore and from Scottish moors and mountains, is as high as 20·999 per cent., and in the free air of towns, and especially during foggy weather, it may sink as low as 20·82. This constancy of composition led some of the early chemists to consider air as a chemical compound of one volume of oxygen and four volumes of nitrogen. That this is not the case appears from the following facts.—

(1) If pure air were a definite compound of oxygen and nitrogen it should be absolutely constant in composition. But it is not quite constant, and, therefore, this fact alone is sufficient proof that the gases are not combined, but only mixed together, as the constituents of a compound always occur in invariable proportions (Art. 17).

(2) This conclusion is confirmed in many ways. Thus, on mixing oxygen and nitrogen in the proportion in which they are found in the atmosphere, none of the phenomena, such as

evolution of heat, and alteration in properties and volume, usually attendant on chemical combination, are perceived; nevertheless, the mixture is actually identical in composition with atmospheric air and possesses all its properties.

(3) Were air a compound it should dissolve in water as such, that is, the proportion of oxygen and nitrogen in dissolved air should be the same as in undissolved air; but if a mixture, the more soluble constituents should dissolve the more readily, and, therefore, more oxygen than nitrogen should dissolve, since oxygen is more soluble than nitrogen. Experiment shows that the latter is the case. If water which has been recently boiled, and then allowed to cool out of contact with air, be shaken with air and the dissolved air be then expelled by boiling, and collected, it will contain 32 instead of 21 per cent. by volume of oxygen

(4) The oxygen and nitrogen in the air do not present a simple ratio to the atomic weights of these elements.

217. Diffusion of Gases.

The uniformity of composition of the atmosphere is partly due to the operation of winds, but it mainly depends upon DIFFUSION, *by which gases mix in opposition to gravitation, and when mixed remain so.* This may be illustrated by the following experiments:—

Exp. 4.—Fill a bottle with carbon dioxide, and leave it mouth upwards. In about half an hour introduce a burning taper; it will not be extinguished, thus proving that the gas has escaped from the vessel. The carbon dioxide, though heavier, has *diffused* into the air above, although the latter is lighter.

Diffusion between gases takes place if they are separated by a porous partition. A partition made of plaster of Paris is very suitable for experiments on diffusion.

Exp. 5.—Take a glass funnel, the mouth of which is about 6 or 8 cm. wide. Place upon the table a plate of glass somewhat larger than the funnel, and pour over it soft plaster of Paris, so as to form a layer 2 or at most 3 mm. thick. Press the funnel upon it, and leave it in this position for half an hour. Remove the plaster round the funnel with a knife, and blow through the tube by this means the funnel may be easily lifted. Leave the glass plate in the sun or in a warm place for an hour, then remove the disc and place it upon three small corks, and leave it for a day or two, so as to dry thoroughly. Now heat the rim of the funnel over the spirit-lamp, till hot enough to melt sealing-wax, and place a layer of the wax round the rim. When this layer has partially hardened, make it thicker with a piece of sealing-wax heated over the spirit-lamp; then heat the whole uniformly, and invert it over the plaster. Now place the funnel, stem upwards, upon a glass plate, and fill it with carbon dioxide by displacement; then lift it together with the glass plate, dip the end in water, and remove the plate. The carbon dioxide will now pass out through the plaster wall, but the lighter air flowing inward with greater velocity, increases the volume of the gas contained in the funnel; the consequence is that bubbles of gas escape from the end of the funnel, and rise through the water. Again, place the funnel in an upright position, and while still covered with the glass plate fill it with hydrogen by displacement, and dip the end into water; the lighter gas will diffuse outwards more rapidly than the heavier enters inwards. The volume of the gas in the interior diminishes and in the course of a few seconds the water rises to about half the height of the funnel.

For the experiment with carbon dioxide, the funnel should dip only a few millimetres into the water, so as not to obstruct unnecessarily the escaping gas-bubbles. With hydrogen the tube must dip somewhat deeper, or the end of it would be above the surface of the water when the latter rises in the funnel.

Graham found that hydrogen diffuses four times as fast as oxygen. Now, the densities of these gases are as 1 : 16; but their diffusive rates are as 4 : 1. This applies to the diffusion of all gases. Hence we have the following law :—

Graham's Law of Gaseous Diffusion.—*The diffusive rates of two masses of gas in contact are inversely proportional to the square roots of their densities.*

It is mainly owing to this gaseous diffusion that gases of such different weights as those which form the atmosphere are kept uniformly distributed instead of forming layers with the heaviest at the bottom. In obedience to this law the heavier gases are compelled to rise, and the lighter ones to fall, until the proportions of them are all the same throughout.

218. Ozone.—The presence of ozone in the atmosphere may be detected by ozone paper, as in Art. 61. The general facts regarding ozone may be stated as follows :—

(1) More ozone is present during the night than during the day, and most of all is found at day-break.

(2) More is found on the sea-coast, and specially when the wind is blowing from the sea.

(3) More is found in the country than in towns. In country air it does not amount to more than to one volume in 700,000 volumes of air.

(4) It is invariably formed when water evaporates, and this is probably the main source of atmospheric ozone; atmospheric electricity and the action of aromatic plants and flowers may be reckoned as minor sources.

219. Aqueous vapor.

Exp. 6—Bring a tumbler of cold water into a warm room, or put a few pieces of ice into a test-tube; moisture is soon deposited on the outside of the glass. A given volume of air cannot take up more than a certain quantity of aqueous vapor and then it is said to be *saturated.* The quantity taken up depends entirely on the temperature, At the temperature of 60°F. there are about 6 grains of aqueous vapor in each cubic foot of air, or one grain to the gallon. Taking a fair-sized room twenty-five feet long and seventeen broad, containing about six thousand cubic feet, it would take about half a gallon or five pounds of water to make the aqueous vapor in this room at 60°F. If the temperature were only 40°F. the aqueous vapor would only be half a grain to the gallon of air,—just half what it is at 60F. Hence the necessity of supplying moisture to the atmosphere when it is raised in temperature.

220. General facts respecting moisture.

(1) Air rarely contains its full saturated amount of moisture except in very cold weather, or in very hot tropical seas.

(2) The most comfortable degree of saturation is from 66 to 70 per cent. More than this checks evaporation from the body, whilst less causes too great evaporation, thereby parching the mouth and drying the skin.

(3) It has been noted that in certain places remarkable as health resorts the degree of saturation is remarkably uniform.

221. Carbon Dioxide.

Exp. 7.—Pour some lime-water into a saucer, leave it exposed to the atmosphere for twenty-four hours, when a thin scum will be found on its surface. Put this scum into a test-tube, add a little hydrochloric acid, and a brisk effervescence will take place, carbon dioxide being given off, which may be collected and its presence indicated in the usual way.

Carbon dioxide is, therefore, a constituent of the atmosphere. The average amount of this gas in free open country air is between 3 and 4 volumes in 10,000 volumes of air, whilst in towns where much coal is burnt the amount may rise as high as 6 or 7 volumes in 10,000. When present in certain quantities it acts most prejudicially on the higher forms of life, nor is the amount which becomes hurtful far removed from the amount at present existing in the air. According to Dr. Parkes, an eminent authority on this subject, air is unhealthy when the carbon dioxide exceeds .06 per cent. or 6 volumes in 10,000. Hence in a sanitary point of view it is exceedingly important to ascertain when this amount is exceeded. For this purpose Dr. Angus Smith gives the following method :—

Exp. 8.—Take a bottle of clear white glass having a well-fitting stopper, and when quite full containing as nearly as possible ten and a-half fluid ounces of water. Fill the bottle with the air to be tested by putting a glass tube to the bottom and sucking out the air. Now pour into it half a fluid ounce of clear lime-water, insert the stopper, and shake vigorously for a short time. Let the bottle stand so that the air bubbles may rise, and observe whether or not turbidity or opalescence is produced. If the liquid does not remain bright and clear the air examined contains more than 6 volumes of carbon dioxide in 10,000 volumes of air, but if no turbidity is produced the sample tested may be considered wholesome, the carbon dioxide being in this case taken as the measure of its general purity. Dr. Smith proposes the following rule as a practical application of this method :—"Let us keep our rooms so that the air gives no precipitate when 10½ oz. bottle is shaken up with half an ounce of clear lime-water." In order that the air in an occupied room may not contain more than ·06 per cent. of carbon dioxide, 3,000 feet of fresh air must be introduced per hour for each person, and about twice this volume for every gas burner that consumes three cubic feet per hour. Fortunately this renewal of air takes place to a considerable extent in most rooms, even when the doors and windows are shut, by the chimney, by cracks and crevices in the doors and windows, and especially through walls. Most building materials are porous when dry but become nearly air-tight when wet. Compact wall-linings, such as ordinary wall-papers, tend to keep walls damp and therefore render rooms much more unhealthy than the old-fashioned whitewash.

222. Ammonia.

Exp. 9.—Fill a bottle with fresh rain water, add to it a small quantity of Nessler's solution, and let it stand for a short time; the water will become of a pale yellow color, indicating the presence of ammonia.

The ammonia present in the atmosphere rarely exceeds one part in a million but this portion, when carried to the soil by means of rain, amounts to between five and six pounds per acre annually. It is from this source that unmanured crops derive the greater part of the nitrogen which they require for the formation of seed and other portions of their structure, plants being unable to assimilate free nitrogen.

223. Nitric Acid.

This substance is always present in the atmosphere in small quantities. It is a source of nitrogen to plants. Sulphurous and sulphuric acids are present in the atmosphere of towns where coal is burnt.

224. Organic Matter.

The atmosphere also contains various organic vapors and particles of dust, which are partly organic and partly inorganic. Amongst the organic substances are the germs of plants and animals. These bodies are the propagators of fermentation and putrefaction, and air which has been freed from them, either by filtration through cotton wool, by ignition, or by subsidence, may be left in contact for any length of time with milk, the juice of meat, etc, without the liquids undergoing the slightest change. When a ray of light is allowed to pass through air thus freed from solid particles, no reflection is noticed, and the space appears perfectly empty, the motes which in ordinary air reflect the light being absent.

QUESTIONS AND EXERCISES ON THE ATMOSPHERE.

1. What are the two chief constituents of air? By whom and when was each discovered?

2. Name the bodies which are *always* present in the air in small quantities, and those which are *occasionally* present.

3. Describe a method for recognizing with certainty each constituent of the atmosphere.

4. Is air a mechanical mixture or a chemical compound? Give reasons in support of your answer.

5. How may the composition of air be ascertained by the eudiometer?

6. How may phosphorus be used to demonstrate approximately the proportion of oxygen and nitrogen present in the air?

7. Describe fully how Dumas and Boussingault determined the composition of air by weight, and sketch the apparatus required for this purpose.

8. In an analysis by Dumas and Boussingault the copper employed gained in weight 1·4094 gram. Find the number of grams of air used and of nitrogen obtained.

9. In a well-ventilated room the amount of carbon dioxide ought not to exceed 6 volumes in 10,000 of air. What weight does this correspond to per litre of air, and how much calcium carbonate would it precipitate from lime-water?

10. A man inhales 18 cubic feet of air per hour. What weight of oxygen does he consume in 24 hours?

11. A man exhales 1,122 cubic inches of carbon dioxide per hour. How many men, at the same rate, would be required to produce 2,240 lbs. of carbon dioxide in 24 hours?

12. Give a short account of what is known about the solid matters, organic and inorganic, found floating in air? What impurities does rain water extract from the air?

CHAPTER VIII.

SECTION I.

METHANE, MARSH-GAS OR LIGHT CARBURETTED HYDROGEN.

Symbol, CH_4. Molecular Weight, 16.

A great many compounds of carbon and hydrogen are known, but they belong to the department of organic chemistry. It will, however, be found convenient to study two of them here. Marsh-gas is the simplest compound of carbon and hydrogen, in which one atom of carbon is united to four atoms of hydrogen; its formula is, therefore, CH_4.

PREPARATION.

225. From Sodium Acetate and Sodium Hydrate.

Exp. 1.—Mix together in a mortar 2 grams of sodium acetate, 4 grams of caustic soda, and 8 grams of slaked lime. Introduce the mixture into a large test-tube, heat, and collect the gas at the pneumatic trough. The lime is merely to prevent the fusion of the mixture, and to prevent it from stopping up the tube. The reaction is—

$$\underset{\text{Sodium acetate.}}{NaC_2H_3O_2} + \underset{\text{Sodium hydrate.}}{NaOH} = \underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Marsh-gas.}}{CH_4}$$

PROPERTIES.

226. Burns but does not support Combustion.

Exp. 2.—Notice that it is tasteless, colorless and inodorous. Hold a bottle of the gas mouth downwards, and apply a flame; the gas burns with a flame a little more luminous than that of hydrogen. Thrust the taper into the bottle; it is extinguished.

227. Lighter than Air.

Exp. 3.—Prove this by pouring from one bottle into another as in Hydrogen, Exp. 5. Since the molecular weight of the gas is 16, its density compared with hydrogen is 8; it is, therefore, the next lightest body known.

228. Products of Combustion.

Exp. 4.—Pour a little lime-water into the bottle used in the last experiment, and shake it up; the lime-water becomes turbid, showing the presence of carbon dioxide; carbon is, therefore, one of the constituents of the gas. While the gas is coming off pass it through a drying-tube, ignite it and hold a cold body in the flame; it soon becomes bedewed with moisture, showing that hydrogen is another of the constituents of the gas.

229. Explosive Mixture with Oxygen and Air.

Exp. 5.—Mix in a soda-water bottle one volume of the gas and twice its volume of oxygen, cork the bottle, wrap a towel tightly round it, and shake it well, take out the cork, and apply a light to its mouth; a sharp explosion occurs, carbon dioxide and water is formed; thus:—

$$\underset{\text{Marsh-gas.}}{CH_4} + \underset{\text{Oxygen.}}{2O_2} = \underset{\text{Carbon dioxide.}}{CO_2} + \underset{\text{Water.}}{2H_2O}$$

If air is used instead of pure oxygen the mixture also explodes, affording the same products, but it will require ten times its volume for complete combustion.

Marsh-gas is very abundant in nature, being produced by the decay of submerged vegetable matter. It is the principal part of the gas which escapes in bubbles when the muddy bottom of a stagnant pool is disturbed. It is found in coal mines, where it is known under the name of "fire damp." When mixed with air and ignited, it causes the violent explosions in coal mines. The product of combustion, the principal part of which is carbon dioxide, is called "after-damp" or "choke-damp."

SECTION II.

ETHENE, OLEFIANT GAS, OR HEAVY CARBURETTED HYDROGEN.

Symbol, C_2H_4. *Molecular Weight, 28.*

This gas is one of the products of the destructive distillation of organic bodies when the temperature is not excessive. It is an important constituent of coal-gas.

PREPARATION.

Exp. 1.—To the hydrogen flask fit a washing bottle from which a delivery-tube leads to the trough. Into this flask pour about 40 c.c. of sulphuric acid, add to it slowly and with frequent stirring 20 c.c. of alcohol, and enough clean sand to convert the mixture into a thick mud. The sand is merely added to prevent frothing and if the sulphuric acid is at all pure it may be dispensed with. Now gently heat and the gas is soon evolved and may be collected over water in the usual way. The sulphuric acid abstracts the elements of water from the alcohol, C_2H_6O; thus:—

$$\underset{\text{Alcohol.}}{C_2H_6O} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Olefiant gas.}}{C_2H_4} + \underset{\text{Dilute sulphuric acid.}}{H_2SO_4,\ H_2O}$$

PROPERTIES.

230. Combustibility.

Exp. 2.—Apply a light to a bottle of the gas standing on the table; the gas burns freely with a strongly luminous flame. After the combustion is over add a little lime-water to the bottle and shake it briskly, the turbidity of the liquid indicates the presence of carbon dioxide. The presence of hydrogen can be shown in the same way as in the case of marsh-gas.

231. Explosive Mixture with Oxygen.

Exp. 3.—Fill a soda-water bottle one-fourth full of the gas and add oxygen till the bottle is full; cork, shake, and surround it with a cloth. Withdraw the cork, and apply a flame to the mouth of the bottle; a loud explosion occurs. The reaction is—

$$\underset{\text{Olefiant gas.}}{C_2H_4} + \underset{\text{Oxygen.}}{3O_2} = \underset{\text{Carbon dioxide.}}{2CO_2} + \underset{\text{Water.}}{2H_2O}$$

232. Union with Chlorine.

Exp. 4.—Invert over a bottle of the gas standing on the table a bottle of chlorine gas. After a short time the color of the chlorine almost disappears, and minute *oily* drops collect at the bottom of the lower jar. Hence the name *olefiant* or *oil-producing* gas.

SECTION III.

COAL-GAS.

PREPARATION.

Exp. 1.—To prepare a small quantity of coal-gas, take a common clay tobacco-pipe, nearly fill the bowl with fragments of bituminous coal, cover the latter with a layer of moistened clay, well pressed down by a cent-piece secured by a piece of wire. Heat the bowl strongly; a gas is given off and passes through the stem, at the end of which it can be lighted.

On a large scale, the coal is distilled in iron retorts. The vapors pass first into the *hydraulic main*, from this the gases pass up and down through a series of vertical pipes called *condensers*. In the hydraulic main and condensers the gas is cooled and the greater part of the tar and ammoniacal liquor are condensed and separated. It then passes through a box filled with perforated trays containing slaked lime. This is called the lime purifier. In passing through the lime the gas loses its carbon dioxide and sulphuretted hydrogen.

233. Composition of Coal-Gas.

Coal-gas thus purified is a mixture of combustible gases and vapors which may be divided into two classes, viz.: (1) Those which burn with a luminous flame, and are termed the *illuminating constituents;* (2) Those which burn with a non-luminous flame, and are termed *diluents.*

The most important of the first class are olefiant gas, and other heavy hydrocarbons. To the second class belong hydrogen, carbon monoxide, and marsh-gas.

234. Water-Gas.

The following process of making illuminating gas has in recent years been successfully carried out:—

Steam from a boiler is carried into a furnace and heated to a high temperature. This "superheated steam" is driven into a retort containing red-hot anthracite coal. The following reaction occurs:—

$$\underset{\text{Carbon.}}{C} + \underset{\text{Steam.}}{H_2O} = \underset{\text{Carbon monoxide.}}{CO} + \underset{\text{Hydrogen.}}{H_2}$$

The carbon monoxide and hydrogen pass into another retort kept intensely hot, into which petroleum is introduced. The petroleum is at once vaporized, and its heavy hydrocarbons decomposed into lighter ones. These mixed with the carbon monoxide and hydrogen constitute the illuminating gas. Great care should be taken to prevent the escape of this gas into a room, as the large quantity of carbon monoxide soon renders it fatal.

SECTION IV.

COMBUSTION.

235. Combustibles and Supporters of Combustion.

We have already defined combustion to be "the chemical combination of two or more bodies accompanied by the development of light and heat." We have also seen that a great many substances, such as phosphorus, sulphur, carbon, etc., burn readily in oxygen. In these cases oxygen is said to be a *supporter of combustion*, while the bodies themselves are said to be *combustible.* But in reality there is no difference in the part played by the two bodies. The chemical change is mutual, and the terms are merely relative, as the following experiments will show :—

Exp. 1.—Place the materials for making oxygen in a test-tube and connect with it a glass tube drawn out to a rather wide jet. Fill a large bottle with hydrogen, and hold it mouth downwards. Now heat the tube gently, and when the oxygen is coming off freely apply a light to the mouth of the bottle containing the hydrogen and thrust the jet into it. The flame will follow the end of the jet in its upward progress into the jar. In this case oxygen is the *combustible* and hydrogen the *supporter* of combustion.

Exp. 2.—Take an alcohol lamp, and by spreading its wick make a large flame. The centre of this flame is dark, and is filled with the vapor of alcohol. Now send a gentle current of oxygen through the jet, and slowly push it into the edge of the dark centre of the alcohol flame. The jet of oxygen will take fire while entering the flame and continue to burn with a very distinct and pretty light. In this little flame the oxygen is the combustible and alcohol the supporter ; while in the large flame around it the alcohol is the combustible and the oxygen of the air the supporter. The fact is, both are equally combustible.

236.—Heat of Combustion.

It has been shown by numerous and careful experiments that when the same weight of the same substance burns to form the same products of combustion a constant amount of heat is invariably evolved, whether the combination takes place slowly or quickly. Thus 12 parts by weight of carbon unite with 32 parts by weight of oxygen to form 44 parts by weight of carbon dioxide, and in this act of union the quantity of heat emitted is sufficient to raise 96960 parts of water from 0°C. to 1°C., and this is the same whether the combination is accomplished in a year or in a few seconds.

237.—The Composition of Fuel.

All fuel is of vegetable origin. Wood and charcoal are evidently so; and all forms of coal are the remains of ancient forests which, under the influence of heat and pressure, have parted with the greater part of their oxygen and hydrogen, while the carbon still remains. In all ordinary fuels carbon and hydrogen are the chief constituents.

238. Combustion in a Stove or Grate.

When air enters through the grate of a stove which is filled with coal, its oxygen combines with the glowing carbon forming carbon dioxide, because the oxygen is in excess. As this carbon dioxide passes upwards it reaches a point where carbon is in excess, and then the carbon dioxide is reduced to carbon monoxide which passes upwards and issues from the top where, if air is present, it combines with its oxygen, forming carbon dioxide. If air is not present it passes up the stove-pipe, and is so much fuel wasted. The combustion in a grate is the same as in a stove except that air is always present to combine with the carbon monoxide as it escapes from the top of the grate. This combination only takes place when the coals at the top of the grate are sufficiently hot to raise the gas to the temperature at which it can combine with the oxygen. The hydrogen of the coal combines with the oxygen of the air, forming steam, which is afterwards decomposed, giving a mixture of hydrogen and carbon monoxide or carbon dioxide according to the temperature.

239. Nature of Flame.

Flame is simple gas in a state of incandescence. When a solid combustible like charcoal burns in the air it produces no flame but simply glows. *Only gases burn with a flame.* In those cases where a solid or liquid seems to burn with a flame, it is either volatilized or decomposed by the heat of the combustion and thus converted into a gas before it burns. We will begin the study of the phenomena of flame by considering the flame of hydrogen flowing from a jet. Just over the opening there is a round column of pure hydrogen. This gas being lighter than air, and being forced out under some pressure, rises and mixes with air. There is, therefore, immediately around the jet a layer of mixed hydrogen and air, which is inflammable. The hydrogen unites with the oxygen of the air, forming steam which is carried away by the current of hot gases. More hydrogen is continually supplied from the jet, and more oxygen from the atmosphere. Thus there is a continuous formation, as fast as it is burned, of this inflammable mixture of hydrogen and oxygen around the central column of hydrogen. Evidently then the flame must be hollow. That this is the case may be shown as follows :—

Exp. 3.—Let the hydrogen issue from the end of a wide jet. Thrust a pin through a lucifer match, about a centimetre from the end, and place it top upwards in the centre of the jet. Apply a flame to the hydrogen some distance above the jet; the hydrogen will burn, but the match will not ignite.

Exp. 4.—Place a few grains of gunpowder on a cork in a saucer which contains a little alcohol, and ignite the alcohol. The gunpowder will not explode till the alcohol is nearly all burnt.

The flame cannot spread inwards, because there is nothing there to support combustion; nor can it spread outwards, because there is nothing there to burn. It is simply that part of the current of gases where the chemical action takes place; and where, consequently, the heat is produced.

240. Luminosity of Flame.

The brightness or illuminating power of a flame depends on the following circumstances :—

(1) *On the temperature of the Incandescent Gases.*

Exp. 5.—Burn a small piece of sulphur, first in air and then in oxygen; the sulphur dioxide produced remains permanently gaseous. When the sulphur burns in the air the illuminating effect is far less than when it burns in oxygen. This is due to the lowering of the temperature by the nitrogen which not only adds nothing to the heat produced by the combination, but on the contrary abstracts a considerable amount of that produced by the union of the sulphur with the oxygen. Burn a small piece of phosphorus, first in air and then in oxygen; the phosphorus pentoxide, the product of the combustion, is gaseous at a temperature far below that of combustion. The increased luminosity is not, therefore, due to solid particles in the flame, but must be due to the increased temperature. Also, under ordinary circumstances, phosphorus burns in chlorine with feeble light, but strongly heated chlorine and phosphorus vapor combine with a dazzling white flame. From these and other experiments, it is inferred that *the luminosity of flame is greater the higher the temperature:* This, however, holds good only with the same kind of flame.

(2) *On the density of the Incandescent Gases.*

Exp. 6.—Burn a jet of hydrogen in a bottle of chlorine gas; it burns with a bright flame, producing hydrochloric acid, HCl, the density of which is 18·2. Now we know that hydrogen burns in oxygen with a very feeble light, forming steam, the density of which is 9. The density of hydrochloric acid is, therefore, twice that of steam.

Alcohol burns with a feebly illuminating flame under ordinary conditions, but its flame becomes very bright when burnt under the receiver of a condensing pump. Frankland has shown that when oxygen and hydrogen are burnt under a pressure of 20 atmospheres they give rise to a luminous flame. Phosphorus and arsenic burn with very luminous flames in consequence of the formation of very dense vapors of phosphorus pentoxide and arsenic trioxide during the combustion. The luminosity of an ordinary flame is materially affected by the pressure of the atmosphere in which it burns, a diminution of pressure causing a loss of illuminating power. The light emitted by a burning candle at the top of a mountain is much less than that given at the bottom. If the light of a given flame, burning in the air when the barometer stands at 30 inches, be represented by 100, each diminution of one inch in the height of the barometer will reduce the luminosity by five. This is not due to any difference in the rate of burning, for the amount of fatty matter, wax, etc., consumed remains constant. It, therefore, appears that a certain relation exists between the luminosity of a flame free from solid particles and the density of its constituents; *the brightest flames contain the densest vapors.*

(3) *On the presence of solid particles within the area of Combustion.*

Exp. 7.—Burn hydrogen from a wide jet, and gently blow into the flame a small quantity of finely divided charcoal, iron filings, or zinc oxide; the flame becomes luminous. Therefore, *a feebly illuminating flame increases in luminosity by injecting solid particles into it.*

On the authority of Sir Humphrey Davy, it was long considered that the luminosity of a flame depended entirely on the presence within it of solid incandescent particles. Now, although there is no doubt that the presence of solid particles may bring about the luminosity of a flame, there are certainly flames which emit an intense light, and yet do not contain any solid particles. Arsenic burns in oxygen with a bright light, and without the immediate formation of a solid substance. No solid matter is separated when a mixture of nitric oxide and carbon disulphide burns. The blinding light of burning phosphorus in oxygen was formerly explained by supposing that the phosphorus flame

itoxide and
iosity of an
sure of the
sure causing
y a burning
hat given at
ig in the air
ted by 100,
rometer will
iy difference
er, wax, etc.,
iat a certain
e from solid
ghtest flames

rea of Com-

gently blow
harcoal, iron
Therefore,
by injecting

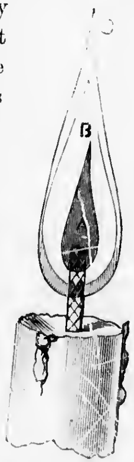
as long con-
tirely on the
ow, although
s may bring
lames which
lid particles.
without the
lid matter is
i disulphide
s in oxygen
phorus flame

---

contained solid particles of phosphorus pentoxide. But this is manifestly impossible; for phosphorus pentoxide is volatile at a temperature far below that of the combustion.

241. The Flame of a Candle.

Exp. 8.—Press down a piece of white note-paper for an instant upon the flame of a candle until it almost touches the wick, then quickly remove it before it takes fire, and observe that it is charred and has received a deposit of soot in the form of a ring. Hold a splint of wood across the flame; it becomes ignited only at the two points where it cuts the outer portions of the flame. The inference is that the dark part consists of unburnt gas. Thrust a straight and rather narrow glass tube into this zone, and hold it nearly vertical; mixed gas and vapor ascend the tube, and may be lighted at the top. *The candle flame, like the hydrogen flame, is hollow.* When the whole flame of the candle, or the flame of any other highly carbonaceous substance, is carefully examined it is seen to consist of four distinct parts: (1) The dark inner cone surrounding the wick. This is made up of the gases or vapors distilled from the wick, together with the products of their partial combustion, and a considerable volume of atmospheric nitrogen (2) Below this a light blue cone, of small area, consisting of the combustible matter of the wick, which becomes mixed with a sufficient amount of oxygen to burn it completely. (3) Above this is a bright, luminous cone, in which the products of the distillation from the wick are still further broken up with the production of dense hydrocarbons, and also with the separation of carbon, which becomes intensely heated and emits a bright light. (4) Surrounding the entire flame is a feebly luminous mantle, which may be easily detected



by burning a little bit of sodium in its vicinity, when it is colored strongly yellow. This cone consists of the final products of combustion of the constituents of the luminous zone mixed with atmospheric air heated to incandescence.

The flame of alcohol appears to be differently constituted from that of a candle; but it has the same number of zones, although their relative areas are very different. The internal cone is very large, in consequence of the great volatility of the alcohol; the luminous portion is very small owing, probably, to the simple manner in which the alcohol is broken up and consumed. From the feebly illuminating power of this cone the mantle appears largely developed.

Frankland considers the luminosity of flame to be due entirely to the dense intensely heated hydrocarbon vapors. Recent researches rather tend to support Davy's hypothesis. Heated surfaces as well as cold ones become covered with soot when held in an ordinary flame, so that the deposit of soot does not result merely from the condensation of dense hydrocarbon vapors. If a rod is held in a luminous gas flame soot is deposited only on that surface against which the issuing gas impinges, and not uniformly on the rod, as it should be if the soot were the vapor of heavy hydrocarbon condensed by its cooling action. The most conclusive evidence for the existence of solid particles in a flame is afforded by the fact that although luminous flames which are free from solid particles cast no shadows on a white screen when exposed to sunlight, ordinary luminous hydrocarbon flames, under similar circumstances, cast distinct shadows in the same way that these flames do which undoubtedly owe their luminosity to solid particles, showing that they too contain solid matters. The luminosity of ordinary flames used for illuminating purposes is therefore probably due not only to the ignition of very dense hydrocarbon vapors, but also to the presence of solid intensely heated particles of carbon.

242. The Bunsen Burner.

If air is let into the interior of a flame of coal-gas, the flame becomes non-luminous, and does not deposit any soot. This principle is seen in the Bunsen burner, which consists of a jet surmounted by a wide brass tube, at the bottom of which are several holes for the admission of air. The air and gas mix in the wide tube in such proportion that they burn at the top with a non-luminous flame. This burner is now universally employed in laboratories for heating purposes, as it gives a high temperature with a smokeless flame. Nearly all gas-stoves, and other arrangements for heating by gas, are essentially Bunsen burners, more or less modified to suit particular purposes. On closing the holes, the gas burns unmixed with air, with the ordinary smoky flame. These facts are usually explained by assuming that the oxygen thus brought into the interior of the flame combines at once with the carbon. This explanation is incomplete if not incorrect. If, instead of mixing air with the coal-gas, such indifferent gases as nitrogen, carbon dioxide, or hydrochloric acid are employed, the flame likewise becomes non-luminous, whilst inflammable gases, such as hydrogen or carbon monoxide, produce the same effect. If the mixed gases are first passed through a red-hot platinum tube, the flame again becomes luminous. It appears, therefore, that a luminous flame may be rendered non-luminous in three ways: (1) by cooling; (2) dilution; (3) by the too rapid oxidation of the separated carbon or carbonaceous vapors.

243. Temperature of Ignition.

In order that any combustible substance may burn, it must first be heated to a certain temperature, and kept at that heat. *The temperature at which any substance takes fire is called the temperature of ignition, or the kindling-temperature of that substance.* This temperature varies widely for different substances. Phosphorus takes fire at about 60°C., whilst a jet of coal-gas cannot be lighted with a piece of red hot iron. Whenever bodies are cooled below the temperature of ignition they go out.

Exp. 9.—Make a small helix of copper wire, and place it over the flame of a burning candle; the flame will be immediately extinguished. The metal so rapidly abstracts the heat from the gases in the flame that their temperature is reduced below their kindling point. If, however, the helix is previously heated to redness, the flame will not be extinguished. Take also a piece of fine wire-gauze, and press it down over the flame of the candle; no flame is seen above the gauze, but instead of flame a cloud of smoke. Hold the gauze over a gas jet, light the gas above it, and gently raise the gauze; it may be removed several inches, and yet the inflammable gas below does not take fire, the flame burning only above the gauze.

244. The Davy Lamp.

The Davy Lamp is an important application of the principle in the preceding experiment. It is simply an ordinary oil lamp surrounded by a cylinder of wire-gauze. If this lamp is introduced into an explosive atmosphere of fire-damp and air, combustion occurs within the lamp, but the flame produced by the union of the gases is prevented from traversing the gauze, owing to the rapidity with which the burning gases are cooled down in contact with the metal. If, however, the gauze should become heated, its action ceases; it allows of the transmission of the flame, and is powerless to prevent an explosion.

245. The Blow Pipe.

The blow-pipe is an instrument which serves to direct a stream of air into a candle or coal-gas flame. In its simplest form it is a tube drawn out to a jet and bent at right angles near one extremity. The introduction of the air within the flame at once destroys its luminosity by effecting a more rapid combustion of the carbonaceous matter.

Exp. 10.—Insert the jet a short distance into the candle-flame just above the wick, and inject a steady blast of air from the mouth; the flame is deflected, its luminosity diminished, and its heating power greatly increased. The three principal zones of the candle-flame are still recognized, but the outer one is much extended, the middle one diminished, and the inner one, extending some distance from the nozzle of the blow-pipe, contains air mixed with other gases. Hold a piece of lead-glass tubing for a short time in the middle zone; it becomes blackened owing to the reduction of the lead to the metallic state. This portion of the flame, therefore, is called the *reducing flame.* Now hold the discolored glass at a short distance beyond the visible point of the flame; the dark color disappears. This effect is due to the oxidation of the lead. This portion is called the *oxidizing flame.* The blow-pipe is an exceedingly useful instrument in the detection of bodies that undergo marked chemical changes under its influences.

QUESTIONS AND EXERCISES ON CHAPTER VII.

1. What are the natural and artificial sources of marsh gas.

2. How much oxygen do you require for the complete combustion of 10 c.c. of marsh gas? How many cubic centimetres of each product of the combustion would be formed?

3. Mention the distinguishing characters of olefiant gas.

4. How many cubic centimetres of oxygen gas are required for the complete combustion of 200 c.c. of olefiant gas? What products are formed, and how many cubic centimetres of each?

5. Explain the structure of flame, and describe how you would support your explanation by experiments.

6. Describe the structure of the flame of a common candle.

7. A current of air passes through a bright coal fire. What gases are contained in the air beyond the fire?

8. What are the products of the combustion of coal gas, and why does it burn, after a large quantity of air is mixed with it, with little or no light?

9. Why does the flame of an ordinary lamp appear brighter with than without a chimney glass ?

10. Explain the nature of the chemical changes effected by the inner and outer flames of the blow-pipe.

CHAPTER IX.

SECTION I.

CHLORINE.

Symbol, Cl. Atomic Weight, 35·5 (35·36). Molecular, Wt. Cl_2, 71.

PREPARATION.

246. From Hydrochloric Acid and Manganese Dioxide.

Exp. 1.—Into the hydrogen flask introduce twenty grams of manganese dioxide, and after fitting the apparatus together pour through the funnel-tube about 60 c. c. of strong hydrochloric acid, and heat the mixture very gently. Since chlorine is 35·5 times heavier than hydrogen, or nearly 2½ times heavier than air, it may be collected by downward displacement, the rate of filling being readily observed owing to the color of the gas. The reaction is expressed by the following equation :—

$$\underset{\text{Manganese dioxide.}}{MnO_2} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Manganous chloride.}}{MnCl_2} + \underset{\text{Water.}}{2H_2O} + \underset{\text{Chlorine.}}{Cl_2}$$

This reaction appears to take place in two stages—the first consists in the formation of manganese tetrachloride ; thus :—

$$\underset{\text{Manganese dioxide.}}{MnO_2} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Water.}}{2H_2O} + \underset{\text{Manganese tetrachloride.}}{MnCl_4} ;$$

but this compound is so unstable that it breaks up into manganous chloride and chlorine :—

$$\underset{\text{Manganese tetrachloride.}}{MnCl_4} = \underset{\text{Manganous chloride.}}{MnCl_2} + \underset{\text{Chlorine.}}{Cl_2}$$

247. From Sodium Chloride, Manganese Dioxide and Sulphuric Acid.

Exp. 2.—Mix 30 grams of sodium chloride with the same amount of manganese dioxide, and transfer the mixture to the flask used in the last experiment. Now mix 38 grams of water with 60 grams of strong sulphuric acid, allow the mixture to cool, and pour it through the funnel-tube into the flask.

In this case hydrochloric acid is first produced by the action of the sulphuric acid on the salt:—

$$\underset{\text{Sodium chloride.}}{NaCl} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Hydrochloric acid.}}{HCl} + \underset{\text{Sodium hydrogen sulphate.}}{NaHSO_4}$$

The hydrochloric acid then enters into reaction with the manganese dioxide, as before, but the manganous chloride is converted into manganous sulphate and hydrochloric acid by the sulphuric acid. The following equation expresses the final result but does not take into account that there are several stages in the reaction:—

$$\underset{\text{Sodium chloride.}}{2NaCl} + \underset{\text{Sulphuric acid.}}{3H_2SO_4} + \underset{\text{Manganese dioxide.}}{MnO_2}$$
$$= \underset{\text{Chlorine.}}{Cl_2} + \underset{\text{Water.}}{2H_2O} + \underset{\text{Manganous sulphate.}}{MnSO_4} + \underset{\text{Sodium hydrogen sulphate.}}{2NaHSO_4}$$

248. Precautions.

(1) Its action on the lungs is extremely irritating and injurious; therefore, the greatest care must be taken to avoid inhaling it.

(2) It should be prepared in a cupboard communicating with a flue, or out of doors; or, if it must be prepared in a room, collect it over warm water and not by displacement, covering every bottle when full with a well-greased glass plate.

(3) As soon as the gas is collected, throw away the water in the trough.

(4) Should a little of the gas get into the lungs, pour a little alcohol into a test-tube, warm it, and inhale the vapor. A little ammonia sprinkled on a warm plate will get rid of the gas in the room.

PROPERTIES.

249. Solubility in Water.

Exp. 3.—Invert a bottle of chlorine in cold water; a little of the water will enter the bottle. Close the mouth of the bottle with the hand, and shake up the gas and water; the hand is drawn in, showing that absorption has taken place, and on removing the hand under water the latter rises in the jar. *Chlorine is, therefore, moderately soluble in cold water.*

250. Is not combustible but supports the Combustion of bodies rich in Hydrogen.

Exp. 4.—Plunge a burning wax-taper into a small bottle of chlorine; the taper burns with a dull smoky flame, depositing abundance of soot, but the gas itself does not take fire. Place a piece of moist blue litmus-paper in the bottle; at the first moment it is reddened, and then rapidly bleached. An acid, therefore, must be one of the products of the combustion. Wax is a compound of carbon, hydrogen, and a little oxygen. The carbon does not unite with the chlorine, and is deposited in a finely divided state on the bottle. The hydrogen unites with the chlorine, forming hydrochloric acid.

Exp. 5.—Take a piece of glass tubing drawn out to a jet, bend it near the end so as to bring the jet parallel to the tube, and connect it by a piece of rubber-tubing with a bottle evolving hydrogen. When the air has escaped, light the jet and plunge it into a bottle of chlorine; the hydrogen continues to burn with a somewhat luminous flame, giving off white fumes of hydrochloric acid.

Chlorine, therefore, must have a great affinity for hydrogen. Indeed this is one of its most characteristic properties, as will appear from the following experiments.

Exp. 6.—Take a strong wide-mouthed bottle of chlorine and another of the same size of hydrogen, bring them mouth to mouth, and keeping them close together invert them several times so as to thoroughly mix them and cover with glass plates. Remove the cover from one of the jars and apply a flame; an explosion results, and acid fumes of hydrochloric acid are produced.

Close the second bottle with a well-greased ground glass plate, and bring it into diffused daylight; the yellow color entirely disappears, and the bottle will be found full of hydrochloric acid gas. If the bottle had been exposed to direct sunlight almost instant combustion would have taken place with explosion.

Exp. 7.—Moisten a strip of blotting paper with warm turpentine and plunge it into a bottle of chlorine; a cloud of black smoke is at once produced, accompanied by a momentary flame. The turpentine is composed of carbon and hydrogen; the chlorine combines with the hydrogen, setting the carbon free; thus:—

$$\underset{\text{Turpentine.}}{C_{10}H_{16}} + \underset{\text{Chlorine}}{8Cl_2} = \underset{\text{Hydrochloric acid.}}{16HCl} + \underset{\text{Carbon.}}{10C}$$

Exp. 8.—Fill a long-necked flask with water and saturate it with chlorine. Place it mouth downwards in a bowl of water and expose it to direct sunlight; in a short time bubbles are seen rising, and a quantity of colorless gas collects above the liquid. Invert the flask and test the gas with a glowing splint of wood; the splint is rekindled and therefore the gas is oxygen. Add litmus solution to the liquid in the flask; the litmus is reddened. The chlorine has combined with the hydrogen of the water, forming hydrochloric acid, and setting the oxygen free.

251. Bleaching power of Chlorine.

Exp. 9.—Moisten a piece of madder-dyed cloth and put it into a bottle of chlorine, covering the mouth of the bottle with a well-greased glass plate; the red color gradually disappears.

Into another bottle of chlorine pour a little strong sulphuric acid, replace the plate and shake well. The sulphuric acid will remove all moisture from the gas. Now dry a piece of the cloth used before and introduce it into the bottle; the color does not disappear.

Hence it appears that the characteristic bleaching action which chlorine exerts on organic coloring matter depends on the presence of moisture. The chlorine decomposes the water combining with its hydrogen and setting free the oxygen, which being in a nascent state readily oxidizes the coloring matter.

Exp. 10.—With ordinary writing ink write across a piece of newspaper and introduce it into a bottle of chlorine; after a short time the written characters disappear while the printed ones are unaffected. Writing ink consists of iron salts of of organic acids, while printers' ink contains carbon in the form of lamp-black. Indigo and logwood are readily bleached, but potassium permanganate and copper sulphate are unaffected by chlorine.

Hence as a rule *Mineral colors are not destroyed by chlorine, nor can it bleach tints produced by carbon.*

Chlorine cannot be used to bleach substances of animal origin, such as silk and wool on account of its destructive action on the tissue. Such substances are bleached by means of sulphurous acid.

252. Combustion of different substances in Chlorine.

Exp. 11.—Dry a small piece of phosphorus, place it in the deflagrating spoon and plunge it into a bottle of chlorine; the phosphorus takes fire and burns brilliantly, showing that chlorine has a great affinity for this substance.

Exp. 12.—Pour enough water into a bottle of chlorine to cover the bottom, place a little powdered metallic antimony in

4.5
5.0
2.8
3.2
2.5
3.6
4.0
2.2
2.0
1.8
1.0

a paper gutter and shake it slowly into the bottle; the metal catches fire and burns brilliantly. Hence antimony and chlorine have a great affinity for each other. Arsenic, copper, and several other metals will burn in a similar manner.

Exp. 13.—Pour a little sulphuretted hydrogen solution into a bottle, gradually add chlorine water, and shake it up; the unpleasant odor disappears, and the smell of chlorine is not detected. The reaction is—

$$\underset{\text{Sulphuretted hydrogen.}}{2H_2S} + \underset{\text{Chlorine.}}{2Cl_2} = \underset{\text{Hydrochloric acid.}}{4HCl} + \underset{\text{Sulphur.}}{S_2}$$

In this case chlorine acts as a *deodorant*, and is commonly used for removing offensive smells.

Chlorine is also a powerful *disinfectant.* It acts either by the direct corrosion of the disease germs, or by the indirect oxidation of the poisonous germs floating in the air. The most convenient source of chlorine for these purposes is *bleaching powder.*

253. Chloride of Lime or Bleaching Powder.

Exp. 14.—Put some moistened slaked lime into a wide-mouthed bottle, shake it well so as to line the inside of the bottle with the lime, and pass a current of chlorine into it; the gas combines with the lime, forming *chloride of lime;* thus:—

$$\underset{\text{Calcium hydrate.}}{2Ca(OH)_2} + \underset{\text{Chlorine.}}{2Cl_2} = \underset{\text{Calcium chloride.}}{CaCl_2} + \underset{\text{Calcium hypochlorite.}}{Ca(OCl)_2} + \underset{\text{Water.}}{2H_2O}$$

On the large scale the lime is spread in thin layers on shelves in large chambers, to which chlorine is admitted.

The article to be bleached is immersed in a clear solution of bleaching powder, and then in a very dilute hydrochloric or sulphuric acid, after which it must be well washed. The decomposition which takes place when hydrochloric acid is used is represented as follows:—

$$\underset{\text{Bleaching powder.}}{CaCl_2 + Ca(OCl)_2} + \underset{\text{Hydrochloric acid.}}{4HCl} = \underset{\text{Calcium chloride.}}{2CaCl_2} + \underset{\text{Water.}}{2H_2O} + \underset{\text{Chlorine.}}{2Cl_2}$$

When chloride of lime is to be used for the sick room, a quarter of a pound should be dissolved in about a gallon of water, and a piece of flannel dipped in the *clear* liquor should then be freely exposed to the air of the room, from which, by the action of the atmospheric carbonic acid, a small but regular evolution of chlorine takes place. Or the bleaching powder may be mixed with half its weight of powdered alum, when a pretty rapid and regular escape of chlorine will ensue.

254. Action of Chlorine on Potassium Hydrate.

Exp. 15.—Pass chlorine into a cold solution of potassium hydrate; when partially saturated with the gas stop the current. The solution is now colorless, and smells somewhat like bleaching powder. The reaction is—

$$\underset{\text{Chlorine.}}{Cl_2} + \underset{\text{Potassium hydrate.}}{2KOH} = \underset{\text{Potassium chloride.}}{KCl} + \underset{\text{Potassium hypochlorite.}}{KOCl} + \underset{\text{Water.}}{H_2O}$$

A few drops of acid added to this solution causes the evolution of chlorine. Potassium hypochlorite was the first compound of chlorine that was employed for bleaching purposes on the large scale. It was called Eau de Javelles from the name of the bleach works where it was prepared.

If the chlorine is passed through a solution of sodium carbonate, sodium hypochlorite is formed; thus:

$$\underset{\text{Sodium carbonate.}}{Na_2CO_3} + \underset{\text{Chlorine.}}{Cl_2} = \underset{\text{Sodium Chloride.}}{NaCl} + \underset{\text{Sodium hypochlorite.}}{NaOCl} + \underset{\text{Carbon dioxide}}{CO_2}$$

Sodium hypochlorite is very useful for removing ink, fruit stains, &c.

Exp. 16.—Pass a current of chlorine through a strong boiling solution of potassium hydrate; on cooling, crystalline plates will separate out. These are crystals of potassium chlorate. The reaction is

$$\underset{\text{Chlorine.}}{3Cl_2} + \underset{\text{Potassium hydrate.}}{6KOH} = \underset{\text{Potassium chlorate.}}{KClO_3} + \underset{\text{Potassium chloride.}}{5KCl} + \underset{\text{Water.}}{H_2O}$$

The potassium chloride is very soluble and therefore remains in solution while the slightly soluble chloride crystallizes out.

SUMMARY AND ADDITIONAL PROPERTIES.

255. History.—Discovered by Scheele, in 1774, when acting on manganese dioxide with hydrochloric acid. Davy, in 1810, proved its elementary character and gave it the name which it now bears (chloros, Gr. green). It was liquified by Faraday by the force of its own elasticity, being one of the first gases he so reduced.

256. Sources—Chlorine does not occur free in nature, but is found combined with potassium, magnesium, and especially with sodium, in sea-water and in the large beds of salt in various parts of the world. The alkaline chlorides occur in the bodies of plants and animals.

257. Properties.—Chlorine is a transparent, greenish-yellow gas with a suffocating, irritating odor. It is about $2\frac{1}{2}$ times heavier than air. A pressure of four atmospheres (60 lbs.) at 15°.5C. reduces it to a yellow limpid liquid, which remains unfrozen even at −140°C. Cold [illegible]er dissolves about twice its volume of chlorine; the solution has the color, odor, and the stringent taste of the gas. It is capable of uniting either directly or indirectly with all the elements with which we are acquainted. For hydrogen its affinity is especially great, but it will not unite directly with oxygen. Chlorine is a powerful bleaching, deodorizing, and disinfecting agent. The compounds of chlorine with metals are called chlorides.

258. Tests—Its odor, color, and bleaching properties constitute the chief tests for chlorine in an uncombined state. When free or combined with a metal *silver nitrate* gives a white, curdy precipitate of silver chloride, $AgCl$, soluble in ammonia, and insoluble in nitric acid.

SECTION II.

HYDROCHLORIC ACID.

Symbol, HCl. Molecular Weight, 36·5.

This gas has already been prepared synthetically by exploding a mixture of equal volumes of hydrogen and chlorine; thus:—

$$\underset{\text{Hydrogen.}}{H_2} + \underset{\text{Chlorine.}}{Cl_2} = \underset{\text{Hydrochloric acid.}}{2HCl}$$

Hydrochloric acid acts energetically on a great many metals, the metals replacing the hydrogen of the acid, forming salts called *Chlorides.*

EXAMPLES OF COMMON CHLORIDES.

SYSTEMATIC NAME.	COMMON NAME.	FORMULA.
Sodium Chloride	Common salt, rock salt......	$NaCl$
Potassium "	Stassfürth salt	KCl
Silver "	Horn-silver...............	$AgCl$
Ammonium "	Sal-ammoniac.............	NH_4Cl
Calcium "	....	$CaCl_2$
Zinc "	Burnett's disinfecting fluid ..	$ZnCl_2$
Mercurous "	Calomel..................	Hg_2Cl_2
Mercuric "	Corrosive sublimate	$HgCl_2$

259. Properties of Chlorides.

(1) All chlorides are soluble in water except silver chloride, $AgCl$, mercurous chloride, Hg_2Cl_2, and a few others.

(2) When strong sulphuric acid, H_2SO_4, is added to a solid chloride or to a strong solution of liquid chloride, acid fumes of hydrochloric acid are evolved.

(3) Silver nitrate added to a solution of a chloride gives a white, curdy precipitate insoluble in nitric acid, but soluble in ammonia.

PREPARATION.

260. From Sodium Chloride and Sulphuric Acid.

Exp. 1.—Into the hydrogen flask put about 20 grams of common salt, and pour upon it 30 c.c. of sulphuric acid, diluted with half this quantity of water and previously cooled, and apply a gentle heat. Collect by downward displacement, testing when the bottle is full by a lighted taper as in the preparation of carbon dioxide. Fill four bottles and take away the flask when the bottles are full as the gas is very irritating to the lungs. The reaction is expressed by the following equation:—

$$\underset{\text{Sodium chloride.}}{NaCl} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Hydrochloric acid.}}{HCl} + \underset{\text{Hydrogen sodium sulphate.}}{NaHSO_4}$$

An ounce or 30 grams of salt yields 700 cubic inches (12 litres) of the gas.

PROPERTIES.

The incombustibility of the gas, its inability to support the combustion of a taper, its weight, and the fumes which it produces in moist air will have been already noticed during its preparation.

261. Acidity.

Exp. 2.—Plunge a piece of blue litmus-paper into a bottle of the gas; it is at once reddened, showing the presence of an acid.

262. Solubility in Water.

Exp. 3.—Bring a bottle of hydrochloric acid, inverted, beneath the pneumatic trough, withdraw the plate and shake the bottle; the water rapidly rises. Or, as in ammonia, Exp. 7, only using blue litmus-solution, which will become red in the flask. 1 c. c. of water at 15° C. dissolves 450 c. c. of the gas at the same temperature. Ordinary hydrochloric acid consists of a solution of the gas in water.

Exp. 4.—Heat a small piece of potassium or sodium in the deflagrating spoon and plunge it into a bottle of dry hydrochloric acid gas; the potassium burns, leaving only hydrogen, which is found to occupy half the original volume; thus:—

$$\underset{\text{Potassium.}}{K_2} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Potassium chloride.}}{2KCl} + \underset{\text{Hydrogen.}}{H_2}$$

This very important acid has received the following names: spirits of salt, muriatic acid, chlorohydric acid, and hydric chloride.

263. **Tests.**—As stated under chlorine, a solution containing even a mere trace of the acid, or of a chloride, gives on the addition of a drop of silver nitrate a white curdy precipitate of silver chloride.

264. **Uses.**—Hydrochloric acid is principally employed in the manufacture of bleaching powder, and in the preparation of sal-ammoniac. It is used as a solvent for tin by the dyer and calico-printer.

265. Aqua Regia.

Exp. 5.—Take two test-tubes, pour a little nitric acid into one and a little hydrochloric acid into the other. Put a few pieces of gold-leaf into each; it is not dissolved. Now pour the acid in one tube into the acid in the other tube; the gold disappears almost immediately. The mixture of the acids is called *aqua regia* because it alone dissolves gold or platinum, which are called royal or noble metals. The reaction is—

$$\underset{\text{Hydrochloric acid.}}{3HCl} + \underset{\text{Nitric acid.}}{HNO_3} = \underset{\text{Chlorine.}}{Cl_2} + \underset{\text{Nisrosyl chloride.}}{NOCl} + \underset{\text{Water.}}{2H_2O}$$

The solution of the gold is due to the action of the chlorine on the metal:—

$$\underset{\text{Gold.}}{2Au} \underset{+}{+} \underset{\text{Chlorine.}}{3Cl_2} \underset{=}{=} \underset{\text{Auric chloride.}}{2AuCl_3}$$

COMPOUNDS OF OXYGEN AND CHLORINE.

OXIDES OF CHLORINE.	CORRESPONDING ACID.
Cl_2O.. Hypochlorous anhydride	Hypochlorous acid..... $HClO$.
Cl_2O_3.. Chlorous anhydride.	Chlorous " ..... $HClO_2$.
ClO_2.. Chlorine peroxide.	
Cl_2O_5.. *Chloric anhydride?*	Chloric " ..... $HClO_3$.
Cl_2O_7.. *Perchloric anhydride.*	Perchloric " ..... $HClO_4$.

Chlorine and oxygen have but slight affinity for each other, hence their compounds when formed are unstable, and many of them highly dangerous. The two anhydrides printed in italics have not yet been prepared, though the corresponding acids are well known. All the acids contain but one atom of hydrogen replaceable by a metal, and are, therefore, monobasic. Their formulæ may be easily remembered by regarding them as successive oxides of hydrochloric acid. They afford a remarkable illustration of the *Law of Multiple Proportions.*

SECTION III.

BROMINE.

Symbol, Br. Atomic Weight, 80 (79·75). Molecular Weight, Br_2, 160.

PREPARATION.

266. From Potassium Bromide, Sulphuric Acid and Manganese Dioxide.

Exp. 1.—Dissolve in water, in a test-tube, a few crystals of potassium bromide and add a few drops of chlorine water to the liquid; the solution changes to a reddish brown. Boil the liquid; the red vapors are given off, and they have a most disagreeable smell. They are the vapors of an element called *bromine* (*Bromos*, Gr., stench) on account of its smell. The reaction is—

$$\underset{\text{Potassium bromide.}}{2KBr} + \underset{\text{Chloride.}}{Cl_2} = \underset{\text{Potassium chloride.}}{2KCl} + \underset{\text{Bromine.}}{Br_2}$$

Bromine may be prepared from potassium bromide, just as chlorine was prepared from sodium chloride; thus:—

$$\underset{\text{Potassium bromide}}{2KBr} + \underset{\text{Sulphuric acid.}}{3H_2SO_4} + \underset{\text{Manganese dioxide.}}{MnO_2} = \underset{\text{Bromine.}}{Br_2} + \underset{\text{Hydrogen potassium sulphate.}}{2KHSO_4} + \underset{\text{Manganese dioxide.}}{MnO_2} + \underset{\text{Water.}}{2H_2O}$$

Dark red fumes are liberated, and a black liquid condenses in the receiver. Bromine is the only non-metallic element known which is liquid at ordinary temperatures. Free bromine gives a yellow tint to starch. Its compounds are used in medicine and photography.

SECTION IV.

IODINE.

Symbol, I. Atomic Weight, 127 (126·53). Molecular Weight, I_2*, 254.*

PREPARATION.

267. From Potassium Iodide, Sulphuric Acid, and Manganese Dioxide.

Exp. 1.—Dissolve a few crystals of potassium iodide, KI, in water in a test-tube, and add a little chlorine water; a brown color is immediately produced, and black particles separate from the liquid. Allow the particles to subside, and pour off the liquid. Apply a gentle heat to the black substance; a violet colored vapor is produced, which condenses in the cool part of the tube in black, shining, metallic-looking scales. This metallic-looking substance is an element called *iodine.*

Iodine is widely distributed through nature, but in small quantities, and always in combination, chiefly with potassium, sodium, and magnesium. It occurs in small quantities in sea-water, from which it is separated and assimilated by various sea-weeds. These sea-weeds, when dried and burned, afford an ash called *kelp*, which contains iodides of sodium, NaI, magnesium, MgI_2, etc. The kelp, after preliminary treatment, is heated with manganese dioxide and sulphuric acid, and yields iodine just as potassium bromide yielded bromine. The reactions are similar.

Iodine acts as an irritant poison, but in small quantities is used in medicine to reduce glandular swellings, etc.

268. Tests.—Dissolve a few grains of potassium iodide in two or three litres of water and a little clear solution of starch paste. To this add a little free chlorine or a few drops of chlorine water, to set free the iodine, stir it, and a beautiful blue tint will be observed. The blue tint will disappear on heating, but will return on cooling.

SECTION V

FLUORINE.

Symbol, F. Atomic Weight, 19 (19·1).

This element has, perhaps, never been obtained, and certainly has not been satisfactorily examined, in the free state; for as soon as it is liberated it attacks and combines with glass and almost every substance with which we are acquainted. It is the only element that has not been combined with oxygen.

It commonly occurs combined with calcium, forming the mineral fluor-spar or calcium fluoride, CaF_2. It is also found in the enamel of the teeth, in the bones of mammalia, and is said to have been detected in the blood and in milk.

269. Hydrofluoric Acid, HF.

Exp. 1.—Make a small cup out of a piece of thin sheet-lead, place in it about 10 grams fluor-spar, add 15 grams of strong sulphuric acid, and apply a gentle heat; fumes of hydrofluoric acid will rise, which are extremely irritating to the eyes. The reaction is—

$$\underset{\text{Calcium fluoride.}}{CaF_2} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Calcium sulphate.}}{CaSO_4} + \underset{\text{Hydrofluoric acid.}}{2HF}$$

This acid is intensely corrosive, attacking glass, earthenware, and most other substances, but does not act upon lead.

Exp. 2.—Warm a small square of glass, rub one side with a piece of wax, and when cold scratch any design upon it, taking care to cut quite through to the glass. Lay the plate upon the leaden cup, waxed side downwards, leave it for a few minutes, remove and clean with a rag dipped in a little turpentine; it will be found corroded wherever the marks have been made, and the design will remain engraved on the glass. Hydrofluoric acid is employed in this way to etch the graduations on thermometers, eudiometers, etc.

270. Comparison of the Group.

These four elements closely resemble each other in many points. In nature they are never found in a free state. Where one is present all are present. They are all monad elements, and the single compound which each one forms with hydrogen is an acid. Their nitrogen compounds are all explosive. They exhibit a remarkable sequence of properties in form and color, in specific gravity, and in atomic weights.

QUESTIONS AND EXERCISES ON CHAPTER IX.

1. Express in symbols the reaction which occurs when hydrochloric acid is heated with manganese dioxide. Give the relative proportions by weight of the elements described by your symbols.

2. Describe (*a*) the physical, (*b*) the chemical properties of chlorine.

3. Chlorine often acts as an oxidizing agent. How do you account for this? Describe some experiments in proof of your explanation.

4. Explain the action of chlorine (*a*) upon slaked lime, (*b*) on caustic potash.

5. What is chloride of lime? Explain how it is prepared, and how it is applied in bleaching cotton fabrics. Would you expect stains of lamp-black or of nitrate of silver to be removed by these means? Why so?

6. Chlorine will bleach vegetable colors. How is this action of chlorine usually explained?

7. Describe the effects produced by chlorine on the following substances:—Indigo, copper foil, powdered antimony, powdered charcoal.

8. State the useful applications of chlorine in the arts.

9. How much salt is required to make 28 litres of chlorine?

10. Sketch the apparatus required for preparing small quantities of hydrochloric acid.

11. How may hydrochloric acid be prepared from its elements? Describe different ways of effecting the union.

12. What volume is occupied by 78 grams of gaseous hydrochloric acid?

13. How many grams of hydrochloric acid can be obtained from 100 grams of salt?

14. How would you prove that hydrochloric acid gas contains chlorine and hydrogene?

15. How is bromine prepared? In what respects do the chemical properties of bromine resemble those of chlorine?

16. Where does iodine chiefly occur? Give a short account of its preparation.

17. In what substances is fluorine found? You wish to test an aqueous solution for hydroflueric acid. Describe how you would do so.

is action of

ne following
, powdered

s.
lorine ?
ll quantities

s elements ?

eous hydro-

tained from

gas contains

cts do the
ine ?
t account of

h to test an
you would

CHAPTER X.

SECTION I.

SULPHUR.

Symbol, S. Atomic Weight, 32 (31·98)

Sulphur occurs native in most volcanic districts; it is also abundantly found in combination with metals, such as iron, copper, lead, and zinc, and is widely distributed in the form of sulphates; sulphur compounds are found widely distributed in the vegetable and animal world.

PREPARATION.

271. From Iron Pyrites.

Exp. 1.—Powder a little iron pyrites, FeS_2, and put it into a narrow test-tube of hard glass, loosely stopping the mouth of the tube with the thumb; a brownish substance condenses on the cool part of the tube, which when heated becomes a reddish-brown liquid. This liquid is the element sulphur. The change that occurs in heating is represented by the equation—

$$\underset{\text{Iron pyrites.}}{3FeS_2} = \underset{\text{Magnetic pyrites.}}{Fe_3S_4} + \underset{\text{Sulphur.}}{S_2}$$

272. From Native Ore.

Almost all the sulphur of commerce comes from Italy, especially from the volcanic districts of Sicily, where it is found mixed with earthy and other impurities. The ore is placed in a pit with a small quantity of fuel, which is then lighted. When the temperature has somewhat increased the mass is covered and allowed to smoulder; the sulphur melts and collects on the floor of the pit, whence it is ladled out.

273. Refining Sulphur.

The crude sulphur is melted in an iron retort, the vapor passing into a large brick chamber, where it is condensed at first into a light crystalline powder termed *flowers of sulphur*, just as aqueous vapor falls as snow when the temperature suddenly sinks below 0°C. After a time the chamber becomes so hot that the sulphur condenses in the liquid form, and collects on the floor, whence it is drawn off and cast into cylindrical moulds, in which state it is known as *roll sulphur* or *brimstone*.

274. Liquefaction.

Exp. 2.—Heat about 20 grams of roll sulphur in a large test-tube; the sulphur begins to melt, and soon forms a clear mobile liquid (120°C.) On continuing the heat it darkens, gets thicker, becoming so viscid (at 220°C.) that the test-tube may be inverted for a moment without spilling any of it. On further heating, it again becomes clear, especially as the boiling point (440°C.) is approached. When this point is reached, slowly pour the sulphur into cold water; the mass of sulphur in the water is translucent, plastic, and can be moulded by the hand into various shapes. In this state it is termed *plastic* or *amorphous sulphur*, and is used for making moulds of coins, etc. When dry, it gradually loses its plastic properties, becoming opaque and brittle.

275. Allotropic Modifications of Sulphur.

(1) *The Crystalline Octohedra.*

Exp. 3.—Put about 1 gram of flowers of sulphur into a test-tube, and pour over them 5 c.c. of carbon disulphide. Loosely cork the tube, and hold it in warm water for a few minutes, shaking occasionally; part of the sulphur is dissolved. Filter the liquid into another tube, and pour a little of it on a glass plate, and leave it to evaporate; small, transparent crystals of sulphur soon appear. When examined through a magnifier these crystals are easily recognized to be right rhombic octohedra.

(2) *The Oblique Prismatic.*

Exp. 4.—Place some sulphur in a common clay crucible, and heat till it is just melted. Allow it to cool till a thick crust has been formed on the surface. Make two holes in this crust with a heated glass rod, and pour out the sulphur, which is still liquid. On removing the crust with a knife, the interior of the crucible is found lined with *needle-shaped* or *prismatic* crystals of sulphur. In a few days these needles lose their transparency and are spontaneously converted into an aggregation of minute octohedra.

(3) *The Amorphous.*

This variety was obtained while examining the effects of heat on sulphur (Exp. 2).

276. Union with other Elements.

Exp. 5.—Boil a little sulphur in a test-tube, and immerse in the vapor a close spiral of fine copper wire; the wire will burn in the sulphur vapor with incandescence, being converted into cuprous sulphide, Cu_2S.

We have already seen that iron combines with sulphur, forming mono-sulphide of iron, FeS (Art. 19). Sodium, when melted and plunged into sulphur vapor, burns with a blinding light. Sulphur also unites with the halogens and with phosphorus.

277. Detection of Sulphur.

Exp. 6.—Mix a small quantity of iron pyrites with an equal quantity of sodium carbonate, fuse the mixture on a piece of porcelain, or on charcoal, before the blow-pipe; sodium sulphide is thus formed, and this may be recognized by moistening the fused mass and placing it on a bright silver coin. The smallest quantity of sulphur can thus be recognized by the formation of a brown stain of sulphide of silver, Ag_2S.

278. Uses.—Sulphur is extensively consumed in the manufacture of matches, and of gunpowder, and still more in the manufacture of sulphuric acid. It is also used in medicine.

SECTION II.

SULPHUR DIOXIDE.

Symbol, SO_2. *Molecular Weight, 64.*

PREPARATION.

279. By burning Sulphur in Oxygen.

The combustion of sulphur in oxygen has already been described (Art. 41), and its combustion in air needs no special description. Neither of these methods affords a convenient source of the gas for experimental purposes, though the combustion in air is very useful when we desire to diffuse a quantity of the gas through the air of a room.

280. By the reduction of Sulphuric Acid.

Exp. 1.—Introduce into the hydrogen flask 30 grams of copper clippings, add through the funnel-tube 60 c. c. of strong sulphuric acid; arrange the flask for collecting the gas by downward displacement, and set it on a piece of wire gauze on the retort stand. Heat the mixture, and when effervescence begins take away the lamp. The gas coming off has the unmistakeable odor of the gas produced by burning sulphur in air or oxygen. The rate of the filling of the bottle may be ascertained by means of a lighted taper, as in carbon dioxide. The reaction is expressed by the following equation :—

$$\underset{\text{Copper.}}{Cu} + \underset{\text{Sulphuric acid.}}{2H_2SO_4} = \underset{\text{Copper sulphate.}}{CuSO_4} + \underset{\text{Water.}}{2H_2O} + \underset{\text{Sulphur dioxide.}}{SO_2}$$

This reaction takes place in three stages :—

(1) The displacement of two atoms of the hydrogen of the sulphuric acid by one atom of copper; thus :—

$$\underset{\text{Copper.}}{Cu} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Copper sulphate.}}{CuSO_4} + \underset{\text{Hydrogen.}}{H_2}$$

(2) The partial deoxidation of a second molecule of sulphuric acid by the nascent hydrogen :—

$$\underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Hydrogen.}}{H_2} = \underset{\text{Sulphurous acid.}}{H_2SO_3} + \underset{\text{Water.}}{H_2O}$$

(3) The breaking up of the sulphurous acid into sulphur dioxide and water :—

$$\underset{\text{Sulphurous acid.}}{H_2SO_3} = \underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O}$$

Have an earthenware dish or a plate under the whole apparatus in case of breakage of the flask.

PROPERTIES.

281. Does not support Combustion.

Exp. 2.—Plunge a lighted taper into a bottle of the gas; it is extinguished and the gas does not take fire.

Advantage is taken of this property to extinguish the fire in a stovepipe or chimney. A handful of sulphur is thrown into the stove or grate, and the sulphur dioxide arising from its combustion quickly checks the fire, especially if the air supply to the chimney be limited as far as possible. It is also essential that the chimney should be closed at the top so that it may be kept full of the gas until the walls shall have cooled below the kindling temperature of the soot.

282. It is heavier than Air.

Exp. 3.—Place a burning taper in an empty bottle and pour a bottle of sulphur dioxide upon it; the taper is extinguished. The molecular weight of the gas is 64, and therefore its density is $\frac{64}{2}=32$ $(H=1)$, or more than twice as heavy as air.

283. Solubility in Water.

Exp. 4.—Invert a bottle of the gas, place it mouth downwards in water and withdraw the plate; the water rises in the gas. Slip the hand beneath the mouth of the bottle, remove it, shake briskly, and again place it mouth downwards in water, repeating the operation till the water has nearly filled the bottle. Water at 0°C. dissolves about 68 times its bulk of the gas. Now add a little of this water to litmus-solution; it immediately becomes red, showing the presence of an acid. Sulphur dioxide is, therefore, an anhydride, as has already been shown in the experiments on oxygen; thus:—

$$\underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O} = \underset{\text{Sulphurous acid.}}{H_2SO_3}$$

This acid is so unstable that it is decomposed completely by boiling into gaseous sulphur dioxide and water.

284. Reducing Action.

Exp. 5.—Add a few drops of potassium permanganate, $KMnO_4$ (Condy's fluid), to some water, and then add sulphurous acid; the beautiful purple color is destroyed. To the bleached liquid add a few drops of barium chloride; a white precipitate insoluble in acids is formed, showing the presence of a sulphate. The sulphurous acid has deoxidized the permanganate, forming sulphuric acid, which has combined with the metals forming sulphates.

Exp. 7.—Place a few violets, pansies, or red roses in a bottle of sulphur dioxide; the color is soon discharged. Now dip the flowers into a very dilute solution of caustic soda or sulphuric acid, and the color reappears if the action of the bleaching agent has not been too long prolonged. The manner in which the sulphur dioxide acts in this case is not well understood. It may be that the dioxide becomes sulphuric acid in the presence of water, the liberated hydrogen combining with the coloring matter to form colorless compounds; thus:—

$$\underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Hydrogen.}}{H_2}$$

Thus the bleaching action of sulphur dioxide, in this case also, is thought to be a reducing one; just the opposite to that of chlorine, which is an oxidizing action.

It is more probable, however, that the bleaching action is due wholly or in part to the direct combination of the acid, with the coloring matters and the forming of colorless compounds. It does not, like chlorine, decompose the coloring matters, for these may usually be restored by acids or alkalies. The reproduction of the yellow color of new flannel, when it is washed with an alkaline soap, is an illustration of this. Ether and benzol also restore the color in some cases. Perhaps the simplest course is to oxidize the sulphurous acid to sulphuric acid. Thus, if a bleached rose be hung momentarily in chlorine gas, or in the vapor of red fuming nitric acid, the sulphurous acid is oxidized to sulphuric acid, and the red color of the rose is restored.

285. Disinfecting and Antiseptic Properties.—Sulphur dioxide is a powerful *antiseptic*, and has been successfully employed for preventing the putrefaction of meat and for stopping fermentation. Sulphur dioxide is also a powerful *disinfectant*, and is often employed to disinfect rooms in which contagious diseases have been treated. For this purpose the apertures are carefully closed, and a quantity of sulphur dioxide is generated in the room by throwing sulphur on red-hot coals contained in an iron vessel. After twenty-four hours the room may be opened and cleaned out.

286. Liquefaction.

Exp. 7.—Pass a stream of sulphur dioxide into a test-tube immersed in a mixture of pounded ice and salt, in equal parts; the gas is condensed into a colorless mobile liquid at a temperature of -10°C. To preserve it in this condition it must be sealed up. For this purpose, take a strong glass tube, close the one end, heat it near the other end, and draw it out till it will just admit a somewhat fine delivery-tube. Place this tube in the freezing mixture, and pass a current of gas into it; the gas condenses as before. Withdraw the delivery-tube, heat the narrow part of the tube with the blowpipe, and draw it out so as to seal it. The sulphur dioxide will remain liquid in the closed tube. The pressure exerted by its vapor is rather less than 2·5 atmospheres; hence it can be liquefied at ordinary temperatures if subjected to such a pressure.

287. Sulphites.

Sulphurous acid, like carbonic acid, contains two atoms of hydrogen, both of which may be replaced by metals; it is therefore a *dibasic* acid. It forms two series of salts called *sulphites*, in one of which half the hydrogen is replaced by a metal, and in the other the whole of the hydrogen is replaced. The first series of salts are termed *acid sulphites*, and the latter *normal sulphites*. Ex.: $NaHSO_3$ is acid sodium sulphite, or hydrogen sodium sulphite, and Na_2SO_3 is normal sodium sulphite.

288. Tests.—(1) All sulphites are decomposed by dilute sulphuric acid with effervescence if the solution is strong. The escaping gas (SO_2) smells of burning sulphur.

(2) When a small quantity of a sulphite or of sulphuric acid has to be recognized, the following is the best test:—

Add to a few pieces of granulated zinc in a test-tube some dilute hydrochloric acid till there is brisk effervescence. Then pour in a little of the liquid to be tested, and hold a lead acetate paper in the escaping hydrogen. Any sulphite present will be reduced by the nascent hydrogen and converted into sulphuretted hydrogen, which will stain the paper; thus:—

$$\underset{\text{Sulphurous acid.}}{H_2SO_3} + \underset{\text{Hydrogen.}}{3H_2} = \underset{\text{Sulphuretted hydrogen.}}{H_2S} + \underset{\text{Water.}}{3H_2O}$$

SECTION III.

SULPHURIC ACID.

Symbol, H_2SO_4. Molecular Weight, 98.

Sulphuric acid is the most important and useful acid known, probably the most important of all the chemical compounds on account of its numerous practical applications and the enormous quantities of it manufactured. It is by the means of it that nearly all the other acids are prepared, and there is scarcely an art or trade in which in some form or other it is not employed.

PREPARATION.

289. **By Oxidizing Sulphurous Acid.**—The method employed in the manufacture of sulphuric acid consists essentially in oxidizing sulphurous acid by atmospheric oxygen, which is accomplished by the aid of nitric oxide.

Exp. 1.—Fill a bottle with nitric oxide, remove the stopper and allow air to mix with the gas; ruddy fumes of N_2O_3 or NO_2 are produced. Now pour into the bottle a small quantity of sulphurous acid solution, insert the stopper and shake; the red fumes rapidly disappear. Remove the stopper and admit the air; the red fumes reappear, and can again be removed by shaking with some more sulphurous acid; and this set of changes can be repeated several times with the single charge of nitric oxide. Now test the liquid in the bottle with barium chloride, followed by hydrochloric acid; a white precipitate is formed, indicating the presence of sulphuric acid. The sulphurous acid has therefore been changed to sulphuric acid; thus:—

$$\underset{\text{Nitric oxide.}}{2NO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Nitric oxide.}}{2NO_2}$$

Then the nitric peroxide reacts with the sulphurous acid; thus:—

$$\underset{\text{Nitric peroxide.}}{NO_2} + \underset{\text{Sulphurous acid.}}{H_2SO_3} = \underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Nitric oxide.}}{NO}$$

The reproduced nitric oxide then takes a further supply of oxygen from the fresh air introduced, and passes it on to another quantity of sulphurous acid, and so on.

On the large scale, sulphur dioxide is formed by burning iron pyrites, FeS_2, in a furnace, sulphur being employed when pure acid is required. The sulphur dioxide formed is conveyed to a large leaden chamber, into which steam and air are also continuously admitted. At the back of the furnace in which the pyrites is burnt, a small vessel is placed containing the materials for making nitric acid. Sulphuric acid is formed in large quantities, and dissolves in the stratum of water with which the bottom of the chamber is covered.

The reactions which occur are but imperfectly understood; the nitric acid is, however, apparently decomposed by the sulphurous acid in such a manner that nitric oxide is produced; thus:—

$$(1)\quad \underset{\text{Nitric acid.}}{2HNO_3} + \underset{\text{Sulphur dioxide}}{3SO_2} + \underset{\text{Water.}}{2H_2O} = \underset{\text{Nitric oxide.}}{2NO} + \underset{\text{Sulphuric acid.}}{3H_2SO_4}$$

The nitric oxide combines with the atmospheric oxygen, which is admitted into the chamber; thus:—

$$(2)\quad \underset{\text{Nitric oxide.}}{2NO} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Nitric peroxide.}}{2NO_2}$$

This nitric peroxide is no sooner produced than it is reduced again to nitric oxide by the sulphurous acid, which is oxidized to sulphuric acid; thus:—

$$(3)\quad \underset{\text{Sulphur dioxide.}}{SO_2} + \underset{\text{Water.}}{H_2O} + \underset{\text{Nitric peroxide.}}{NO_2} = \underset{\text{Sulphuric acid.}}{H_2SO_4} + \underset{\text{Nitric oxide.}}{NO}$$

The nitric oxide again absorbs oxygen to form the higher oxides, which again enter into reactions with the sulphurous acid, and so on continuously, the nitric oxide simply acting as a carrier. Theoretically, therefore, a limited quantity of nitric oxide should suffice to produce an indefinitely large amount of sulphuric acid, but practically this is not the case, as a certain quantity dissolves in the sulphuric acid on the floor of the chamber, and a certain amount escapes with the nitrogen.

The sulphuric acid thus produced is very dilute, and in order to convert it into the sulphuric acid of commerce it is heated in glass or platinum retorts until a large portion of the water is driven off.

PROPERTIES.

290. Affinity for Water.

Exp. 2.—Pour in a fine stream and with frequent stirring four parts of sulphuric acid into one part of water, and stir the mixture with a small test-tube containing a little ether or alcohol colored with aniline red; the liquid in the test-tube will boil violently. A thermometer immersed in the liquid indicates a temperature of 140°C. If the product is cooled to 0°C., crystals separate out, which have the composition H_2SO_4, H_2O; the evolution of heat is, therefore, due to the formation of a definite hydrate.

The avidity with which sulphuric acid absorbs moisture renders it a valuable agent for drying gases; and it is now largely employed in the production of artificial ice. The water to be frozen is placed in a strong flask, which is connected with an air-pump by means of a large tube or cylinder containing strong sulphuric acid. On working the pump, the air is exhausted and the water is cooled by its own evaporation under diminished pressure; and as the vapor produced is at once absorbed by the sulphuric acid, the evaporation becomes so rapid that the temperature of the water is lowered to the freezing point.

291. Action on Compounds containing the Elements of Water.

Exp. 3.—Dip a splint of white pine into sulphuric acid; the pine is blackened. The pine is composed of woody fibre, and the reaction is—

$$\underset{\text{Woody fibre.}}{C_6H_{10}O_5} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Carbon.}}{6C} + \underset{\text{Dilute sulphuric acid.}}{H_2SO_4, 5H_2O}$$

Exp. 4.—Make a strong syrup of sugar and pour into it about one-half as much of strong acid; the sugar is converted into a bulky mass of porous charcoal. The reaction is—

$$\underset{\text{Sugar.}}{C_{12}H_{22}O_{11}} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Carbon.}}{12C} + \underset{\text{Dilute Sulphuric acid.}}{H_2SO_4, 11H_2O}$$

292. Sulphates.

The salts formed by replacing the hydrogen of sulphuric acid by a metal are termed *sulphates.* As this acid, like sulphurous acid, is dibasic, two series of sulphates exist. By displacing one-half of the hydrogen *acid sulphates* are formed, and *normal sulphates* are produced by displacing the whole of the hydrogen.

COMMON SULPHATES.

SYSTEMATIC NAME.	COMMON NAME.	FORMULA.
Sodium Sulphate	Glauber's salt..............	Na_2SO_4
Potassium " ...		K_2SO_4
Magnesium "	Epsom salt........................	$MgSO_4$
Calcium "	Alabaster, gypsum, plaster of Paris ..	$CaSO_4$
Zinc "	White vitriol..............	$ZnSO_4$
Copper "	Blue vitriol................	$CuSO_4$
Ferrous "	Green vitriol or copperas....	$FeSO_4$

293. Properties of Sulphates.

(1) All sulphates are soluble in water except barium sulphate, lead sulphate, strontium sulphate, and calcium sulphate which is a little soluble.

(2) Barium chloride or nitrate added to a solution of a sulphate in water gives a white precipitate insoluble in hydrochloric acid.

SECTION IV.

SULPHURETTED HYDROGEN.

Symbol, H_2S. Molecular Weight, 34.

PREPARATION.

294. By the action of Acids upon Sulphides.

Exp. 1.—Fit up the apparatus for the preparation of hydrogen, attach a small flask as a wash bottle, and place in it a little water. Place in the flask some fragments of ferrous sulphide, FeS, cover them with water, and then add a small quantity of sulphuric acid; the gas is rapidly evolved. Collect over warm water, and fill three or four bottles. The reaction is—

$$\underset{\text{Ferrous sulphide.}}{FeS} + \underset{\text{Sulphuric acid.}}{H_2SO_4} = \underset{\text{Ferrous sulphate.}}{FeSO_4} + \underset{\text{Sulphuretted hydrogen.}}{H_2S}$$

PROPERTIES.

295. Acid Character.

Exp. 2.—Add litmus solution to a bottle of the gas; the color of the litmus is changed into a port wine tint, showing that the gas has feeble acid properties.

296. Combustibility.

Exp. 3.—Attach a jet to the generating apparatus and light the gas; it burns with a feeble blue flame. The reaction is—

$$\underset{\text{Sulphuretted hydrogen.}}{2H_2S} + \underset{\text{Oxygen.}}{3O_2} = \underset{\text{Sulphur dioxide.}}{2SO_2} + \underset{\text{Water.}}{2H_2O}$$

Exp. 4.—Remove a cover from a bottle and introduce a lighted taper; the gas burns but the taper is extinguished, and sulphur is deposited on the inside of the bottle. In this case the supply of oxygen is limited, and the reaction is—

$$\underset{\text{Sulphuretted hydrogen.}}{2H_2S} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Sulphur.}}{S_2} + \underset{\text{Water.}}{2H_2O}$$

297. Solubility in Water.

Exp. 5.—Pour a little water into a bottle of sulphuretted hydrogen, close its mouth with the hand, shake well, invert the bottle under water and withdraw the hand; the water rushes in, showing that the gas is soluble in water. The saturated solution may be prepared by passing a current of the gas through water until the bubbles arising cease to diminish in volume. The water thus saturated smells strongly of the gas, and in many experiments may be used instead of it.

298. Action on Chlorine.

Exp. 6.—Invert a bottle of chlorine over a bottle of sulphuretted hydrogen; sulphur is deposited on the inside of the bottles; thus:—

$$\underset{\text{Sulphuretted hydrogen.}}{2H_2S} + \underset{\text{Chlorine.}}{2Cl_2} = \underset{\text{Sulphur.}}{S_2} + \underset{\text{Hydrochloric acid.}}{4HCl}$$

Sulphuretted hydrogen is almost a constant ingredient in the gas of sewers, and in this way often finds its way into rooms. A little chlorine liberated from chloride of lime deodorizes an atmosphere rendered offensive by this gas.

299. Action on Metals.

Exp. 7.—Place a drop of sulphuretted hydrogen water upon a bright piece of silver; the metal quickly becomes black. The reaction is—

$$\underset{\text{Silver.}}{2Ag} + \underset{\text{Sulphuretted hydrogen.}}{H_2S} = \underset{\text{Silver sulphide.}}{Ag_2S} + \underset{\text{Hydrogen.}}{H_2}$$

It is on this account that silver plate is so easily blackened by the air of towns, which is usually contaminated with sulphuretted hydrogen. An egg spoon is always blackened by the sulphur from the egg. Silver coins kept in the pocket with matches become black. The black sulphide thus formed may be removed by rubbing with strong ammonia or a solution of potassium cyanide.

300. Action on Metallic Salts.

Exp. 8.—Dissolve a small crystal of lead nitrate or acetate in a test-tube half full of water, and to this solution add a few drops of sulphuretted hydrogen water; the solution becomes black; thus:—

$$\underset{\text{Lead nitrate.}}{Pb(NO_3)_2} + \underset{\text{Sulphuretted hydrogen.}}{H_2S} = \underset{\text{Lead sulphide.}}{PbS} + \underset{\text{Nitric acid.}}{2HNO_3}$$

It is in this manner that paints containing white lead (carbonate of lead) are darkened by exposure to the air of towns.

Paintings, whether in oil or water colors, are damaged by exposure to air containing sulphuretted hydrogen. The paintings may be restored by exposure to light, the black sulphide becoming oxidized into the white sulphate.

Most metallic sulphides are insoluble in water; hence sulphuretted hydrogen is peculiarly adapted for precipitating metals from their solutions; and as the sulphides often possess a characteristic color, a solution of the gas in water is much used for detecting many metals. It is also used for separating metals into classes.

Exp. 9.—Prepare a solution in separate test-tubes of each of the following substances:—Arsenic trioxide, copper sulphate, lead acetate, tartar emetic, zinc sulphate, and ferrous sulphate. Add a few drops of hydrochloric acid to each, and pass sulphuretted hydrogen through each solution from the generating apparatus, washing the leading tube between each solution. Observe and tabulate the results.

Through fresh portions of zinc and iron sulphates, *without* the addition of hydrochloric acid, pass sulphuretted hydrogen. Observe and tabulate the results. To these now add a few drops of hydrochloric acid, and observe the results.

301. Poisonous Character.—When inhaled, even when mixed with a considerable volume of air, the gas act as a powerful poison.

302. Tests—A minute trace of sulphuretted hydrogen may be detected by its blackening effect on paper moistened with lead acetate.

QUESTIONS AND EXERCISES ON CHAPTER X.

1. In what condition and in what localities does free sulphur occur? How is native sulphur attracted in Sicily?
2. How are flowers of sulphur and roll sulphur obtained?
3. Describe fully the changes which occur when sulphur is heated from the ordinary temperature to its boiling point.

4. Describe the three modifications of sulphur, and state exactly how each is obtained.

5. What gases are produced when sulphur burns (*a*) in air, (*b*) in oxygen, (*c*) in nitrous oxide!

6. Write in parallel lines corresponding compounds of oxygen and sulphur.

7. Express by equations the reactions which occur when strong sulphuric acid is heated with (*a*) mercury, (*b*) sulphur, (*c*) copper.

8. Explain the bleaching action of sulphur dioxide.

9. How may the process for preparing sulphuric acid be illustrated on the smell scale? Sketch the apparatus.

10. Explain fully the theory of the formation of the sulphuric acid in the leaden chamber. Give equations to represent the reactions supposed to occur.

11. What occurs when sulphuric acid is left exposed to the air?

12. What is the action of sulphuric acid on organic matter, such as sugar?

13. A sample of vinegar is supposed to be adulterated with sulphuric acid. How would you ascertain the presence of sulphuric acid in vinegar?

14. Enumerate the chief properties of sulphuretted hydrogen?

15. How would you prove that sulphuretted hydrogen is a compound of sulphur and hydrogen?

16. 8 grams of sulphur are burnt in air, what volume of sulphur dioxide is formed?

17. How much copper and sulphuric acid are required to make 1000 grams of sulphur dioxide?

18. 16 litres of hydrogen diffuse through an apparatus in 100 minutes, how much sulphur dioxide will diffuse under the same conditions?

19. What weight of iron sulphide will be needed to yield a litre of sulphuretted hydrogen at 0° and 760 mm., and how much air will be required to burn this gas completely to sulphur dioxide and water?

20. An unknown volume of sulphuretted hydrogen requires 110.34 c.c. of chlorine for complete decomposition. What was the volume of the sulphuretted hydrogen?

CHAPTER XI.

SECTION I.

PHOSPHORUS.

Symbol, P. Atomic Weight, 31 (30.96). Molecular Weight, P_4, 124.

PREPARATION.

Phosphorus is prepared from bones, which consist chiefly of gelatin and calcium phosphate. The bones are either boiled with water or treated with superheated steam to extract the gelatin. They are then burnt to a white ash, which is finely powdered and digested with dilute sulphuric acid. The following reaction takes place :—

$$\underset{\text{Bone ash.}}{Ca_3(PO_4)_2} + \underset{\text{Sulphuric acid.}}{2H_2SO_4} = \underset{\text{Calcium Superphosphate.}}{CaH_4(PO_4)_2} + \underset{\text{Calcium Sulphate.}}{CaSO_4}$$

The soluble superphosphate is poured from the insoluble gypsum and evaporated to dryness, after which the solid residue is heated to redness, thus :—

$$\underset{\text{Calcium Superphosphate.}}{CaH_4(PO_4)_2} = \underset{\text{Calcium Metaphosphate.}}{Ca(PO_3)_2} + \underset{\text{Water.}}{2H_2O}$$

This salt is then carefully mixed with charcoal and heated to bright redness in earthenware retorts. The phosphorus passes over and is collected in water :—

$$\underset{\text{Calcium metaphosphate.}}{3Ca(PO_3)_2} + \underset{\text{Charcoal.}}{10C} = \underset{\text{Phosphorus.}}{P_4} + \underset{\text{Calcium phosphate.}}{Ca_3(PO_4)_2} + \underset{\text{Carbon monoxide.}}{10CO}$$

PROPERTIES.

Great care is required in experimenting with phosphorus. It may catch fire even when lying on blotting paper; it should, therefore, be kept and cut under water.

303. Inflammability.

Exp. 1.—Pour a small quantity of carbon disulphide into a test-tube, and add a small piece of phosphorus; it rapidly dissolves. Place a piece of blotting paper on a ring of the retort stand, and pour a little of the liquid upon it; the carbon disulphide evaporates, leaving the phosphorus in a minute state of division in which state it ignites spontaneously.

Exp. 2.—Put a piece of phosphorus the size of a small pea on blotting paper, and sprinkle over it some soot or finely powdered charcoal; it melts after a short time and spontaneously inflames.

304. Phosphorescence.

Exp. 3.—Put into a bottle a quarter of an ounce of ether and a piece of phosphorus the size of a small pea, cork the bottle and let it stand a few days, frequently shaking it. Pour some drops of this solution on the hands and rub them quickly together; the ether rapidly evaporates leaving the phosphorus upon the hands in a state of minute division. The more finely it is divided, so much the more easily does it combine with the oxygen of the air. During this combustion it diffuses a white smoke and a faint light, causing the hands to shine in the dark; hence its name (*Phos*, light, and *fero*, I carry).

Exp. 4.—Moisten a lump of sugar with the solution of phosphorus in ether, and throw it into hot water, the surface of the water will glow in a dark room.

305. Combustion.

Exp. 5.—The combustion of phosphorus in oxygen has already been described (Art. 39). The combustion may also take place under water. Put a piece of phosphorus about the size of a small pea into a wine-glass and pour hot water upon it until the glass is half filled; the phosphorus melts, but does not ignite, as the access of oxygen is prevented by the water. But if air is carefully blown by the mouth through a glass tube upon the bottom of the glass, *combustion* will ensue, which is visible in the dark.

306. Combustion with Chlorine, Bromine, Iodine.

Exp. 6.—Cut a thin slice of phosphorus, dry it gently between the folds of blotting paper and place a few grains of iodine upon it; they unite with the evolution of heat and light. Phosphorus unites directly with these elements.

307. Amorphous or Red Phosphorus.—Phosphorus like sulphur and carbon is capable of existing in several states. Besides the ordinary kind the most interesting is that called red or amorphous phosphorus. This is prepared by heating ordinary phosphorus to a temperature of about 240°C. for some time, in a vessel filled with carbon dioxide, to prevent it from burning. Red phosphorus is insoluble in carbon disulphide, is not poisonous, and inflammable only at a high temperature. It is changed back again to the common yellow variety by the action of heat in the presence of air.

308. Lucifer Matches.—Formerly matches were made by first dipping the ends of the wood into a paste of gum-water, phosphorus, and either nitre or potassium chlorate. By friction the phosphorus is set on fire. The heat of the burning phosphorus is intense enough to ignite the sulphur, which in turn inflames the wood. Instead of sulphur, paraffine is often used. When potassium chlorate is used, the matches burn with an explosive combustion. The "safety matches" contain no phosphorus. This element is put upon the box, instead of upon the match itself. The matter on the top of the match contains potassium, chlorate, and sulphur or antimony sulphide, and sometimes red lead and potassium bichromate. The matter on the side of the box contains red phosphorus. Only by rubbing the match against the surface, can it be easily fired.

309. Action of Phosphorus as a Poison.—Ordinary phosphorus is a powerful poison capable of producing death in a few hours, or when given in small doses, of producing a remarkable train of *poisonous* symptons lasting for many days or even for weeks. There is no good antidote. Giving an emetic and administering at the same time some thick gruel containing chalk or magnesia, the former to suspend the particles, and the latter to neutralize any acid products formed, constitutes the best method of treatment in case of poisoning. Above all, the administering of oils or fatty matters must be avoided.

Tests.—Phosphorus is easily identified by its odor, by its property of fuming in the air, and by shining in the dark.

COMPOUND OF PHOSPHORUS AND OXYGEN.

OXIDES.	CORRESPONDING ACIDS.
	Hypophosphorous acid ...... H_3PO_2
P_2O_3 Phosphorus trioxide..	Phosphorus acid .... H_3PO_3
P_2O_5 Phosphorus pentoxide	Metaphosphoric acid .... HPO_3
" " "	Pyrophosphoric acid .... $H_4P_2O_7$
" " "	Orthophosphoric acid .... H_3PO_4

It will be seen from the above table—

(1) That there is no anhydride for hypophosphorous acid.

(2) That phosphorus pentoxide forms three acids by combination with 1, 2, or 3 molecules of water respectively.

(3) That although hypophosphorous acid, phosphorous acid, and orthophosphoric acid severally contain 3 atoms of hydrogen in the molecules, nevertheless they differ as follows:—

(a) In hypophosphorous acid, only 1 atom of hydrogen can be replaced by a metal (monobasic).

(b) In phosphorous acid, only 2 atoms of hydrogen can be replaced by a metal (bibasic).

(c) In phosphoric acid *all* 3 atoms of hydrogen can be replaced by a metal (tribasic).

SECTION II.

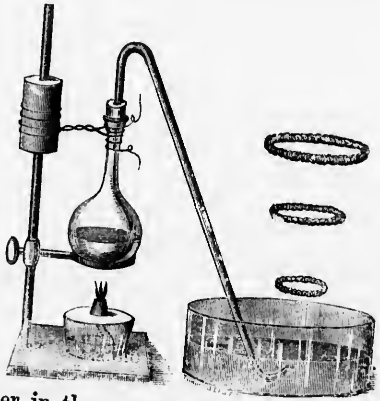
PHOSPHURETTED HYDROGEN.

Symbol, PH_3. *Molecular Weight, 34.*

PREPARATION.

309. By the action of Phosphorus on basic hydrates.

Exp. 1.—Into a flask of about 140 c.c. capacity, put 1 gram of phosphorus and 115 c.c. of a solution of sodium hydrate, obtained by dissolving 40 grains of caustic soda in 110 c. c. of water. Pour a few drops of ether upon the liquid in the flask, then close the flask with a cork carrying a long delivery-tube. Place the flask on the wire-guage, on the retort stand, and immerse the end of the delivery-tube in warm water in the pneumatic trough, and apply a gentle heat; the ether first passes out, carrying with it the air in the flask; a gas in then evolved, each bubble of which, as it passes out of the water into the air, spontaneously inflames, and produces a ring of white smoke. The reaction is—



$$\underset{\text{Phosphorus.}}{P_4} + \underset{\text{Sodium Hydrate.}}{3NaOH} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Sodium hypophospite.}}{3NaH_2PO_2} + \underset{\text{Phosphuretted hydrogen.}}{PH_3}$$

The rings are composed of phosphorus pentoxide, formed by the oxidation of the phosphuretted hydrogen; thus:—

$$\underset{\text{Phosphuretted hydrogen.}}{2PH_3} + \underset{\text{Oxygen.}}{4O_2} = \underset{\text{Phosphorus pentoxide.}}{P_2O_5} + \underset{\text{Water.}}{3H_2O}$$

Phosphuretted hydrogen when pure, is not spontaneously inflammable. When prepared as above it contains a little of the vapor of another compound of phosphorus and hydrogen, P_2H_4, which is spontaneously inflammable.

310. Precautions.

(1) Care must be taken to expel all the air from the flask, otherwise a dangerous explosion will result. This may be done by passing coal-gas, carbon dioxide, or hydrogen, through it, or by a few drops of ether.

(2) If convenient, the flask may be heated in a vessel containing a strong solution of salt. This will guard against the ignition of the phosphorous in case the flask should accidentally break.

(2) The sodium hydrate must not be too concentrated, otherwise there will be a tendency to boil over.

(4) During the experiment, currents of air in the room must be avoided.

(5) When it is desired to stop the evolution of the gas, add some boiling water to the water in the pneumatic trough, withdraw the lamp, and allow the hot water to be slowly drawn into the retort.

QUESTIONS AND EXERCISES ON PHOSPHORUS.

1. What is the chief source of phosphorus and its compounds? What is the most abundant phosphate found in nature? Name other naturally occuring phosphates.

2. Give the details of the process employed for the preparation of phosphorus from bone-ash. Represent all reactions by equations.

3. A piece of bone is heated (*a*) in a covered crucible, (*b*) in an open fire. Describe the substance obtained in each case, and state which would contain the larger proportion of phosphorus.

4. Compare the properties of the two varieties of phosphorus. How would you prove, in spite of the great difference in properties, that the common and red phosphorus are modifications of the same element?

5. The skeleton of a full grown man weighs 24 pounds, and bones contain 58 per cent. of calcium phosphate. Calculate the weight of phosphorus in the skeleton.

6. How are lucifer matches made, (*a*) using common phosphorus, (*b*) using amorphous phosphorus?

7. Describe a process for the detection of phosphorus in case of poisoning.

8. Phosphorus vapor is 4.42 times as heavy as air, find its molecular weight.

9. Calculate the formula of a substance containg calcium 38.72 per cent., phosphorus 20 per cent., and oxygen 48.28 per cent.

10. How is phosphuretted hydrogen prepared? Describe its properties. How much phosphorus is contained in 10 litres of phosphuretted hydrogen measured at 12°C?

SECTION III.

ARSENIC.

Symbol, As. Atomic Weight, 75. Molecular Weight, As_4, 300.

PREPARATION.

311. From Arsenic Compounds.

Exp. 1.—Drop into the point of a drawn-out tube of glass, a small quantity of arsenic trioxide, and above it place a splint of charcoal. Heat the charcoal red-hot, as in Exp. 7, Art. 102; arsenic will be deposited in a ring on the colder parts of the tube, presenting a brilliant metallic appearance. The reaction is—

$$\underset{\text{Arsenic trioxide.}}{2As_2O_3} + \underset{\text{Carbon.}}{6C} = \underset{\text{Arsenic.}}{As_4} + \underset{\text{Carbon monoxide.}}{6CO}$$

Commercial arsenic is prepared by heating its ores in earthenware tubes in a furnace. It is purified by sublimation with charcoal. The arsenic thus obtained is a moderately good conductor of heat and electricity. In nature it occurs in small quantities in the free state, but much more commonly in a state of combination with iron, cobalt, and other metals. Traces of arsenic occur in all mineral waters, and it has been detected in sea-water.

PROPERTIES.

312. Odor and Volatility.

Exp. 2.—Repeat the last experiment and note that on the application of heat, an odor resembling *garlic*, is developed. Close the mouth of the tube and heat the mirror; the arsenic readily volatilizes

Exp. 3.—Break off the end of the tube, hold it at an angle of 45°, and heat the mirror; the arsenic combines with the oxygen, forming arsenic trioxide which condenses in minute octahedral crystals on the colder portions of the tube.

Arsenic readily *oxidizes* in warm, moist air, forming "fly powder," which is probably a mixture of arsenic, and arsenic trioxide. It is used in the manufacture of shot, to faciliate the lead assuming a globular form.

313. Arsenic Trioxide.—This is the white arsenic of the druggist, and is the source of most of the arsenic compounds. It is formed when metallic arsenic or arsenical ores are heated in air.

314. Arsenious Acid.

Exp. 4.—Boil a small quantity of arsenic trioxide in a test-tube; it appears to be insoluble. Add a few drops of litmus solution, it is slightly reddened. Arsenic trioxide is, therefore, soluble to a small extent in water. One grain requires for solution 50 grains of *cold* water, and 12 grains of *boiling* water. The reaction is—

$$\underset{\text{Arsenic trioxide.}}{As_2O_3} + \underset{\text{Water.}}{3H_2O} = \underset{\text{Arsenious acid.}}{2H_3AsO_3}$$

This acid, like carbonic acid, has not been isolated in a pure state.

315. Scheele's Green.

Exp. 5.—Take one part of arsenic trioxide, three parts of potassium carbonate, and fourteen parts of water, and add to a dilute boiling solution of copper sulphate; a green precipitate is formed. The composition of this precipitate is represented by the formula, $CuAsO_3$, commonly known as Scheele's Green or Paris Green. This compound being cheap, and brilliant in color is often improperly used as a pigment for wall-paper, artificial flowers, lamp-shades, and the like. From such paper the poison is mechanically detached and floats in the air as an impalpable powder, a small portion of which is inhaled at every breath, sometimes giving rise to aggravated cases of poisoning.

316. Detection of Arsenic.

Exp. 6.—Take a piece of bright green wall-paper, scrape off some of the pigment, put it into a drawn-out tube, cover with a layer of charcoal and heat as in Exp. 1; a dark ring of metallic arsenic will be obtained, or crystals of the oxide will be deposited on the colder parts of the tube, if the pigment is arsenical.

Exp. 7.—Soak a piece of the paper in a little ammonia, which will dissolve the copper arsenite to a blue liquid. Green colors that contain copper also contain arsenic.

317. Reinsch's test.

Exp. 8.—Introduce some of the green pigment, scraped from wall-paper, into a test-tube, add dilute hydrochloric acid, and place in the tube a slip of bright copper foil; the copper becomes dark colored from the deposition on its surface of metallic arsenic. Dry the copper by pressing it between the folds of blotting paper, and put it in a dry test-tube, a portion of the arsenic is oxidized, and volatilized, and arsenic trioxide deposited as a white sublimate on the cold portions of the tube. This may be more fully examined by Marsh's test, described in the next experiment. This is known as Reinsch's test. The presence of organic matter does not interfere with it, but as the greater part of the arsenic remains behind in combination with the copper, it can be employed only when the quantity of arsenic is considerable. To avoid mistakes, the hydrochloric acid and copper should be examined in precisely the same way, without the suspected paper, so as to render it certain that the arsenic was not derived from them.

318. Marsh's test.

Exp. 9.—Fit up the bottle for generating hydrogen with the drying tube, as in Exp. 11, Art. 76. When the air has escaped, light the hydrogen and hold a piece of white porcelain or stone-ware in the flame; a circle of small drops of water is formed upon it. Dissolve the white sublimate of the last experiment in boiling water, and add through the funnel-tube; the flame becomes bluish-white, and on holding the cold porcelain in it, a black or brown shining mirror is deposited. This is Marsh's test. The presence of the minutest trace of arsenic can be ascertained with perfect certainty by it. The gas is arseniuretted hydrogen, AsH_3. Great care should be taken not to inhale any of it, as it is exceedingly poisonous, and one to which some chemists have already fallen victims.

319. Arsenic distinguished from Antimony.

Exp. 10.—Pour upon the black stains a solution of bleaching powder; the stains will rapidly dissolve. Repeat Exp. 9, substituting a little tartar emetic for arsenic trioxide; black spots are deposited on the porcelain, but they have a sooty appearance. They consist of metallic antimony. Pour upon them a solution of chloride of lime; they remain unchanged. This distinguishes antimony from arsenic.

320. Antidote.—Dilute ferric chloride with about twice its volume of water, and add sodium or ammonium hydrate in slight excess; a bulky brown precipitate forms. Place the precipitate on a cloth, and wash well with water. A teaspoonful may be given at short intervals till five or six are taken. Calcined magnesia is also an antidote. Both these substances form insoluble compounds with arsenic, thus preventing its absorption by the system. An emetic, such as a teaspoonful of mustard in a cup of warm water, should soon follow the antidote, and that in turn should be followed by caster oil.

321. Arsenic and Sulphur.—Arsenic forms two compounds with sulphur, *realgar* As_2S_2, one of the ingredients of white Indian fire, which is often used as a signal light; and *orpiment*, As_2S_3, the coloring ingredient in the pigment called King's yellow.

QUESTIONS AND EXERCISES ON ARSENIC.

1. How does arsenic occur in nature? Name some of its ores, and state how the element is obtained from them?
2. Describe the tests for arsenic, known as "Marsh's test," and "Reinsch's test?"
3. What occurs when arsenic trioxide is heated in contact with charcoal?
4. What is Scheele's green, and how is it prepared? What percentage of arsenic does it contain?
5. What is the best antidote to arsenic poisoning?
6. A room 15 feet long, 10 feet wide, and 10 feet high, is covered with a paper containing 0.78 grains of Scheele's green per square foot. How much arsenic is there in the room?

CHAPTER XII.

SECTION I.

BORON.

Symbol, B. Atomic Weight, 11.

322. **Occurrence.**—Boron does not occur in nature in a free state. It is chiefly found combined as boric acid in the steam which issues from fissures in the earth, in some parts of Tuscany, and bubbles up through small lakes or basins of water. The acid is obtained by evaporating the water of these lakes by an ingenious arrangement whereby the heat required for the purpose is supplied by the hot springs themselves.

323. **Borax.**—A sodium salt of boric acid called *tincal* is obtained from Thibet; when purified by recrystallization, it is sold as *borax*, an irregular substance having the formula $Na_2 B_4O_7$.

324. **Boric Acid.**

Exp. 1.—Add powdered borax to boiling water until it ceases to be dissolved, pour off the solution from any undissolved borax, and mix it with about half its bulk of strong hydrochloric acid; flaky crystals are immediately formed and deposit abundantly as the mixture cools. Collect the crystals in a filter, and wash them with a little cold water. They have the formula $H_3 BO_3$, and an known as *boric or boracic acid.*

Boric acid communicates to its compounds the property of ready fusibility; indeed, it is chiefly on this account that it is valued. Many of the borates are admirably adapted for fluxes, which are used in the glazing of porcelain, and in the melting of gold and silver.

Exp. 2.—Pour alcohol on some boric acid, or on borax which has been moistened with sulphuric acid. Inflame the spirit and in either case it will burn with a greenish flame, from the presence of boric acid.

SECTION II.

SILICON.

Symbol, Si. Atomic Weight, 28.

325. Occurrence.—Silicon is the most abundant solid element. Next to oxygen, it is the chief constituent of the solid crust of the earth. It is always found in combination, and is isolated only with difficulty. It enters into the composition of so many minerals that the greater part of the science of mineralogy is occupied with the history of its compounds. Like carbon and boron, silicon occurs in three modifications, viz.: *amorphous*, *graphoidal*, and *crystalline*.

326. Silicon Dioxide, SiO_2.—This compound is also called *silica*, and is the only known oxide of silicon. It occurs both in a crystalline and in an amorphous form. Rock crystal, amethyst, and cairngorm are illustrations of the crystalline variety; agate, chalcedony, jasper, cornelian, flint, and opal, are illustrations of the amorphous variety.

327. Soluble Glass.

Exp. 1.—Heat a common flint stone in a fire, and when it is red-hot, immerse it in cold water. The sudden cooling so influences and diminishes the internal cohesion, that now the flint can be easily reduced to a fine white powder. Boil in a clean iron ladle four grams of the powder, with eight grams of caustic potash, and sixty grams of water for some hours; supplying of fresh water occasionally as the other evaporates. Pour the mixture into a bottle and let it settle. The silica dissolves in the potash solution and forms with it a thick fluid, *potassium silicate.* Sodium carbonate may be employed instead of caustic potash, but greater heat is required. These soluble silicates are often called *soluble glass,* and are employed for hardening soft building stone, and for the artifical preparation of hard stone from sand.

328. Silicic Acid, H_4SiO_4.

Exp. 2.— Add hydrochloric acid to a strong solution of potassium or sodium silicate. A white gelatinous mass of silicic acid is deposited. If this jelly is collected on a filter, washed well with water, and dried, a white powder, consisting of pure silica in the amorphous state, is obtained.

Almost all springs, as well as plants contain small quantities of soluble silicates. If spring water is evaporated, silica will be found in the insoluble residue; if a plant is burnt, silica will be found in the ashes. Grasses and different kinds of grain are particularly rich in silica, which is to them what bones are to animals. Silica is the substance to which the stalks of plants owe their firmness and stiffness. If there is not enough potash in the soil to render the silica soluble these properties will be wanting in the stalk, and it will readily bend. The horse-tail plant (Equisetum) contains so much silica that it has been used for polishing wood.

In its natural state silica is so hard, as to produce sparks with steel, and is quite insoluble in water and acids, except hydrofluoric acid.

329. Silicates.—The silicates are very numerous and very complicated. Among important silicates may be mentioned felspar, mica, hornblende, common slate, soapstone, and other compounds of frequent occurrence. The different varieties of clay are silicates of aluminium. Glass is a mixture of various silicates, especially those of the alkalies and alkaline earths, formed by the fusion of the ingredients. Window-glass is a silicate of soda and lime. Flint-glass is a potash silicate. Bottle-glass is a silicate of soda and lime, together with alumina and oxide of iron, and it is to this latter that the glass owes its green color.

QUESTIONS ON CHAPTER XII.

1. What is borax, and how is boric acid obtained from it?

2. How does silicon occur in nature? Name its allotropic modifications, and its compounds with oxygen.

3. What is the substance termed soluble glass, and what are its uses?

4. A crystal is given you, and you are requested to ascertain by a chemical experiment whether it is a diamond or a common rock crystal. Describe an experiment which will furnish the required information.

5. A splinter of a colorless, transparent solid body is given you. How would you ascertain (*a*) by physical tests, (*b*) by chemical tests, whether it is glass or diamond.

CHAPTER XIII.

ACIDS, BASES AND SALTS.

330. Properties of Acids.— We have seen that *acids* possess the following properties :—

(1) They have an acid, or sour, taste (Art. 40).

(2) They change blue litmus red (Art. 40).

(3) They act upon metals, hydrogen being evolved, and its place taken by the metal, (Arts. 66, 68, 157, 280).

(4) They act upon metallic hydrates, forming neutral substances and water.

The first characteristic applies only to soluble acids. This may also be said of the second; and besides, copper sulphate, aluminium sulphate, and many other salts of the heavy metals, change blue litmus red. The third may be asserted of alkalies, (Art. 81). In the fourth we have a characteristic which applies to *all* acids.

331. Definition of an Acid. *An acid is a compound, containing one or more atoms of hydrogen, which becomes displaced by a metal when the latter is presented to the compound in the form of a hydrate, water being evolved.*

$$HNO_3 + NaOH = NaNO_3 + H_2O.$$
$$H_2SO_4 + 2KOH = K_2SO_4 + 2H_2O.$$
$$2HCl + Ca(OH)_2 = CaCl_2 + 2H_2O.$$

The hydrogen capable of being so displaced is termed the *displaceable hydrogen.* An acid containing one such atom of hydrogen is said to be *monobasic,* two such atoms, *dibasic,* &c.

332. Bases.—The term base is applied to three classes of compounds :—

(1) Certain compounds of metals with oxygen, as sodium oxide, Na_2O, zinc oxide, ZnO, &c.

(2) Certain metallic hydrates, such as sodium hydrate, $NaOH$, calcium hydrate, $Ca(OH)_2$, &c.

(3) Certain compounds of nitrogen, phosphorus, arsenic, and antimony, such as ammonia, NH_3, &c.

These all possess the property of neutralizing acids.

333. **Definition of a Base.**—*A base is a compound body, capable, completely or in part, of neutralizing an acid.*

334. **Definition of an Alkali.**—*An alkali is a base of a specially active character, soluble in water, to which it imparts a a soapy taste and touch.*

335. **Definition of a Salt.**—*A salt may be considered as an acid in which the hydrogen is replaced by a metal.*

Salts are always produced when a base and an acid act upon each other.

If the salt is free from oxygen and sulphur, it is termed a *haloid salt*; if it contains oxygen, it is termed an *oxysalt*; and if this oxygen is replaced by sulphur, it is termed a *sulphosalt.*

336. **Classes of Salts.**—Oxysalts are divided into *normal, acid*, and *basic.*

337. **Normal Sats.**—*A normal salt is one in which the displaceable hydrogen of the acid is all exchanged for an equivalent amount of the metal.* (Arts. 68, 157).

338. **Acid Salt.**—*An acid salt is one in which the displaceable hydrogen of the acid is only partially exchanged for a metal.* (Arts. 66, 147).

These salts often possess an acid reaction, as is the case with $KHSO_4$; but this reaction is dependent on the nature of the acid and the base. Thus, $KHCO_3$ is neutral, whilst others possess a slightly alkaline reaction.

339. **Basic Salts.**—Basic salts are formed by the combination of a normal salt with a hydrate, thus:—White lead, $2PbCO_3 + Pb(OH)_2$.

CHAPTER XIV.

ATOMICITY OR QUANTIVALENCE.

340. **Atomicity.**—We have seen that an atom of chlorine combines with only one atom of hydrogen, (Art. 260), and that one atom of oxygen combines with two atoms of hydrogen, (Art. 29). An atom of chlorine, therefore, may be said to be equivalent to an atom of hydrogen, while an atom of oxygen is equivalent to two atoms of hydrogen. Similar relations may be observed in the other elements and compounds. We have seen also, that an atom of sodium replaces an atom of hydrogen, (Art. 66), and that an atom of zinc replaces two atoms of hydrogen, (Art. 68). This power is called the *atomicity* of the element. The word *quantivalence* is also used to express the same idea. Hence, the elements may be arranged in groups according to the number of hydrogen atoms with which they combine, or which the replace in a compound, thus :—

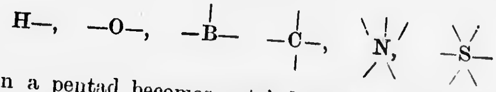
(1) Monads, combine with or replace *one* atom of hydrogen.
(2) Dyads, " " " *two* atoms of hydrogen.
(3) Triads, " " " *three* " "
(4) Tetrads, " " " *four* " "
(5) Pentads, " " " *five* " "
(6) Hexads, " " " *six* " "

341. **Variation in Atomicity.**—The degrees of atomicity of the elements are not invariable; thus, sulphur in SO_3 is a hexad, in SO_2, a tetrad, and in H_2S, a dyad. Nitrogen in NH_4Cl, is a pentad, in NH_3, a triad, and in N_2O, a monad. The circumstances under which the variation takes place cannot well be defined, but the extent or rate of variation is governed by the following very simple law, to which there are only a few exceptions :—

When an element changes its atomicity, either increasing or diminishing, the change is by two degrees at a time.

Thus, a pentad may become a triad or a monad, but not a tetrad or a dyad; so a tetrad may become a dyad, but not a triad or a monad. Nitric oxide, NO, and nitric peroxide, NO_2, are exceptions to this law.

To give a concrete conception to the foregoing facts, the atom of hydrogen may be represented as having only one point of attachment, or *bond*, by which it can be united with any other element; zinc, as having two such bonds; boron, three, and so on. Thus the atoms of the elements may be graphically represented in the following manner:—

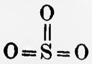


When a pentad becomes a triad, it is assumed that two bonds unite and satisfy each other. The maximum number of bonds of an element is called its *absolute atomicity;* the number of bonds united together, its *latent atomicity,* and the number of bonds actually engaged in linking it with the other elements of a compound, its *active atomicity.* When all bonds are engaged, the compound so formed is more stable than when some of the bonds are latent; thus CO_2 is more stable than CO.

TABLE OF ATOMICITIES.

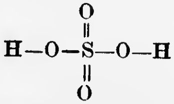
Monads.	Dyads.	Triads.	Tetrads.	Pentads.	Hexads.
Hydrogen	Oxygen.	Boron	Carbon Silicon	Nitrogen Phosphorus Arsenic	Sulphur Selenium Telurium
Fluorine Chlorine Bromine Iodine	Barium Strontium Calcium Magnesium Zinc Cadmium Copper Mercury	Gold	Tin Lead Aluminium Platinum Palladium	Antimony Bismuth	Chromium Manganese Iron Cobalt Nickel Molybdenum Tungsten
Potassium Sodium Lithium					
Silver					

342. Graphic Formulæ. — Graphic formulæ are employed to express the manner in which it is assumed that the constituent atoms of compounds are associated together; for example, sulphur trioxide is regarded as an atom of hexad sulphur with three atoms of dyad oxygen, and this is illustrated by the graphic formula—



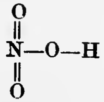
$$O=\overset{\overset{O}{\|}}{S}=O$$

When this oxide is brought in contact with water, it combines with it, forming sulphuric acid, H_2SO_4, and it is supposed that in this compound only two of the oxygen atoms are wholly associated with the sulphur atom; each of the remaining oxygen atoms being united by one of its bonds to the sulphur atom, and by the remaining bond to an atom of hydrogen; thus :—



$$H-O-\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{S}}-O-H$$

The graphic formula of nitric acids is—



$$\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{N}}-O-H$$

Rational formulæ of a much simpler description than these graphic formulæ are generally employed. For instance, sulphuric acid is usually represented by the formula $SO_2(OH)_2$, and nitric acid by the formula $NO_2(OH)$. These rational formulæ, besides expressing the molecular composition of the compounds to which they apply, convey more or less information with regard to the nature of the compound itself, viz. :— the class to which it belongs, the manner in which it is formed, and the behavior it will exhibit under various circumstances. Formulæ which merely express the relative number of atoms present in a compound are termed *empirical formulæ.*

CHAPTER XV.

CHEMICAL EQUATIONS AND MODES OF CHEMICAL ACTION.

343. **Chemical Equations.**—When dilute nitric acid is poured on copper clippings, nitric oxide is given off, and copper nitrate is left in the flask. As no hydrogen is evolved, we infer that the hydrogen, and part of the oxygen of the acid have combined to form water. Hence, without reference to the quantities, the equation is—

$$Cu + HNO_3 = Cu(NO_3)_2 + NO + H_2O, \text{ (Art. 172).}$$

Now since matter is indestructible, the number of atoms of each element must be the same on each side of the equation. By means of this principle the coefficients are determined usually by inspection. When this is too difficult they may be determined as follows :—

Denoting the number of molecules of the factors by a and b and the number of molecules of the product by x, y, z, we have—

$$aCu + bHNO_3 = xCu(NO_3)_2 + yH_2O + zNO.$$

Now, equating the number of atoms on each side of the equation we have—

$a=x$, the coefficients of the atoms of copper.
$b=2y$ " " " hydrogen.
$b=2x+z$ " " " nitrogen.
$3b=6x+y+z$ " " " oxygen.

From these equations we find that—

$$3b = 6a + \tfrac{1}{2}b + b - 2a;$$

$$\therefore b = \frac{8a}{3}.$$

$$x = a.$$

$$y = \frac{4a}{3}.$$

$$z = \frac{2a}{3}.$$

Substituting these values in the equation and dividing by a, we have—

$$3Cu + 8HNO_3 = 3Cu(NO_3)_2 + 4H_2O + 2NO.$$

As another example of a still more difficult character take the ordinary test for a nitrate, viz., add ferrous sulphate to the nitrate, then add sulphuric acid, and a black ring will be formed. (Art. 176.) The products are potassium sulphate, K_2SO_4 ferric sulphate, $Fe_2(SO_4)_3$, water, H_2O, and the black brown liquid, $(FeSO_4)_2NO$. Hence, the equation may be written as follows :—

$$aKNO_3 + bH_2SO_4 + cFeSO_4 = \omega K_2SO_4 + xFe_2(SO_4)_3 + z(FeSO_4)_2NO$$

Equating the coefficients of the molecules on each side of this equation, and for convenience equating the coefficients of the oxygen molecules last, we have—

$a = 2\omega$.... the coefficients of the atoms of potassium.
$a = z$ " " nitrogen.
$2b = 2y$ " " hydrogen.
$b+c = \omega+3x+2z$ " " sulphur.
$c = 2x+2z$ " " iron.
$3a+4b+4c = 4\omega+12x+y+9z$ " " oxygen.

$$\therefore 3a+4b+4(2b+a) = 2a+12\left(b-\frac{a}{2}\right)+b+9a;$$

$$\therefore b = 2a,$$

$$c = 5a,$$

$$\omega = \tfrac{1}{2}a.$$

$$x = \frac{3a}{2},$$

$$y = a,$$

$$z = a.$$

Substituting these values in the equation dividing by a and simplifying, we have—

$$2KNO_3 + 4H_2SO_4 + 10FeSO_4 = K_2SO_4 + 3Fe_2(SO_4)_3 + 4H_2O + 2(FeSO_4)_2NO.$$

344. Modes of Chemical Action.

(1) By the direct combination of the elements or compounds with each other. (Arts. 39, 40, 41, 189, etc.)

(2) By the replacement of one element or group of elements in a body by another element or group of elements. (Arts. 66, 68, 172 etc.)

(3) By an exchange of elements or groups of elements in two or more bodies. (Arts. 125, 153.)

(4) By the resolution of a compound into its elements, or into two or more less complex compounds. (Arts. 32, 34, 123.)

(5) By the re-arrangement of the molecules, as red phorphorus, from the common variety.

QUESTIONS ON CHAPTER XV.

1. What is meant by the atomicity or equivalence of an element? Give the atomicity of the non-metallic elements.

2. What do you mean by absolute, latent, and active atomicity? Illustrate your meaning by examples. Mention some chemical reactions in which the active atomicity of an element undergoes changes.

3. Give the absolute, latent, and active atomicity of each element in the following compounds :—

CO, CO_2, NH_3 H_2S, AS_2O_3, C_2H_4, $MnCl_2$

4. What are empirical, rational, and graphic formulæ? Give examples of each.

5. Draw the graphic formula of the following compounds and mention the atomicity of each element therein :—

Water, nitric acid, sulphurous acid, carbonic acid.

6. Describe some experiments to show the modes of chemical action.

APPENDIX.

SUGGESTIONS FOR THE CONSTRUCTION OF APPARATUS, &c.

345. **Pneumatic Trough.** — An earthenware pan, or wooden box, covered with three or four coats of paint, will answer the purpose very well. Two pieces of wood of equal thickness, loaded with lead, or a flower-pot saucer may be used for the shelf. If the latter, a hole may be made in the centre by chopping away the material with a file, taking care that the blows of the hammer are light. The notch at the side may be made in the same way.

346. **Retort Stand.**—The base and upright may be made of hard wood instead of iron. The upright passes through two large corks, around which passes copper wire as in Arts. 35, 66, 73. The ends of one of these wires can be made into rings for supporting flasks and retorts, while the ends of the other wire may be twisted round the flask to keep it steady. These wires will be found far more useful than the ordinary rings.

347. **Cork-borers.**—The tube of a steel pen may be made to answer, but will not last long. Holes may be made in corks by a rat-tail file alone, the file being worked into the cork like a gimlet.

348. **Deflagrating Spoon.**—The bowl of a tobacco pipe, partially filled with Plaster of Paris will answer pretty well. A piece of copper wire should be twisted round the bowl, and the end bent up and passed through a piece of tin on which a cork is fixed with Plaster of Paris.

349. **Funnel-tubes.**—In place of these, ordinary funnels attached to glass tubes by India-rubber tubing or by corks, may be used.

350. Spirit-lamp.—A small bottle fitted with a cork, through which passes a bit of glass tubing or barrel of a steel pen, will make a good spirit-lamp. A nick is made in the side of the cork to admit air. A short wide test-tube is fitted over the cork to prevent evaporation when the lamp is not in use.

351. Bottles.—All reagents should be kept in well stoppered bottles, those containing liquids having glass stoppers. These may be of 200 to 250 c.c. (6 to 8 oz.,) capacity for acids, ammonia etc. Dry reagents should be kept in small, 30 to 60 c.c. (1 to 2 oz.) wide-mouthed bottles. Every bottle should be plainly labelled with the name and formula of the substance contained in it.

352. Bottles, pouring from.—In pouring from a bottle, hold the bottle in the right hand with the label towards the palm, remove the stopper with the third and fourth fingers of the left hand, and pour the liquid into the test-tube held between the thumb and index finger of the same hand; then catch the drop adhering to the lip of the bottle, by touching it with the stopper, and replace the latter in its place.

353. Corks.—All corks should be softened by rolling under a board upon the table, or rolling under the foot upon the clean floor. Steaming also softens the hardest corks. If the cork is to be used in an apparatus for chlorine, or other corrosive gas, it must be soaked in melted paraffine to close all the pores.

354. Glass, Manipulation of.—To bend a tube, hold it in the lamp flame, turning it constantly to secure uniform heating till it begins to yield at the heated point, then remove it from the flame, and slowly bend it to the required angle. If the tube is large and thin walled, or if it is to be bent through a large angle, it should be bent at several contiguous points in succession, till the desired angle is obtained.

Glass tubing is cut by making a sharp scratch with the edge of a file at the desired point, and then steadily drawing the two portions apart, and at the same time, applying the lateral pressure as if attempting to bend the tube where the file mark was made.

To close a tube at one end, heat it as before, and removing it from the flame, draw apart the portions on each side of the heated point; finally, melt down the pointed ends in the flame; on blowing in at the open end of the tube while the other end is red hot, a bulb will be made on that end.

To cut glass, take an iron rod about 10 inches long, and $\frac{1}{4}$ inch in diameter, and insert one end in a wooden handle. With a pointed piece of soap trace a line where the cut is to be made, and make a notch in the glass with a file. Heat the iron red-hot and apply it about 1mm. from the end of the notch, till a fine crack begins to creep towards the rod; then slowly move the rod along the traced line and the crack will follow it. If it is a bottle that is to be cut, apply the rod to one end of the notch for about 5 seconds, and then quickly apply it to the other end and hold it till the crack begins to creep towards it, and then move it as before.

To bore a hole in glass, make a thick paste by partially dissolving camphor in turpentine; nip off a short piece from a rat-tail file, and keeping the ragged end wet with the paste, a hole can be bored by strong pressure and a twisting motion.

355. **Litmus Solution.**—Boil the litmus in water, adding more water from time to time to replace what is boiled away; add a few drops of sodium hydrate solution, and a drop or two of carbolic acid to prevent decomposition.

356. **The Metric System**—For an explanation of this system, see the Canadian edition of Hamblin Smith's Arithmetic. The following rules are useful in chemistry:—

A centimetre = ·392708 in. = $\frac{4}{10}$ in. nearly.
Hence, to convert centimetres into inches,
Multiply by 4 and reject the last figure.
A millimetre = ·0393708 in. = $\frac{4}{100}$ in. nearly.
Hence, to convert millimetres into inches,
Multiply by 4 and point off the last two figures.

eral pressure
k was made.
removing it
side of the
n the flame;
e other end

long, and ¼
ndle. With
to be made,
he iron red-
notch, till a
lowly move
w it. If it
of the notch
e other end
it, and then

partially dis-
iece from a
the paste, a
motion.
ater, adding
oiled away;
drop or two

tion of this
Arithmetic.

TABLE OF ELEMENTARY BODIES WITH THEIR SYMBOLS AND ATOMIC WEIGHTS.

The first column of numbers contains those used in this book. The second contains those used in Roscoe's Chemistry.

Names of Elements.	Symbols.	At. Wt.	Roscoe.	Names of Elements.	Symbols	At. Wt.	Roscoe.
Aluminium	Al	27·5	27.3	Molybdenum	Mo	96	95·6
Antimony	Sb	122	122·0	Nickel	Ni	59	58·6
Arsenic	As	75	79·9	Niobium	Nb	99	94·0
Barium	Ba	137	136·8	Nitrogen	N	14	14·01
Bismuth	Bi	210	210·0	Osmium	Os	199	198·6
Boron	B	11	11·0	Oxygen	O	16	15·96
Bromine	Br	80	79·75	Palladium	Pd	106·5	106·2
Cadmium	Cd	112	111·6	Phosphorus		31	30 96
Cæsium	Cs	133	133 0	Platinum	K	197·1	196·7
Calcium	Ca	40	39·9	Potassium	R	39·1	39·04
Carbon	C	12	11·97	Rhodium	Ro	104·3	104·1
Cerium	Ce	92	141.2	Rubidium	Rb	85·4	85·2
Chlorine	Cl	35·5	35·37	Ruthenium	Ru	104·2	103·5
Chromium	Cr	52·5	52·4	Selenium	Se	79·5	78·0
Cobalt	o	59	58·0	Silicon	Si	28	28·0
Copper	Cu	63·4	63·0	Silver	Ag	108	107·66
Didymium	Di	96	147·0	Sodium	Na	23	22 99
Erbium	E	112·6	169·0	Strontium	Sr.	87·6	87·2
Fluorine	F	19	19·1	Sulphur	S	32	31 98
Gallium	Ga			Tantalum	Ta	182	182 0
Gold	Au	196 6	196·2	Tellurium	Te	129	128·0
Hydrogen	H	1	1·0	Thallium	Tl	203·6	203·6
Indium	In	113·4	113·4	Thorinum	Th	115·7	231·5
Iodine	I	127	126·53	Tin	Sn	118	117·8
Iridium	Ir	198	196·7	Titanium	Ti	50	48·0
Iron	Fe	56	55·9	Tungsten	W	184	184·0
Lanthanum	La	90·2	139·0	Uranium	U	120	240·0
Lead	Pb	207	206·4	Vanadium	V	51·3	51·2
Lithium	L	7	7·01	Yttrium	Y	59·7	93·0
Magnesium	Mg	24	23·94	Zinc	Zn	65	64·9
Manganese	Mn	55	54·8	Zirconium	Zr	89·5	90·0
Mercury	Hg	200	199·8				

LIST OF CHEMICALS.

The quantities here given are such as will enable the teacher or student to perform the experiments in this work several times. Substances which may be readily obtained, such as sugar, starch, marble, etc., do not appear on the list :—

Alcohol, methylated	12 ozs.
Ammonia	8 ozs.
Ammonium nitrate	2 ozs.
" chloride	1 oz.
Arsenious oxide	¼ oz.
Alum	½ oz.
Barium chloride, a few grains.	
Bleaching powder	2 ozs.
Calcium chloride	2 ozs.
Copper sulphate	¼ oz.
" turnings	4 ozs.
Charcoal, wood, in lumps	1 oz.
" animal	4 ozs.
Carbon disulphide	¼ oz.
Carbolic acid	¼ oz.
Ferrous sulphate	½ oz.
Flour-spar	½ oz.
Gold-leaf, 2 square inches	
Hydrochloric acid	1 lb.
Iron filings	4 ozs.
Indigo solution	2 ozs.
Lead acetate	½ oz.
Lime	1 lb.
Litmus	⅛ oz.
Manganese dioxide	8 ozs.
Magnesium wire	4 ft.
Mercuric oxide	½ oz.
Mercuric chloride	¼ oz.
Nitric acid	4 oz.
Oxalic acid	1 oz.
Phosphorus	1 oz.
Potassium	Small bottle.
" iodide	½ oz.
Potash caustic	½ oz.
Potassium permanganate	⅛ oz.
" nitrate	4 ozs.
" bromide	⅛ oz.
" chlorate	8 ozs.
" tartrate	1 oz.
(*Cream of tartar.*)	
Pyrogallic acid	¼ oz.
Plaster of Paris	1 lb.
Paraffine	2 ozs.
Sulphuric acid	1½ lbs.
Sodium	¼ oz.
Sodium bicarbonate	4 ozs.
Sulphur	4 ozs.
Silver nitrate	⅛ oz.
Zinc granulated	1 lb.

APPARATUS.

The following list includes the apparatus required for performing the experiments described in this work. The principal articles of steady consumption are glass-tubing flasks, test-tubes, corks, and india rubber tubing. Many of the other articles once obtained, last a long time. Other apparatus may be added with great advantage, but it is better not to begin with too much as the beginner is almost sure to order many thing which he will not really require.

6 bottles for gas receivers, 4 of 1 pint, and 2 of 1 quart.
A set of 5 beakers.
A Bunsen burner, if gas is used.
3 dozen corks assorted.
Set of 3 cork borers, with iron rod.
12 feet copper wire, medium size.
2 strips of copper foil, 4 inches by ½ inch.
A Deflagrating spoon.
6 flasks from 1 litre to ¼ litre; 2 strong, flat bottom.
2 Funnels, 2 inches and 3 inches diameter.
2 funnel-tubes.
2 strips of zinc 4 inches by ½ in.
Triangular file, with handle.
Round file, 5 inches long, with handle.
Filter paper.
1 lb. glass tubing assorted.
Graduated glass measure, 50 cu. cent.
6 assorted ground glass plates.
Galvanic battery, 2 cells, bichromate of potash.
4 feet india rubber tubing, $\frac{3}{16}$ inch and ¼ inch.
Small mortar and pestle.
6 inch platinum wire.

Platinum foil, 2 inch by 1 inch.
2 Retorts.
1 Retort Stand.
A spirit lamp, 4 ounces capacity.
Pair of small scales, with gram weight, 4,2,5,10,20 and 50.
A strong soda water bottle.
1 doz. test-tubes 6 inch by ¾ inch.
A piece of wire gauze 5 inches square.

INDEX.