An Introduction to Practical Chemistry

BY

G. B. NEAVE, M.A., D.Sc.

AND

J. WATSON AGNEW, F.I.C.

Lecturers and Demonstrators in the Department of Chemistry,
Glasgow and West of Scotland Technical College

BLACKIE AND SON LIMITED
50 OLD BAILEY, LONDON
GLASGOW AND BOMBAY
1911

PREFACE

This little book has been written primarily for the use of First Year Students of Chemistry in this College, before they proceed to a course of qualitative and quantitative analysis, but it may also serve as a guide for those who wish to have only an elementary knowledge of the subject.

The experiments, which have been selected from various sources, are such as, in our opinion, are most instructive and at the same time fairly easily performed by a student of average ability. It may be thought that the explanations of experiments are in many cases very full, but our experience has shown that it is not a wise thing to leave too much to the student, because from many experiments he can, and often does, draw erroneous conclusions.

We desire to express our thanks to Mr. J. F. Smith for the care he has taken in preparing the illustrations, and also to Professor G. G. Henderson for the interest he has shown during the preparation of the book, and for a number of valuable suggestions.

G. B. NEAVE.
J. WATSON AGNEW.

Department of Chemistry,
Technical College,
Glasgow.

CONTENTS

INTRODUCTORY

Before performing an experiment, the student must carefully read and fully understand the directions.

When the experiment has been completed, he should write an account of it, *in his own words,* in *a scroll notebook, and state any conclusions which* can be drawn from it. The report should be written afterwards in his Laboratory Report Book.

Experiments marked with an asterisk are to be performed only by students taking a three or four years' course of chemistry.

Experiments marked [D] must be performed in a draught-chamber.

PRACTICAL CHEMISTRY

PRELIMINARY DIRECTIONS

Cutting of Glass Tubing

Glass tubing or rod is cut by laying it on the bench and making a deep scratch at the required place with a triangular file. The glass is then held in both hands, and a pull and a gentle pressure exerted on it as in the attempt to break it.

The sharp edges of a freshly cut tube must always be rounded before use. This can be done by heating in a Bunsen flame until the glass softens.

Bending of Glass Tubing

The tube is held by both hands in the upper part of the ordinary *fish-tail* flame, so as to heat about 2 in. of the glass. In order to heat all sides equally the tube is slowly rotated on its axis. In a few minutes the glass becomes soft and is removed from the flame and bent to the required angle. After cooling, the soot may be removed by a piece of cloth or paper.

N.B.—The Bunsen flame must never be used for bending glass tubing.

Boring of Corks

A cork borer—a metal tube sharpened at one end—is selected, of a diameter slightly less than that of the tube which is to be inserted in the cork. The latter, after being softened in a cork-press, is held firmly against a wooden surface, and the hole made by gently pushing the borer through it with a constant rotating movement. A round file may be used to smooth the walls of the hole or, if necessary, to enlarge it. Care must be taken that the glass tube fits the hole accurately. The cutting is facilitated if the borer is moist.

Heating of Glass and Porcelain Vessels

In heating a liquid in a vessel, care must be taken that the flame does not touch the vessel above the surface of the liquid.

Porcelain dishes, which are generally used for boiling and evaporating liquids, are supported by wire gauze on a tripod.

Porcelain crucibles are employed when solid substances are to be strongly heated, and are supported by a triangle on a tripod. They are heated by a naked flame, at first gently, then more strongly. The crucible must always be allowed to cool slowly. Crucible tongs are used for lifting hot crucibles and lids.

Glass vessels must always be heated cautiously. A test tube is heated in the Bunsen flame, but to prevent local heating it is kept continually in motion.

Flasks and **beakers** are heated by placing them on wire gauze on a tripod over a Bunsen, or sometimes on a sandbath.

The Measuring of Liquids

In order to gain an idea of volume, and also to graduate for himself some vessels, the student should first carry out the following exercises:—

1. The capacity in cubic centimetres of the flasks and beakers in his stock of apparatus is determined. This is best done by filling each vessel with water and pouring into a measuring cylinder, care being taken always to read the lower edge of the meniscus. The volume of each vessel may be marked by means of a writing diamond.

2. Volumes of 20 cub. cm. and 50 cub. cm. are marked on a small beaker, and 5 cub. cm., 10 cub. cm., and 15 cub. cm. on a test tube.

PHYSICAL AND CHEMICAL CHANGE

Experiment 1.—A small quantity of mercury—about the size of a pea—is heated in a test tube. The liquid boils and is converted into a vapour, which condenses on the upper and colder part of the tube.

It is evident that mercury can exist in the two states—liquid and vapour.

Experiment 2.—A small quantity of sulphur is heated as in Experiment 1.

It will be seen that sulphur can exist in three states—solid, liquid, and gas.

Experiment 3.—Two or three crystals of iodine are very gently heated in a test tube, the latter being held some distance above the flame. The iodine does not melt, but is converted directly into a gas, which condenses to a solid on the colder part of the tube.

The process of conversion from solid to gas and then to solid is known as "sublimation", and the substance is said to "sublime".

Experiment 4.—A small piece of platinum wire or foil is heated in the Bunsen flame, and the changes, if any, carefully noted.

In these four experiments it has been observed that heat produces certain changes in the substances investigated. These changes are, however, not permanent. On removal of the source of heat, the substances are restored to their original state or condition. The changes are said to be **physical.**

Experiment 5.—A piece of magnesium ribbon about 2 in. long is held by means of crucible tongs in the flame. The magnesium burns with a white, dazzling light, and the white powder which results does not resemble the original magnesium. This change is quite unlike those which occurred in the first four experiments. The magnesium has been converted into an entirely different substance, and the change is said to be **chemical.**

In a **physical change** the substance merely changes its state or condition, while in a **chemical change** a different substance is formed.

Experiment 6.—A small quantity of mercury is mixed with a small quantity of flowers of sulphur and the mixture heated in a test tube. There results a black substance, which resembles neither mercury nor sulphur. The change, therefore, is chemical. Two substances have combined to form a new substance, a process which is known as "synthesis".

Experiment 7.—A small quantity of powdered lead nitrate, which has been dried by pressing between sheets of filter paper, is heated in a test tube.

From one substance two are obtained, a reddish-brown gas and a yellow solid. The lead nitrate is said to be "decomposed", the process being known as "decomposition".

Experiment 8.—The action of heat on water, sodium chloride, copper carbonate, and mercuric oxide is next observed, and in each case the student should note the changes which occur and state whether they are chemical or physical.

ELEMENTS AND COMPOUNDS

Definition of a Compound.—A substance which can be decomposed into two or more simpler substances is called a compound. When two or more substances combine to form a third substance the product is a compound.

Experiment 9.—The student should write down a list of the substances already met with, which, in his opinion, are compounds.

Experiment 10.—The action of heat on sugar and copper nitrate is observed in order to decide, if possible, if these substances are compounds.

Definition of an Element.—A substance which cannot be decomposed by any known method is called an element.

A list of the elements is given in the Appendix, p. 94.

MIXTURES AND PURE SUBSTANCES OR COMPOUNDS

Experiment 11.—A small quantity of flowers of sulphur is intimately mixed with about half its bulk of fine iron filings, and the mixture divided into three equal portions.

(*a*) The first portion is shaken in a test tube with about 10 cub. cm. of water. The iron, having a specific gravity much greater than that of water and than that of sulphur, sinks quickly, so that a partial separation of the iron and sulphur can be effected.

(*b*) The second portion is placed in a dry test tube and a little carbon disulphide added. [As carbon disulphide catches fire very easily, it must always be kept at a distance from the flame.] On shaking the test tube, it will be seen that the sulphur disappears, being dissolved by the carbon disulphide, while the iron remains unaffected. Thus, again, the iron and sulphur are separated.

(*c*) The third portion is heated in a dry test tube, and the changes, if any, are noted. The end of the test tube is then broken in a mortar, and attempts made, as with the first and second portions, to separate the iron and sulphur.

The components of a mixture can be mechanically separated, while a compound can be decomposed only by a chemical method.

THE BALANCE

The construction and use of the balance will be explained by a demonstrator.

The following instructions and rules must be carefully observed:—

1. The case and pans of the balance must be kept scrupulously clean.

2. When the beam is released it should rest properly on the agate knife edges, and should swing evenly.

3. The object whose weight is to be determined is placed on the left-, and the weights on the right-hand pan.

4. Substances to be weighed must not be placed directly on the pans, but on a watchglass or dish whose weight has been previously found. The substance after being weighed must not be carried from the balance room on an open watchglass, but in a desiccator, or on a watchglass covered by another watchglass.

5. Deliquescent or volatile substances are weighed in a weighing bottle, and a sufficient quantity is shaken out into a beaker, or other suitable vessel, and the weighing bottle and contents again weighed. The difference in weight is the weight of the quantity taken.

6. The beam must be at rest when anything is placed on or removed from the pans.

7. The weights must always be lifted by the forceps, never by the hand.

8. Crucibles, dishes, &c., must always be quite cold before being weighed.

OXYGEN

Oxygen exists in the atmosphere to the extent of one-fifth by volume, but it is not a simple matter to obtain the gas from that source. For laboratory purposes it can be obtained by heating certain compounds containing oxygen, for example, mercuric oxide.

Experiment 12.—A small quantity of mercuric oxide is placed in a test tube, and heated in the Bunsen flame. The mercuric oxide first of all turns black, and thereafter globules of mercury are deposited on the upper part of the tube. A glowing splinter of wood inserted into the tube is rekindled, owing to the presence of the gas oxygen. The mercuric oxide on heating is decomposed into the liquid metal mercury and oxygen:

$$2\,HgO = 2\,Hg + O_2.$$

Preparation of Oxygen

Experiment 13.—To prepare oxygen in quantity 15 grm. of potassium chlorate are mixed with about 5 grm. of manganese dioxide. The mixing should be done on a sheet of paper with a spatula, since oxidizing agents, like potassium chlorate, must never be mixed in a mortar. The mixture is then loosely packed in a hard-glass test tube, leaving a clear space between the mixture and the edge of the tube. The tube is then fitted with a cork and delivery tube for collecting the evolved oxygen over water (fig. 1). On gently heating the mixture from the front to the back of the tube the oxygen is given off, and after all air has been displaced from the apparatus four gas-collecting bottles are filled with the gas, each bottle being closed with a greased glass plate when filled. At the end of the experiment the delivery tube must be removed from the trough before the heating is stopped. The soluble portion of the residue is dissolved in water, the solution filtered and tested for chloride with silver nitrate (see Experiment 88):

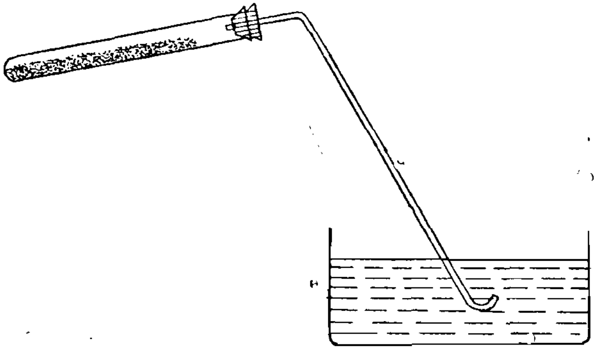

Fig. 1

$$2\,KClO_3 = 2\,KCl + 3\,O_2.$$

The student should note that there is no fusion of the potassium chlorate, and that the oxygen is evolved with gentle heating. The presence of the manganese dioxide causes the decomposition of the potassium chlorate to take place at a much lower temperature than if the potassium chlorate were used alone. The student can confirm this when carrying out Experiment 18.

An action of this kind is known as a "catalytic action", and the manganese dioxide is the "catalyst" or "catalytic agent", and can be recovered unchanged at the end of the experiment.

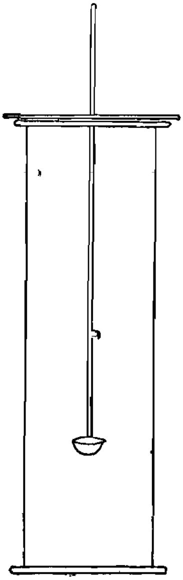
Fig. 2

Experiment 14.—About 5 cub. cm. lime water are placed in the first bottle containing oxygen, the greased plate is immediately replaced, and the contents vigorously shaken. The lime water is unaffected by the oxygen. A piece of wood charcoal is now placed in a deflagrating spoon (fig. 2), whose brass cover has been adjusted so that the spoon when placed in the gas bottle will be about 2 in. from the foot of the bottle. The spoon is now placed in the Bunsen until the charcoal glows, and then quickly placed in the bottle containing oxygen. The charcoal burns more brightly in oxygen than in air. When the combustion has ceased, the spoon is removed, and the bottle is immediately covered with the greased plate and vigorously shaken. The lime water will turn turbid, owing to the presence of the gas carbon dioxide, which has been produced by the combustion of the charcoal in oxygen:

$$C + O_2 = CO_2$$
$$CO_2 + Ca(OH)_2 = CaCO_3 + H_2O.$$

Experiment 15 [D].—A little sulphur is now placed in the deflagrating spoon, ignited in the Bunsen, and placed in the second jar of oxygen. Again the combustion is more vigorous than in air. On shaking up the products of combustion with a little water coloured with litmus, the latter is turned bright red, showing the presence of an "acid". The sulphur first of all burns to form the gas sulphur dioxide:

$$S + O_2 = SO_2.$$

The sulphur dioxide is then dissolved by the water, giving a solution of sulphurous acid:

$$SO_2 + H_2O = H_2SO_3.$$

Experiment 16 [D].—The spoon is now carefully cleaned, and a small piece of phosphorus is burned in the same manner as the sulphur. The phosphorus must be cut under water, transferred to a piece of filter paper with the crucible tongs, and, after drying, placed in the spoon by means of the tongs. The phosphorus should not be removed from the water until it is required, and must on no account be touched with the fingers, as the heat of the hand is sufficient to ignite it. In this case the combustion is attended by the production of copious white fumes. On shaking up with water this white solid dissolves, and when tested with litmus the solution has an acid reaction. The phosphorus first of all burns to form phosphorus pentoxide:

$$P_4 + 5\,O_2 = 2\,P_2O_5.$$

The phosphorus pentoxide then dissolves in the water, forming phosphoric acid:

$$P_2O_5 + H_2O = 2\,HPO_3.$$

Experiment 17.—A piece of sodium[1] about the size of a pea is cut off with a penknife, carefully dried between filter paper, and placed in a deflagrating spoon, which is heated until the sodium melts, and then placed in the fourth bottle of oxygen. The sodium burns brilliantly, and when the residue is dissolved in water and tested with litmus a blue coloration is produced, showing the presence of an "alkali". The first product of combustion is sodium oxide, which dissolves in the water, with formation of the alkali sodium hydroxide. The latter is an example of a class of substances known as "bases". A "base" is an oxide or hydroxide of a metal.

From these experiments it is seen that elements combine with oxygen to form oxides. Some of these oxides combine with water to form "acids", others to form "bases". When an acid acts on a base, water and a "salt" are formed.

*** Experiment 18.—Determination of the Volume of Oxygen which can be obtained by heating Potassium Chlorate.**—A hard-glass tube is accu-

---

[1] In this case also, the sodium must not be touched with the fingers, and the filter paper must be perfectly dry, since sodium reacts violently with water.

rately weighed. About 2 grm. potassium chlorate are then placed in the tube, and the weight of the tube and potassium chlorate is again accurately found. The apparatus shown in fig. 3 is then fitted up. *a* is the hard-glass test tube containing the potassium chlorate and fitted with a cork, through which passes the delivery tube *g*, connected by rubber tubing to the tube leading to the 2-litre bottle *b*. Through the cork in the 2-litre bottle there also passes the tube *d*, which reaches to the bottom of the bottle, while the other end is connected to the rubber tubing *f* leading to the beaker *c*. On the rubber tubing is a screw clip *h*. The bottle *b* is filled with water, and the apparatus connected together. The apparatus must now be tested to see if it is perfectly air-tight, and this is done by gently

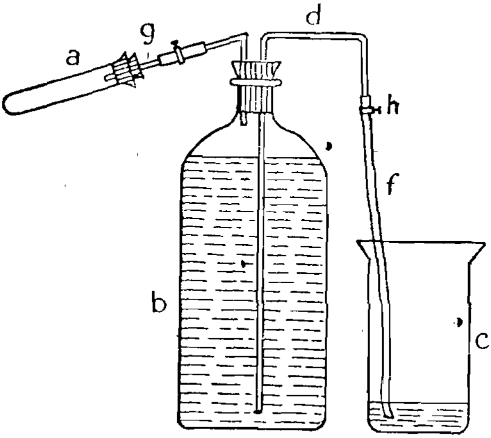

Fig. 3

sucking the tube *f*. If no water runs out, then the experiment may be proceeded with; but should there be a flow of water the leak must be found and remedied.

The potassium chlorate is first gently heated, and finally at the full strength of the Bunsen, until no more gas is evolved. The water which is expelled from the bottle by the evolution of the oxygen is collected in the beaker. When all the oxygen has been evolved the tube is allowed to cool, while the rubber tube dips below the surface of the water in the beaker. The object of this is to allow water to flow back from the beaker to the bottle as the tube cools down. When the tube is quite cold the levels of the water in the bottle and beaker are adjusted by raising the latter. When they are the same the screw clip is closed, and the volume of water in the beaker is accurately measured. At the same time the barometric pressure and the temperature of the water are noted. The test tube is now removed, and the delivery tube is closed by a piece of rubber tubing about 3 in. in length fitted with a screw clip. The weight of the test tube and potassium chloride is now found. The oxygen is kept for the next experiment.

The following data have now been obtained:—

Weight of tube empty =
Weight of tube + potassium chlorate =
Weight of potassium chlorate =

After heating—

Weight of tube + potassium chloride =
Weight of tube =

Weight of potassium chloride =
Weight of oxygen = weight of potassium
chlorate − weight of potassium chloride.

Volume of water in beaker = V cub. cm.
Barometric pressure = P mm.
Temperature of water = $t°$ C.
Tension of aqueous vapour at $t°$ = p mm.
(See Tables, p. 95.)

The volume of water collected in the beaker is equal to the volume of oxygen evolved by heating the potassium chlorate. This volume must now be reduced to normal temperature and pressure—0° cm. and 760 mm. barometric pressure. This is done by employing the formula:

$$\frac{PV}{T} = \frac{P'V'}{T'}, \text{ or } V' = \frac{V \times (P - p) \times 273}{760 \times (273 + t)},$$

where P, V, and T are the original pressure, volume, and absolute temperature, and P′, V′, and T′ the final pressure, volume, and absolute temperature. The absolute temperature = temperature Centigrade + 273.

From the weight of potassium chlorate taken a certain volume of oxygen has been obtained, then the volume which can be obtained from 100 grm. of potassium chlorate is calculated. The weight of a litre of oxygen is also to be calculated from the data found in the experiment, the loss of weight of the tube being the weight of the volume of oxygen collected.

***Experiment 19.—Determination of the Weight of a Litre of Oxygen by Direct Weighing.**—A 500 cub. cm. round-bottomed flask (fig. 4) is fitted with a tight-fitting rubber cork, through which passes a piece of glass tubing about 3 in. in length. The lower end of the tube is flush with the cork, and the upper end is closed with rubber tubing and a screw clip. The cork is now fitted tightly into the flask, and the position to which it sinks is marked on the neck with a file. This is done so as to ensure that the cork will always be thrust in to the same distance. The flask is now completely filled with water and the cork is pushed in to the mark, the clip being left open. The clip is now closed, and, after removing any adhering drops of water, the contents of the flask are poured into a measuring cylinder and the volume of water is accurately noted. This gives the capacity of the flask. About 50 cub. cm. of water are placed in the flask and the cork is put in to the mark, leaving the clip open. The water in the flask is now boiled, with the flame of the Bunsen entirely under the level of the water in the flask. After the water has boiled for about five minutes, the clip is closed, and at the *same time* the Bunsen is removed. When the flask is cold and dry it is suspended by means of a wire to the hook on the balance and weighed. The suspension wire must

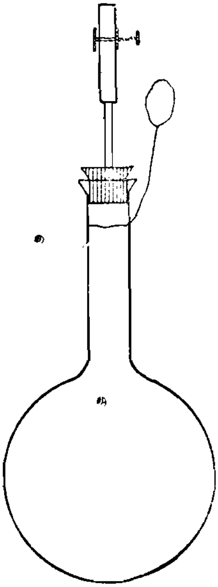

Fig. 4

be kept on the flask till the end of the experiment.

The flask has now to be filled with oxygen, and this is done from the supply in the aspirator. The end of the outflow tube from the aspirator is again placed in a beaker of water. The rubber tubing and clip are removed from the other tube of the aspirator and the tubing on the flask immediately slipped on. The screw clips are opened and the beaker raised until the level of the water in the beaker is above the level of the water in the aspirator, when both screw clips are closed. The flask is removed from the aspirator and the screw clip is momentarily opened, in order to allow the oxygen in the flask to come to atmospheric pressure. The flask and its contents are now accurately weighed, and the temperature of the balance-room and the barometric pressure are carefully noted. Finally the water which still remains in the flask is measured by pouring into a small measuring cylinder.

The volume of the oxygen will be the volume content of the flask, less the volume of water remaining in the flask at the end of the experiment. Now, besides oxygen, we also have present water vapour in the flask, and the weight of this water vapour must be deducted from our weight of oxygen. In practice it is more convenient to assume that the presence of the water vapour affects the barometric pressure, and this effect has been calculated for all ordinary temperatures. The value of this pressure of aqueous vapour at the temperature of the balance-room (see table of vapour pressure of water, p. 95) is deducted from the barometric pressure, and the difference is taken as the pressure under which the oxygen is measured. The volume of the oxygen is now reduced to normal temperature and pressure. The difference in the weight, before and after admission of the oxygen to the flask, is the weight of the oxygen. The volume which this weight of oxygen occupies at normal temperature and pressure having been already found, the weight of 1000 cub. cm. of oxygen can now be calculated.

HYDROGEN

Hydrogen occurs in nature combined with oxygen in the compound water, and can be prepared from water:—

1. By electrolysis of acidulated water.
2. By the action of some metals on water or steam.

$$\begin{aligned} 2\,Na + 2\,H_2O &= 2\,NaOH + H_2 \\ Mg + H_2O &= MgO + H_2 \\ 3\,Fe + 4\,H_2O &= Fe_3O_4 + 4\,H_2. \end{aligned}$$

Since hydrogen is an essential constituent of all acids, it can be prepared from an acid. In the laboratory the gas is usually prepared by the action of a metal on dilute sulphuric acid or hydrochloric acid.

Experiment 20.—Some iron filings are placed in a test tube, and a few cubic centimetres dilute hydrochloric acid are added. A vigorous reaction begins, and the effervescence shows that a gas is being evolved. On placing the mouth of the test tube to a flame it is seen that the gas is combustible. The gas evolved is hydrogen, and the liquid remaining in the tube is ferrous chloride solution.

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

Preparation of Hydrogen

Experiment 21.—A 300 cub.-cm. Erlenmeyer flask is taken and about 25 grm. granulated zinc are placed in it. The flask is then fitted with a two-holed cork, carrying a thistle funnel and a delivery tube (fig. 5). The thistle funnel must reach almost to the bottom of the flask, and the delivery tube is to be bent for collecting the gas over water. Great care must be taken that the apparatus is perfectly air-tight, and should there be any leakage a fresh cork must be taken. When the apparatus is quite ready, *dilute* sulphuric acid is poured down the thistle funnel until the zinc and the end of the thistle funnel are covered. Should the evolution of the gas slacken during the subsequent experiments a further

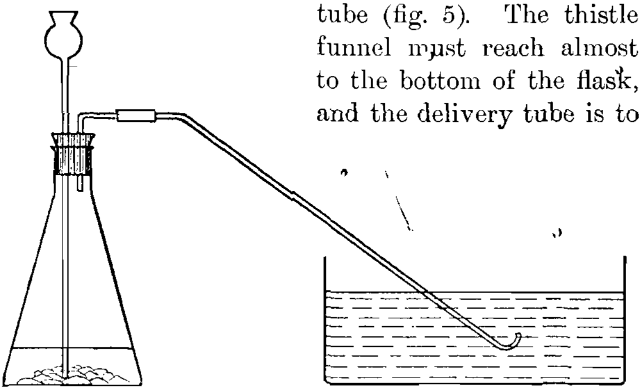

Fig. 5

quantity of dilute sulphuric acid is added. When sufficient time has been given for the expulsion of the air in the apparatus, a test tube is filled with water and then with hydrogen by displacement over water. When the tube is full of the gas, the end is closed with the thumb and the tube is removed to a flame. If the gas in the tube burns quietly, then all air has been displaced from the apparatus and the gas may now be collected in the bottles. If, however, there is a slight explosion, the gas must be tested again until it is shown to be free from air. Two bottles of the gas are collected.

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

Experiment *22.*—After collecting two bottles of the gas, the delivery tube is removed from the apparatus and a calcium chloride drying tube, fitted with a piece of hard-glass tubing *a*, is substituted

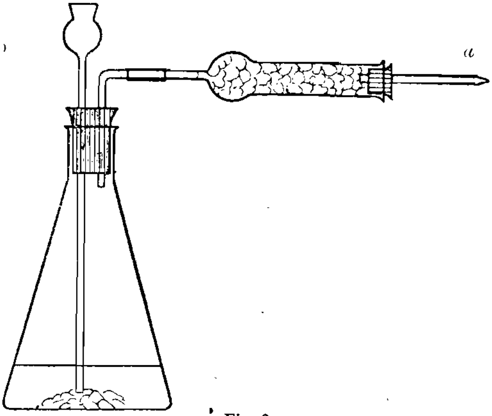

Fig. 6

(see fig. 6). The gas issuing from the calcium chloride tube is collected by upward displacement in a test tube, and tested until *perfectly* pure. When pure the gas is ignited, and a cold dry beaker is held over the flame for a few seconds. A slight film will be noticed on the inside of the beaker. This film is water, and is produced by the combination of the hydrogen with the oxygen present in the air.

$$2\,H_2 + O_2 = 2\,H_2O.$$

Experiment 23.—One of the bottles containing the gas is held mouth downwards, and on slipping away the cover a lighted taper is pushed right up into the bottle. The taper will be extinguished but the gas will burn at the mouth of the bottle, showing that hydrogen is not a supporter of combustion but is combustible.

Experiment 24.—The second bottle of the gas is placed mouth upwards on the bench and the glass plate is removed. At the end of a short time a taper is thrust into the bottle and it will be seen that the taper continues to burn and that there is no sign of the gas in the jar burning. Hydrogen being a very light gas has rapidly escaped out of the bottle, and the gas in the bottle is now air.

Experiment 25. — Preparation of Zinc Sulphate. — When all action has ceased in the flask, the contents are filtered, the filtrate being caught in a porcelain basin. The contents of the basin are now slowly evaporated, until the volume is about 30 cub. cm. The basin is then set aside, and on cooling, white crystals of zinc sulphate will separate out, which may be filtered off and dried between filter paper.

WATER

Pure water is a colourless, odourless liquid. Natural waters are never pure. Rainwater has dissolved some of the gases in the air—nitrogen, oxygen, and carbon dioxide. Water which rises from a spring always contains some soluble salts which it has dissolved when passing through the soil. Water can be purified from suspended solids by filtration and from dissolved solids by distillation.

Purification of Water by Distillation

Experiment 26.—A crystal of potassium dichromate is dissolved in about 200 cub. cm. of water, and the resulting solution is placed in a flask, fitted with a cork and long delivery tube (see fig. 7). On boiling

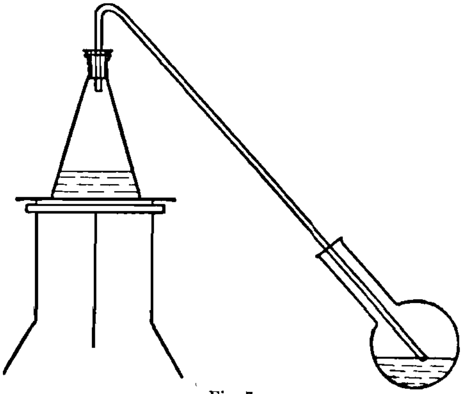

Fig. 7

the water in the flask it is converted into vapour, which passes into the tube where it "condenses" to the liquid state and runs into the "receiver" at the end of the tube. After collecting about 50 cub. cm. of the "distillate" the "distillation" is stopped. The distillate is perfectly colourless and only contains gaseous impurities, all the solid impurities remaining in the water in the distilling flask. The distillate is now evaporated to dryness in a porcelain evaporating basin, when no residue is obtained, showing that the water has been freed from dissolved solids by distillation.

Solution

It is well known that when certain substances, such as salt and sugar, are mixed with water, they change their appearance and state; they are said to be dissolved by the water. The name "solution" is given to the process and to the resulting mixture. The liquid is called the "solvent" and the dissolved substance the "solute". The solute and solvent can be separated again by distillation, the residue which remains in the distilling flask being the solute, and the liquid which collects in the receiver, the solvent.

Comparison of Solubilities

Experiment 27.—Small quantities of copper sulphate, sand, and potassium nitrate are shaken with distilled water in separate test tubes. The liquids are then filtered through separate filters into evaporating basins. On evaporating the contents of the basins to dryness, those basins which contain a residue have evidently contained a solution of a soluble substance. Where there is no residue, the substance is insoluble in water.

Experiment 28.—Some substances which are insoluble, or sparingly soluble, in water are readily soluble in other liquids. For example, if a small quantity of flowers of sulphur be shaken with carbon disulphide it is found that the sulphur readily dissolves. Similarly, iodine will readily dissolve in ether. When working with ether and carbon disulphide all flames in the vicinity of the experimenter should be extinguished, as both liquids and their vapours are very inflammable.

Experiment 29.—A mixture of 1 grm. each of *sand, sugar, and sulphur is made up. The* constituents of the mixture are then to be separated and recovered, the method employed depending upon their solubilities in different solvents.

A definite quantity of water can only dissolve a maximum definite quantity of a certain substance at a fixed temperature. Such a solution is said to be a *saturated* solution of the solid at that temperature. If the temperature is raised, in general more of the solid can be dissolved, giving a saturated solution at the higher temperature. With a few exceptions it has been found that increase of temperature increases the solubility of solids. The reverse is true of dissolved gases. This increase of solubility with increase of temperature furnishes us with a means of purification for soluble solids, known as "purification by recrystallization".

Experiment 30.—Crude sodium nitrate is added to 50 cub. cm. boiling water in a beaker, until no more will dissolve. This hot saturated solution is then rapidly filtered, to free it from suspended impurities, and the filtrate set aside to cool to room temperature. The excess of the sodium nitrate in solution over the quantity necessary to produce saturation in the cold is deposited in geometrical shapes characteristic of the substance and known as "crystals". The greater part of the soluble impurities remains dissolved in the liquid surrounding the crystals. This liquid is usually called the "mother-liquor". The crystals are filtered off, and then dried between filter paper.

In some cases, notably where the solid contains a large amount of water of crystallization, the excess of dissolved solid does not crystallize out on allowing a hot saturated solution to cool down to room temperature. In such a case we obtain what is known as a "supersaturated" solution. If a crystal of the dissolved solid be added to a supersaturated solution, the excess at once crystallizes out.

Experiment 31.—A hot saturated solution of sodium thiosulphate is prepared by adding the finely powdered salt, in small quantities at a time, to 5 cub cm. of boiling water in a test tube. When the solution is saturated the mouth of the tube is lightly closed with a plug of cotton wool, and the solution is then set aside to cool. Since no solid crystallizes out on cooling, the solution is now in a supersaturated condition. A crystal of sodium thiosulphate is added to the solution, when the contents of the tube will practically solidify, due to the excess of sodium thiosulphate suddenly crystallizing.

Experiment 32.—Determination of the Solubility of Potassium Chlorate at the Temperature of the Laboratory.—A small quantity of water, is shaken with powdered potassium chlorate for about fifteen minutes in order to make a saturated solution. When the potassium chlorate has settled to the foot of the flask about 10 cub. cm. of the clear solution is drawn off by means of a small pipette and run into a small weighed porcelain basin. The temperature of the laboratory is now noted, using a Centigrade thermometer. The basin containing the solution is weighed and then placed on a water-bath and evaporated to dryness. The basin is transferred to an air oven and kept there for about fifteen minutes at a temperature of 105° C., and then allowed to cool in a desiccator and weighed. The process of heating in the air oven and cooling in the desiccator is repeated until the weight is constant.

Weight of basin =
Weight of basin + solution =
Weight of solution =

Weight of basin + salt =
Weight of basin =
Weight of salt =

Weight of solution =
Weight of salt =
Weight of water =

$$\frac{\text{Weight of water}}{\text{Weight of salt}} = \frac{100}{x},$$

where x is the solubility in water at t° C.

Water of Crystallization

Many salts contain "water of crystallization". The presence of this water of crystallization very often affects the crystalline structure of the salt, and in some cases the colour of the salt is dependent on the presence or absence of this water of crystallization.

Experiment 33.—A substance which on exposure to the air loses its water of crystallization is said to be "efflorescent". A *clear* crystal of washing soda (sodium carbonate with water of crystallization), $Na_2CO_3 . 10\,H_2O$, is exposed to the air on a watch-glass for about a week. The sodium carbonate will "effloresce" and gradually crumble down to a white powder, in appearance quite unlike the original clear crystal. The falling down to a powder and loss of crystalline structure are due to the washing soda losing part of the water of crystallization.

Experiment 34.—A substance which on exposure to the air absorbs moisture from the air is said to be "hygroscopic". A crystal of copper sulphate, $CuSO_4 . 5\,H_2O$, is ground to a powder and then heated in a crucible for about fifteen minutes. The original blue colour of the copper sulphate will disappear, and the residue in the crucible is white. The ignition in the crucible has caused the copper sulphate to lose the five molecules of water of crystallization, and the salt is now "anhydrous". The contents of the crucible are now left exposed to the air. The anhydrous copper sulphate is a hygroscopic substance, and by absorption of atmospheric moisture gradually regains the blue colour of the "hydrated" salt.

temperature to redness. The crucible is heated at redness for about ten minutes, then cooled in the desiccator and weighed. This ignition, cooling and weighing are repeated until the weight is constant. The final loss of weight represents the weight of water in the weighed sample, from which the percentage of water in barium chloride can be calculated.

Weight of crucible =
Weight of crucible + barium chloride =
Weight of barium chloride =

After ignition—
Weight of crucible + barium chloride =
Weight of crucible =
Weight of barium chloride =

Weight of water = loss in weight on ignition.

***Experiment 38.—Determination of the Gravimetric Composition of Water.**—Water can be synthesized by passing hydrogen over heated copper oxide:

$$H_2 + CuO = H_2O + Cu.$$

The copper oxide is reduced by the hydrogen to copper with formation of water (fig. 8). A Jena glass tube about 12 in. in length is filled with granular copper oxide, and since the copper oxide contains traces of water the tube and contents must be dried. To dry the copper oxide the tube is placed on a Ramsay burner, and one end is connected to an aspirator and the other to a wash-bottle containing concentrated sulphuric acid. The burner is now lighted, and, starting with a very small flame, the temperature is gradually raised, whilst a slow stream of dry air is drawn through the tube by means of the aspirator. It is advisable to protect the corks by slipping on to the tube two pierced shields of asbestos. After about fifteen minutes' heating the burner is extinguished, and the air is allowed to pass while the tube cools. When the tube is quite cold it is securely corked and weighed. The two corks must be kept, as they will

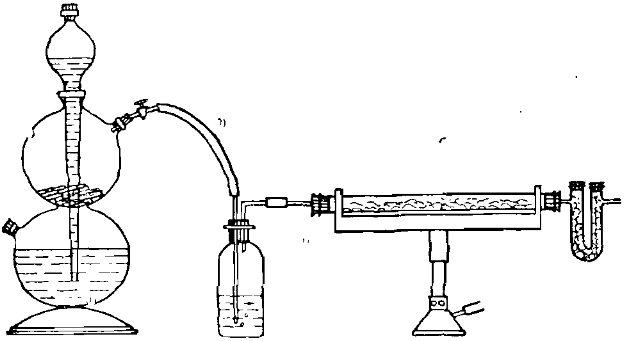

Fig. 8

be used in the next weighing. At the same time a **U** tube containing granular calcium choloride is weighed. The two side tubes on the **U** tube should be sealed by pieces of rubber tubing plugged with glass rod, which are only removed when the tube is being weighed. The tube is now replaced on the burner, and one end is again connected to the drying-bottle, which in turn is connected to a Kipp's hydrogen generating apparatus. The other end is closed with a cork through which passes one of the side tubes of the **U** tube, the end of the small tube being flush with the end of the cork. A second calcium chloride tube may be connected to the weighed calcium chloride tube to prevent absorption of atmospheric moisture. A *slow* stream of hydrogen is now passed through the apparatus, and after some time the purity of the gas is tested (see Experiment 21). When the hydrogen issuing at the end of the apparatus is pure the burner is lighted and the temperature is gradually raised. The copper oxide is reduced, and when the metallic appearance has spread to about half the length of the tube the temperature is gradually lowered, and finally the tube is allowed to cool in a stream of hydrogen. The **U** tube is now detached and the stoppers put on. At the same time the tube containing the reduced copper oxide is attached to the aspirator, and a slow stream of dry air is drawn through the tube for three or four minutes, in order to displace the hydrogen. The tube is now corked and weighed, and the weight of the calcium chloride **U** tube is also found. The increase in weight of the calcium chloride tube gives the weight of water formed, and the loss of weight in the tube which was heated gives the weight of oxygen. The weight of hydrogen is found by subtracting the weight of oxygen from the weight of water:

Weight of calcium chloride tube before experiment =
" " " " after experiment =
Weight of water =
Weight of copper oxide tube before experiment =
" " " " after experiment =
Weight of oxygen =
Weight of hydrogen = weight of water
− weight of oxygen.

The student should now calculate the weight of oxygen which combines with 1 grm. of hydrogen.

NITROGEN

Air is a mixture of oxygen and nitrogen. The latter gas may be obtained from air by removing the oxygen by means of substances which combine easily with it. If, for example, phosphorus be burned in a closed vessel containing air, it combines with the oxygen, forming phosphorus pentoxide, P_2O_5 ($P_4 + 5\,O_2 = 2\,P_2O_5$), while the nitrogen remains unchanged. A more convenient method, however, is shown in the following experiment.

***Experiment 39.—Preparation of Nitrogen by passing Air over Red-hot Copper.**—The apparatus shown in fig. 9 is fitted together. *a* is a Jena glass tube of ½-in. bore and about 12 in. long, supported on a Ramsay burner. *a* is connected to the bottle *c* (of about 2 litres capacity) by the right-angled tube *b*. The tube *d* reaches to the foot of the bottle *c*, which

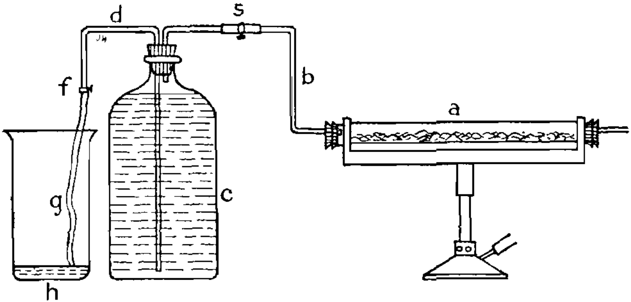

Fig. 9

is filled with water, and is joined to the rubber tubing *g*, which carries a screw clip *f*. Before beginning the experiment the cork in the bottle *c* must be shown to be air-tight. A quantity of copper turnings is placed in the Jena tube and heated, at first gently and then at a red heat. When the copper is quite hot the screw clip *f* is opened and the water made to flow from *c* by sucking at the open end of the rubber tube. By means of the screw clip the flow is regulated so that about two drops of water leave the tube each second. The copper is oxidized ($2\,Cu + O_2 = 2\,CuO$), and the nitrogen collects in the bottle, taking the place of the water which has flowed into the beaker *h*. When about a litre of nitrogen has been collected the heating is discontinued, the screw clip *s* is closed, and, after cooling, the tube *a* is removed. The gas is then used for the following experiment.

***Experiment 40.—Determination of the Weight of a Litre of Nitrogen.**—The method and the calculation are the same as those already employed in finding the weight of a litre of oxygen (Experiment 19).

Experiment 41.—Preparation of Nitrogen by heating Ammonium Nitrite.

$$NH_4NO_2 = N_2 + 2\,H_2O,$$

$$\text{or}\quad NH_4Cl + NaNO_2 = NaCl + 2\,H_2O + N_2.$$

10 grm. ammonium chloride and 15 grm. sodium nitrite with 20 cub. cm. water are placed in a small flask fitted with a cork and delivery tube dipping under water in a pneumatic trough. The mixture in the flask is *gently* heated until the reaction sets in,

when the flame is removed and the action allowed to proceed without further heating. Should the evolution of gas become violent, the flask is cooled by means of water. One bottle of the gas is collected and tested with a burning taper.

THE ATMOSPHERE

* **Experiment 42.—Determination of the Weight of a Litre of Air.**—The method is the same as that employed in the determination of the weight of a litre of oxygen, while the calculation is similar. (Experiment 19.)

From Experiments 19, 40, and 42 the specific gravities of oxygen and nitrogen relative to air can be calculated. Further, the density of air (H = 1) is 14·44, and the densities of oxygen and of nitrogen and their molecular weights (equal to twice the density) can be found.

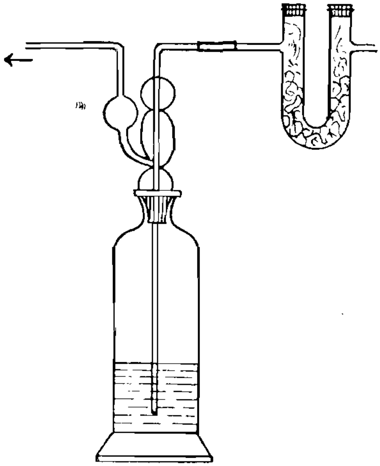

Fig. 10

Experiment 43.—To show the Presence of Moisture and of Carbon Dioxide in Air.—A current of air is drawn through a **U** tube (fig. 10) containing anhydrous calcium chloride, and through a bottle containing lime water, by means of a Bunsen pump. It will be seen that water collects in the **U** tube and dissolves the calcium chloride, while in the bottle a precipitate of calcium carbonate appears.

***Experiment 44.—Estimation of Oxygen and Nitrogen in Air.**—It has already been mentioned that phosphorus easily combines with oxygen to form phosphorus pentoxide, while nitrogen remains unchanged. By allowing this reaction to take place in a graduated tube over water the proportion of oxygen in the air can be determined.

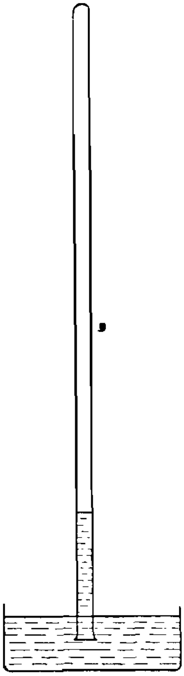

Fig. 11

A graduated gas measuring tube of 150 cub. cm. capacity is taken (fig. 11), and about 10 cub. cm. water placed in it, the open end closed with the thumb, and the tube inverted in a tall cylinder of water. After standing for five minutes the tube is fixed in a clamp and raised until the levels of the water outside and inside are equal and the volume of the enclosed air is noted. The temperature of the water and the barometric pressure are also noted. The tube is again closed with the thumb and transferred to a beaker or shallow dish containing water, and is supported by the clamp on a retort stand. A small piece of phosphorus, ¼ in. in length, is fastened to a copper wire about 20 in. long and passed up the tube until it is about 12 in. above the surface of the water. The phosphorus is allowed to remain in contact with the air for at least twenty-four hours. At the end of that time the water will have risen inside the tube. After withdrawing the phosphorus from the tube, the end is closed with the thumb, the tube transferred to the tall cylinder, and the volume of the remaining gas read off, as was done in the case of measuring the initial volume of air. The temperature and pressure are again noted.

The initial volume of air, v_1, was measured at temperature $t°_1$, and barometric pressure P_1. This must be reduced to 0° and 760. The final volume of gas (nitrogen), v_2, at $t°_2$ and P_2, must also be reduced to N.T.P. If these volumes be v_3 and v_4 respectively, the percentage of nitrogen is easily calculated: $v_3 : v_4 :: 100 : x$, where x = percentage of nitrogen. Then $100 - x$ gives the percentage of oxygen.

NITRIC ACID

Nitric acid is prepared by the action of concentrated sulphuric acid upon sodium or potassium nitrate:

$$NaNO_3 + H_2SO_4 = NaHSO_4 + HNO_3.$$

Preparation of Nitric Acid

Experiment 45.—About 20 grm. sodium nitrate are placed in a 4-oz. tubulated retort (fig. 12), which is fitted with a glass stopper. The retort is then placed on a sand bath, with the end of the neck leading into a flask which floats on water. The receiving flask is kept cool during the experiment by placing a pad of wet filter paper on it. About 20 cub. cm. concentrated sulphuric acid are now added through the tubulus, and the contents of the retort heated. The nitric acid, being volatile, distils over, and is collected in the cooled receiver. The distillate is coloured yellow through the presence of oxides of nitrogen, but may be rendered colourless by blowing a stream of air from the blowpipe through the acid while still warm. The residue in the retort is allowed to cool thoroughly, and is then washed out with water.

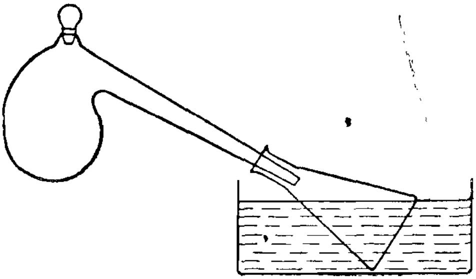
Fig. 12

Since nitric acid contains more than three-fourths of its weight of oxygen, in combination with hydrogen and nitrogen, with which it readily parts to oxidizable substances, it forms a very powerful oxidizing agent.

Experiment 46 [D].—A small quantity of sawdust is placed in a porcelain basin, and then gently heated just to charring point in the fume chamber. The

Bunsen is now withdrawn, and about 1 cub. cm. of the concentrated acid is quickly poured on to the sawdust. The nitric acid oxidizes the sawdust, and in this case the oxidation is so violent that the sawdust is inflamed.

Experiment 47 [D].—About 1 grm. of flowers of sulphur is placed in a conical flask, and covered with about 5 cub. cm. of concentrated nitric acid. The mixture is then heated over wire gauze in the fume chamber for about five minutes. The nitric acid gradually oxidizes the sulphur to sulphuric acid, and the presence of sulphuric acid may be shown by diluting the solution, filtering, and adding barium chloride solution, when a white precipitate of barium sulphate is thrown down:

$$H_2SO_4 + BaCl_2 = BaSO_4 + 2\,HCl.$$

The precipitate is insoluble in hydrochloric acid.

Experiment 48.—Some sulphuretted hydrogen water is placed in a test tube, and a few drops of strong nitric acid are added. The sulphuretted hydrogen gas which is present in solution is oxidized by the nitric acid to sulphur, which is thrown down as a fine dirty-yellow precipitate:

$$H_2S + O = H_2O + S.$$

Owing to the fact that nitric acid is such a powerful oxidizing agent, we rarely obtain hydrogen as a product of the reaction of a metal upon nitric acid (with the exception of the metal magnesium). In most cases the metal acts on nitric acid, with formation of the nitrate and liberation of oxides of nitrogen. In some cases the metal is oxidized, and oxides of nitrogen are evolved.

Experiment 49 [D].—**Preparation of Nitrates.**—A small piece of lead is placed in one porcelain basin, and a small piece of copper foil in another. The metals are then covered with dilute nitric acid (sp. gr. = 1·2), and gently heated in the fume chamber. A reddish gas is given off while the metals are dissolving, and when solution is complete the contents of the basins are evaporated almost to dryness. On allowing to cool, white crystals of lead nitrate separate in one basin, and blue crystals of copper nitrate are formed in the second basin:

$$3\,Pb + 8\,HNO_3 = 3\,Pb(NO_3)_2 + 4\,H_2O + 2\,NO$$
$$3\,Cu + 8\,HNO_3 = 3\,Cu(NO_3)_2 + 4\,H_2O + 2\,NO.$$

The nitric oxide that is first formed is a colourless gas, but on coming into the air combines with the oxygen to form the reddish-coloured gas nitrogen peroxide:

$$2\,NO + O_2 = 2\,NO_2.$$

Experiment 50 [D].—The oxidizing effect of nitric acid on metals may be shown with tin. A small piece of tin is placed in a crucible, and then covered with strong nitric acid. On *gently* warming the crucible on a pipeclay triangle a vigorous reaction sets in, and copious fumes of a reddish gas are thrown off and the tin is gradually converted into a white powder. When the reaction ceases the contents of the crucible are gently evaporated and finally ignited. The first action of the nitric acid is to convert the tin into metastannic acid with evolution of nitrogen peroxide:

$$5\,Sn + 20\,HNO_3 = H_{10}Sn_5O_{11} + 5\,H_2O + 20\,NO_2.$$

On ignition the metastannic acid loses water, and is changed into stannic oxide:

$$H_{10}Sn_5O_{15} = 5\,H_2O + 5\,SnO_2.$$

The salts of nitric acid also act as oxidizing agents, and are frequently employed in the laboratory as such.

Experiment 51.—About 2 grm. potassium nitrate are placed in a test tube and heated until the crystals melt. A small piece of charcoal is heated in the Bunsen until it glows, and is dropped on the molten potassium nitrate, when it will be seen to react vigorously with the nitrate, moving about on the surface in a state of incandescence:

$$4\,KNO_3 + 5\,C = 2\,K_2CO_3 + 3\,CO_2 + 2\,N_2.$$

Experiment 52 [D].—The previous experiment is repeated, using sulphur instead of charcoal. The 2 grm. of potassium nitrate are mixed on a sheet of paper, using the spatula, with about 0·5 grm. sulphur. On warming the mixture the sulphur is oxidized by the nitrate, potassium sulphate being formed:

$$2\,KNO_3 + 2\,S = K_2SO_4 + SO_2 + N_2.$$

Experiment 53.—A crystal of copper nitrate is moistened with water, wrapped in tinfoil, and then placed on a watchglass. After standing for a little a violent reaction sets in, steam is evolved, and the tin is oxidized by the copper nitrate.

Experiment 54. — Test for Nitric Acid or a Soluble Nitrate. — The presence of nitric acid or a soluble nitrate may be shown by the "ring" test. About 1 cub. cm. of sodium nitrate solution is placed in a test tube, and then an equal volume of concentrated sulphuric acid is added. The mixture is *thoroughly* cooled by shaking the tube in a stream of cold water. A few drops of a *freshly prepared* solution of ferrous sulphate are now cautiously poured down the side of the tube. At the junction of the two liquids a dark-brown ring is produced. On warming the contents of the tube the ring gradually disappears. The sulphuric acid first of all acts on the nitrate with liberation of nitric acid, which is reduced by the ferrous sulphate to nitric oxide:

$$2\,NaNO_3 + 2\,H_2SO_4 = 2\,NaHSO_4 + 2\,HNO_3$$

$$6\,FeSO_4 + 3\,H_2SO_4 + 2\,HNO_3 = 3\,Fe_2(SO_4)_3 + 2\,NO + 4\,H_2O.$$

The nitric oxide then combines with the excess of ferrous sulphate, forming a dark-brown substance which is supposed to have the formula $FeSO_4 . NO$. On heating, this compound is decomposed with loss of nitric oxide, and consequent disappearance of the dark-brown colour.

Experiment 55.—Another test may be carried out by adding a small quantity of concentrated sulphuric acid to a nitrate in the solid state, and then placing a strip of copper foil in the mixture. On gently warming, a reddish gas is formed in the tube, and the solution turns a deep-blue colour. The nitrate and sulphuric acid interact with formation of nitric acid, which then reacts with the copper, forming copper nitrate and nitric oxide. The colourless gas nitric oxide combines with the oxygen of the air, forming the red gas nitrogen peroxide.

OXIDES OF NITROGEN

The oxides of nitrogen are prepared either from nitric acid or a nitrate. The pentoxide (N_2O_5) and the trioxide (N_2O_3) cannot be readily prepared in the laboratory.

Experiment 56.—Preparation of Nitrous Oxide.—About 20 grm. ammonium nitrate are placed in a 250 cub.-cm. round-bottomed flask, which is fitted with a delivery tube for collecting the gas over water. Care should be taken to prevent any small pieces of cork falling into the flask. The pneumatic trough is now filled with *warm* water, since the gas is moderately soluble in cold water. The flask is *gently* heated with a small flame, and, when the contents of the flask fuse, the flame must be kept below the level of the fused salt. Three full bottles of the gas and one bottle half-full are collected:

$$NH_4NO_3 = N_2O + 2\,H_2O.$$

Experiment 57.—A glowing splinter of wood is thrust into one of the bottles of the gas. The nitrous oxide rekindles the glowing splinter, and in this respect resembles oxygen.

Experiment 58.—The second bottle of the gas is placed mouth upwards on the bench. On withdrawing the covering plate no red fumes are formed, since nitrous oxide does not combine with the oxygen of the air, as nitric oxide does. The bottle which is half-full is to be kept for an experiment with nitric oxide.

Experiment 59.—The third bottle is transferred from the trough containing warm water to a beaker containing cold water, and the greased plate is removed. The water will slowly rise in the bottle, showing that nitrous oxide is soluble to some extent in cold water.

Experiment 60.—Preparation of Nitric Oxide.—A flask containing about 15 grm. copper turnings is fitted with a thistle funnel and a delivery tube for collecting the gas over water. Dilute nitric acid (sp. gr. = 1·2) is now poured down the thistle funnel until the turnings are completely covered. Usually the action begins at once, but it may be necessary to start it by *gentle* heating. The gas that is formed is at first coloured red, but gradually turns colourless. When the gas in the flask is colourless, three full bottles and one half-full bottle of the gas are collected. The copper and nitric acid interact with formation of nitric oxide and copper nitrate which forms a blue solution in the flask:

$$3\,Cu + 8\,HNO_3 = 3\,Cu(NO_3)_2 + 4\,H_2O + 2\,NO.$$

Experiment 61 [D].—One of the bottles of the gas is placed mouth upwards in the fume chamber.

On withdrawing the greased plate, dense red fumes are formed. The nitric oxide combines with the oxygen of the air to form nitrogen peroxide:

$$2\,NO + O_2 = 2\,NO_2.$$

Experiment 62.—Another bottle is tested with a burning taper. Although nitric oxide contains more oxygen than nitrous oxide, it does not support the combustion of a taper.

Experiment 63.—A few drops of a freshly prepared solution of ferrous sulphate are added to the third bottle, and, after replacing the greased plate, the contents are shaken. The ferrous sulphate solution turns almost black, and if this solution is poured into a test tube, diluted and then boiled, the dark colour will disappear. This dark-brown substance is the same as was formed in the "ring" test for nitric acid, namely $FeSO_4.NO$, which is decomposed on heating.

Experiment 64.—The half-bottle of nitrous oxide is now placed in the trough beside the half-bottle of nitric oxide, and the contents of one bottle are allowed to bubble into the other bottle. The nitrous oxide does not combine with the nitric oxide, and no red fumes are produced, as in the case of oxygen and nitric oxide. In this way it is possible to distinguish between nitrous oxide and oxygen.

Experiment 65.—Preparation of Nitrogen Peroxide.—A small quantity of lead nitrate (5 grm.) is ground to a fine powder in a mortar and then transferred to a porcelain basin. The porcelain basin is heated on the sand bath with constant stirring of the lead nitrate, until the salt is quite dry. This preliminary drying is necessary, as the crystals of lead nitrate are apt to "decrepitate" when heated. The dried lead nitrate is now strongly heated in a hard-glass test tube, and the evolved gases are collected over water in a test tube. The lead nitrate on heating is decomposed with formation of lead oxide, nitrogen peroxide, and oxygen:

$$2\,Pb(NO_3)_2 = 2\,PbO + 4\,NO_2 + O_2.$$

The reddish colour of the gas in the test tube and delivery tube is due to the presence of the nitrogen peroxide, but on passing through the water the nitrogen peroxide is dissolved and the collected gas is oxygen, which can be tested for with a glowing splinter. The nitrogen peroxide on dissolving in the *water forms nitric and nitrous acids:*

$$2\,NO_2 + H_2O = HNO_3 + HNO_2.$$

The test for nitric acid is applied to a portion of the water in the trough.

* **Experiment 66.—Determination of the Gravimetric Composition of Nitrous Oxide.**—A 250 cub.-cm. round-bottomed flask containing about 25 grm. ammonium nitrate is fitted to a wash bottle containing concentrated sulphuric acid. The wash bottle is attached to one end of a weighed hard-glass tube containing copper turnings.[1] The other end of the tube is connected to an aspirator. The copper is

---

[1] Copper turnings which are to be used in gravimetric work should always be washed in benzene, so as to remove traces of grease.

first heated to redness, and then the ammonium nitrate is gently warmed. At the same time the screw clip on the aspirator is opened, and the water is allowed to drop slowly from the exit tube. The water which is formed is absorbed by the sulphuric acid, and the nitrous oxide is decomposed on passing over the red-hot copper, forming copper oxide and nitrogen, which passes to the aspirator:

$$N_2O + Cu = CuO + N_2.$$

When about 250 cub. cm. of nitrogen have collected in the aspirator the experiment is stopped. The increase in weight of the hard-glass tube gives the weight of oxygen in the gas. The volume of nitrogen is calculated in the usual way, and by multiplying the corrected volume by 0·001255 the weight of the nitrogen is found.

The student should now calculate what weight of nitrogen will combine with 16 grm. of oxygen.

* **Experiment 67.—Determination of the Weight of a Litre of Nitrous Oxide.**—A tube for collecting by downward displacement is now attached to the drying bottle, and by gently warming the ammonium nitrate the usual specific-gravity flask can be filled with dry nitrous oxide. When the specific-gravity flask is full of nitrous oxide, which can be shown by means of a glowing splinter, it is corked and weighed. The temperature and pressure of the atmosphere are also carefully noted.

The capacity of the flask being known, this volume is first of all reduced to 0° C. and 760 mm. The weight of air which would fill the flask is then calculated, and this weight is subtracted from the weight of the flask when full of air, giving the weight of the vacuous flask. (One litre of air weighs 1·293 grm. at 0° C. and 760 mm.) The weight of the vacuous flask is then subtracted from the weight of the flask when full of nitrous oxide, giving the weight of a known volume of nitrous oxide, from which the weight of 1 litre of the gas can be found. Having found the weight of 1 litre of nitrous oxide, the density relative to hydrogen is calculated. (One litre of hydrogen weighs 0·0896 grm. at 0° C. and 760 mm.)

* **Experiment 68.—Determination of the Gravimetric Composition of Nitric Oxide.**—The apparatus is similar to that used in determining the gravimetric composition of nitrous oxide, and the method of calculating the results is the same.

* **Experiment 69.—Determination of the Weight of a Litre of Nitric Oxide.**—After finishing the gravimetric experiment, the flask and wash bottle are transferred to the fume chamber. A delivery tube for collecting by downward displacement is fitted to the wash bottle. This tube should be long enough to reach to the bottom of the specific-gravity flask, and the gas must be passed in until the contents of the flask are perfectly colourless, when the flask is quickly corked and weighed. The weight of a litre and the density relative to hydrogen are then calculated.

The four preceding experiments enable us to calculate the molecular weights of nitrous and nitric oxides. In the case of nitrous oxide, the density relative to hydrogen is 22. The molecular weight is found by multiplying the density relative to hydrogen by 2. The molecular weight of nitrous oxide is therefore 44. From the gravimetric experiment we found that 28 parts by weight of nitrogen combine with 16 parts by weight of oxygen. If two atoms of nitrogen, atomic weight = 14, were to combine with one atom of oxygen, atomic weight = 16, the compound would have a molecular weight = $14 \times 2 + 16 = 44$. Similarly, nitric oxide with a density 15 has a molecular weight 30. From the gravimetric experiment we found that 14 parts by weight of nitrogen combine with 16 parts by weight of oxygen. One atom of nitrogen combining with one atom of oxygen would give a compound whose molecular weight is $14 + 16 = 30$.

These experiments also illustrate the Law of Multiple Proportions, the different proportions of nitrogen combining with 16 parts of oxygen standing in the ratio of 2 : 1.

AMMONIA

This gas can readily be prepared by the action of an alkali on an ammonium salt. When a large supply of the gas is required, it is convenient to warm the concentrated solution of the gas (0·880 ammonia).

Preparation of Ammonia

Experiment 70 [D].—About 25 grm. ammonium chloride are intimately mixed with an equal weight of slaked lime, and *the mixture placed* in a round-bottomed flask fitted with a delivery tube for upward displacement (fig. 13). The apparatus is then placed in the fume chamber, and on gently warming the flask, *ammonia is evolved.* When all air has been displaced from the flask, three bottles are filled with the gas and left standing mouth downwards:

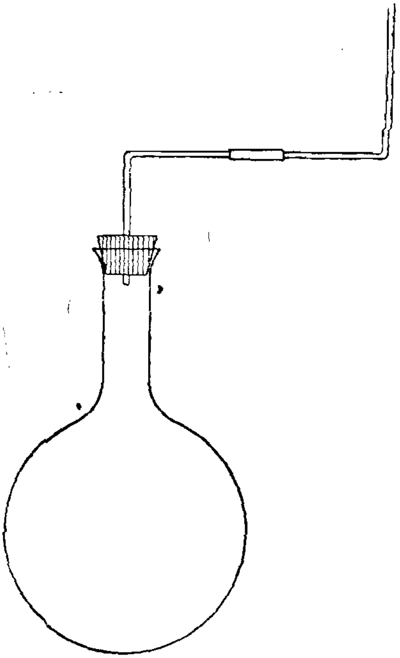

Fig. 13

$$2NH_4Cl + Ca(OH)_2 = CaCl_2 + 2NH_3 + 2H_2O.$$

Experiment 71 [D].—A taper is extinguished on being thrust into one of the bottles, and the ammonia does not burn. At the moment of entry of the taper into the bottle a slight greenish-yellow flame may be noticed playing round the taper flame. When ammonia is heated, it may be made to burn in air, and sometimes a slight combustion takes place, due to the heating caused by the flame of the taper.

Experiment 72.—The second bottle is placed in a trough of water, and the greased plate is removed. The water rapidly rises in the bottle, showing that ammonia is very soluble in water. When all absorption has ceased, the plate is replaced and the bottle is removed. On testing the solution in the bottle with litmus it will be found to have an alkaline reaction. On dissolving ammonia in water, besides the simple solution of the gas, there is formed the substance ammonium hydroxide:

$$NH_3 + H_2O = NH_4OH.$$

Experiment 73 [D]. — A drop of strong hydrochloric acid, on the end of a glass rod, is placed in the third bottle of ammonia. Dense white clouds are produced, which gradually condense on the sides of the bottle in the form of minute crystals readily soluble in water. The ammonia and hydrochloric acid react together, forming the crystalline solid ammonium chloride:

$$NH_3 + HCl = NH_4Cl.$$

The ammonium radicle NH_4 forms a large number of salts, all of which can be prepared by neutralizing the free acid with ammonia gas, or with ammonium hydroxide solution:

$$NH_3 + HNO_3 = NH_4NO_3$$
(ammonium nitrate).

$$2NH_3 + H_2SO_4 = (NH_4)_2SO_4$$
(ammonium sulphate).

$$NH_3 + H_2S = NH_4HS$$
(ammonium hydrosulphide).

$$NH_4HS + NH_3 = (NH_4)_2S$$
(ammonium sulphide).

All ammonium salts on heating with an alkali evolve ammonia, a reaction which forms the test for the ammonium radicle:

$$(NH_4)_2SO_4 + 2NaOH = Na_2SO_4 + 2NH_3 + 2H_2O.$$

Experiment 74. — Test for the Ammonium Radicle.—A solution of an ammonium salt is heated with caustic soda solution. Ammonia is at once given off, and can be identified by the characteristic smell, the alkaline action towards litmus paper, and the turning brown of moist turmeric paper.

Experiment 75. — One drop of ammonium hydroxide solution is added to about 100 cub. cm. of water in a flask. Two drops of Nessler's solution are now added to this dilute solution of ammonia, and a yellow coloration is immediately produced. This reaction forms a very delicate test for the presence of ammonia.

* **Experiment 76.—Determination of the Gravimetric Composition of Ammonia.** — The composition of ammonia by weight may be determined by passing the dry gas over heated copper oxide, when

the ammonia is decomposed with liberation of nitrogen and formation of water:

$$3\,CuO + 2NH_3 = 3\,H_2O + N_2 + 3\,Cu.$$

A Jena glass tube, about 1 ft. in length, containing copper oxide is heated on the Ramsay burner to expel moisture, then securely corked, allowed to cool, and weighed accurately. A 250 cub.-cm. Erlenmeyer flask, containing about 100 cub. cm. of 0·880 ammonia, is fitted with a **U** tube containing stick caustic soda

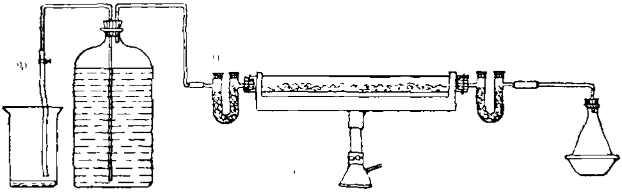

Fig. 14

and the whole is then accurately weighed, the **U** tube being sealed to prevent absorption of atmospheric moisture. A calcium chloride tube is also weighed, with the usual precautions. The whole apparatus is then connected together as shown in fig. 14. The copper oxide is now gradually raised to redness, and at the same time the screw clip on the aspirator is opened. When the copper oxide is red-hot, the water bath is *gently* heated, with the result that ammonia gas is evolved and passes over the copper oxide, reducing it to the metallic state. The water that is formed is absorbed in the calcium chloride, and the liberated nitrogen collects in the aspirator.

When about half the copper oxide is reduced, the screw clip on the aspirator is closed and the Ramsay burner extinguished. At the same time the Erlenmeyer and drying U tube are removed from the bath and sealed, the calcium chloride U tube is sealed and the copper oxide tube is corked. All these operations require to be performed very speedily.

The volume of water which has been expelled from the aspirator into the beaker is accurately measured, at the same time the temperature and pressure are noted, and the parts of the apparatus which were originally weighed are again weighed.

The volume of nitrogen is reduced to 0° and 760 mm., and the corrected volume is then multiplied by 0·001255, which gives the weight of the nitrogen. The weight of hydrogen can be obtained either from the weight of water obtained in the calcium chloride tube, or from the weight of oxygen lost in the copper oxide tube.

The student should then calculate what weight of hydrogen will combine with 14 grm. of nitrogen.

*** Experiment 77.—Determination of the Volumetric Composition of Ammonia.**—About 250 cub. cm. dilute caustic soda solution are placed in a porcelain basin in the fume chamber and bromine is gradually added to the solution until the liquid has a permanent straw colour. The bromine and caustic soda react together according to the following equation:

$$2\,NaOH + Br_2 = NaBrO + NaBr + H_2O.$$

A perfectly dry eudiometer is now completely filled with ammonia by means of the Erlenmeyer and drying **U** tube used in the previous experiment. When the eudiometer is filled, it is gradually withdrawn from the delivery tube, closed with the thumb, and transferred to the basin.[1] On removing the thumb, the liquid gradually rises in the tube, and this rise is accelerated by gentle shaking until no further absorption takes place. When absorption ceases the tube is again closed with the thumb and quickly placed in the levelling trough, where the levels are adjusted and the residual volume of gas noted. The action of the hypobromite is to oxidize the ammonia with liberation of nitrogen:

$$3\,NaBrO + 2\,NH_3 = N_2 + 3\,H_2O + 3\,NaBr.$$

The residual volume in the tube is therefore nitrogen, and it will be found that this is half the original volume. Two volumes of ammonia on decomposition give 1 volume of nitrogen, or from the equation:

$$2\,NH_3 = N_2 + 3\,H_2,$$

1 volume of nitrogen would combine with 3 volumes of hydrogen to form 2 volumes of ammonia.

HYDROCHLORIC ACID

Experiment 78.—Two grams common salt (sodium chloride) are placed in a test tube and about 3 cub. cm. concentrated sulphuric acid added. A gas is evolved which forms white fumes in moist air, and also when a glass rod which has been dipped in a

---

[1] The hands should be washed immediately after being in contact with any corrosive liquid, e.g. an alkali or acid.

solution of ammonia is held at the mouth of the test tube. *It will also be observed that the gas reddens* litmus paper:

$$NaCl + H_2SO_4 = NaHSO_4 + HCl.$$

Experiment 79 [D].—Preparation of Hydrochloric Acid.—About 30 grm. salt are placed in a conical flask fitted with a thistle funnel and delivery tube for downward displacement. Through the thistle funnel 30 cub. cm. concentrated sulphuric acid are poured in small quantities at a time, and three bottles of the gas are collected, each being then covered with a greased glass plate.

A burning taper is plunged into the first bottle.

The second is inverted in a basin of water, when the gas dissolves, forming a solution which reddens litmus.

Into the third bottle 2 or 3 cub. cm. concentrated ammonia solution are poured. White fumes of ammonium chloride are formed:

$$NH_3 + HCl = NH_4Cl.$$

*** Experiment *80*.—*Volumetric* Composition of Hydrochloric Acid Gas.**—*To the* delivery tube of the generating apparatus is attached by means of rubber tubing a bottle containing concentrated sulphuric acid (fig. 15), and to the latter a tube long enough to reach to the foot of a gas-measuring tube, whose *total* capacity in cubic centimetres is known. The latter is completely filled with the gas and the open end closed by the thumb. About 10 cub. cm.

of sodium amalgam[1] are now poured into the gas-measuring tube, which is again closed immediately and shaken for a minute or two. It is then inverted in a tall cylinder of water, and the volume of the remaining gas (hydrogen) read off, when the levels of the water outside and inside are equal. The mercury is collected and its volume measured.

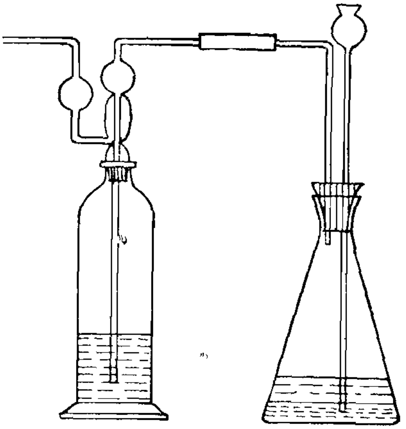

Fig. 15

The volume of hydrogen is equal to half the volume of hydrochloric acid used, which is equal to the total capacity of the tube less the volume of mercury.

Experiment 81.—Small quantities of potassium chloride, ammonium chloride, and calcium chloride are each acted on by concentrated sulphuric acid,

---

[1] Sodium amalgam is prepared by pressing by means of a pestle five or six small pieces of freshly cut sodium, one by one, under the surface of some mercury contained in a porcelain mortar. During the operation, which should be carried out in a draught-chamber, the hand should be protected by a glove or duster.

and the evolved gas is shown to be hydrochloric acid. The equations for these reactions should be written.

Experiment 82.—Action of Hydrochloric Acid on some Metals.—Small pieces of copper, lead, zinc, iron, aluminium, and magnesium are placed in separate test tubes and a few cubic centimetres hydrochloric acid added to each.

With copper and lead the action is very slight. In the other cases, chlorides of the metals are formed and hydrogen is evolved.

Formation of Metallic Chlorides

The chlorides of the metals are formed by the action (*a*) of chlorine on the metal; (*b*) of hydrochloric acid solution on the metal, with or without heating; (*c*) of the acid on the oxide, the hydroxide, or the carbonate of the metal. Chlorides which are insoluble in water are prepared by the method of precipitation.

Experiment 83.—The solution formed by the action of tin on hydrochloric acid is evaporated almost to dryness. The residue is stannous chloride, $SnCl_2, 2\,H_2O$.

Experiment 84.—One gram of copper oxide is dissolved in hydrochloric acid and the solution evaporated, leaving cupric chloride as the residue:

$$CuO + 2\,HCl = CuCl_2 + H_2O.$$

Experiment 85.—Ten cubic centimetres sodium hydroxide solution are measured into a porcelain basin, and dilute hydrochloric acid is added drop by drop from a burette until the solution turns litmus paper faintly red. On evaporation to dryness there remains sodium chloride. This can be recognized by its taste, and by the action of concentrated sulphuric acid:

$$NaOH + HCl = NaCl + H_2O.$$

Experiment 86.—About 1 gr. sodium carbonate is treated with dilute hydrochloric acid until evolution of gas ceases, and the solution is evaporated to dryness. The residue is shown to be sodium chloride:

$$Na_2CO_3 + 2\,HCl = 2\,NaCl + H_2O + CO_2.$$

Experiment 87.—A few cubic centimetres of solutions of *silver nitrate, lead nitrate, and mercurous* nitrate are placed in separate test tubes, and a few drops of dilute hydrochloric acid are added to each. The three chlorides are precipitated:

$$\begin{aligned} AgNO_3 + HCl &= AgCl + HNO_3. \\ Pb(NO_3)_2 + 2\,HCl &= PbCl_2 + 2\,HNO_3. \\ HgNO_3 + HCl &= HgCl + HNO_3. \end{aligned}$$

Heat is now applied to the contents of each test tube to find out if the chlorides are soluble in *hot* water. Only one, it will be seen, dissolves. A solution of ammonia is added to the other two, and the effect noted in each case.

Experiment 88.—Test for a Soluble Chloride.—Silver nitrate solution added to a solution of a chloride produces a white precipitate of silver chloride, which is soluble in ammonia, and is reprecipitated on addition of nitric acid.

CHLORINE

Chlorine is prepared by the oxidation of hydrochloric acid, the oxidizing agent generally employed being manganese dioxide.

Note. — All experiments with chlorine should be performed in a draught-cupboard.

Preparation of Chlorine

Experiment 89 [D].—In a round-bottomed flask fitted with a thistle funnel and delivery tube for downward displacement are placed 20 grm. granular manganese dioxide. Fifty cubic centimetres concentrated hydrochloric acid are added through the thistle funnel, and the mixture heated on a small water bath. Four cylinders of the gas are collected, each being covered with a greased glass plate as soon as it is full, and the following experiments are performed.

Experiment 90 [D].—A lighted taper is plunged in the first cylinder, and burns with a smoky flame. After the flame has been extinguished, a little water is poured into the cylinder, and hydrochloric acid tested for by means of silver nitrate solution. The wax of the taper is a mixture of compounds of carbon and hydrogen, the latter of which combines with chlorine, while the carbon is deposited in the form of soot.

Experiment 91 [D].—In a test tube fitted with a cork and delivery tube, with the end bent upwards (fig. 16), some granulated zinc and dilute sulphuric acid are placed. The hydrogen — which of course must be free from air—is lighted at the end of the delivery tube, and the burning jet lowered into the second cylinder of chlorine, and allowed to burn for about a minute. The contents of the bottle are then

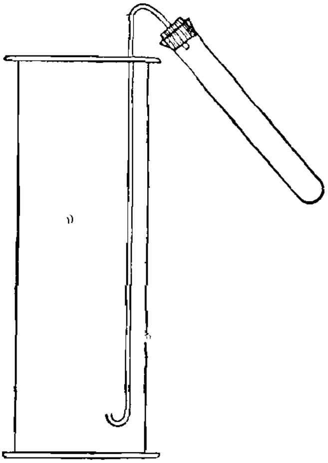

Fig. 16

shaken with a small quantity of water and tested for hydrochloric acid.

Experiment 92 [D]. — About ½ grm. of finely powdered antimony thrown into the third cylinder, at once catches fire, antimony chloride being formed.

Experiment 93 [D]. — A few cubic centimetres concentrated sulphuric acid are poured into the fourth cylinder of chlorine, for the purpose of removing moisture from the gas. A piece of Turkeyred cloth[1] is dried by heating in an air oven for a few minutes. The dried cloth is then suspended in the dry chlorine, when it will be seen that no change takes place. After wetting the cloth with water, and again suspending in chlorine, the colour is removed, a result which shows that chlorine bleaches only in presence of water.

Chlorine can also be prepared by acting on a mixture of sodium chloride and manganese dioxide with concentrated sulphuric acid.

Experiment 94.—In a test tube 2 grm. of salt, 2 grm. of manganese dioxide, and a few cubic centimetres concentrated sulphuric acid are gently heated. Chlorine is evolved and is tested for by its bleaching action on moist litmus paper.

Note.—If the mixture be heated strongly, hydrochloric acid will be given off and will turn the litmus red.

Experiment 95 [D].—A solution of chlorine in water (known as chlorine water) is prepared by allowing the gas from the apparatus in Experiment 89 to bubble for some minutes through water.

Experiment 96.—To a few cub. cm. chlorine water some sulphuretted hydrogen water is added. The white precipitate is filtered off, and the clear filtrate tested for hydrochloric acid:

$$H_2S + Cl_2 = 2\,HCl + S.$$

Experiment 97.—A few cub. cm. chlorine water

---

[1] The cheap red cloth dyed with para-nitraniline red is not bleached by chlorine.

and a few cub. cm. of a solution of potassium bromide are shaken together in a test tube. About 1 cub. cm. of carbon disulphide is added, and the mixture, after again shaking, allowed to stand. The chlorine displaces the bromine, which is then dissolved by the carbon disulphide:

$$2\,KBr + Cl_2 = 2\,KCl + Br_2.$$

Experiment 98.—In the same way it can be shown that chlorine displaces *iodine* from potassium iodide:

$$2\,KI + Cl_2 = 2\,KCl + I_2.$$

Other oxidizing agents can be used in place of manganese dioxide to oxidize hydrochloric acid.

Experiment 99.—A mixture of 1 grm. potassium dichromate and 2 or 3 cub. cm. concentrated hydrochloric acid are gently heated in a test tube, and the chlorine tested for.

$$K_2Cr_2O_7 + 14\,HCl$$
$$= 3\,Cl_2 + 2\,CrCl_3 + 2\,KCl + 7\,H_2O.$$

Experiment 100.—Similarly, it can be shown that potassium permanganate oxidizes hydrochloric acid:

$$2\,KMnO_4 + 16\,HCl$$
$$= 8\,H_2O + 2\,KCl + 2\,MnCl_2 + 5\,Cl_2.$$

Bleaching powder is manufactured by allowing chlorine to act on slaked lime:

$$Ca(OH)_2 + Cl_2 = CaOCl_2 + H_2O.$$

Experiment 101.—The properties of a sample of bleaching powder are noted, its appearance, the action *of water and of dilute acids, and its bleaching action in presence of acids.*

EQUIVALENTS

Definition: The equivalent of an element is that weight of the element which combines with or displaces 1 part by weight of hydrogen.

Experiment 102.—Determination of the Equivalent of a Metal by Displacement of Hydrogen from an Acid.—An accurately weighed quantity of the metal (0·3 grm. zinc, 0·1 of magnesium, or 0·08 of aluminium) is placed in a glass tube 2 in. long, $\frac{3}{16}$ in. in diameter, closed at one end, and filled with water, so that no air remains in it. A gas-measuring tube of 150 cub. cm. capacity is half-filled with concentrated hydrochloric acid, then filled to the top with water, the open end covered with the thumb, and inverted in a beaker containing water. Next, the small tube with the metal is placed under the open end of the gas-measuring tube. When all the metal has dissolved, the gas tube is transferred to a tall cylinder of water, and allowed to stand for ten minutes. *At the end of that time the hydrogen* will have attained the temperature of the water, and its volume is read. Finally, the temperature of the water and the height of the barometer are noted. From the latter is subtracted the tension of aqueous vapour. (See p. 95.)

Calculation.—The observed volume in cubic centimetres is reduced to normal temperature and pressure.

Since 22·32 litres of hydrogen at 0° and 760 mm. weigh 2 grm., the weight of the hydrogen is then found.

This weight having been displaced by the weight of metal taken, the weight of the latter which would displace 1 grm. of hydrogen, that is the equivalent of the metal, is then calculated.

* **Experiment 103.—Determination of the Equivalent of Tin in Stannous Compounds.** — The apparatus necessary is similar to that employed in Experiment 18, the only change being that the hard-glass test tube is replaced by a round flask (fig. 17). About 1 grm. of tin (accurately weighed) is placed in the flask with about 50 cub. cm. hydrochloric acid. The flask is then fitted tightly to the delivery tube. On heating gently the action begins, and hydrogen passes over into the aspirator, expelling water into

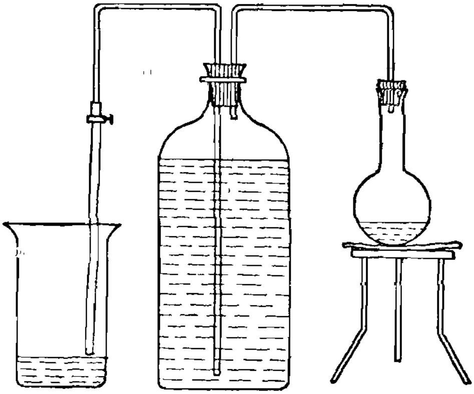

Fig. 17

the beaker. When the tin has all dissolved, the hydrogen is allowed to cool, and the beaker raised until the levels in the aspirator and the beaker are equal. The screw clip is closed, and the volume of expelled water measured. The barometric pressure and the temperature of the water are noted.

The calculation is similar to that in Experiment 102.

In Experiments 102 and 103 the equivalent of the metal has been found by measurement of the hydrogen displaced from an acid. In the following experiments it is found by determination of the weight of oxygen which combines with a known weight of the metal. The equivalent of oxygen is 8.

Experiment 104 [D].—Determination of the Equivalent of Tin in Stannic Compounds.—About 0·5 grm. of tin (accurately weighed) is placed in a previously weighed porcelain crucible, supported on a triangle and tripod. Concentrated nitric acid is added drop by drop from a glass rod as long as action continues. The crucible and its contents are heated, at first gently, and then strongly, and after cooling in a desiccator are weighed. The tin has been converted into stannic oxide, and the weight of this, *minus* the weight of tin taken, is the quantity of oxygen which has combined with the tin. The weight of metal which would combine with 8 grm. of oxygen is then calculated.

***Experiment 105 [D].—Determination of the Equivalent of Copper.**—0·2 grm. copper is placed in a weighed crucible, about 1 cub. cm. dilute nitric acid added, and the crucible gently heated. After the copper has dissolved, the solution is evaporated to dryness on a water bath, and the solid residue heated with the Bunsen flame at first gently, and finally at the full heat of the flame. The crucible and its contents are then cooled in the desiccator and weighed. The operations of heating, cooling, and weighing are repeated until the weight is constant.

The calculation is similar to that in Experiment 104.

CARBON

Carbon is known in three modifications, two of which, diamond and graphite, are crystalline, while the third, charcoal, is amorphous. The crystalline forms are found in nature. The following are the important forms of charcoal, the names serving to indicate their origin: Wood charcoal, bone charcoal, and lamp-black.

Experiment 106. — Preparation of Charcoal from Sugar. — Two grams sugar are heated in a dry test tube. Gases are evolved and charcoal remains in the tube.

Experiment 107. — Acids and alkalis have no action on charcoal in the cold. This may be shown by allowing the common acids and alkalis to come into contact with wood charcoal.

Experiment 108.—Decolorizing Action of Bone Charcoal. — A small quantity of dilute litmus is boiled with a little bone charcoal and the solution filtered. The colour is entirely removed.

Experiment 109. — The Reducing Action of Charcoal. — In a hard-glass test tube, fitted with a cork and delivery tube, about 2 grm. powdered copper oxide mixed with ½ grm. of powdered wood charcoal are heated. The evolved gas is passed into lime water:

$$2\,CuO + C = CO_2 + 2\,Cu.$$

CARBON DIOXIDE

It has been shown in Experiment 14 that carbon dioxide is formed when carbon burns in oxygen, and that it is found in the atmosphere (Experiment 43).

Preparation of Carbon Dioxide

Experiment 110.—The most convenient method for the preparation of carbon dioxide is by acting on a carbonate—*usually marble*—with *dilute* hydrochloric acid. The apparatus consists of a conical flask fitted with thistle funnel and delivery tube for collecting the gas by downward displacement. Some marble chips are placed in the flask and dilute hydrochloric acid is poured down the thistle funnel. Three cylinders of the gas are collected:

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

Experiment 111.—The first cylinder of gas is tested with a lighted taper.

Experiment 112.—Into the second cylinder some lime water is poured and shaken. A white precipitate of calcium carbonate appears:

$$Ca(OH)_2 + CO_2 = CaCO_3 + H_2O.$$

This precipitate is soluble in hydrochloric acid.

Experiment 113.—The gas in the third cylinder is poured (as if it were a liquid) into another cylinder containing a few cubic centimetres of lime water and the latter is then shaken. This experiment illustrates the high specific gravity of carbon dioxide.

Experiment 114.—The end of the delivery tube is dipped under a few cubic centimetres of water in a test tube, to which two or three drops of litmus have been added. Carbon dioxide is passed through the water until the litmus is turned reddish in colour, owing to the presence of carbonic acid (H_2CO_3). On boiling the solution the litmus is restored to its original colour. The carbon dioxide is expelled from the solution so that carbonic acid cannot be obtained by evaporating off the solvent.

Experiment 115.—In the same way carbon dioxide is passed into some caustic soda solution (1 cub cm. caustic soda diluted with 5 cub. cm. water). It will be seen that the gas is absorbed. If the gas is passed until the solution is saturated, sodium bicarbonate is formed:

$$NaOH + CO_2 = NaHCO_3.$$

That the solution contains a carbonate may be shown by adding hydrochloric acid, when carbon dioxide will be evolved:

$$NaHCO_3 + HCl = NaCl + CO_2 + H_2O.$$

Experiment 116.—Carbon dioxide is passed into about 30 cub. cm. lime water. Calcium carbonate is precipitated, but on further addition of carbon dioxide the solution again becomes clear. This is due to the formation of calcium bicarbonate [$CaH_2(CO_3)_2$], which is soluble in water:

$$CaCO_3 + H_2O + CO_2 \\ = CaH_2(CO_3)_2, \text{ or } CaCO_3.H_2CO_3.$$

HARD AND SOFT WATERS

Experiment 117.—Five cubic centimetres of the clear solution of calcium bicarbonate are placed in a test tube, and to it soap solution is added drop by drop from a burette. After the addition of each drop, the test tube and its contents are vigorously shaken, until a lather is obtained which remains for three or four minutes. The number of drops of soap solution which has been added is noted.

Experiment 118.—A second 5 cub. cm. are boiled for some minutes, filtered, and tested with soap solution as in Experiment 117. The quantity of soap solution necessary for the formation of a lather should be less in this case.

Experiment 119.—Five cubic centimetres of a solution of calcium sulphate are treated in the same way with soap solution, and the number of drops of the latter, which are required to give a lather, noted.

Experiment 120.—Experiment 119 is repeated, but the calcium sulphate solution is boiled before the addition of the soap solution.

Experiment 121.—To 5 cub. cm. of calcium sulphate solution, sodium carbonate solution is added until no more precipitate is thrown down. This is filtered off, and the clear filtrate treated with soap solution as in the previous experiments.

Water is said to be "hard" when it does not readily form a lather with soap. "Hardness" is caused by the presence of calcium and magnesium salts in solution. Soap is a mixture of the sodium or potassium salts of certain organic acids, and the precipitate or scum produced when soap is added to a hard water consists of the calcium and magnesium salts of these acids. A lather will not be formed until all the calcium and magnesium are removed from solution. There are two kinds of hardness: temporary hardness, which can be removed by boiling, and permanent hardness. The former is caused by the presence of bicarbonates of calcium and magnesium, the latter by the sulphates or other soluble salts of these metals. On boiling water containing the bicarbonates, normal carbonates are precipitated:

$$CaH_2(CO_3)_2 = CaCO_3 + H_2O + CO_2.$$

With sulphates, on the other hand, boiling does not soften the water, and the calcium and magnesium are precipitated by addition of sodium carbonate:

$$CaSO_4 + Na_2CO_3 = CaCO_3 + Na_2SO_4.$$

CARBONIC ACID AND CARBONATES

It has been shown that carbonic acid exists only in solution. The carbonates, however, are well-known substances, and two of these, sodium bicarbonate and calcium carbonate, have already been considered.

The usual methods for the preparation of carbonates are:

1. The action of carbon dioxide on a hydroxide:

$$Ca(OH)_2 + CO_2 = CaCO_3 + H_2O.$$

2. The addition of sodium carbonate solution to the solution of a salt. Compare Experiment 121:

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

Experiment 122. — To a few cubic centimetres barium chloride solution, sodium carbonate solution is added until no more barium carbonate is precipitated. After filtration the barium carbonate is treated with hydrochloric acid. Compare Experiment 112:

$$\begin{aligned} BaCl_2 + Na_2CO_3 &= 2\,NaCl + BaCO_3. \\ BaCO_3 + 2\,HCl &= BaCl_2 + H_2O + CO_2. \end{aligned}$$

Experiment 123. — Action of Heat on Carbonates.—The action of heat on the following carbonates is examined: (1) Magnesium carbonate; (2) zinc carbonate; (3) sodium bicarbonate; and (4) sodium carbonate.

Carbon dioxide is tested for, and in each case where it is evolved, the residue is moistened with water and tested with litmus paper, and, finally, hydrochloric acid is added:

$$
\begin{aligned}
MgCO_3 &= MgO + CO_2. \\
MgO + H_2O &= Mg(OH)_2. \\
ZnCO_3 &= ZnO + CO_2. \\
2\,NaHCO_3 &= Na_2CO_3 + H_2O + CO_2.
\end{aligned}
$$

Experiment 124. — Estimation of Carbon Dioxide in Calcium Carbonate.—About 1 grm. (accurately weighed) of calcium carbonate is placed in a porcelain crucible which has been previously weighed, and is heated at the blowpipe for about fifteen minutes. The crucible and its contents are then cooled in a desiccator and weighed. It is again heated until the weight is constant, and the percentage loss of carbon dioxide is calculated.

CARBON MONOXIDE

This oxide of carbon is not formed when carbon burns in air or in oxygen, but when carbon dioxide is passed over carbon heated to redness:

$$C + CO_2 = 2\,CO.$$

Experiment 125 [D].—**Preparation from Sodium Formate.**—Three grams sodium formate and 5 cub. cm. concentrated sulphuric acid are placed in a test tube, which is fitted with a cork and delivery tube as shown in fig. 1. On heating the mixture carbon monoxide is evolved and one cylinder of the gas is collected over water.

A few cubic centimetres lime water are poured into the cylinder. No change occurs. A lighted taper is next applied to the gas, which burns with a blue flame. The lime water is turned milky, showing the presence of carbon dioxide:

$$2\,CO + O_2 = 2\,CO_2.$$

SULPHUR

The element sulphur can exist in four modifications, two of which are crystalline and two amorphous:—

1. Rhombic sulphur, which gets its name from the form of the crystals. Sulphur is found in nature in this form.
2. Prismatic or monoclinic sulphur.
3. Plastic sulphur.
4. Amorphous sulphur.

Of these four modifications the rhombic is the most stable at ordinary temperatures, and the other forms on standing tend to revert to it.

Preparation of the Different Modifications of Sulphur

Experiment 126. — Rhombic Sulphur. — About 1 grm. of roll sulphur is dissolved in a small quantity of carbon disulphide (all lights must be kept away), the solution filtered and allowed to evaporate spontaneously in a well-ventilated draught-cupboard. Well-defined rhombic crystals separate.

Experiment 127. — Prismatic or Monoclinic Sulphur. — A small quantity of sulphur is melted in a crucible at the lowest possible temperature and then allowed to cool slowly. The crust which forms on the surface is pierced with a file, and the remaining liquid is poured out. After a few minutes the crust is cut round the edge and removed, when the interior of the crucible is seen to contain long needles.

Experiment 128 [D].—**Plastic Sulphur.**—About 3 grm. of sulphur are melted in a test tube, clamped to a retort stand, and the temperature gradually raised until the sulphur boils. Care must be taken that the sulphur vapour does not catch fire. During the heating, changes in the colour and fluidity of the sulphur occur. When it melts, a yellow mobile liquid is produced, which at a higher temperature becomes darker in colour, and so viscous that though the test tube be inverted the liquid does not run out. At a still higher temperature the sulphur again becomes mobile, and on boiling is converted into a yellow gas. When the liquid sulphur begins to boil, it is poured into cold water. There is formed a solid resembling rubber. This is the so-called "plastic" sulphur. On standing for some hours it loses its elasticity and changes into the rhombic modification.

Experiment 129.—**Amorphous Sulphur.**—This modification is formed when certain sulphur compounds are decomposed. A few grams sodium thiosulphate are dissolved in water, and hydrochloric acid added to the solution. The precipitated sulphur is collected on a filter, and, after drying, a test is made of its solubility in carbon disulphide.

Experiment 130.—Sulphur is insoluble in water, but if shaken with a little water and allowed to stand it is oxidized to sulphuric acid. After removal of the unchanged sulphur by filtration, the acid is tested for by adding barium chloride solution. (See test for sulphuric acid, p. 92.)

Sulphur on burning in air or in oxygen forms sulphur dioxide. ($S + O_2 = SO_2$.)

Experiment 131 [D].—A small quantity of sulphur on the end of a copper wire is ignited by heating in the Bunsen flame. The sulphur dioxide can be detected by its suffocating odour.

Nitric acid oxidizes sulphur to sulphuric acid. (See Experiment 47.)

Sulphur combines directly with some metals to form sulphides. The blackening of the copper wire in Experiment 131 is due to the formation of copper sulphide. ($Cu + S = CuS$.)

Experiment 132.—Preparation of Ferrous Sulphide.—About 1 grm. of fine iron filings and ½ grm. of flowers of sulphur are gently heated in a test tube. ($Fe + S = FeS$.) The ferrous sulphide is kept for the next experiment.

SULPHURETTED HYDROGEN

Experiment 133.—Preparation [D].—Ferrous sulphide is treated in a test tube with a few cubic centimetres dilute hydrochloric acid. ($FeS + 2\,HCl = FeCl_2 + H_2S$.) The disgusting odour of the gas is observed, also its action on a silver coin, and on a piece of filter paper moistened with a solution of lead acetate. Silver sulphide and lead sulphide are formed. It is also shown that the gas burns. With a free supply of air, water and sulphur dioxide are formed ($2\,H_2S + 3\,O_2 = 2\,SO_2 + 2\,H_2O$). When the quantity of air is limited, water and sulphur are produced. ($2\,H_2S + O_2 = 2\,H_2O + 2\,S$.)

Experiment 134.—Preparation of Sulphuretted Hydrogen Water. — The gas obtained from the Kipp's apparatus in the H_2S-room is led into 50 cub. cm. of water, contained in a small flask, until the water smells strongly of the gas. (As sulphuretted hydrogen is a poison, care must be taken not to inhale it.)

The action of the solution on litmus paper is observed.

Sulphuretted hydrogen is regarded as an acid, and the sulphides as its salts, formed by the replacement of the hydrogen by metals.

Experiment 135.—Action of Sulphuretted Hydrogen on Solutions of Metallic Salts. — A few cubic centimetres of the sulphuretted hydrogen water prepared in Experiment 134 are added to each of the following solutions in separate test tubes: Copper sulphate, antimony chloride, ferrous sulphate, ferric chloride, barium chloride.

The following reactions occur:—

1. **Copper Sulphate.** — A black precipitate of cupric sulphide is formed:

$$CuSO_4 + H_2S = CuS + H_2SO_4.$$

2. **Antimony Chloride** gives an orange precipitate of antimony sulphide:

$$2\,SbCl_3 + 3\,H_2S = Sb_2S_3 + 6\,HCl.$$

3. **Ferrous Sulphate.**—The reaction represented by the equation—

$$FeSO_4 + H_2S = FeS + H_2SO_4$$

does not occur, because ferrous sulphide is soluble in dilute acids, and it is seen from the equation that sulphuric acid would be produced. If, however, the solution is made just alkaline with ammonia, the ferrous sulphide separates as a black precipitate.

For the precipitation of metallic sulphides, which are dissolved by dilute acids, ammonium sulphide [$(NH_4)_2S$] is used.

4. **Ferric Chloride.** — In this case a white precipitate appears on the addition of sulphuretted hydrogen water. That this precipitate is sulphur can be shown by filtering and testing it by burning and observing the smell of sulphur dioxide. (See Experiment 131.) Further, the ferric chloride is reduced to ferrous chloride:

$$2\,FeCl_3 + H_2S = 2\,FeCl_2 + 2\,HCl + S.$$

Then, as with ferrous sulphate, the sulphide is precipitated only after the solution has been made alkaline:

$$H_2S + 2\,NH_4OH = (NH_4)_2S + 2\,H_2O.$$
$$FeCl_2 + (NH_4)_2S = 2\,NH_4Cl + FeS.$$

5. **Barium Chloride.**—No reaction takes place in this case.

The behaviour of sulphuretted hydrogen towards solutions of metallic salts, and the different colour and solubilities of the metallic sulphides, is employed in the detection and separation of the metals in analytical chemistry.

SULPHUR DIOXIDE

Sulphur dioxide (SO_2) has already been formed by burning sulphur in oxygen (Experiment 15) and in the air (Experiment 131). It is, however, most conveniently prepared by acting on concentrated sulphuric acid with copper.

Experiment 136 [D].—**Preparation of Sulphur Dioxide.**—Twenty grams copper turnings are placed in a round flask fitted with a thistle funnel and delivery tube for collecting the gas by downward displacement. The flask is placed on a small sand bath in a draught-cupboard and 20 cub. cm. concentrated sulphuric acid are poured down the thistle funnel. The flask and its contents are gently heated, and as soon as the action commences the flame is withdrawn.

The gas issuing from the end of the delivery tube is allowed to come in contact with a piece of filter paper soaked in a solution of potassium dichromate, when the latter is turned green.

Two cylinders of the gas are now collected, the fact that they are full being shown by applying the potassium dichromate test just given. When filled, they are closed with greased glass plates.

Experiment 137. — Into the first cylinder is brought a lighted taper, which is immediately extinguished. Sulphur dioxide is non-inflammable and does not support combustion.

Experiment 138.—The second cylinder is inverted in a basin of water, and the glass plate removed. The water ascends in the cylinder, since sulphur dioxide is soluble in water. The action of the solution on litmus is now observed.

Experiment 139.—Reducing Action of Sulphur Dioxide.—Sulphur dioxide is passed into a few cubic centimetres sulphuretted hydrogen water. A white precipitate of sulphur appears:

$$2\,H_2S + SO_2 = 2\,H_2O + 3\,S.$$

The sulphur can be tested for in the usual way. The action of sulphur dioxide on potassium dichromate has been seen in Experiment 136.

Experiment 140.—The action of sulphur dioxide on a solution of potassium permanganate is also examined. The solution is rendered colourless:

$$\begin{aligned} 2\,KMnO_4 + 5\,SO_2 + 2\,H_2O \\ = K_2SO_4 + 2\,MnSO_4 + 2\,H_2SO_4. \end{aligned}$$

Sulphurous Acid and Sulphites

Sulphur dioxide has been shown to be an acidic oxide, since with water it forms an acid, sulphurous acid (H_2SO_3). The acid itself is unknown in the pure state, but the salts—sulphites— are well known, and can be prepared by the action of sulphur dioxide on bases.

Experiment 141.—Sulphur dioxide is passed into about 10 cub. cm. of a solution of sodium hydroxide in a test tube until no more of the gas is absorbed, and the solution is then divided into three portions:

(*a*) The first portion is evaporated to dryness in a porcelain basin. The residue is sodium hydrogen sulphite, which was formed according to the equation:

$$NaOH + SO_2 = NaHSO_3.$$

Normal sodium sulphite (Na_2SO_3) can be formed by saturating 10 cub. cm. with sulphur dioxide and adding another 10 cub. cm. of the sodium hydroxide solution:

$$NaHSO_3 + NaOH = Na_2SO_3 + H_2O.$$

The action of dilute hydrochloric acid on the sulphite is examined, and the gas which is evolved detected:

$$NaHSO_3 + HCl = NaCl + H_2O + SO_2.$$

(*b*) To the second part of the solution of sodium hydrogen sulphite an equal volume of sodium hydroxide solution is added to form sodium sulphite, and then a few drops of barium chloride solution. A white precipitate of barium sulphite appears:

$$Na_2SO_3 + BaCl_2 = BaSO_3 + 2\,NaCl.$$

It is shown that the precipitate is dissolved by hydrochloric acid:

$$BaSO_3 + 2\,HCl = H_2O + SO_2 + BaCl_2.$$

(*c*) Bromine water is added drop by drop to the third portion until the solution remains yellow. On the addition of barium chloride solution there appears a white precipitate which is insoluble in hydrochloric acid:

$$\begin{aligned} Na_2SO_3 + Br_2 + H_2O &= Na_2SO_4 + 2\,HBr. \\ Na_2SO_4 + BaCl_2 &= BaSO_4 + 2\,NaCl. \end{aligned}$$

The sodium sulphite is oxidized to sodium sulphate, which with barium chloride produces barium sulphate, insoluble in hydrochloric acid.

Experiment 142 [D]. — Preparation of Copper Sulphate.—The contents of the flask used in preparing sulphur dioxide are transferred to a porcelain basin in a draught-cupboard and gently heated until all action has ceased. A small quantity of water is added, the solution filtered, and the filtrate evaporated to small bulk and allowed to cool. Blue crystals of copper sulphate ($CuSO_4 . 5\,H_2O$) separate.

SULPHURIC ACID

Sulphur trioxide (SO_3), the anhydride of sulphuric acid cannot be prepared by a simple laboratory method, nor can sulphuric acid itself. The latter, however, is a very important compound and its principal properties are shown in the following experiments:

Experiment 143.—The appearance, viscosity, and high specific gravity of the concentrated acid are noted.

Experiment 144.—A splinter of wood is dipped in concentrated acid in a test tube and writing traced on a piece of paper. Both wood and paper are charred, especially on gently warming.

Experiment 145.—Action of Concentrated Sulphuric Acid on Metals.—A small quantity of concentrated acid is added to small pieces of copper, tin, zinc, and lead in separate test tubes. If no action takes place in the cold, heat is applied:

$$\begin{aligned} Cu + 2\,H_2SO_4 &= CuSO_4 + SO_2 + 2\,H_2O. \\ Sn + 2\,H_2SO_4 &= SnSO_4 + SO_2 + 2\,H_2O. \end{aligned}$$

Zinc and lead are practically unacted upon by concentrated sulphuric acid. (Compare Preparation of Hydrogen, Experiment 21.)

Experiment 146 [D].—**Action of the Concentrated Acid on Carbon and on Sulphur.**—As in Experiment 145, small pieces of carbon and sulphur in separate test tubes are treated with concentrated sulphuric acid:

$$C + 2\,H_2SO_4 = 2\,SO_2 + CO_2 + 2\,H_2O.$$

The usual tests for the two gases are applied.

$$S + 2\,H_2SO_4 = 3\,SO_2 + 2\,H_2O.$$

Experiment 147.—Sulphuric Acid Soluble in Water.—About 10 cub. cm. of the concentrated acid are slowly poured into about 50 cub. cm. water contained in a beaker. During the addition the water is constantly stirred. The sulphuric acid dissolves in the water with evolution of heat.

Experiment 148. — Action of the Dilute Acid on some Metals.—Small pieces of copper, zinc, and iron in separate test tubes are treated with some of the dilute acid prepared in Experiment 147. The evolved gas is tested in each case, and the equation written down.

Sulphates

Sulphates can be prepared by the general methods for preparing salts. Zinc sulphate and copper sulphate have already been made (Experiments 25 and 142) by action of the acid on the metals.

By acting on an oxide, hydroxide, or carbonate of a metal with sulphuric acid, sulphates can also be obtained. (Compare chlorides, p. 66.)

Experiment 149. — Preparation of Sulphates which are *Insoluble or only Slightly Soluble.*— To a few cubic centimetres of a fairly concentrated solution of calcium chloride, a few drops dilute sulphuric acid are added. The *white precipitate* is calcium sulphate:

$$CaCl_2 + H_2SO_4 = CaSO_4 + 2\,HCl.$$

In a similar manner lead sulphate and barium sulphate are obtained:

$$Pb(NO_3)_2 + H_2SO_4 = PbSO_4 + 2\,HNO_3.$$
$$BaCl_2 + H_2SO_4 = BaSO_4 + 2\,HCl.$$

If lead sulphate is not precipitated, a few cubic centimetres alcohol are added.

Sulphuric acid is a dibasic acid, and forms normal and acid salts. With sodium, for example, it forms sodium sulphate (Na_2SO_4), and sodium bisulphate or sodium hydrogen sulphate ($NaHSO_4$).

* **Experiment 150. — Preparation of Sodium Sulphate and Sodium Hydrogen Sulphate.**—Five cubic centimetres sodium hydroxide solution are measured into a porcelain basin which has been weighed, and dilute sulphuric acid gradually added from a burette until the solution is neutral. The reaction of the solution is tested during the addition of the acid by taking out a drop on a glass rod and bringing it in contact with litmus paper. Litmus should not be added to the sodium hydroxide. The volume of acid added is carefully noted.

Five cubic centimetres of the sodium hydroxide solution are measured into a second weighed porcelain basin, and twice the volume of sulphuric acid which was required in the first case is added.

Both solutions are evaporated almost to dryness on a water bath, then to complete dryness on a sand bath, and, after cooling in a desiccator, are weighed. During the evaporation, care must be taken to avoid loss by spirting:

$$2\,NaOH + H_2SO_4 = Na_2SO_4 + H_2O.$$
$$NaOH + H_2SO_4 = NaHSO_4 + H_2O.$$

Experiment 151. — The previous experiment is repeated, using hydrochloric acid instead of sulphuric acid. Only one sodium salt is formed.

Experiment 152.—Test for Sulphuric Acid and Sulphates.—To a few cubic centimetres of a solution of any sulphate a few drops of barium chloride solution are added. The precipitated barium sulphate is insoluble in hydrochloric acid. (Compare carbonates, p. 79, and sulphites, p. 88.)

Experiment 153.—Test for an Insoluble Sulphate.—A small quantity of an insoluble sulphate, for example barium sulphate, is mixed with sodium carbonate and heated on charcoal in the reducing flame of the blowpipe. Sodium sulphide is produced, and can be tested for by placing it on a silver coin and adding a drop of water. A black stain of silver sulphide appears on the coin.

APPENDIX

International Atomic Weights (1911)

		O = 16.
Aluminium	Al	27·1
Antimony	Sb	120·2
Argon	Ar	39·88
Arsenic	As	74·96
Barium	Ba	137·37
Beryllium	Be	9·1
Bismuth	Bi	208·0
Boron	B	11·0
Bromine	Br	79·92
Cadmium	Cd	112·4
Cæsium	Cs	132·81
Calcium	Ca	40·09
Carbon	C	12·0
Cerium	Ce	140·25
Chlorine	Cl	35·46
Chromium	Cr	52·0
Cobalt	Co	58·97
Copper	Cu	63·57
Dysprosium	Dy	162·5
Erbium	Er	167·4
Europium	Eu	152·0
Fluorine	F	19·0
Gadolinium	Gd	157·3
Gallium	Ga	69·9
Germanium	Ge	72·5
Gold	Au	197·2
Helium	He	3·99
Hydrogen	H	1·008
Indium	In	114·8
Iodine	I	126·92
Iridium	Ir	193·1
Iron	Fe	55·85
Krypton	Kr	82·9
Lanthanum	La	139·0
Lead	Pb	207·1
Lithium	Li	6·94
Lutetium	Lu	174·0
Magnesium	Mg	24·32
Manganese	Mn	54·93
Mercury	Hg	200·0
Molybdenum	Mo	96·0
Neodymium	Nd	144·3
Neon	Ne	20·2
Nickel	Ni	58·68
Niobium	Nb	93·5
Nitrogen	N	14·01
Osmium	Os	190·9
Oxygen	O	16·0
Palladium	Pd	106·7
Phosphorus	P	31·04
Platinum	Pt	195·2
Potassium	K	39·1
Praseodymium	Pr	140·6
Radium	Ra	226·4
Rhodium	Rh	102·9
Rubidium	Rb	85·45
Ruthenium	Ru	101·7
Samarium	Sm	150·4
Scandium	Sc	44·1
Selenium	Se	79·2
Silicon	Si	28·3
Silver	Ag	107·88
Sodium	Na	23·0
Strontium	Sr	87·63
Sulphur	S	32·07
Tantalum	Ta	181·0
Tellurium	Te	127·5
Terbium	Tb	159·2
Thallium	Tl	204·0
Thorium	Th	232·4
Thulium	Tu	168·5
Tin	Sn	119·0
Titanium	Ti	48·1
Tungsten	W	184·0
Uranium	U	238·5
Vanadium	V	51·06
Xenon	X	130·2
Ytterbium	Yb	172·0
Yttrium	Y	89·0
Zinc	Zn	65·37
Zirconium	Zr	90·6

Tension of Aqueous Vapour at Different Temperatures

t = temperature in degrees Centigrade.
p = tension in millimetres of mercury.

t.	p.	t.	p.	t.	p.	t.	p.
0	4·6	7	7·5	14	11·9	21	18·4
1	4·9	8	8·0	15	12·6	22	19·6
2	5·3	9	8·5	16	13·5	23	20·8
3	5·6	10	9·1	17	14·4	24	22·1
4	6·0	11	9·8	18	15·3	25	23·5
5	6·5	12	10·4	19	16·3	26	24·9
6	7·0	13	11·1	20	17·3	27	26·4

British and Metric Systems of Measurement

Length—

1 m. = 10 dm. = 100 cm. = 1000 mm. = 39·37 in.
1 in. = 2·54 cm. = 25·4 mm.

Area—

1 sq. cm. = 0·15 sq. in.
1 sq. in. = 6·45 sq. cm.

Volume—

1 cub. cm. = 0·06 cub. in.
1 litre = 61·02 cub in.
1 cub. in. = 16·4 cub. cm.
1 gal. = 4·54 litres.
1 litre = 1·76 pint.

Weight—

1 grm. = 15·43 grains.
1 grain = 0·0648 grm.
1 oz. (avoir.) = 28·35 grm.
1 lb. ,, = 453·5926 grm.
1 kilo = 2·2 lb.

Temperature—Centigrade and Fahrenheit

To convert degrees Centigrade into degrees Fahrenheit, and vice versa, the following formulæ are used:—

$$F^\circ = \frac{9}{5}C^\circ + 32.$$

$$C^\circ = \frac{5}{9}(F^\circ - 32).$$

A BRIEF LIST of BLACKIE & SON'S SCIENTIFIC, TECHNICAL, AND MATHEMATICAL PUBLICATIONS

PHYSICS AND CHEMISTRY

ARCHIBALD AND RANKIN—ELECTRICAL LABORATORY COURSE. For Junior Students. By R. D. Archibald, B.Sc., A.M.Inst.C.E., and R. Rankin, A.G.T.C., Stud.I.E.E. 1*s.* 6*d.* net.

BERNTHSEN—A TEXTBOOK OF ORGANIC CHEMISTRY. English Translation from the German of A. Bernthsen, Ph.D. *New Edition.* Edited and revised to date by Prof. J. J. Sudborough, Ph.D., D.Sc., F.I.C. 7*s.* 6*d.* net.

CAVEN—SYSTEMATIC QUALITATIVE ANALYSIS. For Students of Inorganic Chemistry. By R. M. Caven, D.Sc.(Lond.), F.I.C. 3*s.* 6*d.* net.

CAVEN AND LANDER — SYSTEMATIC INORGANIC CHEMISTRY from the Standpoint of the Periodic Law. A Textbook for Advanced Students. By R. M. Caven, D.Sc.(Lond.), and Prof. G. D. Lander, D.Sc. 6*s.* net.

CHEETHAM—ELEMENTARY CHEMISTRY. *Practical and Theoretical.* By T. A. Cheetham, F.C.S. First Year's Course. 1*s.* 6*d.* Second Year's Course. 2*s.*

DESCHANEL—DESCHANEL'S NATURAL PHILOSOPHY. *An Elementary Treatise.* By Professor A. Privat Deschanel, of Paris. Edited by Professor J. D. Everett, M.A., D.C.L., F.R.S. *Sixteenth Edition.* Medium 8vo, cloth, 18*s.* The following parts may be had separately, limp cloth, 4*s.* 6*d.* each. Part II. Heat; Part III. Electricity and Magnetism; Part IV. Sound and Light.

DRAPER—A TEXTBOOK OF HEAT. By Charles H. Draper, B.A., D.Sc.(Lond.). 4*s.* 6*d.*

———LIGHT, HEAT, AND SOUND. By Charles H. Draper, B.A., D.Sc. 2*s.*

ELLIS—ELEMENTARY PRINCIPLES OF ALTERNATING-CURRENT DYNAMO DESIGN. By A. G. Ellis, A.C.G.I., A.Am.I.E.E. 12*s.* net.

ENDECOTT—**A SCHOOL COURSE IN PHYSICS.** LIGHT AND SOUND. By F. C. ENDECOTT. A Notebook, Exercise Book, and Textbook in one. With or without Answers. 2*s.* 6*d.*

EVERETT—**ELEMENTARY TEXTBOOK OF PHYSICS.** By Professor J. D. EVERETT, M.A., D.C.L., F.R.S. 3*s.* 6*d.*

FERGUS—**ELEMENTARY OPHTHALMIC OPTICS.** By FREELAND FERGUS, M.D., F.R.S.E. 3*s.* 6*d.* net.

GUNN—**THE ARITHMETIC OF MAGNETISM AND ELECTRICITY.** With numerous Examples and Exercises. By ROBERT GUNN. 2*s.* 6*d.*

HARRISON—**CHEMISTRY FOR ALL: or, Elementary Alternative Chemistry.** By W. JEROME HARRISON, F.G.S., and R. J. BAILEY. 1*s.* 6*d.*

HENDERSON AND PARKER—**AN INTRODUCTION TO ANALYTICAL CHEMISTRY.** By Professor GEORGE G. HENDERSON, D.Sc., M.A., F.I.C., and Professor MATTHEW A. PARKER, B.Sc. With Coloured Plate. 5*s.*

JAMIESON—**HEAT SHADOWS.** Some new experiments in conduction and radiation of heat. By WALTER JAMIESON, B.Sc., A.M.I.E.E. 6*d.* net.

KERR AND BROWN—**ELEMENTARY PHYSICS.** *Practical and Theoretical.* First Year's Course. By JOHN G. KERR, M.A., LL.D. 1*s.* 6*d.* Second Year's Course. By JOHN G. KERR, M.A., LL.D., and JOHN N. BROWN, A.R.C.S. Illustrated. 2*s.* Third Year's Course. By JOHN N. BROWN, A.R.C.S. Illustrated. 2*s.*

LOW—**AN ELEMENTARY TEXTBOOK OF APPLIED MECHANICS.** By Prof. DAVID ALLAN LOW (Whitworth Scholar), M.I.Mech.E. 2*s.*

PARR—**ELECTRICAL ENGINEERING MEASURING INSTRUMENTS.** By G. D. ASPINALL PARR, M.Inst.E.E., A.M.I.Mech.E. With 370 Illustrations. Demy 8vo, 9*s.* net.

PINKERTON—**ELEMENTARY TEXTBOOK OF DYNAMICS AND HYDROSTATICS.** By R. H. PINKERTON, B.A. 3*s.* 6*d.*

———**HYDROSTATICS AND PNEUMATICS.** By R. H. PINKERTON, B.A. 4*s.* 6*d.*

———**THEORETICAL MECHANICS.** (SOLIDS and LIQUIDS.) By R. H. PINKERTON, B.A.(Oxon). 2*s.*

RICHARDSON—**MAGNETISM AND ELECTRICITY.** AND THE PRINCIPLES OF ELECTRICAL MEASUREMENT. By S. S. RICHARDSON, B.Sc. 5*s.* net.

SEXTON—**ELEMENTARY INORGANIC CHEMISTRY.** *The retical and Practical.* With a course of chemical analysis and series of Examples in Chemical Arithmetic. By Prof. A. Humbol Sexton, F.I.C., F.C.S. 2*s.* 6*d.*

SUDBOROUGH AND JAMES—**PRACTICAL ORGANI CHEMISTRY.** By Prof. J. J. Sudborough, Ph.D., D.Sc F.I.C., and T. Campbell James, M.A., B.Sc. 5*s.* net.

TIMPANY—**THE ARITHMETIC OF ELEMENTAR PHYSICS AND CHEMISTRY.** By H. M. Timpany, M.Sc. 1

WELCH—**CHEMISTRY LECTURE NOTES.** By G. E. Welc B.Sc. 1*s.* 6*d.*

BIOLOGY

Including Botany, Physiology, Hygiene, &c.

CLARK—**AN ELEMENTARY TEXTBOOK OF ANATOM** By Professor Henry Edward Clark. Illustrated. 5*s.*

CLAXTON—**METHODICAL NATURE STUDY.** By Wm. Claxton. Fully Illustrated with coloured and black-and-whi plates. 6*s.*

DAVIS—**ELEMENTARY PHYSIOLOGY.** By Prof. Ainswort Davis. With Appendix for Agricultural Students. Cloth, 2*s.*

DOUGLAS—**THE LAWS OF HEALTH.** A Handbook School Hygiene. By Carstairs C. Douglas, M.D., D.Sc. 3*s.* ne

ELKINGTON—**HEALTH IN THE SCHOOL: or, Hygiene f Teachers.** By J. S. C. Elkington, M.D., D.P.H. Illustrate 2*s.* net.

ELLIOT—**A FIRST COURSE IN PRACTICAL BOTAN** By G. F. Scott Elliot, M.A., B.Sc., F.L.S., F.R.G.S. With ov 150 Illustrations. 3*s.* 6*d.*

———**NATURE STUDIES (Plant Life).** By G. F. Scot Elliot, M.A., B.Sc., F.L.S. Fully Illustrated. 3*s.* 6*d.*

HARRIS—**SEASONAL BOTANY.** By M. O'Brien Harri D.Sc. Interleaved. 8*d.*

KERNER—**THE NATURAL HISTORY OF PLANTS: Thei Forms, Growth, Reproduction, and Distribution.** From th German of the late Anton Kerner von Marilaun, by F. W Oliver, M.A., D.Sc., with the assistance of Lady Busk, B.Sc and Mrs. M. F. Macdonald, B.Sc. 2000 Illustrations. In tw volumes, 30*s.* net.

KNIGHT—**FOOD AND ITS FUNCTIONS.** A Textbook f Students of Cookery. By James Knight, M.A., B.Sc., D.Sc Illustrated. 2*s.* 6*d.*

M'GREGOR-ROBERTSON—**AN ELEMENTARY TEXT-BOOK OF PHYSIOLOGY** (Elementary and Advanced). By J. M'GREGOR-ROBERTSON, M.A., M.B., C.M. 4*s.*

MURCHÉ—**ELEMENTARY PHYSIOLOGY.** By VINCENT T. MURCHÉ. 2*s.*

OLIVER—**ELEMENTARY BOTANY.** By JOSEPH W. OLIVER. *New Edition.* Revised and Enlarged by W. B. GROVE, M.A. 2*s.*

———***THE STUDENT'S INTRODUCTORY TEXTBOOK*** **OF SYSTEMATIC BOTANY.** By J. W. OLIVER. 4*s.* 6*d.*

PEDLEY AND HARRISON—**OUR TEETH: How Built Up, How Destroyed, How Preserved.** Described and Illustrated by R. DENISON PEDLEY, F.R.C.S.Ed., L.D.S.Eng., and FRANK HARRISON, M.R.C.S.Eng., L.D.S.Ed. Richly Illustrated from Microscope Photographs. Demy 8vo, 5*s.* net.

STOPES—**ANCIENT PLANTS.** A simple account of the past vegetation of the Earth and of recent important discoveries made in this realm of Nature Study. By MARIE C. STOPES, D.Sc., Ph.D., F.L.S., Lecturer in Fossil Botany, Manchester University. 4*s.* 6*d.* net.

———**THE STUDY OF PLANT LIFE.** *Second Edition.* Fully Illustrated. By M. C. STOPES, D.Sc., &c. 3*s.* 6*d.*

THOMSON—**THE SCIENCE OF LIFE.** An outline of the History of Biology and its recent advances. By Prof. J. ARTHUR THOMSON, M.A. Cloth, 2*s.* 6*d.*

WAKEFIELD—**ELEMENTARY HYGIENE.** By H. ROWLAND WAKEFIELD. 2*s.*

———**ELEMENTARY HYGIENE AND PHYSIOLOGY.** By H. ROWLAND WAKEFIELD. 2*s.* 6*d.*

AGRICULTURE, ETC.

THE AGRICULTURIST'S CALCULATOR. A Series of Forty-six Tables for Land Measuring, Draining, Manuring, Planting, Weight of Hay and Cattle by Measurement, Building, &c. Forming a complete ready-reckoner for all engaged in Agriculture. F'cap 8vo Bound in roan, 9*s.*

FLEISCHMANN—**THE BOOK OF THE DAIRY.** A Manual of the Science and Practice of Dairy Work. Translated from the German of W. FLEISCHMANN, Ph.D., by C. M. AIKMAN, M.A., Sc.D., F.R.S.E., F.I.C., and R. PATRICK WRIGHT, F.H.A.S., F.R.S.E. Fully Illustrated. Demy 8vo, 10*s.* 6*d.*

MACEWEN—**THE PUBLIC MILK SUPPLY.** A Practical Handbook. By HUGH A. MACEWEN, M.B., Ch.B.(Glas.), D.P.H. (Lond.), D.P.H.(Cantab.). Illustrated. 2*s.* 6*d.* net.

WEBB—**THE PRINCIPLES OF HORTICULTURE.** A series of Practical Scientific Lessons. By WILFRED MARK WEBB, F.L.S. Illustrated. F'cap 4to, 2*s.* net.

WRIGHT—**THE PRINCIPLES OF AGRICULTURE.** By Prof. R. PATRICK WRIGHT, F.H.A.S., F.R.S.E. F'cap 8vo, 1*s.* 6*d.*

DAVIS—**ELEMENTARY PHYSIOLOGY.** With Appendix for Agricultural Students. By Prof. J. R. AINSWORTH DAVIS, M.A. 2*s.*

GEOLOGY, MINING, ETC.

DRON—**THE COAL-FIELDS OF SCOTLAND.** By ROBERT W. DRON, A.M.I.C.E., M.I.M.E. Demy 8vo, cloth, 15*s.* net.

FORBES—**FIRST STEPS IN COAL-MINING.** For use in Supplementary and Continuation Classes. By ALEXANDER FORBES, M.Inst.M.E. Illustrated. 2*s.* 6*d.*

GREGORY—**GEOGRAPHY: Structural, Physical, Comparative.** A Textbook for Advanced Students. By Prof. J. W. GREGORY, D.Sc., F.R.S. With a series of specially constructed Coloured Maps, and Sketch Maps and Plans in Black-and-White. 6*s.* net.

———**PHYSICAL AND STRUCTURAL GEOGRAPHY.** By Professor J. W. GREGORY, D.Sc., F.R.S. Being the Introductory Part of the above, printed separately with Plates and Coloured Maps. 2*s.* 6*d.*

HARRISON—**A TEXTBOOK OF GEOLOGY.** An Introduction to the Study of Rocks and their Contents. By W. JEROME HARRISON, F.G.S. 3*s.* 6*d.*

HARRISON—**ELEMENTS OF METALLURGY.** By W. JEROME HARRISON, F.G.S., and WILLIAM J. HARRISON. Illustrated. 2*s.* 6*d.*

PEEL—**ELEMENTARY TEXTBOOK OF COAL-MINING.** By ROBERT PEEL, Mining Engineer. Illustrated. 2*s.* 6*d.*

SCHWARZ—**CAUSAL GEOLOGY.** By Professor E. H. L. SCHWARZ, A.R.C.S., F.G.S. Illustrated. Demy 8vo, 7*s.* 6*d.* net.

SEXTON—**FUEL AND REFRACTORY MATERIALS.** By Prof. A. HUMBOLDT SEXTON, F.I.C., F.C.S. *Second Edition, Revised and Enlarged.* 5*s.* net.

STOPES—**ANCIENT PLANTS.** A simple account of the past vegetation of the earth and of recent important discoveries in that realm of nature-study. By MARIE C. STOPES, D.Sc., Ph.D., F.L.S. Demy 8vo, 4*s.* 6*d.* net.

ENGINEERING

CLARK—**A MANUAL OF RULES, TABLES, AND DATA FOR MECHANICAL ENGINEERS.** By DANIEL KINNEAR CLARK, M.Inst.C.E., M.I.M.E. Demy 8vo, cloth, 16*s.*; half-morocco, 20*s.*

ELLIS—**ELEMENTARY PRINCIPLES OF ALTERNATING-CURRENT DYNAMO DESIGN.** By H. G. ELLIS, A.C.G.I., A.Am.I.E.E. 12*s.* net.

GOUDIE—**THE GEOMETRY OF THE SCREW PROPELLER.** By WILLIAM J. GOUDIE, B.Sc. F'cap 4to, paper, 1*s.* 6*d.* net.

PARR—**ELECTRICAL ENGINEERING MEASURING INSTRUMENTS.** By G. D. ASPINALL PARR, M.Inst.E.E., A.M.I.Mech.E. With 370 Illustrations. Demy 8vo, 9*s.* net.

SHAXBY—**ELEMENTARY ELECTRICAL ENGINEERING.** By J. W. SHAXBY, B.Sc. Illustrated. 3*s.* net.

ARCHITECTURE, BUILDING, ETC.

CROSSKEY—**ELEMENTARY PERSPECTIVE.** Arranged to meet the requirements of architects, draughtsmen, and art students. By LEWES R. CROSSKEY. Fully Illustrated. Imperial 8vo, 3*s.* 6*d.*

CROSSKEY AND THAW—**ADVANCED PERSPECTIVE.** Involving the drawing of objects when placed in oblique positions, shadows, and reflections. Arranged to meet the requirements of architects, daughtsmen, and art students. By LEWES R. CROSSKEY and JAMES THAW. Fully Illustrated. Imperial 8vo, 4*s.* 6*d.*

GOURLAY—**ELEMENTARY BUILDING CONSTRUCTION AND DRAWING FOR SCOTTISH STUDENTS.** By Prof. CHARLES GOURLAY, B.Sc. With 24 plates and full descriptive letterpress., 6*s.* net. Plates of the above, with introduction, done up in portfolio. 3*s.* 6*d.* net.

M'INTYRE—**A TEXTBOOK OF SCIOGRAPHY.** By J. H. A. M'INTYRE, M.I.Mech.E., Whitworth Scholar. Imperial 8vo, 3*s.* 6*d.*

PEDDIE—**THE PRACTICAL MEASURER: or, Tradesman and Wood Merchant's Assistant.** A series of tables for the use of wood merchants, builders, carpenters, joiners, sawyers, plasterers, painters, glaziers, masons, bricklayers, &c. With a treatise on the mensuration of superficies and solids. By ALEXANDER PEDDIE. Illustrated with diagrams and examples. F'cap 8vo, bound in roan, 7*s.* 6*d.*

PUBLIC HEALTH, ETC.

DOUGLAS—**THE LAWS OF HEALTH.** A Handbook of School Hygiene for Teachers. By CARSTAIRS C. DOUGLAS, M.D., D.Sc., F.R.S.E. Illustrated. 3*s.* net.

ELKINGTON—**HEALTH IN THE SCHOOL: or, Hygiene for Teachers.** By J. S. C. ELKINGTON, M.D., D.P.H. Illustrated. 2*s.* net.

MACEWEN—FOOD INSPECTION. A Practical Handbook. By Hugh A. Macewen, M.B., Ch.B.(Glas.), D.P.H.(Lond.), D.P.H.(Cantab.). Illustrated. Crown 8vo, 5*s.* net.

———THE PUBLIC MILK SUPPLY. A Practical Handbook. By Hugh A. Macewen. Illustrated. 2*s.* 6*d.* net.

PEDLEY AND HARRISON—OUR TEETH: **How Built Up, How Destroyed, How Preserved.** Described and Illustrated by R. Denison Pedley, F.R.C.S.Ed., L.D.S.Eng., and Frank Harrison, M.R.C.S.Eng., L.D.S.Ed. Richly Illustrated from Microscope Photographs. Demy 8vo, 5*s.* net.

MATHEMATICS

BARR—PRACTICAL MATHEMATICS. Embodying a Preparatory Technical Course for Craftsmen. By Thomas Barr, B.A. 2*s.* net.

BRIDGETT AND HYSLOP—TRIGONOMETRY. *Theoretical and Practical.* Embodying Graphic and other Solutions. By R. C. Bridgett, M.A., B.Sc., F.C.S., and William Hyslop, M.A., B.Sc. 4*s.* 6*d.*

BRIDGETT—EXPERIMENTAL TRIGONOMETRY. By R. C. Bridgett, M.A., B.Sc., F.C.S. 1*s.*

CAIN—A BRIEF COURSE IN THE CALCULUS. By Prof. William Cain, A.M., C.E. 6*s.* net.

DAVISON—EASY MATHEMATICAL PROLBEM PAPERS. By Charles Davison, Sc.D. 2*s.* With Answers and Hints for Solution. 2*s.* 6*d.*

DILWORTH—A NEW SEQUEL TO EUCLID. By Professor W. J. Dilworth, M.A. 2*s.* 6*d.* In parts: Part I, 1*s.* Part II, 2*s.*

DOUGALL—TEST PAPERS IN MATHEMATICS. Scottish University Preliminary—Higher and Intermediate Standard. Arranged by John Dougall, M.A. With Answers, 1*s.*

HAWKINS—ELEMENTARY GEOMETRY OF THE STRAIGHT LINE, CIRCLE, AND PLANE RECTILINEAL FIGURES. By Cecil Hawkins, M.A. 3*s.* 6*d.* *Key*, 5*s.* net.

KERR—ALGEBRA. Up to and Including Progressions and Scales of Notation. By J. G. Kerr, M.A., LL.D. 2*s.* With Answers, 2*s.* 6*d.*

———CONSTRUCTIVE GEOMETRY. Being steps in the synthesis of ideas regarding the properties and relations of geometrical figures arranged for the First Year's Course in Science. By J. G. Kerr, LL.D. 1*s.* 6*d.*

KNIGHT—**MATHEMATICAL WRINKLES FOR MATRICULATION AND OTHER EXAMS.** Six Sets of London Matriculation Papers, with full Solutions. By Dr. W. T. KNIGHT. 2*s.* 6*d.*

———**ALGEBRAIC FACTORS.** HOW TO FIND THEM AND HOW TO USE THEM. By Dr. KNIGHT. 2*s.* *Key*, 3*s.* 6*d.* net.

LAYNG—**A GENERAL TEXTBOOK OF ELEMENTARY ALGEBRA.** Covering the ground of a school or ordinary college course. By A. E. LAYNG, M.A. Complete, with or without Answers, 4*s.* 6*d.*; or in two volumes, with or without Answers, 2*s.* 6*d.* each. BOOK I.—To Quadratic Equations. BOOK II.—To Binomial Theorem, Exponential and Logarithmic Series. EXERCISES ONLY.—Exercises to Book I. 1*s.* With Answers, 1*s.* 6*d.* Exercises to Book II. 1*s.* 6*d.* With Answers, 2*s.*

———**A FIRST COURSE OF ALGEBRA.** Embodying Graphic and other Solutions and Applications to Arithmetic and Geometry, and forming a brief First Course. By A. E. LAYNG, M.A. With or without Answers. 2*s.* 6*d.*

———**EUCLID'S ELEMENTS OF GEOMETRY.** Edited by A. E. LAYNG, M.A. With Exercises, Appendix, and Examination Papers. BOOKS I to VI, and XI. 3*s.* 6*d.* BOOK I, 1*s.*; II, 6*d.*; III, 1*s.*; IV, 6*d.*; V and VI together, 1*s.*; XI, 1*s.* 6*d.* BOOKS I–II, 1*s.* 3*d.*; I–III, 2*s.*; I–IV, 2*s.* 6*d.* *Key* to BOOK I, 2*s.* 6*d.* net; complete, 5*s.* net.

LEIGHTON—**INTERMEDIATE GEOMETRY.** Experimental, Theoretical, Practical. By A. LEIGHTON, M.A., B.Sc. 1*s.* 6*d.*

———**ELEMENTARY MATHEMATICS.** Algebra and Geometry. By A. LEIGHTON, M.A., B.Sc. 2*s.*

LOGARITHMS—**A HANDY BOOK OF LOGARITHMS.** With Practical Geometrical Appendix. Demy 8vo, cloth, 2*s.*

LYSTER—**MATHEMATICAL FACTS AND FORMULÆ.** Being Hints to Junior Students in Arithmetic, Algebra, and Plane Trigonometry. Compiled by A. E. LYSTER, M.A. Cloth limp, 9*d.*; cloth boards, 1*s.*

MILLAR—**AN INTRODUCTION TO THE DIFFERENTIAL AND INTEGRAL CALCULUS.** With examples of applications to Mechanical Problems. By W. J. MILLAR, C.E. 1*s.* 6*d.*

PINKERTON—**ELEMENTARY TEXTBOOK OF TRIGONOMETRY.** By R. H. PINKERTON, B.A.(Oxon). 2*s.*

SCHMALL—**A FIRST COURSE IN ANALYTICAL GEOMETRY—Plane and Solid.** With numerous Examples. By CHARLES N. SCHMALL. 6*s.* net.

TOMES—**A FIRST COURSE IN PRACTICAL MATHEMATICS.** By B. A. TOMES. 1*s.* 6*d.*

TURNBULL—**ELEMENTARY PLANE GEOMETRY.** By V. M. TURNBULL, M.A. 2*s.*