NIOR GRADE SCIENCE

EORGE A. WATSON

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JUNIOR GRADE SCIENCE

AN INTRODUCTORY COURSE OF PHYSICS AND CHEMISTRY FOR IRISH INTERMEDIATE SCHOOLS

BY

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PREFACE

In the compilation of this book the object has been to provide in convenient form a course in Physics and Chemistry suitable for use in connection with the new syllabus prescribed for Junior Grade students by the Department of Agriculture and Technical Instruction for Ireland. It is interesting to note, in connection with the new Irish syllabuses in science, that they anticipate by about a year the conclusions of Sir J. J. Thomson's Committee appointed to enquire into the Position of Natural Science in the Educational System of Great Britain.

The report states that "A general course in science should fulfil two functions; (*a*) it should train the mind of the student to reason about things he has observed for himself and develop his powers of weighing and interpreting evidence; (*b*) it should also make him acquainted with the broad outlines of great scientific principles, with the way in which these principles are exemplified in familiar phenomena and with their applications to the service of man." Further, the report points out that the time allotted to the study of science, even if on the most liberal scale, does not permit the various branches of science to receive the detailed treatment possible where only one or two branches are attempted. "There must necessarily be great gaps in the student's knowledge of these branches, but in the general course we think these should be distributed, rather than that one or more of these branches should escape consideration altogether."

The new syllabuses of the Department aim at carrying out this object, and in so doing they introduce certain changes in the matter and methods of science teaching in Irish Intermediate Schools. In the Junior Grade (the needs of which this book are intended to meet) students are now required to possess the knowledge hitherto comprised in the two introductory courses. The present volume is so arranged that the necessary ground may be covered in the course of the school year if five hours per week be devoted to the subject. It is possible, however, by judicious selection of the matter to be dealt with by demonstration, to cover the ground in the course of the year when demands upon the time-table will permit only four hours per week to be given to science teaching.

The questions at the ends of the chapters are taken from the Intermediate Examination papers; and acknowledgment is gratefully made of the courtesy of the Assistant Commissioners in permitting these to be used.

I have to express my gratitude to Prof. R. A. Gregory and Mr. A. T. Simmons for permission, of which I have fully availed myself, to use the subject-matter and diagrams in their numerous text-books, and also for much valuable advice and criticism while the book was passing through the press.

GEORGE A. WATSON.

September 1918.

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SECTION I.: PHYSICS.

CHAPTER I.

1. FORMS OF MATTER.

Matter.—Every person is being continually brought by means of his senses into contact with things of all kinds. These objects, substances, stuffs, or whatever they may be called, become known in different ways, appealing as they do to different senses. Some are smelt, some are seen, others are felt, while others again become known by the powers of tasting and hearing. But by whichever of the five gateways of knowledge—as the senses have been called—the information reaches the brain, the result is an acquaintance with *material* things, or, as they may also be called, **forms of matter**, in the universe. The general term used to speak of all kinds of material things is *matter*, which may be defined as **all things existing in or out of the world, which are become aware of by the help of the senses.**

It is worth while to bear in mind that all phenomena of physical or natural science are examined by the help of the senses. When, therefore, any information is gained by an accurate use of the carefully trained senses, it may be regarded as a fact of science which may be reasoned about in order to construct a theory or arrive at a generalisation. To use an expression of Huxley's, " science is organised common sense."

Different kinds of matter.—Though matter exists in an infinite variety of forms, living and dead, the different kinds may be arranged roughly into three classes, according to certain of the properties they possess. These classes are :

(1) Solid things, or **solids** ;
(2) Liquid things, or **liquids** ;
(3) Gaseous things, or **gases**.

Sometimes the last two are made into one class and called **fluids**.

Solids.—Things which have a size and shape of their own, and remain of the same size and shape so long as they are not interfered with, are called **solids.** Some solids are harder than others, and some can have their shape altered more easily than others. But none of them change by themselves. In fact, **a solid body does not readily alter its size or shape. It keeps its own volume and the same form unless acted upon by a considerable force.**

Another way of expressing the chief facts contained in the definition is to say that *solids possess rigidity.* Because of its rigidity every solid has a definite free surface and a definite shape.

Liquids.—If a wine-glassful of water is poured into a tumbler the water fills the lower part of the tumbler. Originally the water had the shape of the wine-glass, but after the transfer it takes the shape of the lower part of the tumbler. And, if it be poured successively into vessels of different shapes, the water always assumes the shape of the lower part of the containing vessel. Throughout all the changes of shape there is no alteration of size ; for if at the end of the experiment the water is poured back into the wine-glass it exactly fills it.

The surface of a liquid at rest is always horizontal if the vessel containing it is not very small. This is one of the results of the fluidity, or power of flowing, possessed by liquids in common with gases. When not held by the sides of the containing vessel a liquid at once flows.

These facts provide a definition which includes all the important characters of liquids. **A liquid at rest has a horizontal surface and takes the shape of the part of the vessel below the surface, but the conditions remaining the same, the liquid keeps its own size or volume, however much its shape may vary.**

Not only is the power of flowing possessed by fluids, but also by some powders. Fine sand may be made to flow, but the particles do not move freely over one another, so the surface is left uneven. Indeed the horizontal surface of liquids at rest provides a clear means of distinguishing between a solid and a liquid.

Gases.—A liquid always adapts itself to the shape of the containing vessel, and presents a level surface at the top ; a gas, on the other hand, will, however small its volume, immediately spread out and do its best to fill the vessel, however large ; and it does not present any surface to the surrounding air. We can never say exactly where the gas leaves off and the air begins.

Gases possess fluidity to a much more marked degree than liquids. Liquids and gases are, as has been seen, both classified as fluids. But, whereas liquids are almost incompressible, gases are very easily compressed into a much smaller space; their chief characteristics can be stated thus: **Gases are easily compressible and expand indefinitely.**

Changes of state in same kind of matter.—To the fact that there are three kinds of material things, we must add another idea, viz., that the **same matter can exist in all three forms.** The change in the state of matter, whether from the solid to the liquid condition, or from the liquid form to the gaseous state, is most easily brought about by heat. Reverse changes, viz., from gas to liquid and from liquid to solid, can be effected by cooling.

The degree of heating required to bring about the above changes varies very greatly with different substances. Iron must be heated very much more than ice before it can become a liquid. Alcohol, again, has to be cooled to a much greater extent than water before the liquid condition gives place to that of a solid.

When a substance has, as a result of heating, successively passed through the solid, liquid, and gaseous states, then, if the conditions are reversed and the gas is continuously cooled, the liquid form is first reassumed, and subsequent cooling causes the liquid to change again into the original solid.

Sudden and gradual changes.—The circumstances attending the change from the solid to the liquid, or from the liquid to the gaseous state, are not always the same as in the case of water. When a crystal of iodine is heated, it appears to pass suddenly from the condition of a solid to that of gas. Camphor is another instance of this sudden transition from solid to vapour. When, on the other hand, sealing-wax is heated, it very gradually passes into the liquid condition, and may be obtained in a kind of transition stage—neither true solid nor true liquid.

CHAPTER II.

MEASUREMENT OF LENGTH.

2. LENGTH.

i. Simple measurements.—Obtain a rule divided into inches and eighths or sixteenths of an inch on one edge, and tenths on the other edge. Measure the length, or width, of any convenient object, such as the top of a table or desk. Write down the length in feet, inches, and fractions of an inch, thus :

Width of desk, 2 feet, $3\frac{5}{8}$ inches.
Length of sheet of foolscap paper, 1 foot, $1\frac{1}{4}$ inches.

ii. Decimal fractions.—Look at the part of a rule having inches divided into tenths. How many tenths are there in half an inch ? The tenths of an inch can be written like common fractions, for instance, $\frac{1}{10}$ signifies one-tenth, and $\frac{3}{10}$ means three-tenths. A more convenient way is to separate the inches from the tenths by means of a dot. Thus, a length of $6\frac{3}{10}$ in. is written on this *Decimal System* as 6·3. The two ways of writing tenths in common and decimal fractions may be compared as follows :

Common fractions	$\frac{1}{10}$	$\frac{2}{10}$	$\frac{3}{10}$	$\frac{4}{10}$	$\frac{5}{10}$	$\frac{6}{10}$	$\frac{7}{10}$	$\frac{8}{10}$	$\frac{9}{10}$	$\frac{10}{10}$
Decimal fractions	0·1,	0·2,	0·3,	0·4,	0·5,	0·6,	0·7,	0·8,	0·9,	1

iii. Measurements in tenths.—Measure one or two small objects as before, using the part of your rule divided into inches and tenths. Write down the number of inches and tenths ; thus, length of a new lead pencil, 7·1 inches.

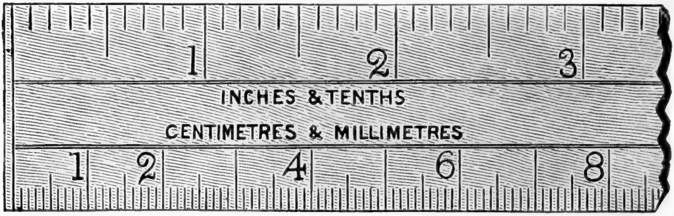


FIG. 1.—A ruler showing Inches and tenths, Centimetres and Millimetres.

iv. Metric measures of length.—(*a*) Look at a rule divided into inches and parts of an inch on one edge, and metric measures on the other, as in Fig. 1. The smallest divisions upon the metric scale are *millimetres* (*mm.*); 10 of these millimetres make 1 *centimetre* (*cm.*), 10 centimetres make a *decimetre* (*dm.*), and 10 decimetres make 1 *metre*. A metre thus contains 10 decimetres, 100 centimetres, and 1000 millimetres.

(*b*) Find how many millimetres there are in the length of this page. Write down the result in (1) millimetres, (2) centimetres and tenths of a centimetre, (3) decimetre and tenths and hundredths of a decimetre.

v. Relation between Metric and British measures of length.—(*a*) Measure the length of this page both in inches and centimetres; also determine other lengths in the two systems of measurement.

Put down the results in parallel columns, as shown below, and from them calculate the number of centimetres in an inch.

Length in centimetres.	Length in inches.	$\frac{\text{No. of centimetres}}{\text{No. of inches}}$.

(*b*) Upon the back of an ordinary tape measure, mark off a length equal to 100 centimetres, that is, 1 metre, starting from the point where the inches begin on the other side. Prick a hole through the measure at the point where you mark the length of a metre, then turn over the measure and notice where the hole occurs on the inch scale.

Number of inches in one metre $= 39{\cdot}3$.
Number of centimetres in one yard $= 91{\cdot}4$.

British measures of length.—To secure uniformity, and provide a permanent length with which others may be compared if necessary, a standard of length has to be agreed upon, and this has been described as follows in a publication of the Standard Department of the Board of Trade: "In the **Imperial system of measures** the **yard** is the unit of length, and from this measure are derived all other measures of extension based on that system, whether linear, superficial, or solid. It is represented by the distance between two fine lines traced on a metal bar which is deposited with the Board of Trade. This bar is made of bronze or gun metal; it is 38 inches long, and one square inch in section, the defining lines being traced on gold plugs or pins inserted at the bottom of a cylindrical hole near to each end of the bar."

Metric measures of length.—When new standards of measurement were being considered in France, in 1795, French geometricians decided that an arbitrary standard such as our yard, in view of the chance of its loss or destruction, was undesirable, and they suggested that if a fraction of the circumference of the earth were taken, it would be possible in the event of the loss of the standard, to replace it by an exact copy. They proposed the one ten-millionth part of the distance from the earth's equator to the pole—that is, of the earth's quadrant—as a suitable length, and this they called the **metre**. After bars had been prepared of this length it was unfortunately found that the length of the quadrant had not been exactly determined, and consequently the length of the standard metre at Sèvres is arbitrary. An exact copy of this standard is kept in this country; and for our purposes the metre may be defined as "the length, at the temperature of 0° C., of the iridio-platinum bar, numbered 16, deposited with the Board of Trade." It is equal to 39·37079 inches, or briefly, 3 feet 3·3 inches. The metre is subdivided into ten equal parts, each of which is called a **decimetre**; the tenth part of the decimetre is called a **centimetre**, and the tenth part of the centimetre is known as a **millimetre**. Thus, we get

$$\begin{array}{r}\text{10 millimetres}\end{array} = 1\text{ centimetre}$$

$$\left.\begin{array}{r}\text{10 centimetres}\\ \text{100 millimetres}\end{array}\right\} = 1\text{ decimetre}$$

$$\left.\begin{array}{r}\text{10 decimetres}\\ \text{100 centimetres}\\ \text{1000 millimetres}\end{array}\right\} = 1\text{ metre}$$

The multiples of the metre are named **deka-**, **hekto-**, and **kilo-metres.** Their value is seen from the following table:

10 metres = 1 dekametre
100 metres = 1 hektometre
1000 metres = 1 kilometre

The kilometre is equal to about five-eighths of a mile, or eight kilometres are equal to five miles.

3. RATIO OF THE DIAMETER OF A CIRCLE TO ITS CIRCUMFERENCE.

Measurement of the length of curved lines.—All curved lines can be regarded as made up of a very large number of small straight lines. The larger the number of small straight lines into which we suppose the line to be divided, the more nearly shall we get to the

FIG. 2.—To illustrate how a curve may be considered to be made up of an infinite number of short lines.

actual length of the curved line. If we could take an infinite number of such small straight lines, there would be no difference between the sum of their lengths and the length of the curved line. Fig. 2 helps to explain these statements.

i. **Dividers.**—Draw (1) a circle of 1·5 cm. radius; (2) the quadrant of a circle of 4 cm. radius. Measure the lengths of these curves by the help of dividers as follows: Make a mark on the curve with a fine pointed pencil as a starting point. Procure a fine pair of dividers and open them a small distance, using your scale to determine this distance: 3 mm. will be a convenient amount for the ends to be apart. Starting from the mark, place one leg of the dividers upon it, and see that it does not slip; then bring the other leg up till it is exactly on the curve. Put down 1 on a piece of paper, or keep count in your mind. Now keeping the second leg of the dividers firm, bring round the first until it is exactly on the curve and fix it there. Put down 2 on the piece of paper. Repeat this process until you have passed right round the curve. The number of times you have moved the dividers, multiplied by 3 mm., will give you the length of the curve in mm.

ii. **Thread.**—Use a fine piece of thread instead of the dividers, and measure the length of the same curves with it.

Starting at the mark as before put one end of the thread upon it and hold it there with the nail of the first finger of your left hand. Make the thread coincide as nearly as you can with a small part of the curve, and place the nail of the first finger of your right hand upon it. Now release your left-hand finger and carefully place it at the point where your right-hand finger is held; then, using your right hand, go on to make some more of the string exactly coincide with another small length of curve. Repeat this until you have completed the whole curve. Measure the length of thread with your mm. scale. Compare the result with that obtained using the dividers.

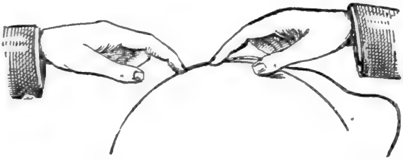


FIG. 3.—Measurement of a curved line with thread.

Stick a number of pins at close intervals along a curved line the ends of which meet; wrap cotton round these pins, and when the cotton overlaps cut it; measure the cut piece of cotton.

iii. **Tracing wheel.**—Procure a tracing wheel, such as is used by dressmakers. Or, obtain a small toothed wheel with sharp teeth from a disused watch or clock, and mount it with a thin nail as an axle upon a strip of wood. Run the wheel across a page of your note book, and measure the distance in millimetres between twenty-

FIG. 4.—A tracing wheel for measuring curved lines.

one of the impressions of the teeth. By dividing the result by twenty you can find the distance between two successive teeth. Repeat the determination. Write in your note book:

Distance between twenty impressions - - - ...... mm.
Therefore, distance between successive teeth - - ...... mm.

Now determine the length of a curved line by running the tracing wheel along the line and counting the number of spaces between the impressions it makes. This, multiplied by the distance between two teeth, gives the length of the curve.

iv. Opisometer.—Obtain and examine an instrument called an *Opisometer*. A small wheel with fine teeth fits upon a screw of fine pitch which forms an axis, and as the wheel is turned it moves towards one end of the axis. Turn the wheel until it reaches one end of the screw. Make a mark upon it at the bottom. Place the mark at the starting point on the curve you wish to measure, and carefully cause the wheel to travel along the whole length of the curve. When it has completed the journey along the curve, make a mark on the rim of the wheel touching the end of the curve and lift the wheel off. Place this last mark over some division on your millimetre scale, and run the wheel along the scale in the opposite direction to which it formerly turned, so making it retrace its course along the screw axis. When it arrives at the end of the axis notice the division on the mm. scale. Record the length, and compare it with those obtained by the string and dividers.

v. Circumference of a cylinder.—(*a*) Wrap a strip of paper closely round a wooden cylinder (Fig. 5), and make a small hole with a pin at a place where the paper overlaps. Unroll the paper and measure the distance between the two holes. This gives the distance round the cylinder, that is, its **circumference**.

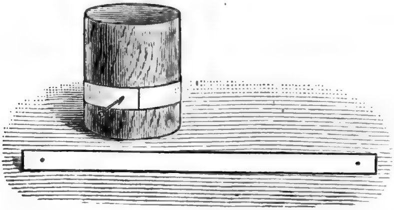


FIG. 5.—Measurement of the circumference of a cylinder.

Measure the circumference in another way by winding a piece of thread round the cylinder several times side by side, and dividing the total length of the thread by the number of coils.

(*b*) Measure the diameter of the cylinder in several places by applying the rule to one end and taking the greatest distance across.

Repeat the preceding exercise with cylinders of different sizes, such, for instance, as a tin canister, or a short piece of wide glass tubing, or a round bottle. Make the measurements in centimetres. In each case divide the length of the circumference by that of the diameter and notice the result obtained.

vi. Diameter of a sphere.—To measure the diameter of a ball an indirect method must be used, because of the difficulty of measuring the diameter of a sphere directly with a scale.

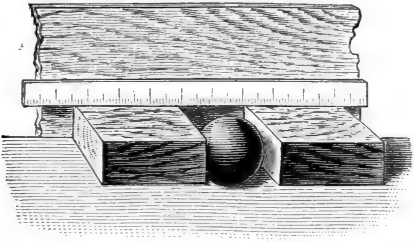


FIG. 6.—Measurement of the diameter of a sphere.

Procure two rectangular wooden blocks larger across than the sphere. Push the two blocks against a third, or against any upright surface, so as to keep them parallel, and then place the wooden sphere between them and measure its diameter, as in Fig. 6. The scale may be placed lying flat on the blocks, and if it is a thin metal rule it can be used in this way. But if a wooden scale is used it should be placed on the blocks "edge on," as in Fig. 7. Make measurements with the sphere in different positions, and find the mean or average result.

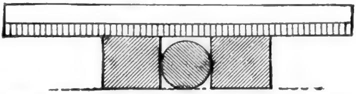


FIG. 7.—How to use a scale "edge on" to prevent inaccurate measurement.

Note. In the previous experiments, each student should take a circle differing in diameter from the others; so that a set of results can be obtained for circles from, say, one cm. diameter to twenty cms. These results should be collected under the headings of **mean diameter** and **mean circumference.** From the results a graph should be obtained.

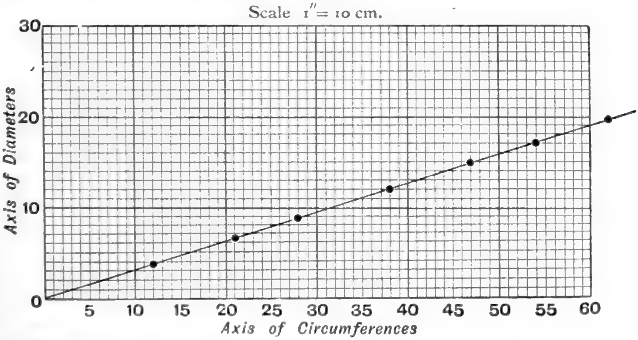


FIG. 8.—Graph showing the relation between the diameter and circumference of a circle.

Ratio of the diameter of a circle to its circumference.—When the results of careful experiments on the diameters of circles and their respective circumferences are collected and plotted on squared paper it is found that all the points lie on a straight line passing through the origin. This shows that the circumferences increase in size at the same rate as the corresponding diameters. For example, from the graph it is seen that a circle having a diameter of 5 cm. has a circumference of 15·7 cm., while a circle with a diameter of 10 cm. has a circumference of 31·4 cm., or, when the diameter is doubled, the circumference is also doubled. From the last figure it is evident that a circle with a diameter of 1 cm. would have a circumference of 3·14 cm. Accurate measurements show that *the circumference of any circle is equal to* 3·1416 *times its diameter.*

QUESTIONS ON CHAPTERS I. AND II.

1. Explain how you have found the length of a centimetre in inches, and the length of an inch in centimetres.

2. Describe three methods you have used in measuring the lengths of curved lines and state how any two of them may be adapted to finding the length of a winding road.

3. The centres of two wheels are 12 feet apart, and the diameter of each wheel is 15 inches. Find the least length of belt $\frac{1}{4}$ inch thick which will pass right round the wheels.

4. Describe any methods for measuring lines which are not straight. Show how the principles of these methods may be used for large distances.

5. What do you understand by parallel blocks? Give two examples of measurements for which such apparatus may be used. Point out the likely sources of error in making the measurements and explain how you would guard against them.

6. The wheel of a certain bicycle revolves 800 times in running a mile. What is the diameter of the wheel?

CHAPTER III.

MEASUREMENT OF AREA.

4. UNITS OF AREA.

i. **Area of a square.**—Draw a square each side of which is one inch long, and also a square with sides two inches long. Divide the latter into four square inches. Then draw a square containing nine square inches.

Notice that in each of the cases the number of squares is equal to the number obtained by multiplying the length of a side by itself.

Thus $1 \times 1 = 1$;

$2 \times 2 = 4$;

$3 \times 3 = 9$.

Multiplying a number by itself is called *squaring* it, and the result is known as the *square* of the number. When a small 2 is written near the top to the right of a number in this way, 6^2, it means that the square of the number must be used; in fact, the small figure, which is called an *index*, tells you how many sixes must be multiplied together; thus,

$$6^2 = 6 \times 6 = 36.$$

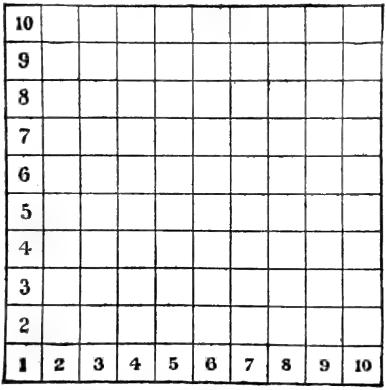


FIG. 9.—To illustrate how to divide a square decimetre into square centimetres.

(*b*) Draw a square of a decimetre side and divide two adjacent sides into centimetres. Draw lines through the divisions parallel to these sides and count how many square centimetres thus formed there are in a square decimetre.

(*c*) Draw a square of any convenient size; find the number of square centimetres in it. Then measure the length of a side in inches, and calculate the number of square inches there are in the figure.

The class results should be collected, tabulated, and from them the number of square centimetres in a square inch can be found.

Area in square centimetres.	Area in square inches.	$\frac{\text{Square centimetres}}{\text{Square inches}}$.

Measurement of area.—By the term area is meant the amount of surface which the thing referred to has. For the purpose of expressing this it is necessary to have a unit or standard. In the English system the unit of area is taken from the corresponding unit of length. For example, the length of this page would be stated in inches, while the length of Ireland would be stated in miles. The area of this page would be stated in square inches; that is, the number of squares of one inch side into which it could be divided. In the same way, in the metric system the unit would be the square centimetre. By measuring a number of square figures, in both square centimetres and square inches, and dividing the latter measurement into the former, the relation between these units of area can be obtained. Careful work shows that one square inch is equal to 6·451 square centimetres.

5. AREA OF RECTANGLES.

i. Area of a rectangle.—On the squared paper draw a rectangle of any convenient size. Find the area of it by counting the number of units of area enclosed in the figure. Also count the number of units of length in two adjacent sides.

Different members of the class should select rectangles varying in size and shape. The results should then be collected and tabulated.

Area.	Length.	Breadth.	Length × Breadth.

Area of rectangles.—When a rectangular figure is drawn on squared paper its area can be found by counting the number of units of area inside its boundaries. The most convenient way to do this is to count the number of units of area in one row, and then multiply by the number of rows of units in the figure.

In effect the same result is obtained by multiplying the number of units of length by the number of units of breadth. And it is found that this is true for rectangles of different sizes and shapes.

Rule. **The area of a rectangular figure is found by multiplying the length by the breadth.**

It is important to note that whatever the unit of length used in the measurements, the area will be found in the corresponding unit of area.

6. AREA OF PARALLELOGRAMS.

i. Area of parallelograms.—Draw two parallel lines a short distance apart upon squared paper. Mark off equal lengths upon one of the lines and construct two parallelograms as in Fig. 10. (Every four-sided figure the opposite sides of which are parallel is a parallelogram.) Count the number of squares in each parallelogram, and either estimate the fractions of squares near the sides, or, what is easier and generally just as satisfactory, count all parts of squares included in the figure equal to or greater than a half-square as one, and neglect all fractions less than a half square.

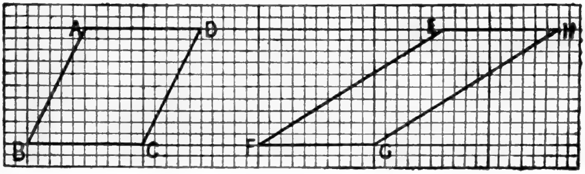


FIG. 10.—Parallelograms on equal bases and between the same parallels.

Measure in your book the height of each figure in units, and also the number of units in the base.

Class results should be collected as before.

ii. Area of parallelograms (*another method*).—Cut out two cardboard parallelograms $ABCD$, $EFGH$ (Fig. 11) and draw a line from D perpendicular to BC, and from G perpendicular to EF. Cut off the two triangles DLC, GMF, and place them so as to convert each parallelogram into a rectangle.

Area of parallelograms.—If a parallelogram is cut out in cardboard, a triangle can be cut off it, such that it can be fitted exactly at the opposite side to form a rectangle (Fig. 11). This is true whatever the shape of the parallelogram. Evidently the area of each complete figure is the same whether the triangle is in one position or the other. In other words, a parallelogram has the same area as a rectangle on the same base and having the same altitude, or perpendicular height. As the area of a rectangle is equal to the base multiplied by the height, the same rule evidently holds good for determining the area of a parallelogram.

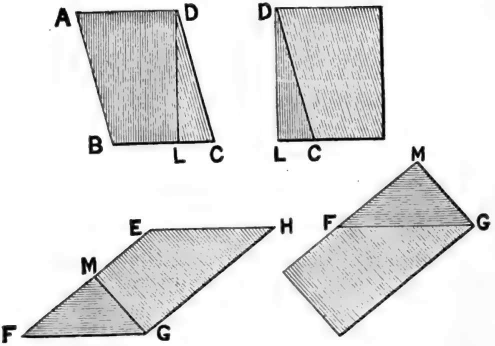


FIG. 11.—Every parallelogram can be considered as a rectangle.

A confirmation of this rule is obtained by drawing a number of parallelograms on squared paper. It is found that the area of each, as obtained by the method of counting squares, is equal to the product of the base and perpendicular height.

Rule. **To find the area of a parallelogram multiply the base by the perpendicular height.**

It does not matter which side is considered as base.

7. AREA OF TRIANGLES.

i. **Area of triangles.**—Draw two parallel lines upon squared paper, and between them draw two triangles of different shape, but having equal bases. Count the number of small squares in each triangle as before. Measure the number of units of length in the base, and also the perpendicular height. The class results should be collected and tabulated.

Base.	Height.	Base × Height.	Area.

Area of a triangle.—When the area of a number of triangles, of different size and shape, is found, and the results tabulated as shown above, inspection shows that the area of each triangle, as found by the method of counting squares, is just one half the product of the base and perpendicular height.

Rule. **To find the area of a triangle multiply the base by the perpendicular height and divide by two.**

It does not matter which side is taken as base.

8. AREA OF A CIRCLE.

i. Area of a circle.—Draw a circle four inches in diameter on gummed paper. (A gummed jam-pot cover may be used.) Divide into small triangles as in Fig. 12. Cut out the triangles neatly and arrange them as in Fig. 13. Moisten the gummed side of the paper and paste into the laboratory notebook. Note the approximate shape of the new figure and express its area in terms of the radius and circumference of the circle.

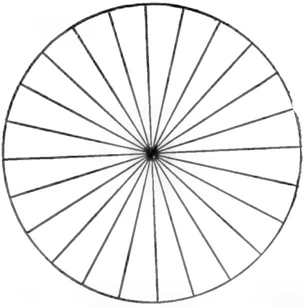


FIG. 12.—Circle divided into triangles.

The area of a circle.—The area of a circle may be found by dividing the circle up into small triangles as shown in Fig. 12, and taking the sum of them. If the triangles are arranged as shown in Fig. 13 a figure closely resembling a parallelogram is formed. The height of this figure is equal to the radius, and the base to half the length of the circumference of the circle. Therefore, since the area of the figure equals that of the original circle, **the area of a circle is equal to the radius multiplied by half the circumference.** The figure is not a perfect parallelogram, but it is evident that by making the segments much smaller the approximation becomes more perfect. Theoretically there is no limit to the smallness of the triangles, and it may therefore be taken that the above rule is accurate. If for the circumference

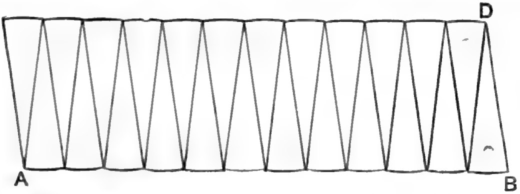


FIG. 13.—Figure formed by triangles cut from a circular disc.

we substitute its value in terms of the radius as found in Section 3, the rule may be simplified in the following way.

$$
\begin{aligned}
\text{Area of circle} &= \frac{\text{circumference}}{2} \times \text{rad.} \\
&= \frac{\pi \times 2\ \text{rad.}}{2} \times \text{rad.} \\
&= \pi \times \text{rad.} \times \text{rad.} \\
&= \pi \times \text{rad.}^2
\end{aligned}
$$

CHAPTER IV.

MEASUREMENT OF VOLUME.

9. UNITS OF VOLUME.

i. Cubic inch and cubic foot.—Procure a box or a block one cubic foot in size. Divide the top and the other faces of the cube into square inches (Fig. 14). Notice that the area of each face of the cubic foot is one square foot. Count the number of square inches marked on one face. Notice that 144 cubic inches could be cut out of a slab of the cube one inch thick. How many slabs having a thickness of one inch could be cut from one cubic foot, and how many cubic inches are there altogether in such a cube ?

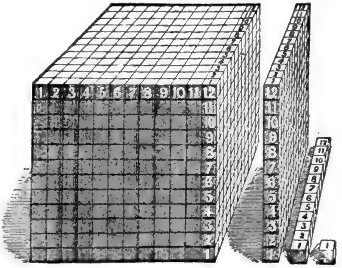


FIG. 14.—To show the relation between 1 cubic inch and 1 cubic foot.

ii. Cubic centimetre and cubic decimetre.—Using a block one cubic decimetre in size, divide the top and faces into square centimetres. How many square centimetres are there in each face ? How many cubic centimetres are there in a slab one centimetre thick ? How many cubic centimetres go to make one cubic decimetre ? (Fig. 15.)

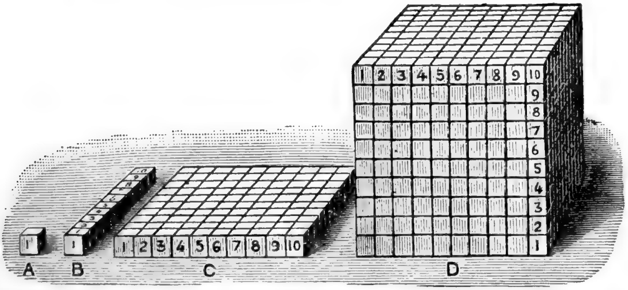


FIG. 15.—To show the relation between 1 cubic centimetre and 1 cubic decimetre.

iii. Fluid measure.—Obtain a half-pint graduated glass measure, such as is used by photographers; obtain also a smaller one (Fig. 16). Examine the divisions upon them: they represent fluid ounces, or parts of a fluid ounce, in Apothecaries' measure.

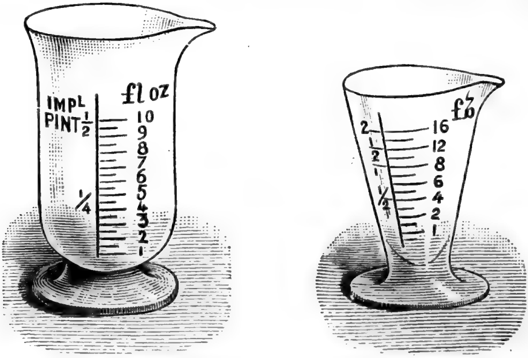


FIG. 16.—Glasses graduated for British fluid measure.

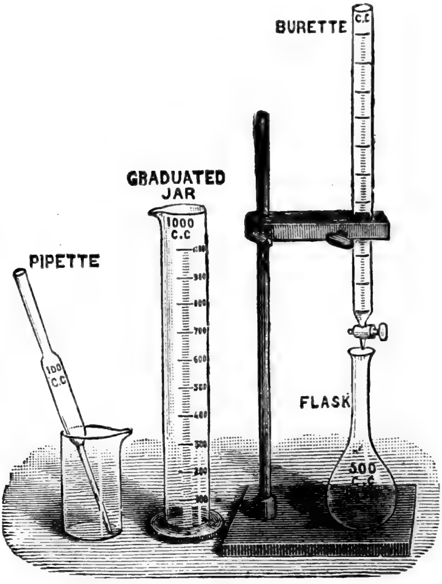


FIG. 17.—Graduated measuring vessels.

iv. Graduated measuring jar.—(*a*) Obtain a jar graduated into cubic centimetres, like that in Fig. 17. Notice the graduations on it, and the method of reading them. *When supplied with a piece of graduated apparatus, an examination of it should always be made so as to understand the principle on which it has been marked.*

v. Relation between British and metric measures of capacity.—(*a*) Use graduated jars to determine: (1) The number of cub. cms. in 1 fluid ounce; (2) the number of ounces and drachms in 100 cub. cms.; (3) the number of ounces and drachms in 1000 cub. cms.

(*b*) Fill a measure up to the half-pint mark with water. Pour the water into the metric graduated jar, and so find the number of centimetres in a half-pint. Repeat the operation, so as to determine the number of cubic centimetres in one pint.

(*c*) Pour 1000 c.c., that is, 1 litre, of water into a vessel. Then measure this amount of liquid in pints, so as to find the value of 1000 c.c. in British measure.

vi. Volumes of solids determined by displacement.—(*a*) Fill a metric measuring jar about half way up with water. Notice the level. Hold a cubic inch of wood in the water by means of a long pin. Notice the rise of level. The difference between the two levels evidently represents the number of cubic centimetres equal to one cubic inch.

(*b*) Determine the volume of a stone or any irregular solid by observing the difference of level when it is immersed in water (Fig. 18).

vii. Burette.—Examine a burette (Fig. 17). Notice that the divisions are numbered from the top downwards. Fix the burette upright in a clamp, and about half fill it with water. Notice the level. Find the volume of a piece of glass tubing by gently lowering it into the water and noticing the rise of level produced. (The glass must, of course, be completely immersed in water.)

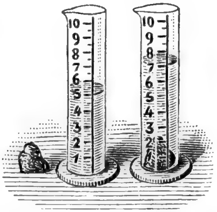


FIG. 18.—The rise of level of the water when the stone is put in shows the volume of the stone.

Measurement of volume.—The volume of a thing is its size, or bulk, expressed in proper units. In dealing with volumes, three dimensions have to be considered. Just as a plane surface, or area, measuring one foot in each of the directions—length and breadth—is called a square foot from the name of the figure which it forms, so a solid which is obtained by measuring a foot in three directions—length, breadth, and thickness—is called a cubic foot, from the name **cube** given to the solid so formed. Similarly, using the **Metric system**, the term cubic metre, or cubic decimetre, is used.

A hollow vessel is capable of holding a certain volume, and this is usually referred to as the **capacity** of the vessel. In the metric system a special name is given to the capacity of a hollow cubic decimetre, that is, a hollow cube having a decimetre edge. It is called a **litre**, and is equal to about one and three-quarters English pints. The sub-multiples and multiples of a litre are named in a similar way to those of the metre. There is no such simple relation between the measure of length and volume in the English system, though the gallon is defined as a measure which contains 10 lbs. of pure water at a certain temperature and pressure. A gallon has a volume of $277\frac{1}{4}$ cubic inches.

In measuring the volumes of irregular solids, advantage is usually taken of the fact that when immersed in a fluid they displace a volume of fluid equal to their own volume. This may be measured as a difference of level (Fig. 18). The water displaced could also be poured into a cubic inch box or into a cubic centimetre box, and the number of cubic inches or cubic centimetres could be thus found. Or, we could get a glass measure having cubic inches or cubic centimetres marked upon it, and pour the displaced water into it. But the best plan of all is to use a vessel having cubic centimetres marked upon it. Water can be put in such a vessel up to a certain mark, and the number of cubic centimetres of water displaced by the solid can be seen at once by noticing the number of divisions between the levels of the water before and after the solid is put in.

10. VOLUME OF PRISMS AND CYLINDERS.

Simple application of volume measurement.—The student cannot have failed to notice that the

volume of a cube = area of the base × height.

The volume of any rectangular block, or of a cylinder, can be calculated by the application of the same rule. This can be easily understood by referring to the accompanying diagrams (Fig. 19, A), which may be taken to represent solids divided into slabs one centimetre thick, while the lowest slab of each block illustrated is divided into small cubes representing cubic centimetres. The number of cubic centimetres in the bottom slab of each solid is evidently the same as the

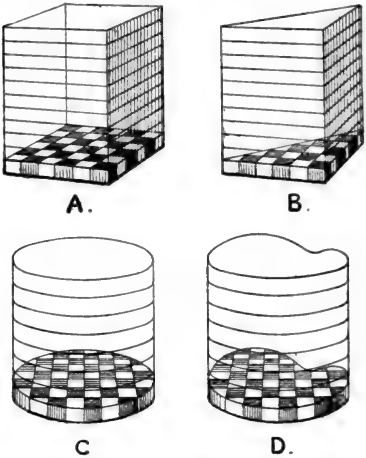


FIG. 19.—To illustrate how the volumes of bodies having the same sectional area all the way up are determined by multiplying the area of the base by the height.

number of square centimetres in the base. Hence, if the base had an area of 30 square centimetres, a slab 1 centimetre high would contain 30 cubic centimetres. Two such slabs would therefore contain 60 cubic centimetres, three slabs 90 cubic centimetres, and so on for any number of slabs. The volume of any block having parallel ends, and the same width and breadth all the way up may evidently be reasoned out in the same manner, being equal to the area of the base multiplied by the vertical height.

Rule. **The volume of a prism or cylinder = area of the base × height.**

Metric Measures of Length and Capacity.

Fractions and Multiples.	Length.	Capacity.
$\frac{1}{1000}$	Milli-metre	Milli-litre
$\frac{1}{100}$	Centi-metre	Centi-litre
$\frac{1}{10}$	Deci-metre	Deci-litre
1	**Metre**	**Litre**
10	Deka-metre	Deka-litre
100	Hekto-metre	Hekto-litre
1000	Kilo-metre	Kilo-litre

British and Metric Equivalents.

1 inch = about $2\frac{1}{2}$ centimetres.
1 foot = ,, 3 decimetres.
1 yard = ,, $\frac{9}{10}$ metre.
1 mile = ,, $1\frac{3}{5}$ kilometre.
1 gallon = ,, $4\frac{1}{2}$ litre.

1 cm. = about $\frac{2}{5}$ in.
1 dm. = ,, 4 in.
1 m. = ,, 3 ft. $3\frac{1}{3}$ in.
1 km. = ,, $\frac{5}{8}$ mile.
1 litre = ,, $1\frac{3}{4}$ pints.

QUESTIONS ON CHAPTERS III. AND IV.

7. A rectangular plot of ground is 100 yards in length and 60 yards in breadth, a walk 4 feet broad going round the ground close to and inside the boundaries of the plot.

What is the area of the walk ?

What percentage of the total area is occupied by the walk ?

8. Using squared paper, prove that the area of any triangle equals the altitude multiplied by half the base.

9. Prove by the use of squared paper :

(*a*) that triangles of different shapes but of equal bases and altitudes are equal in area ;

(*b*) that the area of a triangle equals half the product of the lengths of the base and altitude.

10. A solid circular cylinder the diameter of which is 7 cm. and the height 12 cm. is cut into two equal parts by a plane through the axis.

Calculate the total surface area and also the volume of one of the parts.

11. Plot a curve showing the relation between the area of a circle and its radius for values of the radius up to 12 cms.

12. Equilateral figures having 3, 4, 5, 8, 12 and 16 equal sides, respectively, were described in a circle. Their areas were found to be 1·31, 2·00, 2·40, 2·80, 3·00 and 3·08 sq. in. respectively. Plot a curve showing the relation between area and number of sides. What indication does it give you as to the probable area of the circle ?

13. A circle has a greater area than any other plane figure of equal perimeter. How would you verify this statement by two practical methods ?

14. You are given a steel cylinder, an egg, and several very small spheres of equal size. Explain carefully, mentioning precautions necessary to secure accuracy of measurement, how you would proceed to find the volume (*a*) of the cylinder, (*b*) of the egg, (*c*) of one of the smaller spheres.

15. Given metre and inch rules, a balance and a burette measuring cubic centimetres, how would you determine the weight of a cubic foot of water ?

16. The internal diameter of a circular conical glass is 5 cm. at the bottom and 7 cm. at the top, and the depth is 10 cm.
Calculate the volume of water it will hold.

17. A piece of paper in the form of a quarter of a circle of three inches radius is made into a conical funnel by putting its two bounding radii together. Find the height of the cone and the diameter of its base.

18. Sketch in detail a burette ready for use. Describe how you would use it in finding the density of a long iron nail.

CHAPTER V.

INERTIA. FORCE. GRAVITATION.

11. THE NATURE OF FORCE.

i. To show that a mass at rest tends to remain so. Tie a piece of cotton or thread to the handle of a heavy door and then attempt to move the door by pulling the thread sharply to one side. The thread will break and the door will hardly be moved. The door can, however, easily be pulled to one side if the pulling force is gradually applied.

ii. A moving mass requires force to stop it.—Swing the heavy door used in the preceding exercise to and fro, keeping loose the thread attached to it. While it is swinging away from your hand, stop your hand suddenly. The thread will break and the door will proceed on its swing.

Inertia.—Common experience tells everyone that things do not move of themselves. An object at rest remains at rest until it is forced to move. Moreover, if it is moving it tends to go on moving in the same direction and with the same velocity until made to change by the application of force. In a word, dead matter is helpless and conservative. It does nothing by itself, and objects to alter its condition whatever that condition may be. **The inability shown by a material body to change by itself its condition of rest or of uniform motion in a straight line is called its inertia.** It is exemplified when a cyclist is suddenly stopped, for the tendency to continue moving is so great that the cyclist, if he is travelling quickly, is shot over the handle-bar of his machine. This law of inertia is often spoken of as Newton's **First Law of Motion.**

First law of motion.—**Every object remains at rest or moves with uniform velocity in a straight line until compelled by force to act otherwise.**

This law, which Newton first stated as being always obeyed by bodies in nature, means, first, that if a body is at rest, it will remain still until there is some reason for its moving—until some outside influence, which is called a **force**, acts upon it. In fact, the law really supplies us with a definition of force. Nobody finds any difficulty in understanding the rule so far. But it is not so easy to see the meaning of the words referring to uniform motion in a straight line. An example will make this clear. Consider a ball moving uniformly along ice. After a time the ball comes to rest, and therefore it does not continue in a state of uniform motion. But it moves for a longer time on ice than it would do on a road. The ice is smoother than the road, and there is a connection between the roughness or smoothness and the length of time during which the ball moves. If the ice could be made smoother and smoother, the ball would move for a longer and longer time, and if both the ball and the ice were perfectly smooth, there is no reason why the ball should ever stop. The roughness or friction represents, then, the force which causes the ball to change its state of uniform motion for one of rest. If a body in a state of uniform motion could be placed outside the influence of what Newton has called "impressed forces" it would afford us an example of *perpetual motion.* But because these impressed forces cannot be eliminated perpetual motion is impossible.

Definition of force.—Newton's first law enables force to be defined. **Force is that which produces, or tends to produce, motion in matter; or alters, or tends to alter, the existing motion of matter.** It must, however, be clearly understood that by defining force we do not get to know anything more about it. Nobody can tell what force is. All we can know are the effects produced by force.

The force of gravitation.—Experiments and observations made by Newton led him to the conclusion that it was the rule of nature for every material object to attract every other object, and that this force of attraction is proportional to the masses of the bodies; a large mass exerts a greater force of attraction than a small mass. But the farther these bodies are apart the less is the attraction between them, though it is not less in the proportion of this distance, but in that of the square of the distance. This diminution of a force according to the inverse proportion of the square of the distance applies to so many cases that it ought to be clearly understood before proceeding. Suppose two bodies of equal mass are one foot away from one another and attract each other with a certain force. If the distance between the masses is doubled, the strength of the attraction between them is only one-quarter of what it was; for the square of 2 is $2 \times 2 = 4$ and the inverse of 4 is $\frac{1}{4}$. In the same way, if the bodies are placed three feet apart, the force of attraction is $\frac{1}{9}$ of the original force. Putting Newton's law together it stands thus: **Every body in nature attracts every other body with a force directly proportional to the product of their masses**

and inversely proportional to the square of the distance between the bodies; and the direction of the force is in the line joining the centres of the bodies.

Consider the case of a cricket ball on the top of a house. The earth attracts the ball, and, by Newton's law, the ball attracts the earth. The ball, if free to move, falls to the earth; to be correct, however, we must think of the ball and the earth moving to meet one another along the line joining their centres. But the ball moves as much farther than the earth as the earth's mass is greater than that of the ball; and for practical purposes this is the same as saying that only the ball moves and that the earth remains still.

This force of attraction between all material bodies is called the force of gravitation, but we must point out that this is only a name. Calling this force "gravitation," and the rule according to which it acts the "law of gravitation," does not teach anything about the nature of the force itself.

CHAPTER VI.

MASS AND WEIGHT.

12. MASS AND ITS MEASUREMENT.

1. **Measurement of mass.**—(*a*) Take two pieces of iron or brass, called in ordinary language " pound " and " half pound weights " ; or a " pound " and a " two pound " will do. Lift the two pieces of metal. One feels heavier than the other ; that is, the masses are different.

(*b*) Examine examples of British masses, *e.g.* an ounce, a pound, a half-hundredweight. Also examine a box of metric masses, generally spoken of as a box of " weights."

(*c*) Compare a pound with a kilogram. Hang the 100-gram mass from a spring balance, and notice that the downward pull or its *weight* is equal to the weight of $3\frac{1}{2}$ ounces. What, then, is the British equivalent of the weight of a kilogram ? It is evidently equal to the weight of $3\frac{1}{2}$ ounces $\times 10$ $=$weight of 35 ounces $=$weight of $2\frac{1}{5}$ lbs. (roughly).

The mass of any body is the quantity of matter it contains.—In our country the standard or *unit* of mass is the quantity of matter contained in a lump of platinum of a certain size which is deposited with the Board of Trade. This amount of matter is called the **Imperial Standard Pound Avoirdupois**, and the mass of any other body is spoken of as being a certain number of times more or less than the standard pound, that is, as containing so many more times as much (or as little) matter than that contained in the imperial standard pound. Unfortunately, this is not a universal standard ; in France they have a standard of their own. The French standard is kept at Sèvres and is called a **kilogram**, and the system of masses founded upon it is used in all scientific work throughout the world (Fig. 20).

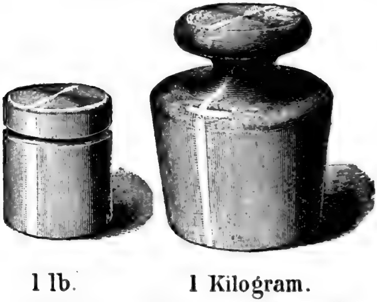


FIG. 20.—Comparative sizes of the British Standard Pound and the Metric Kilogram.

In the metric system a name is given to the mass of water which will exactly fill a *cubic centimetre* at a temperature of 4° C. It is called a **gram.** The same prefixes are used to express fractions and multiples of a gram as have been employed in the case of the metre and litre. The kilogram, or unit of mass, is one thousand times greater than that of a gram, and is the unit in use for ordinary purposes.

METRIC MEASUREMENT OF MASSES.

10 milligrams = 1 centigram.
10 centigrams = 1 decigram.
10 decigrams = 1 gram.
10 grams = 1 dekagram.
10 dekagrams = 1 hektogram.
10 hektograms = 1 kilogram.

Mass is not weight.—If a mass of 1 lb. is dropped from the hand it falls to the ground. If the same mass is hung upon the end of a coil of iron wire, the coil is made longer by the downward pull of the mass fixed to its end. The amount by which a steel spring is lengthened, as the result of such downward pull of masses attached to its end, is used to measure their **weights** in the instrument called a **spring balance.** If a delicate balance of this kind like those used in weighing letters, is used, the *weight* of a small piece of iron hung on to the balance can be made to appear greater by holding a strong magnet beneath it. But, though the weight may appear greater, the mass or quantity of matter is, of course, the same whether the magnet is under the iron or not (Fig. 21).

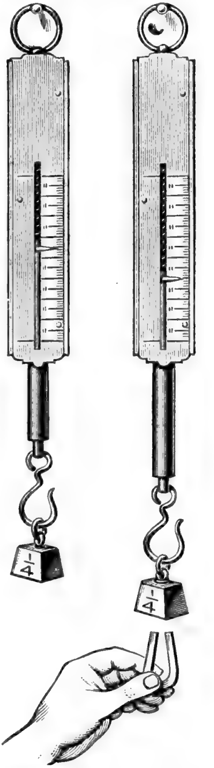


FIG. 21.—The iron weight may be made to appear heavier by a magnet, but the mass does not change.

There is thus a clear distinction between mass and weight; for mass signifies quantity of substance, while weight is a force. In other words, **the mass of a thing is the quantity of matter in it, and its weight is the force with which the earth pulls it.** Unsupported things fall to the ground; a fact which can also be expressed by saying that they are attracted to the earth. Now, even when they are supported, like the objects on a table, the earth attracts them just as much, only the table prevents them from falling, as they would do if there were no table there. The force with which a body is attracted by the earth is its weight. But it must be remembered that this force is just the same whether things actually fall to the ground or not. A spring balance shows the weight of a thing, and masses are compared by means of the ordinary balance or pair of scales.

The attraction of gravity.—Bearing in mind that weight is really a measure of the attraction between an object and the earth, it will be clear from Newton's law of gravitation that since a thing up in a balloon is farther away from the earth (which acts exactly as if its whole mass were collected at its centre) than when at the sea-level, the weight of this thing ought to be less than it is at the sea-level. This is found to be the case, but, actually to demonstrate the difference in weight, the weight must be measured by a spring balance as in Experiment 12 (*c*).

Similarly, because the earth is not a perfect sphere, but is flattened at the poles, points at the surface of the earth in the region of the tropics are at a greater distance from the centre than points similarly situated in the neighbourhood of the poles. Consequently, the weight of a mass situated on the earth in the tropics should be less than the weight it would have if it were moved into the polar regions. This has been found to be the case.

The rotation of the earth is another disturbing influence. While places on the equator are carried round with a velocity of over a thousand miles an hour, those near the poles have but a small velocity of rotation, while the pole itself is at rest. It is clear that if we consider a mass at the equator its tendency is to obey the first law of motion (p. 23), and to fly off at a tangent, and part of the force of gravitation is expended in preventing this flight—the remainder of the force of gravity is operative as the weight of the mass under consideration. At the pole there is no tendency to move off tangentially, and the whole of the force of gravitation is felt as the weight of the body. For this reason alone the mass would weigh less at the equator. At places intermediate between the poles and the equator the diminution in the weight of the body, or the diminution in the acceleration due to gravity, is less; it diminishes as the nearness to the pole is increased.

CHAPTER VII.

PARALLEL FORCES. CENTRE OF GRAVITY. THE LEVER.

13. PARALLEL FORCES.

i. Example of parallel forces.—Place the ends of a stiff lath or rod of uniform thickness upon two letter balances, or support the rod by hanging each end from a spring balance. Notice the load borne by each balance; then weigh the rod, and so determine the fraction of the load supported at each end.

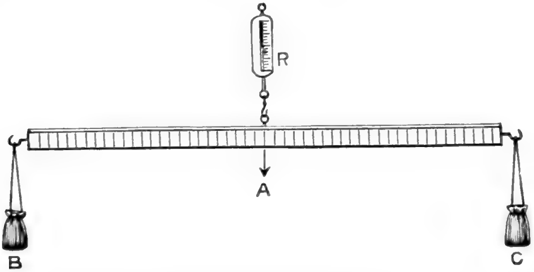


FIG. 22.—The reading of the spring balance is equal to the total weight of the lath and the loads.

ii. Resultant of parallel forces.—(*a*) Suspend a light lath from a spring balance by a ring above its centre (Fig. 22). Notice the reading of the balance. Hang two *equal weights* in bags from the ends of the lath, and again observe the reading of the spring balance. Repeat the experiment with *unequal weights* placed in bags arranged on the lath so as to counterpoise one another. Record your results thus:

Weight of lath *A*.	Weight *B*.	Weight *C*.	Total $A+B+C$.	Reading of Spring Balance, *R*.

By comparing columns 4 and 5 it will be seen that the three downward forces A, B, and C, acting upon the lath, are kept in equilibrium by one upward force R.

(*b*) Hang the two weights B and C together from the middle of the lath suspended from the spring balance. Notice that the reading of the balance is the same as when the two weights are hung from the ends of the lath.

Parallel forces.—It has been seen that the earth exerts a downward pull upon all objects on its surface, and that in consequence of this all things fall to the ground if unsupported. It follows, therefore, that every mass which is supported above the earth's surface is constantly being pulled downwards, even though it does not fall. If a beam, for instance, is supported horizontally by resting the ends upon two posts, each particle of it may be regarded as being pulled earthwards by an attractive force. The direction of the pull is everywhere towards the centre of the earth, so for any one spot on the earth's surface we may consider the attractive forces due to gravity to be parallel to one another.

When a stiff lath or rod of uniform thickness rests upon two letter balances or is supported by hanging each end from a spring balance, the experiment represents on a small scale the case of a beam referred to before; and by using spring balances it can be proved that the weight supported at its ends is equally divided between the two supports. In other words, the two upward forces exerted by the balances are together equal to the downward force represented by the weight of the beam.

If a load be placed anywhere upon the lath, the balances still show that when the lath is in equilibrium the sum of the upward forces is equal to the sum of the downward forces.

Principle of parallel forces.—The principle of parallel forces demonstrated by the experiments referred to may now be definitely expressed as follows: **The resultant of a number of parallel forces is numerically equal to the sum of those which act in one direction, less the sum of those which act in the opposite direction.** In other words, the resultant is equal to the *algebraic* sum of the forces.

If two *equal parallel forces* act in the same direction upon a body, the total force will be obtained (as might be expected) by adding the two individual forces together. In like manner, if two *unequal parallel forces* act in opposite directions the net effect can be found by subtracting the smaller of the two forces from the greater; the direction of the resultant is that of the greater force.

14. DETERMINATION OF CENTRES OF GRAVITY.

i. Experimental methods of determining centres of gravity.—(*a*) Procure a disc of sheet cardboard and find by trial the point on which it may be balanced, that is, the centre of gravity of the disc. Make a hole in the card near the edge, and take a plumb-line consisting of a thread with a piece of lead tied at one end and a hook of thin wire at the other. Hang the disc from the hook, and then suspend both as shown in Fig. 23 (*a*), so that the disc and lead are both suspended and the thread passes over the point of suspension. The thread also passes through the centre of gravity. Do this for various holes in the edge of the disc, and see that in all cases the *vertical line through the point of suspension passes through the centre of gravity.*

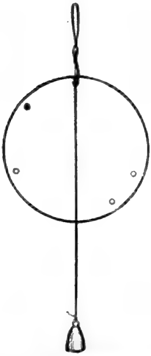


FIG. 23 (*a*).—Determination of centre of gravity of a disc.

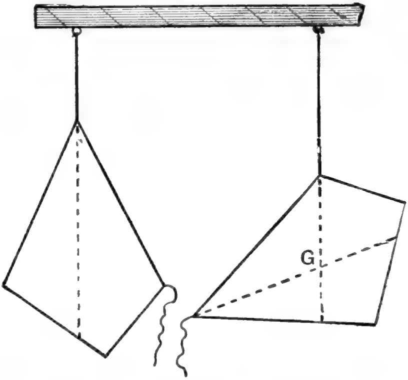


FIG. 23 (*b*).—A method of determining centre of gravity.

(*b*) Another way to find the centre of gravity of a plane figure, such as a triangular plate of metal or card, is to tie strings at each corner of the plate. Then hang the plate by one of the strings to a support such as one of the rings of a retort stand. Allow it to come to rest, and, using a straight edge, draw a chalk mark across the plate in the same straight line with the string as shown by the dotted line in the figure. Now attach the same plate by one of the other threads exactly as before, and again make a mark in continuation of the string. The two chalk marks intersect at a point marked *G*, Fig. 23 (*b*). Untie and do the same with another string, the third line passes through the intersection of the first two. Obtain a similar point for each of the other plates. Also determine in the same way a similar point for irregular plates of wood, zinc, or cardboard.

ii. Centres of gravity of skeleton solids.—(*a*) Procure a skeleton cube or tetrahedron, and suspend it as in the preceding experiments. Mark the verticals through the point of suspension by light wires attached by wax, and thus find the position of the centre of gravity. The centre of gravity will not be on any of the bars of the skeleton solid used.

(*b*) Find the centre of gravity of an open wickerwork basket, such as a waste-paper basket. To do this, suspend the basket, and hang a plumb-line from the point of suspension. Tie a piece of thread across the basket in the direction of the plumb-line ; then suspend the basket from another, point, and notice where the plumb-line crosses the thread. The point of intersection, which need not be actually on the framework itself, is the centre of gravity.

Centre of gravity.—Consider a large number of weights, some heavier than others, suspended from a horizontal rod. A certain position can be found at which a spring balance would have to be attached in order to keep the rod in equilibrium. When the rod is hung from this point the tendency to turn in one direction is counteracted by the tendency to turn in the other, so the rod remains horizontal. The weights may be regarded as parallel forces, and the pull of the spring balance as equal to their resultant. Now consider a stone, or any other object, suspended by a string. Every particle of the stone is being pulled downwards by the force of gravity, as indicated in Fig. 24. The resultant of these parallel forces is represented by the line *GF*, and the centre of the forces is the point *G*. The point *G*, through which the resultant (*GF*) of the parallel forces due to the weights of the individual particles of the stone always passes, is known as the *centre of gravity*. For the stone to be in equilibrium, the string must be attached to a point in the line *GF*, produced upwards.

FIG. 24.—Parallel forces due to gravity.

Every material object has a centre of gravity, and the position of this point for a particular object is the same so long as the object retains the same form.

Experimental methods of determining centres of gravity.—The centre of gravity of such geometrical figures as circles, squares, and parallelograns is really the centre of the figures, and can therefore be determined geometrically. In the case of unsymmetrical figures, however, the centre of gravity cannot be so easily found by geometry, and is best determined by experiment.

The experimental method adopted for determining the centre of gravity of any material body depends upon the considerations set forth in the preceding paragraphs. The body, the centre of gravity of which is required, is allowed to hang quite freely, either by means of a cord or on a smooth peg, and when it has come to rest a vertical line through the point of support is marked upon it. If a string is employed, this vertical line will be a continuation of the string, and at once drawn by the help of a ruler. If the body which is being experimented with is hung from a smooth nail, by means of a hole bored in the body, the vertical line must be drawn with the help of a plumb-line. The point of support is then shifted and the operation repeated. Since the centre of gravity of the plate is in both straight lines it must be located at their intersection.

Plates of all shapes balance about their centres of gravity.—After the centre of gravity of a sheet of metal, or other stiff material, has been determined by hanging it from a support in the manner described in Experiment 14 i. (*a*) and (*b*), it will be found that if this sheet be so arranged that a pointed upright is immediately under the centre of gravity, the plate will be supported in a horizontal position. This affords a convenient means of checking the correctness of the experiment performed.

Geometrical determination of centres of gravity.—It has been sufficiently explained that the centres of gravity of straight lines, circles, squares, and other regular figures is at their geometrical centres. Hence, the geometrical constructions for determining these central points also locate the positions of their centres of gravity.

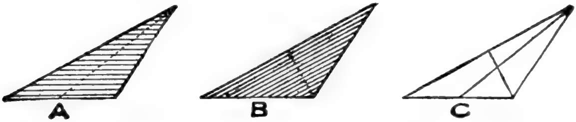


FIG. 25.—Geometrical illustration of centre of gravity of a triangular plate.

The centre of gravity of a **parallelogram** is at the intersection of its diagonals.

The centre of gravity of a **triangle** is determined by bisecting any two sides and joining the middle points so obtained to the opposite angles. The intersection of the lines so drawn gives the centre of gravity. The centre of gravity is found, by measuring, to be one-third the whole length of the line drawn from the middle point of the side to the opposite angle, away from the side bisected.

In fact, a triangular plate may be considered as made up of a number of narrow strips of material which decrease in length from the base to the apex. The centre of gravity of each strip is the middle of the strip; hence the line drawn from the apex to the middle of the base passes through each centre of gravity (Fig. 25, *A*). By taking another side as base, a similar line can be drawn from the middle to the opposite angle (Fig. 25, *B*). These lines intersect at one-third the distance up the line so drawn, measured from the base, and the point of intersection is the centre of gravity of the triangular plate (Fig. 25, *C*).

To find the centre of gravity of a quadrilateral by construction, the plan is to divide it into two triangles by drawing a diagonal. By the method just described the centre of gravity of each triangle is found, and the points so obtained are joined. The centre of gravity of the quadrilateral lies on this line. Repeat the process drawing the other diagonal. Join the centres of gravity of the second pair of triangles, the centre of gravity of the quadrilateral lies on this line. Hence, it is situated at the point of intersection of this line and the first one obtained in the same way.

Centres of gravity of other bodies.—The method of drawing lines across the surface of a thin plate is not practicable in the case of bodies such as blocks or skeleton solids having three dimensions. The experiments described in 14 ii. explain two methods which can be employed usefully in cases of this kind. The ingenuity of the student will provide other equally suitable plans for particular cases.

15. EQUILIBRIUM.

i. **Conditions of equilibrium.**—Place upon a square edged table or board one of the cardboard figures of which you have found the centre of gravity. Gradually slide the figure near the edge until it would just topple over; keeping it in this position, draw a line along the under side of the cardboard where the edge of the table touches it. Then place the cardboard in another position and again mark where the edge of the table touches it when it would just topple over. The intersection of these lines is the centre of gravity, and it will be noticed that the cardboard would just topple over when the centre of gravity falls outside the edge of the table.

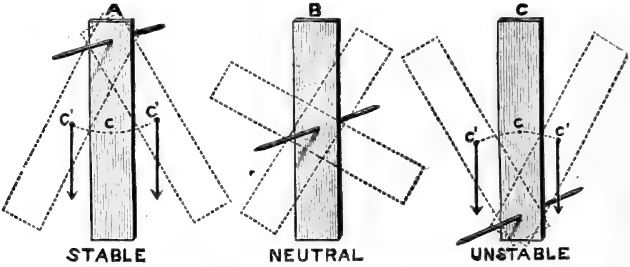


FIG. 26.—Relative positions of centre of gravity and point of support for stable, neutral, and unstable equilibrium.

ii. **Stable, neutral, and unstable equilibrium.**—Procure an oblong strip of wood or cardboard (Fig. 26). Support the strip as at *A* by a long pin pushed through it; it is then in stable equilibrium, for the slightest turn either to right or left raises the centre of gravity. When supported as at

B, the strip is in neutral equilibrium; and when supported as at *C*, it is in unstable equilibrium, for the slightest movement lowers the centre of gravity.

Relation of centre of gravity to base of support.—A circular disc, in which it will be remembered the centre of gravity coincides with the geometrical centre, will not rest upon a table if the centre is beyond the edge of the table, but will topple over. In a similar way, if any plain figure lies flat upon a table, the centre of gravity of the figure must be within the edge of the table. The same conditions apply to any object resting upon a support. For an object resting upon a base to be in equilibrium, a vertical line drawn from the centre of gravity downward must fall within the base. When this vertical line falls outside the base the body topples over.

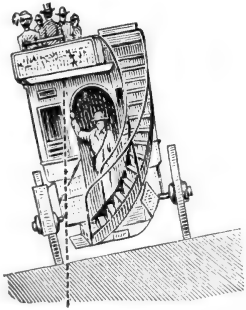


FIG. 27.—If the vertical line from the centre of gravity should fall outside the base of support, the omnibus would topple over.

Consider the case of an omnibus on level ground. The centre of gravity is somewhere inside the omnibus, and a vertical line drawn from it downward would fall within a line traced around the omnibus upon the ground. But, if the outside of the omnibus is filled with people and the vehicle happens to be running across a sloping road, it might topple over, for a jerk might cause so great a change of position of the centre of gravity as to make the vertical line from the centre fall outside the base of support, and in such a case an accident must happen (Fig. 27).

Equilibrium.—When a body is at rest, all the forces acting upon it balance one another (or, what is the same thing, any force is equal and opposite to the resultant of the remaining forces), and it is said to be in equilibrium. It is in **stable equilibrium** when any turning motion to which it is subjected raises the centre of gravity; in **unstable equilibrium** when a similar movement lowers the centre of gravity, and in **neutral equilibrium** when the height of the centre of gravity is unaffected by such movement. Consequently, if a body in stable equilibrium is disturbed, it returns to its original position; if in unstable equilibrium, it will, if disturbed, fall away from its original position; while if the condition of equilibrium is neutral it will, in similar circumstances, stay where it is placed.

Conditions of stability of suspended and resting objects.—The centre of gravity must in every case be below the point of support for a suspended object to be in equilibrium. The greater the distance between the point of support and the centre of gravity the greater is the tendency to return to the position of equilibrium.

When the centre of gravity and the point of support of a suspended object are close together the equilibrium of the object is easily disturbed. A good balance partly owes its sensitiveness to this condition, the centre of gravity and point of support being designedly brought close together.

It has been shown that in the case of a freely suspended object the centre of gravity is at its lowest point when the object is in equilibrium. Let us see how this applies to a body supported upon a surface below the centre of gravity.

A body is least liable to be upset when the centre of gravity is at a considerable distance from all parts of the edge of the base ; for, when this is the case, the body has to be tilted through a large arc before the centre of gravity falls outside the vase.

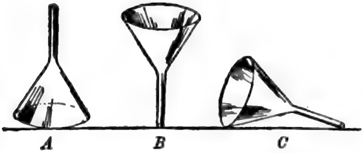


FIG. 28.—A funnel in (*A*) stable equilibrium, (*B*) unstable equilibrium, (*C*) neutral equilibrium.

A funnel standing upon its mouth is an example of a body which cannot be easily overturned on account of the low centre of gravity and its distance from the edge of the base (Fig. 28, *A*). It is then in stable equilibrium. If the funnel is stood upon the end of the neck it can be easily overturned, because very little movement is required to bring the centre of gravity outside the base. It is thus in unstable equilibrium. When the funnel lies upon the table it is in neutral equilibrium, for its centre of gravity cannot then get outside the points of support.

QUESTIONS ON CHAPTERS V. VI. AND VII.

19. Define the *mass* and *weight* of a material body, carefully distinguishing between the terms.

Give the British and metric measures of mass.

20. What can you learn from the direction in which a sheet of cardboard hangs when freely suspended in turn from each of a number of points near the edge ? When suspended in this way what determines the position in which the cardboard remains at rest ?

21. The top of a deal table measures 4 ft. by 2 ft. and is $\frac{3}{4}$ inch thick. It is supported by legs at the corners $2\frac{1}{2}$ ft. long, whose cross section is a square of 2-inch side. How would you find by calculation *or* by experiment the position of the centre of gravity of the whole table ?

22. Explain how the centre of gravity of a metal plate can be found by balancing it in different positions on the horizontal edge of a ruler. Show that the ruler does not always divide the plate into two parts whose weights are equal.

23. Explain fully what you understand by the centre of gravity of a body.

A sheet of cardboard of uniform thickness in the form of an equilateral triangle of 10 inches side is folded along the line joining the middle points of two sides, so as to bring one vertex to the middle point of the opposite side. Through what distance is the centre of gravity of the cardboard moved by the folding ?

CHAPTER VIII.

THE LEVER AND BALANCE.

16. THE LEVER.

i. **Balancing equal weights on a lever.**—(*a*) Make or obtain a lever consisting of a strip of light wood graduated in centimetres, and having a thin ring screwed into one edge, above the central point, and a hook screwed into each end (Fig. 29). Hang the lever from a nail by the middle ring. If it does not exactly balance, plane off a little wood from the end which sinks, or slightly unscrew the hook at the end which rises above the horizontal, until the lever does set itself horizontally.

FIG. 29.—A simple lever.

Hang two pill-boxes or small linen bags by their strings from the lever, one on each side of the fulcrum or pivot, at equal distances from it. Into one of the bags place say 50 grams, and find how many grams you must place into the other bag in order to restore equilibrium. Repeat the experiment with the bags at a different distance.

It will be found always that equilibrium is obtained when equal weights are at equal distances from the turning point or fulcrum.

ii. **Principle of moments.**—(*a*) Place some pieces of lead in each of four linen bags, and adjust by means of shot or small bits of lead until they weigh respectively 50 grams, 100 grams, 200 grams, and 300 grams. Place the 100 gram bag about 12 cm. from the fulcrum of the lever, and balance it with a 50 gram bag on the other side. Record the distance from the fulcrum in each case. Repeat the experiment by balancing 50 grams against 100 grams, 50 grams against 200 grams, 100 grams against 300 grams and other combinations.

Record your observations in columns as below :

LEFT SIDE OF LEVER.		Load × Distance.	RIGHT SIDE OF LEVER.		Load × Distance.
Load. (1)	Distance from Fulcrum. (2)	(3)	Load. (4)	Distance from Fulcrum. (5)	(6)

Compare the numbers in columns 3 and 6, and state in words the law or rule indicated by the results.

The experiments show that there is a definite proportion between the weights on the two sides of a lever, and their distances from the fulcrum The proportion is:

$$\begin{array}{c}\text{left}\\\text{load}\end{array} : \begin{array}{c}\text{right}\\\text{load}\end{array} :: \begin{array}{c}\text{right}\\\text{distance}\end{array} : \begin{array}{c}\text{left}\\\text{distance.}\end{array}$$

Or, expressed in another way, the loads are inversely proportional to their distances from the fulcrum.

The turning effect of any force acting upon a lever, as each load did in the experiment, is termed the *moment* of the force. The comparisons of columns 3 and 6 prove that the *moment* or turning effect is measured by the product of load into distance from the fulcrum.

(*b*) Hang two bags on the same side of the lever at different distances and one bag on the other side. Move the single bag until equilibrium is obtained. Do this several times with the bags in different positions, and compare the sum of the moments of the forces acting on one side with the moment of the force on the other side.

(*c*) Hang a small book or a bag with shot or nails in it on one side of the fulcrum, and the 100 gram bag from the other. Move this bag along the lever until equilibrium is obtained. Then, remembering that

$$\text{load}_1 \times \begin{array}{c}\text{distance}\\\text{from fulcrum}\end{array} = \text{load}_2 \times \begin{array}{c}\text{distance}_2\\\text{from fulcrum,}\end{array}$$

calculate the weight of the book or bag of nails. Repeat the experiment, using the 200 gram bag.

In the preceding experiments the fulcrum has been between the forces due to the loads hanging from the lever. The forces may, however, both act on one side of the fulcrum. It is convenient to call one of the forces the effort and the other the resistance.

Machines.—A machine is a contrivance by means of which a given force is made to resist or overcome another force acting in a contrary direction.

What are termed "the mechanical powers" are really simple machines which can be used to overcome resistance. The mechanical principles underlying the action of such a simple machine as the lever have now to be considered.

The lever.—A lever is a rigid bar which can be turned freely about a fixed point. The **fulcrum** of a lever is the fixed point about which the lever can be turned. The force exerted when using a lever is often described as the *power* and the body lifted or resistance overcome as the *weight.* These words are convenient, but they are not correctly used in connection with levers, as their true meanings are confused by so doing. It is better to substitute the word *effort* for power, and *resistance* or *load* for weight. It should be borne in mind, that, so far as mechanical principles are concerned, there is no difference between the power and the weight; both represent forces, and as such they must be considered in the action of levers.

The perpendicular distances from the fulcrum to the lines of actions of forces acting upon a lever, are known as the **arms of the lever.** In Fig. 30 the distance *AC* is the arm at one end of which the "weight" acts, and *BC* is the arm at one end of which the "power" acts.

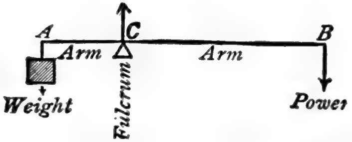


FIG. 30.—Terms used in connection with levers.

Principle of the lever.—It is easy to show by experiment, that when a lever is in equilibrium the following equation holds good:

$$\text{force on one side} \times \text{distance from fulcrum} = \text{force on other side} \times \text{distance from fulcrum.}$$

This principle of moments applies to all levers, so that all that need be remembered when considering the action of a lever of any kind, are the forces working in one direction and their distances from the fulcrum, compared with the forces or resistances which oppose them and the distances of these from the fulcrum.

Moments.—Refer to the diagram (Fig. 31), where *F* represents the point of support, or fulcrum, of a lath or other straight lever, and M_1 is a weight at a distance *AF* in equilibrium with a smaller weight M_2, at a greater distance *FB*.

The force acting at *A* is the weight of M_1, acting vertically downwards; and the force at *B* is the weight of M_2, acting in the same direction. Each force tends to turn the lever in a particular direction, and this turning effect is called the **moment of the force.** The moment of the force acting vertically downwards at *A* is the product of the force equal to the weight of M_1 into the distance *AF*, which, as the diagram shows, is measured at right angles to the direction in which the force acts. Similarly, the moment of the force equal to the weight of M_2, about the point *F*, is equal to the product of this force and the perpendicular distance *BF*.

FIG. 31.—To illustrate moments of forces.

This rule of universal application for taking moments will have to be used several times later on, and should be well borne in mind.

The moment of a force about any point is the product obtained by multiplying the force by the perpendicular distance between the point and the line of action of the force.

17. THE BALANCE AND ITS USE.

i. The balance.—(*a*) Uncover the balance and identify the different parts by reference to Fig. 32. Raise the beam, *AB*, of the balance, off the supports by turning the handle *C*. Notice whether the pointer *F* swings equally on both sides of the middle of the scale *G*: if it does the balance is ready for use; but if not let down the beam and turn the small screw at *B* then try again. Repeat this adjustment until the swings to right and left are equal.

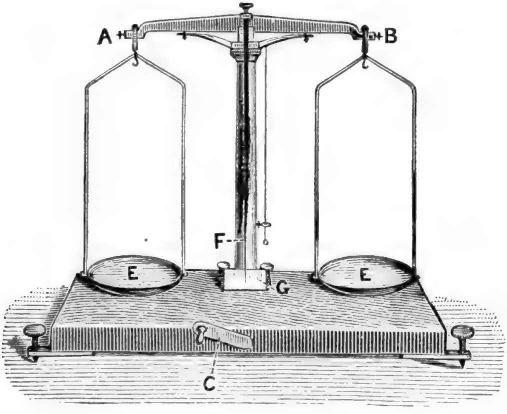


FIG. 32.—The Student's Balance.

(*b*) When using the balance *always* place the substance to be weighed in the *left-hand pan*, *E*, of the balance. Examine the box of weights, and, selecting a weight which you estimate to be about the same as that of an object placed in the pan, *E*, take hold of it *with the forceps* and place it in the *right-hand pan*. Now raise the beam slightly to see whether the estimated "weight" is nearly equal to the mass of the substance in *E*. If the weight appears a little below what is wanted, pick up, with the forceps, the next heaviest weight in the box, and try it in the pan with the other. If the two together are too heavy, take out the smaller weight and put in the one below it, and so on, adding one weight after the other, *without missing any* until the correct weight is found. When you have completed the weighing, write down and add up the weights that are missing from their places, and check the figures as you put each weight back in its place.

(*c*) Examine examples of British masses, *e.g.* an ounce, a pound, a half hundredweight. Also examine a box of metric masses, generally spoken of as a box of "weights" (Fig. 33).

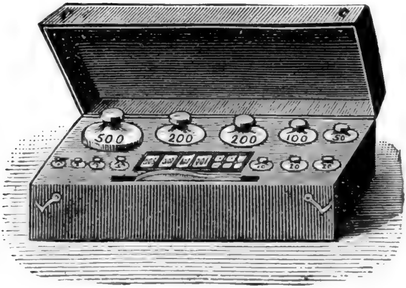


FIG. 33.—Box of metric weights.

Masses are determined by means of the balance.—The instrument generally used for the purpose of estimating masses is the balance or pair of scales. This simple apparatus is best understood with the help of the simple experiments already performed which have shown that when equal masses are hung from points on a lath supported at its centre, then, when the lath is at rest in a horizontal position, or is in a condition of equilibrium, the distance of the masses from the fulcrum or pivot will be found the same in each case. Similarly, when the pans are at equal distances from the supports there is equilibrium only when the masses are equal. When unequal masses are at different distances from the fulcrum, there is equilibrium only when

$$\text{mass on one side} \times \text{distance from fulcrum} = \text{mass on other side} \times \text{distance from fulcrum.}$$

The balance.—The form of balance shown in Fig. 34 is evidently similar in principle to the supported lath with pill-boxes. This kind of balance is good enough for ordinary purposes, but when exact weighings are wanted, a better form, such as that already described (Fig. 32) is used. All the parts in this balance are very carefully made, and the greatest possible pains are taken to have very delicate supports and accurate adjustments. Instead of the wooden lath described just now, a brass beam is employed. This is supported at its middle line on a knife edge of hard steel or agate, which, when the balance is in use, rests on a true surface of similar steel. The hooks to which the pans are attached are likewise provided with a V-shaped groove of hard steel, which also, when the balance is in use, rests upon knife edges on the upper parts of the beam. To the middle of the beam is attached a pointer, the end of which moves in front of an ivory scale, fixed at the bottom of the upright which carries the beam. When not in use, the beam and hooks are lifted off the knife edges by turning a handle.

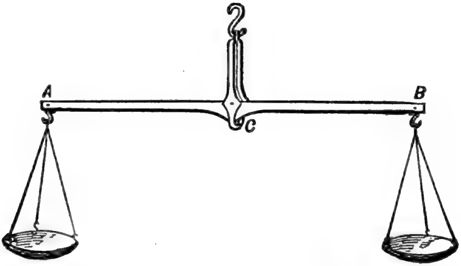


FIG. 34.—A simple balance, or pair of scales.

QUESTIONS ON CHAPTER VIII.

24. A uniform plank used as a see-saw is 16 feet long and weighs 90 lbs. Two boys, whose weights are 60 lbs. and 100 lbs. respectively, sit astride it one foot from each end. Where must the plank be supported in order that it may balance ?

25. It is easier to balance a lever when weights are hung from it than when the weights are placed on top of it. How do you explain this ?

26. A scale whose length is a metre is balanced on a fulcrum at its centre. A weight of 30 gm. is hung on the scale at 10 cm. from one end. Calculate where each of the following weights should be hung in order to balance the 30 gm., namely, 40, 50, 60, 75, 80 gm. Point out the relation between the weights and their distances from the fulcrum. What is the smallest weight which you could use to balance the 30 gm. ?

27. State the rules which you have been taught to observe in weighing a body on a balance, and give reasons for them.

28. Sketch and explain a simple apparatus for weighing bodies, making use of only one known weight.

29. Being given a metre scale and a known weight, show how these might be employed to weigh bodies.

30. A wooden lath used as a lever is first balanced on a fulcrum. On one side of the fulcrum, 15 cm. from it, a weight of 42 gm. is then hung from the lath, and on the other side of the fulcrum, 24 cm. from it, a weight of 35 gm. is hung.

Find whether the weights would or would not balance one another. If they do not balance, find where the weight of 35 gm. must be hung from the lath that the two weights may balance.

CHAPTER IX.

MEASUREMENT OF DENSITY.

18. DENSITY.

i. (*a*) **The mass of 1 c.c. of various bodies.**—Determine, by measurement and calculation the number of cubic centimetres in a wooden block. Then find the mass of the block in grams. From these results calculate the mass of one cubic centimetre.

(*b*) Find the volume and mass of a number of bodies, each made of different material; say, lead, brass, iron, cork and stone. If regular solids are available the volume should be found by calculation from measurements. The volume of irregular solids can be found by the method of displacement.

FIG. 35.—How a pipette is used.

ii. The mass of 1 c.c. of various liquids.—Weigh a small beaker, and into it run from a pipette 10 c.c. of water. Weigh again. Repeat with other liquids; such as Methylated spirit, Turpentine.

Note. In using a pipette care should be taken that the liquid is not drawn into the mouth. The *forefinger* and *not* the thumb should be placed on the top (Fig. 35). Tabulate results.

Substance.	Volume in c.c.	Mass in grams.	Mass of 1 c.c.

(*c*) Counterpoise a small beaker with lead shot. On the pan with the shot place a 50 gm. weight. Run water into the beaker till a counterpoise is again obtained. Note the difference in sizes of the water in the beaker and the 50 gm. weight. Measure the water in a graduated cylinder, and assuming that the weight is a cylinder, find its volume.

The meaning of density—1. Pieces of different substances of the same size or volume may have unequal weights.

2. Pieces of different substances which have equal weights may have very different sizes or volumes.

It is usual to speak of these facts by saying that things have different **densities.** A pound of feathers, or cotton-wool, has exactly the same weight as a pound of lead, but the feathers (or the cotton-wool) take up much more room, or have a larger volume than the piece of lead. The matter in the lead may be packed more closely than in the feathers, which accounts for its taking up less room. The shortest way of saying all this is to say that lead is **denser** than either cotton-wool or feathers.

If a small thing is comparatively heavy, then it is called a *dense* thing, or it is said to have a *high density.* If, on the other hand, a large thing has a small weight, it is said to have a *low density.* Moreover, the densities of things having equal volumes are in the same proportion as their weights.

Standard of density.—But to compare densities it is better to have a standard, just as there is a standard of length, the yard, with which to compare other lengths ; or a standard of area with which to compare other areas. The density of water at a certain fixed temperature is taken as the standard. Why it is necessary to specify the temperature will be understood later.

The weight of one cubic centimetre of water at a temperature of 4° C. is one gram, and its density is taken as the standard of density, expressed in grams per cub. cm. Similarly, a substance, the weight of a cubic centimetre of which is two grams, would be said to have a density of 2 grams per cub. cm., for it must contain twice as much matter as water does, packed into one cubic centimetre. The weight of a cubic centimetre of quicksilver is 13·6 grams, *i.e.* it contains 13·6 times as much matter in one cubic centimetre as there is in one cubic centimetre of water. Its density is therefore 13·6 grams per cub. cm.

Density is the weight of unit volume of a substance.—Suppose a cube of soap and a cube of lead be cut so as to have equal weights ; the soap will evidently be larger than the lead, and it will be just as many times larger as its density is less than the density of the lead. With equal weights, the greater the density the smaller is the volume. It follows from this that if the volume of a body is multiplied by its density, we obtain its weight. Or expressed as an equation,

$$\text{volume} \times \text{density} = \text{weight}.$$

From this it follows that

$$\text{density} = \frac{\text{weight}}{\text{volume}}.$$

In using this relation between the volume and weight care must be taken that the values of weight and volume are reduced to the proper units. In all scientific work it is customary to adopt the cubic centimetre and gram as the units of volume and weight respectively.

The ratio of the weight of *any volume* of a substance to the weight of the same volume of water is equal to the **relative density** of the substance, or, as it is frequently called, the *specific gravity.*

19. DENSITY OF LIQUIDS.

i. Use of a burette for finding the density of a liquid.—Examine a burette, noticing especially its division into cubic centimetres. Fill it with liquid.

Notice that the surface of the liquid is not flat but curved downwards, being higher near the sides of the burette than at the centre.

In reading the level of a liquid having a surface of this kind, the lowest part of the curved surface should be observed, and the eye should be brought on a level with it, as shown in Fig. 36.

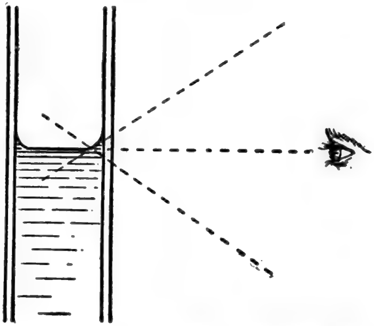


FIG. 36.—Showing that an inaccurate reading of the level of liquid in a burette is obtained, if the eye is not level with the surface of the liquid.

A piece of paper held behind the burette, opposite the eye, at an angle of 45°, is an assistance in reading the graduations.

Fill the burette with methylated spirit. Make sure that the nozzle is also filled, by running a little of the liquid out. Weigh a small beaker. Having noted the level of the liquid in the burette, *slowly* run into the beaker 10 c.c. of the liquid and again weigh. Then run in a further 10 c.c. and weigh. If the weight of each quantity of liquid is not the same, run in a further 10 c.c. and weigh. This weight should agree with one of the others. If so it may be taken as accurate.

Find the weight of 1 c.c. from two concordant results.

Determination of the density of a liquid by means of a burette and balance.—It is sometimes desirable to be able to find the density of a liquid without allowing it to touch the hands or even a duster, as must be done in wiping the density bottle. In such a case a burette may be used with advantage. For example, the density of strong nitric acid may be determined, using a burette with a glass tap, by running a known volume of the acid into a weighed stoppered bottle, and finding the weight of it. By calculation the weight of 1 c.c. can be found.

To ensure accuracy in reading the burette, it is advisable to find the weight of a number of equal volumes. With accurate manipulation the weights are equal.

20. USE OF A DENSITY BOTTLE.

i. Take a narrow mouth stoppered bottle of about 4 oz. capacity, and file a vertical groove upon the stopper. Weigh the bottle, Fill it with water and insert the stopper so that the bottle is completely filled. Wipe away from the rim the water that has escaped through the file mark, and again weigh. What mass of water does the bottle hold? Hence what *volume* of any liquid?

Again fill the bottle with methylated spirit, vinegar, turpentine, ink, or any other liquid, and find the mass of each liquid the bottle contains when full. As you know the *volume*, calculate in each case the *density*.

ii. Weigh some small pieces of lead and place them in the bottle. Fill the bottle with water and shake it in order to get rid of air bubbles; then insert the stopper, wipe the bottle, and weigh. How much less is the mass than the total mass of the lead and the bottle of water? This is obviously the mass of water forced out by the lead. What is its volume? and what then is the volume of the lead? Hence find the density of the lead.

FIG. 37.—A bottle for determining relative density.

iii. Find by the method used in the preceding experiment, the densities of such things as tin tacks (these are really made of iron), bits of slate pencil, brass wire or brass nails.

In all the above experiments it has been assumed that the mass of 1 c.c. of water is exactly 1 gram. As this will be later found to be only the case at 4° C., the experiments have really only given the relative density with respect to water at the temperature of the experiment. The difference, however, will be very small.

Determination of relative densities by relative density bottle.—A simple method of determining the relative densities of substances is to use a specific gravity or relative density bottle. Such a bottle often consists of a small glass flask, holding about fifty grams of water. It is provided with a nicely-fitting ground stopper, with a very small bore through it or a vertical groove cut upon it (Fig. 37). To use the bottle in determining the relative density of liquids and powders, the weight of the empty bottle and stopper must first be known. The bottle is then filled with pure water, the stopper inserted, and the water which is forced through the hole in the stopper wiped off, and the bottle and its contents weighed. In this way the weight of water which just fills the bottle is found. If now the bottle is emptied and carefully dried inside and out, and filled with the liquid of which the density is required, say spirits of wine, and weighed again, we have the weight of the liquid which just fills the bottle, or the weight of equal volumes of the liquid and water. The ratio—weight of liquid divided by weight of water—gives the relative density of the liquid.

21 TO FIND THE DENSITY OF A SMALL QUANTITY OF A LIQUID.

i. Use of U-tube.—For the determination of the density of a liquid, of which only a small quantity is available, a tube of the form shown in Fig. 38 can be used.

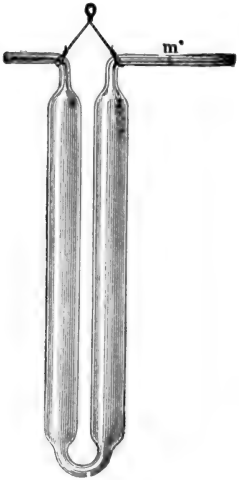


FIG. 38.—A tube with fine bore near ends, for the determination of the densities of liquids.

FIG. 39.—Method of filling a density tube having a fine bore.

The tube is filled with liquid by putting one of the open ends in a vessel containing the liquid, and attaching to the other end a pipette, by means of which suction can be applied (Fig. 39). When the tube is full the pipette is taken off, and the tube is then ready to be suspended from the balance. This method of finding density will be used later.

Determination of relative density with small U-tube.—When only a small quantity of a liquid is available, a smaller vessel must be used than the density bottle previously described. A tube of the form shown in Fig. 38 can be obtained, or can be made from a piece of quill tubing. The capillary tubes may be dispensed with, if the tube is drawn out at one side, and a file mark put on the other tube as shown at *m*, Fig. 38. Such an apparatus is filled as described above, but the liquid is allowed to run out till the file mark is reached, as in the case of a graduated pipette.

CHAPTER X.

PRINCIPLE OF ARCHIMEDES.

22. LOSS OF WEIGHT OF BODIES IMMERSED IN LIQUIDS.

i. Loss of weight of a body immersed in water.—Suspend a brick from a spring balance and note the weight of it. Immerse the brick in a pail of water, and again note its weight.

(*a*) Using a more sensitive spring balance, repeat this experiment with a large glass stopper, and again with a smaller stopper. The bodies lose in weight, and the amount of loss depends on the size of the body.

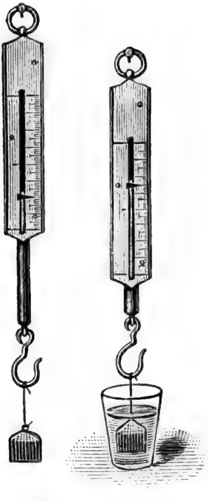


FIG. 40.—The block weighs less when immersed in water than when suspended in air.

ii. Repeat the experiment with the stoppers, using (*a*) Brine, (*b*) Methylated spirit. The loss of weight in spirit is less than the loss for the same body in water. The loss is greater when the body is weighed in brine than when weighed in water.

iii. Accurate determination of loss of weight of a body immersed in water.—Place a platform or bridge across one pan of the balance, as shown in Fig. 41. Suspend by means of a piece of thread a cube or cylinder of metal, letting it hang in an empty beaker, standing on the bridge, *H*, *H*, Fig. 41. Find the weight of the object. Then pour water into the beaker, and find the weight of the object when immersed in water. Make certain that there are no bubbles of air adhering to the object, before weighing in water. Find accurately by measurement and calculation the volume of the object. Repeat the experiment, using liquids of which the density has already been found, instead of water.

Tabulate the results.

Liquid.	Loss of wt. in liquid.	Volume of body in c.c.	Wt. of liquid displaced. Vol. × Density.

Buoyancy.—Most people have noticed when in a bath that if there is water enough, and they take hold of no support, the water buoys them up, or they experience a tendency to rise up through the water. It is as if the water resists being displaced and presses the displacing object upwards. In the case of things which float, such as a wooden rod or a lead pencil, the results of this upthrust which the water exerts can easily be seen, by pushing either the rod or pencil down into the water and then letting go, when the solid rises up through the water. Even in the case of bodies which sink, there is the same upthrust on the part of the water, but it is not enough to float them. The effect which the water has upon such bodies can, however, be seen in the loss of weight which they experience if they are weighed by a spring balance when hanging in water (Fig. 40).

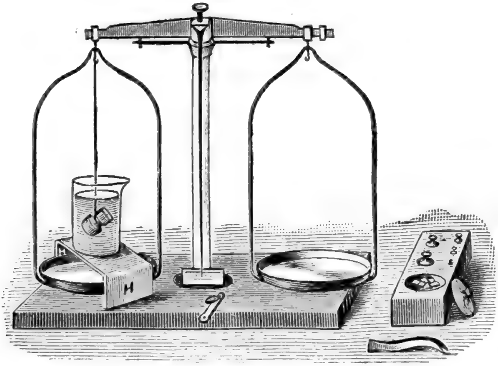


Fig. 41.—How to find the weight of an object suspended in water.

Loss of weight of things immersed in water.—It is easy to prove by experiment that an object weighs less in water than out of it. If a cubic centimetre of lead or any other heavy material, is hung from a spring balance and then suspended in water, it will be found to weigh the weight of one gram less in water than out. If two cubic centimetres are suspended from the balance, the loss of weight is the weight of two grams. In every case the loss of weight measured in this way is equal to the number of cubic centimetres of the solid immersed in the water. The loss is thus equal to the **weight of the water displaced.** This fact leads to a highly important conclusion, known after its discoverer as the **Principle of Archimedes.**

The Principle of Archimedes.—**When a body is immersed in water it loses weight equal to the weight of the water displaced by it.**—If the body be immersed in any other liquid, then the loss of weight is equal to the weight of an equal volume of that liquid. It does not matter what substance the thing is made of; the amount of loss of weight depends upon the *volume* of the part immersed, and not upon the material.

This principle explains many interesting facts. For instance, a ship made of iron, and containing all kinds of heavy things, is able to float in water although the material of which it is made is denser than water. This is because the ship and all its contents only weigh the same as the volume of the water displaced by the immersed part of the hull. Or, the ship as a whole weighs less than a quantity of water the same size as the ship would weigh.

Now, too, it can be seen why some solids float and some sink. When an object weighs more than an equal volume of water it sinks. When an object weighs less than an equal volume of water it floats. When an object weighs the same as an equal volume of water it remains suspended in the water.

A balloon rises in the air because the gas in it, together with the car and tackle, weigh less than an equal volume of air. If the balloon were free to ascend it would rise to a height where its weight would be equal to the weight of an equal volume of the air around it.

23. APPLICATION OF THE PRINCIPLE OF ARCHIMEDES TO THE DETERMINATION OF VOLUMES AND DENSITIES.

i. Accurate determination of the volume of a small body.—(*a*) Using the method described in Section 22, iii., find the loss of weight of a small glass stopper when weighed in air and in water. Since the loss of weight is equal to the weight of the water displaced it follows that the loss of weight in grams gives the volume in cubic centimetres, because one gram of water has a volume of one cubic centimetre.

(*b*) Find the volume, by the displacement of water in a graduated cylinder. Compare with the previous result.

ii. Density of glass.—Find the density of the stopper from the results of the last experiments.

iii. Relative density of liquids.—Determine the loss of weight of the stopper already used when weighed in methylated spirit. As the volume of spirit displaced by the stopper is the same as the volume of water displaced by it, the loss of weight in spirit represents the same volume of liquid as the loss of weight in water. Find the relative density of the spirit.

iv. Density of a cork.—As a cork floats in water, a sinker must be used. Obtain a piece of lead sufficient to sink the cork. Suspend the cork in water as in previous experiments, and weigh. Place the cork on the pan of the balance below the bridge, and, with the sinker still in water, weigh. Now tie the cork to the lead; lower into the water, taking care to remove air bubbles, and find weight. As the lead is in the water throughout the weighings the result is not affected by it. Find the weight, volume, and density of the cork.

The volume of small solids of irregular shape.—The volume of a solid can be found by observing the amount of water which it displaces in a graduated vessel. As the graduations on cylinders seldom read by less than 1 c.c. it is evident that this method does not admit of a degree of accuracy much beyond the whole number. The Principle of Archimedes provides a means of finding the volume of a small solid, correct to the second place of decimals. By weighing the body in air and in water as described in Section iii. the loss of weight is found. This gives the weight of a volume of water equal to the volume of the body. And since one gram of water has a volume of one cubic centimetre, the number of grams of water displaced gives the number of cubic centimetres also, and hence the volume.

Relative density of solids.—The fact that when a body is immersed in water it loses weight equal to the weight of the water displaced by it, provides the means of determining the density of a solid compared with water.

All it is necessary to know is:

1. The weight of the object; this can be determined by weighing it in air.
2. The weight of an equal volume of water.

The Principle of Archimedes enables this to be done in the following manner:

The object is suspended by means of a fine thread, from one side of the beam of a balance, in such a way that it is completely immersed in water. Then by weighing it is found that the weight is less than when hanging in air, on account of the loss of weight in the water. The buoyancy of the water acting upwards overcomes part of the pull of the earth downwards. The difference in the weight of the object in air and its weight when immersed in water gives the weight of a volume of water equal to the volume of the object. From these numbers the density of the solid compared with water as a standard can be calculated at once.

$$\text{Relative density of the solid} = \frac{\text{weight of the object in air}}{\text{weight of an equal volume of water}}$$

$$= \frac{\text{weight in air}}{\text{loss of weight in water}}.$$

Relative density of a substance which floats in water.—In the case of a body which is lighter than water, this method can be used with the assistance of a sinker, of sufficient weight to sink the body. The method is exactly the same as before, with the exception that the sinker is suspended in water throughout the experiment. Its weight does not in any way affect the results, as it is for the time being a part of the balance, just as the pans are.

Relative density of a liquid.—If a stopper be weighed in air, in a liquid, and in water, the loss of weight in each case is equal to the weight of a portion of liquid of the same volume as the stopper. From these numbers the density of the liquid can be calculated.

$$\left.\begin{matrix}\text{The relative density of the}\\ \text{liquid}\end{matrix}\right\} = \frac{\text{weight of a given volume of liquid}}{\text{weight of an equal volume of water}}.$$

The volume of a solid which is soluble in water.—This method can be used to find the volume of a solid which dissolves in water. Instead of weighing the solid in water it can be weighed in some liquid in which it does not dissolve. The loss of weight gives the weight of a volume of the liquid equal to the volume of the body. From this the weight of an equal volume of water can be obtained, by dividing by the relative density of the liquid. In this way the density of solids soluble in water may also be obtained. Sugar may be weighed in alcohol. Many substances which are soluble in water are insoluble in petroleum.

CHAPTER XI.

LAW OF FLOATING BODIES.

24. FLUID DISPLACEMENT AND FLOATING BODIES.

i. Some things sink others float in water.—Fill a pneumatic trough with water, and carefully place lumps of different things, *e.g.* pieces of lead, iron, oak, pine and cork, one after another, into the water. Observe that (1) some sink and others float, (2) of those which float some sink further into the water than others. Take the objects which sink in water and place them in mercury. Notice that they float. (Fig. 42).

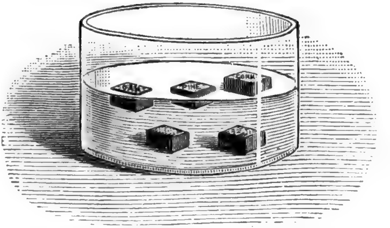


FIG. 42.—When substances are denser than water they sink in water; and if less dense, they float.

ii. Volume of water displaced by solids which float.—(*a*) Obtain a rectangular rod of wood, 1 square cm. in section and about 15 cm. long, with lines around it 1 cm. apart. Gouge a small piece of the wood out of one end, and put lead into the hole; flatten the end by filling in with wax.

Put some water in the graduated jar and notice its level. Find the weight of the rectangular rod, and then place it in the jar with the leaded end downwards. Notice how many cubic centimetres of the rod are immersed, and also how many cubic centimetres of water are displaced (Fig. 43). Since the weight of 1 cub. cm. of water is 1 gram, the number of cubic centimetres of water displaced is also the weight in grams of the water displaced. This weight will be found equal to the weight of the whole rod.

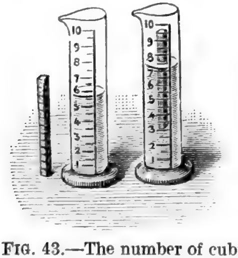


FIG. 43.—The number of cubic centimetres in the part of the rod under water is equal to the number of cubic centimetres of the water displaced.

(*b*) Weigh one of the larger wooden cubes. Slightly oil the cube and place it in a beaker of water. Mark upon the cube with a pencil where the surface of the water touches it. Then take out the cube and determine the volume of the immersed portion in cubic centimetres. From this will be obtained the volume, and therefore the mass of water displaced.

(*c*) Fill the divided glass cylinder with water up to a certain mark. Notice the level of the water.

Draw lines at equal distances apart across a narrow strip of paper and fix the paper inside a test tube, as in Fig. 44. Float the test tube in water in the graduated jar and put mercury or shot into it until a certain mark upon the strip of paper inside it is on a level with the surface of the water. Notice the number of cubic centimetres of water displaced when the test-tube is thus immersed.

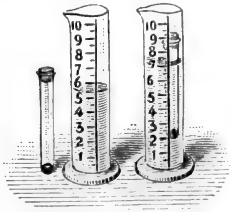


FIG. 44.—The weight of the test-tube and contents is equal to the weight of water displaced.

Then take out the test-tube, dry it, and weigh it together with the mercury it contains. The total weight of the test-tube and contents will be found equal to the weight shown by the number of cubic centimetres of water displaced. Repeat the experiment with the test-tube immersed to a different mark.

Float the test-tube and mercury in spirits of wine and milk in succession. Notice that in the former case it sinks deeper than the mark, while in the other not so deep.

(*d*) Place the loaded test tube or a hydrometer (1) in milk, (2) in water, (3) in a mixture of milk and water. Observe the depth to which it sinks in each case.

Water displaced by solids.—If a solid one cubic centimetre in size sinks in water it pushes aside one cubic centimetre of water to make room for itself. If its size is two cubic centimetres, it makes two cubic centimetres of water rise above the level the water had at first. Whatever the size of the solid it must have room, and this room is obtained by displacing an amount of water of exactly the same size.

Floating bodies.—A solid which sinks in water or any liquid, displaces a volume of liquid equal to its own volume. When a solid floats, the case is slightly different. Part of the solid is in water and part out of the water, and, of course, only the part immersed is pushing the water aside in order to make room for itself. **In the case of a floating object, therefore, the volume of liquid displaced is equal to the volume of the part of the solid below the surface.**

When any object is floating in water, a certain volume of it is under water, and a certain volume is above the surface. The depth at which it floats depends upon its density. A rod of heavy wood sinks deeper in water than a rod of light wood of the same size. The water displaced by the heavy wood has therefore a greater volume, and consequently a greater weight, than that displaced by the light wood. But there is one important fact which applies to both cases and should be kept well in mind. It is that **the weight of the water displaced by the**

immersed part of a floating object is equal to the whole weight of the object. If, therefore, the question is asked, how far does an object which floats sink into water the answer is—it goes on sinking until it has displaced an amount of water having a weight equal to the whole weight of the floating object.

Since the depth at which an object floats in water is decided by this rule, we have a ready way of deciding whether an object will sink further in another liquid or not so far. If the liquid into which it is put is less dense than water, like spirits of wine, it is clear that to make up a given weight we shall want more of the liquid. Consequently, to make up a weight equal to the weight of the floating body, the object will have to sink further into the spirit than into the water. If, on the other hand, the object is placed in a liquid such as mercury, which is denser than water, it will not sink so far, because it will not take so much of this denser liquid to have a weight equal to that of the floating body.

The hydrometer.—The construction of a simple instrument called the **hydrometer** is based upon these conclusions. The instrument is made in various shapes, and graduated differently according to the special use for which it is intended. All hydrometers are, however, used to measure the densities of fluids by observing the depth to which they sink when immersed in them. The **lactometer** is a form of hydrometer employed for measuring the density of milk. When placed in pure milk a lactometer should float with the mark *P* (Fig. 45) on a level with the surface of the liquid. In a mixture of milk and water the lactometer floats with some other division level with the surface of the liquid. Thus, in milk 10 per cent. below the average density, the 10 above the *P* mark is level with the surface.

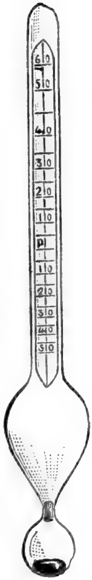


FIG. 45.—A hydrometer for determining the quality of milk. It is called a lactometer.

An experienced observer is, therefore, able from the readings of a lactometer to tell whether a sample of milk has a correct density, or whether it is heavier or lighter than it should be. At the same time it must be clearly understood that it is not possible to decide at once from the reading of a lactometer whether a sample of milk has been adulterated or not. There are other considerations to be taken into account.

QUESTIONS ON CHAPTERS IX., X., AND XI.

31. How would you test the accuracy of the readings of a burette ?

32. Enumerate the different methods by which you have found the specific gravity of a liquid, and mention any difficulties which you have had with any two of them.

33. Describe fully how to determine the density of milk without using any form of hydrometer.

34. Enumerate any special properties which distinguish water from mercury, and illustrate your statements from results of your own experiments or observation.

35. A bottle when filled with water weighs 211 gm. Fifty gm. of a solid are then added to the bottle, and the water which flows over is removed, leaving the bottle still filled. The bottle then weighs 255 gm. Find the specific gravity of the solid.

36. How would you verify by experiment that when a body floats in water the upward pressure on the body is equal to the weight of the displaced water ?

37. How would you prove that the weight of methylated spirit displaced by a floating body is equal to the weight of the body ?

38. On what does the apparent change in the weight of a body when immersed in a liquid depend ? How would you verify your answer by experiment ?

A piece of wood weighs 215 gm., and when fastened to a piece of metal and both immersed in water the joint weight is 166 gm. The metal alone weighs 280 gm. when immersed in water. Calculate the specific gravity of the wood, explaining your method.

39. Describe fully an accurate method of determining the volume of a small piece of metal of irregular shape.

40. Describe carefully how you would determine accurately the volume of a small piece of cork.

41. How would you find the volume of a piece of cork of irregular shape ? What sources of error might affect the measurement ?

42. The specific gravity of cork is 0·25, of iron 7·5. Find what weight of iron can be supported in water by 1 lb. of cork, (1) when the iron is on top of the cork and entirely out of the water ; (2) when the iron is fastened below the cork.

43. A gold coin is counterpoised on a balance by means of brass weights. The whole balance is then lowered into water. How will the equilibrium be effected ? Explain your answer.

44. A vessel containing water is placed on the scale pan of a balance, and equilibrium produced by weights on the other pan. A solid suspended from a string is lowered into the water, but not allowed to touch the vessel, and it is found that 5·3 grams added to the other pan restore equilibrium.

The solid is then allowed to rest on the bottom of the vessel, the string being slack, and a further 16·8 grams have to be added to the other scale pan.

What is the specific gravity of the solid ?

Explain your method of calculation.

45. In what circumstances will a solid body float in a liquid ? Describe experiments to establish definitely the law of flotation.

46. One gram of cork (Sp. gr. =0·25) is tied to 2 grams of brass (Sp. gr. =8). What is the specific gravity of the heaviest liquid in which they will sink ?

47. Find the ratio of the volume of an iceberg which is under water to the volume above the surface. The specific gravity of the sea water is 1·03, and of the ice 0·92.

48. A pound of iron and a pound of lead are hung on the two arms of a lever, the distance of each from the fulcrum being 1 foot. If both weights are allowed to hang completely immersed in water, find the distance through which the lead weight must be shifted to restore the equilibriun of the lever. The specific gravities of iron and lead are 7·5 and 11·3 respectively.

CHAPTER XII.

PROPERTIES OF GASES.

25. WEIGHT OF AIR.

i. Air has weight.—Fit up an eight ounce flask with a closely-fitting rubber stopper, having a hole in it through which a short piece of glass tubing passes. To this is fixed a small piece of rubber tubing closed with a clip. Suck out as much air as possible by means of an air pump; close the clip before removing from the pump and then weigh. Open the clip; air rushes in. Weigh again. There is a decided increase in weight. Air has weight!

ii. To find the weight of a litre of air.—Use the flask from the last experiment. Put 50 c.c. of water in it. Fit as shown in Fig. 47, *taking care that the clip is open.* Heat till steam is issuing freely, and continue heating for a further five minutes;

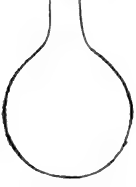


FIG. 46.

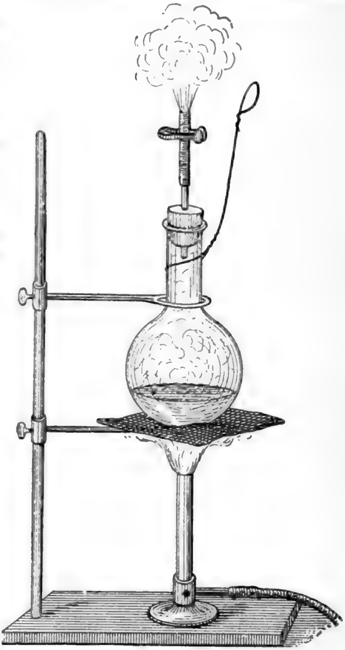


FIG. 47.

by that time the steam will have expelled all the air, then remove the flame and close the clip *immediately after.* Holding the flask firmly, place it under running water (note that the water in the flask boils). Wipe the outside of the flask thoroughly and weigh. Open the clip and note the inrush of the air. Weigh again.

Find the volume of air in the flask by taking the volume of water remaining in the flask from the volume of water required to fill the flask.

Calculate the weight of 1000 c.c. Note the readings of the barometer and thermometer. The weight of air under "standard" conditions, will be found later.

Air is a kind of matter.—There is abundant evidence that air is matter. It can be felt by the body as wind, and the results of its motion are seen when it moves trees or loose objects on the ground. Even on the calmest day the air can be felt by swinging the open hand to and fro.

Air has weight.—As air is a kind of matter it must have weight. By weighing a flask from which air has been removed, and then again after the admission of air, it is found that air has undoubtedly got weight. Careful experiments show that under "standard" conditions a litre of air weighs 1·293 gm.

26. PRESSURE EXERTED BY THE ATMOSPHERE.

i. Take a U-tube such as shown in Fig. 48, and half fill it with water. Attach a piece of rubber tubing to one limb, and suck air out. Notice that air rises in that limb.

If a good air pump is available, repeat this experiment, using a large U-tube made by connecting two three-foot lengths of glass tubing by means of pressure tubing. Half fill the tube with mercury and connect one limb to the air pump. Remove as much air as possible. Note that it is impossible to get the mercury to rise above a certain point.

ii. Place in water one tube of the Hare's apparatus for determining density (Fig. 49), and the other tube in mercury. Suck out the air. The liquids rise. Why? Notice the difference in the level of the mercury and the water. Explain the cause of this difference. Let one of the tubes of the Hare's apparatus be much wider than the other. Place the ends of the tubes in mercury, and suck out the air. Is there any difference in the height of the mercury in the large and small tubes?

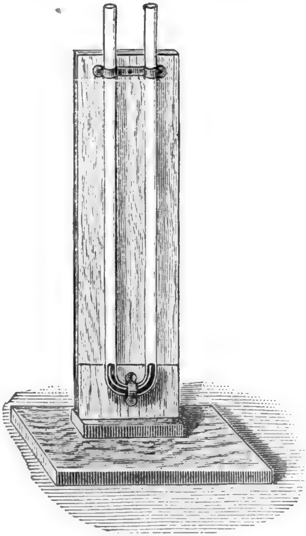


FIG. 48.

Pressure exerted by the atmosphere.—It has been seen that air has weight, and as there is a very considerable quantity of it, extending as it does to well over 100 miles from the surface of the earth, it may be expected to exert a pressure on bodies beneath it. That this is the case is shown by placing water in a U-tube and sucking the air out of one limb: the water rises in that limb, because the air on the water in the open limb forces it down, the air in the other limb which previously counterbalanced this pressure being removed by suction.

The same kind of effect is noticed when air is sucked out of a Hare's apparatus, one limb being inverted over water and the other over mercury. In this case the water rises further than the mercury, and measurements of the lengths of the respective columns show that their lengths are inversely proportional to their densities, so that per unit of cross-section the same weight is supported in each column. The surface of the liquid in the beaker in each case corresponds to the surface of the liquid in the open limb of the U-tube, and the air pressing down on it supports the liquid in the limb.

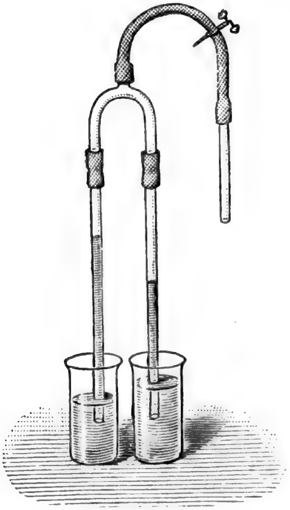


FIG. 49.—Hare's apparatus for comparing the densities of liquids.

26A. MEASUREMENT OF THE PRESSURE OF THE AIR.

i. The principle of the mercurial barometer.—Procure a barometer tube and fit a short piece of india-rubber tubing upon its open end. Tie the free end of the tubing to a glass tube about six inches long, open at both ends. Rest the barometer tube with its closed end downwards and pour mercury into it (being careful to remove all air bubbles) until the liquid reaches the short tube. Then fix the arrangement upright as in Fig. 50. The mercury in the long tube will be seen to fall so as to leave a space of a few inches between it and the closed end. The distance between the top of the mercury column in the closed tube and the surface of that in the open tube will be found to be about 30 inches.

ii. The cistern barometer.—Procure a thick glass tube about 36 inches long and closed at one end. Fill the tube with mercury; place your thumb over the open end; invert the tube; place the open end in a cup of mercury and take away your thumb (Fig. 51). Measure the distance from the surface of the mercury in the basin to the top of the mercury column.

Slant the barometer and measure the vertical height of the mercury by means of a plumb-line. Compare it with the previous reading.

The mercurial barometer.—The apparatus represented in Fig. 50 calls to mind the **U**-tube method of determining densities by balancing columns of liquid (p. 61). Here, however, the columns balancing the mercury in the tube *A* is not apparent. The experimental fact that the column *A* does not rush up through *O* to adjust the liquid levels to equality demands explanation. Remember that the top of the column *A* is closed and protected while *O* is open. Something acts at *O* and is capable of sustaining the column *A*. This something is the weight of the air. In fact, a

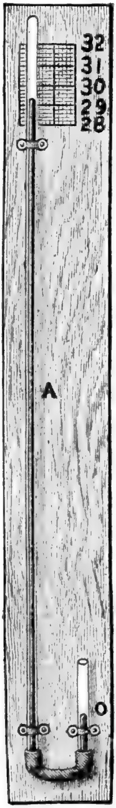


FIG. 50.—The pressure of the air on the mercury in the small tube is able to keep up the column of mercury in the long tube.

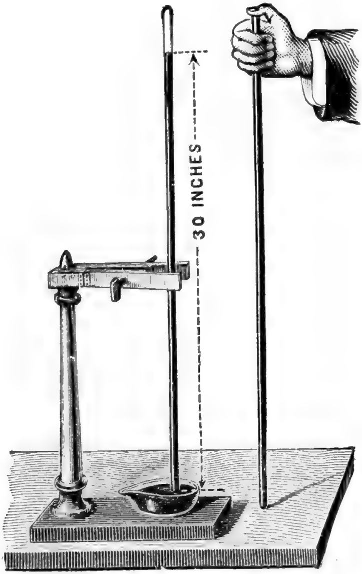


FIG. 51.—Construction of a barometer.

column of air many miles high balances a column of mercury about 30 inches high.

If, however, a hole were made in the closed end of the tube the balance would be disturbed, and the mercury would fall down the long tube and flow out of the short tube *O*.

The column of mercury is kept in its position by the weight of the atmosphere pressing upon the surface of the mercury in the short open tube. The weight of the column of mercury and the weight of a column of the atmosphere with the same sectional area is exactly the same ; both being measured from the level of the mercury in the short stem of the apparatus shown in Fig. 50, the mercury column to the top of the column in the long tube the air to its upper limit, which, as has been seen, is a great distance from the surface of the earth. If for any reason the weight of the atmosphere becomes greater, the mercury will be pushed higher to preserve the balance ; if it should become less, then similarly the amount of mercury which can be supported will be less, and so the height of the column of mercury is diminished.

The height must in every case be measured above the level of the mercury in the tube or cistern open to the atmosphere. In the usual arrangement, illustrated by Fig. 50, a line is drawn at a fixed point *O*, and the short tube is shifted up or down until the top of the mercury in it is on a level with this line.

It will now be understood why it is so necessary to remove all the air bubbles in Experiment 26A, i. If this is not done, when the tube is inverted the enclosed air would rise through the mercury and take up a position in the top of the longer tube, above the mercury. The reading would not then be thirty inches, for instead of measuring the whole pressure of the atmosphere, what we should really be measuring would be the difference between the pressure of the whole atmosphere and that of the air enclosed in the tube. In a properly constructed barometer, therefore, there is nothing above the mercury in the tube except a little mercury vapour.

An arrangement like that described constitutes a barometer, which may be defined as an instrument for measuring the pressure exerted by the atmosphere.

The cistern barometer.—Other forms of barometer are often employed for the determination of the pressure of air. A very common arrangement is that of Experiment 26A, ii., which is a repetition of one by an Italian physicist, Torricelli. The principle of its action is precisely that of the barometer just described, except that the U-tube principle is not immediately apparent. There is, however, the same balance maintained between the column of mercury in the tube and a column of air outside it, pressing down upon the mercury in the basin.

A column of mercury will be supported in the tube by the pressure of the atmosphere. The distance between the top of the column and the surface of the mercury in the cup will be about 30 inches, or 76 cm., when the tube is vertical (Fig. 52, *b*). If the tube is inclined so that the closed end of it is less than this height above the mercury in the cup (Fig. 52, *c*) the mercury fills it completely; and if the tube is less than 30 inches long, it is always filled by the mercury whether it is inclined or not (Fig. 52, *a*). On an average the atmosphere at sea-level will balance a column of mercury 30 inches in height. No matter if the closed tube is 30 feet long, the top of the mercury column will only be about 30 inches above the level of the mercury in the basin.

The empty space above the column of mercury in the tube is often referred to as the **Torricellian vacuum.**

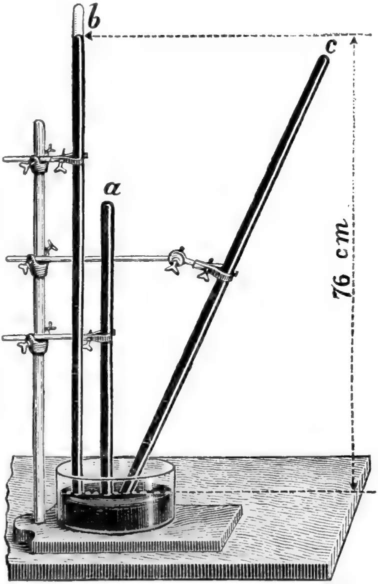


FIG. 52.—The atmosphere at sea-level will support a column of mercury up to 30 inches, or 76 cm. in length.

The aneroid barometer.—Owing to its necessary size and weight, the mercury barometer is not a portable instrument. The **aneroid barometer** has been designed to get over this difficulty. It consists essentially of a small box closed on top with a thin sheet of corrugated metal, and partially exhausted of air (Fig. 52*a*). This metal top is very sensitive to variations of pressure, and by means of levers the movement of the metal is magnified and transmitted to a circular dial, where a hand or pointer is rotated. The dial is graduated to read the same as a mercury barometer, by comparison with one of the latter.

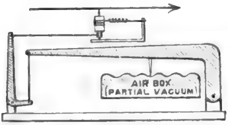


FIG. 52*a*.—Section of an aneroid barometer.

27. BOYLE'S LAW.

i. Relation between pressure and volume of gases.—(*a*) Select a glass tube about 20 cm. long, and neatly closed at one end (Fig. 53, *A*). Tie a piece of stout india-rubber tubing about a metre long upon the open end of the tube, and fix the other end of the tubing upon a glass tube *B*, about 20 cm. long, open at both ends. Now carefully fill the tubes with mercury until the level of the liquid is about 10 cm. from the open end. Afterwards fix the closed tube upright in a retort stand with the sealed end upwards, place your finger upon the open end of the other tube, and lower the open end so as to make the air pass into the closed tube. This apparatus will enable you to measure the expansion and compression of air.

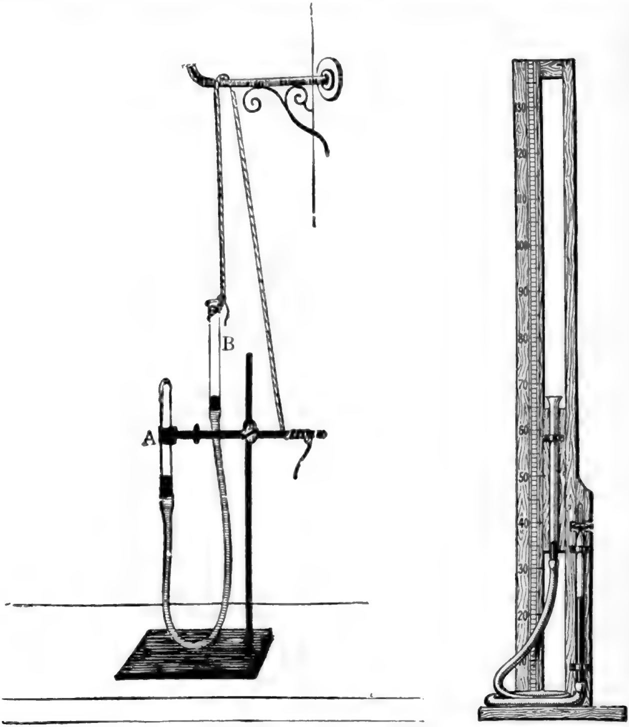


FIG. 53.—Two forms of apparatus to show the relation between the pressure and volume of a gas.

Support the tube *B* with its open end upwards, and at such a height that the mercury stands at the same level in the closed and open tubes. The imprisoned air is then at the same pressure as the air outside.

If the closed tube is uniform in bore, and the inside of the sealed end is nearly flat, the volume of the imprisoned air is proportional to the length of tube occupied by the air, so that if the air is made to occupy one-half the original length of tube its volume is one-half the original volume. The pressure upon the enclosed air is equal to the pressure due to the column of mercury between the level of mercury in the closed tube and that in the open tube, *plus* the pressure of the atmosphere. Observe the height of the barometer, and make the difference of level of the mercury in the two tubes equal to it by lifting the open tube. Then record as follows:

Height of the barometer, - - - - - - - ...... cm.
Length of air column when the mercury is at the same level in both tubes, that is, when the imprisoned air is at atmospheric pressure, - - ...... cm.
Height of barometer ...... cm. + equal height in tube, ... cm.
Length of air column under pressure of two atmospheres, ... cm.

The pressure to which the enclosed air is subjected in the second case is double that in the first case; find the amount by which the volume of the air, represented by the length of the air column, has been diminished.

(*b*) Lower the open tube until the air in the closed tube almost reaches the india-rubber junction. Measure the length of the air column and the difference of level of the mercury in the two tubes. Repeat the experiment by reading the volume of air and the head of mercury at every few centimetres up to the highest point you can raise the open tube. Record your results as indicated below:

Height of Barometer in cm.	Difference of Level of Mercury in cm.	Total Pressure on the Air, P.	Volume of Air, V.	Volume × Total Pressure. ($P \times V$.)

ii. A simple form of Boyle's Law apparatus.—Procure a length of thermometer tubing, AB (Fig. 54), about 75 cm. long and 1 mm. bore. Seal it at B and expand the end A somewhat. Clamp AB in a vertical position by the side of a metre scale, and connect a small funnel to A by means of a short piece of rubber tubing. Pour a little pure, clean mercury into the funnel and induce it to run down the bore of the tube by inserting a thin, clean, steel wire. In this way any desired volume of air can be enclosed.

The length of the column of enclosed air may be taken to represent its *volume* (V). If H = the height of the barometer, and h = the length of the mercury thread (both expressed in the same units), then the total pressure on the enclosed air = $(H + h)$.

Introduce more mercury in the same manner, and in this way alter the values of V and h. The volume of the air under the pressure of the atmosphere alone can be observed by laying the glass tube flat on the table.

Perform several experiments and record your results in the following way:

Volume (V).	Pressure ($H+h$).	Volume × Pressure.

Boyle's Law.—To understand clearly how and why the density of the atmosphere varies, it is necessary to become acquainted with the rule expressing the relation between the volume and pressure of a gas. This can be satisfactorily done by one of the forms of apparatus employed in Experiments 27, i. and 27 ii., which provide a means of subjecting an enclosed quantity of air to varying pressures, by the addition of smaller or larger quantities of mercury. When in the apparatus shown in Fig. 53 the mercury in both tubes stands at the same level, the enclosed air is at the same pressure as the air of the room, but as the tube *B* is raised, the mercury in it stands higher than that in *A*, and the air enclosed in *A* is under a total pressure equal to the sum of that due to the atmosphere and that due to a column of mercury equal in length to the difference of levels of the mercury in *A* and *B*. In these circumstances the volume of the air in *A* decreases, and it decreases more and more as the total pressure is increased. When the results of experiments with any satisfactory form of apparatus are tabulated, certain very important relations between the volume of a gas and the pressure to which it is subjected become evident. It is found that the volume regularly diminishes as the pressure is increased, and in the same ratio. The converse is also found to be true, viz., that when the volume of a gas increases the pressure upon it has diminished, and exactly at the same rate.

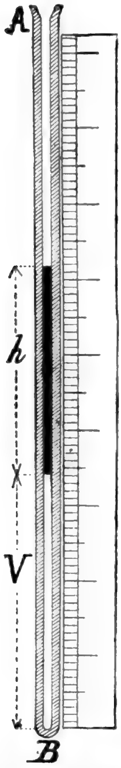


FIG. 54.—Simple form of apparatus for verifying Boyle's Law.

But, in both these cases, it is understood that the temperature of the gas remains the same; that is, the temperature of the gas under different pressures must not alter.

The tabulated results of the experiments reveal another important relation, which is, however, another way of expressing those already noticed. It is found that, when there is no alteration of temperature, the product obtained by multiplying the volume of a given quantity of gas by the pressure to which it is subjected is always the same, or remains constant.

These facts were discovered by Boyle, and are included in what is known as Boyle's Law. It can be expressed by saying that when **the temperature remains the same, the volume of a given mass of gas varies inversely as its pressure.** Or, what is the same thing, **the temperature remaining the same, the product of the pressure into the volume of a given mass of gas is constant.**

But it has been seen that if the volume occupied by a given weight of a substance is increased, its density is decreased, and if the volume is decreased, its density is increased. Therefore, by decreasing the volume of the enclosed air in the above experiment, its density is increased. The increase of density and the increase of pressure are proportional to one another. It is not difficult to apply these facts to the case of the atmosphere. It has been learnt that the pressure of the atmosphere decreases as we ascend, and we are now able to add that its density decreases also and at the same rate. Therefore the densest atmosphere is that at the surface of the earth, leaving out, of course, the air of mines and other cavities below the surface, where the air is denser still. The air gets less dense, or rarer, as we leave the earth's surface, until eventually it becomes so rare that its existence is practically not appreciable.

QUESTIONS ON CHAPTER XII.

49. How could you prove that air has weight ? How would you find the weight of the air contained in a flask ?

50. Distinguish between "mass" and "weight." Explain how you would find, by experiment, the weight of a cubic metre of the air in a room

51. Describe briefly two methods for obtaining the weight of a litre of air, and show which is the more likely to give an accurate result.

52. Explain how you would find the weight of air contained in a room. How would you expect this to vary with the height of the barometer and thermometer ?

53. Give an account of the construction of a barometer, and explain how it measures the pressure of the atmosphere.
Calculate the pressure of the atmosphere when the mercury in a barometer stands at 76·2 cm.

54. Describe how you would proceed to construct and set up a simple barometer, indicating the chief precautions to be observed.

55. Describe a simple barometer. Why is mercury used ? Would any other liquid serve ? How may a barometer be used for measuring heights of mountains ?

56. Explain as clearly as you can why the height of the barometer gives a measure of the pressure of the atmosphere.

57. How would the position of the mercury in a simple barometer be affected (1) if the tube were sloped ? (2) if the tube had a very irregular section, wide in some places and narrow in others ? Give full reasons for your answers.

58. A faulty barometer contains some air above the mercury. Will the extent of its rise and fall be greater or less than the corresponding movements of a correct barometer ? Give the reason for your answer.

59. A barometer tube stands vertically with its open end below the surface of mercury. The mercury stands at the same level inside and outside the tube, and the closed end of the tube is 20 inches above this level. If the tube is now raised until the mercury stands 4 inches higher inside than outside, find the height of the closed end above the outside level, the height of the barometer being 29 inches.

60. State the law which governs the change of volume of a given mass of gas under varying pressure, at constant temperature. Who discovered this law ?

61. State Boyle's Law.
A U-tube of uniform bore has one end closed and the other open. Mercury is poured in, and when it stands at the same level on the two sides the length occupied by air on the closed side is 12 inches. More mercury is then poured in till the level on the closed side rises by 1 inch. What is now the difference of level on the two sides and how much mercury was poured in between the two observations ? The height of the barometer is 30 inches.

62. Describe a way of finding out how the pressure of a quantity of air changes when the air is compressed into smaller volume. Draw a sketch of the apparatus used, and state the result which is arrived at.

63. If you were supplied with a barometer tube, a deep reservoir full of mercury and a measuring scale, how would you study the relation between the pressure and volume of a quantity of air at constant temperature ?

64. Being provided with a U-tube of uniform bore with one end closed, a measuring scale, and some mercury, how could you determine the atmospheric pressure ?

65. How would you examine the relation between the volume and pressure of a gas, the temperature being constant ? Give a sketch of the apparatus used.

CHAPTER XIII.

EFFECTS OF HEAT.

28. EXAMPLES OF EXPANSION BY HEAT.

i. Expansion of solids.—(*a*) Take a metal ball suspended by a chain as shown in Fig. 55, and suspend it by the side of a metal ring, through which it just passes easily. Heat the ball in a laboratory burner for a few minutes and then try to drop it through the ring. It is too large and rests on the ring. Now allow it to cool slowly and notice that after a short time it gets smaller and will slip through.

FIG. 55.—Apparatus for showing that a metal ball is larger when hot than it is when cold.

FIG. 56.—Compound strip of ebonite and wood to show the greater expansion of ebonite when heated.

(*b*) Solder a strip of wire or brass, about two feet long, to one of iron of the same length. Straighten the compound strip by hammering ; then heat it. Notice that the strip bends, because the brass expands more than the iron. The same effect can be shown by means of a strip of ebonite glued to a strip of wood, on account of the ebonite expanding more than the wood (Fig. 56).

(*c*) Place a heavy mass on one end of the iron bar resting upon one of two blocks, as in Fig. 57. Let the other end bear upon the needle placed upon the other block and having the straw pointer fixed to it. Heat the bar with a flame and notice that the pointer moves on account of the expansion of the iron.

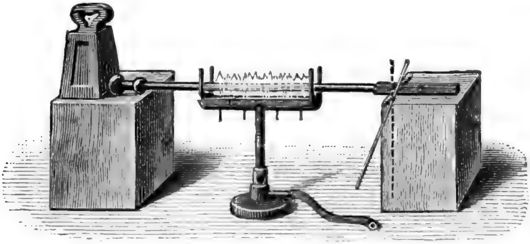


FIG. 57.—A flat bar of metal having one end kept from moving by a heavy mass is heated and the other end moves the pointer, because the bar gets longer.

ii. Expansion of liquids.—(*a*) Procure a 4-oz. flask and fit it with a cork. Bore a hole through the cork and pass through it a long glass tube which fits tightly. *Fill* the flask with water coloured with red ink. Push the cork into the neck of the flask and so cause the coloured water to rise up the tube. See that there is no air between the cork and the water. Now dip the flask in warm water, and notice that the liquid soon gets larger and

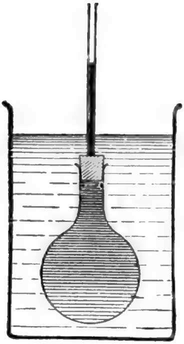
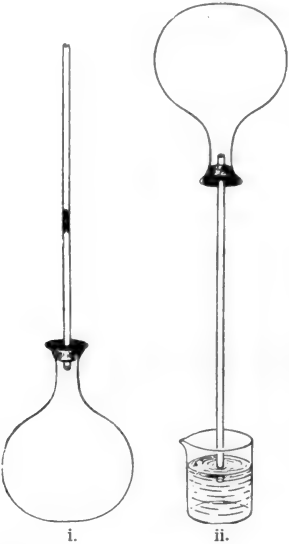


FIG. 58.—The expansion of a liquid.



i. ii.

FIG. 59.—Experiments to show the expansion of air by heat.

rises up the tube (Fig. 58). Take the flask out of the warm water, and see that the coloured water gets smaller as it cools, and that it sinks in the tube.

(*b*) Arrange two other flasks as in the last experiment, but filled respectively with alcohol and turpentine. Push in the corks till the liquid stands in each tube at the same height. Put the flasks to the same depth into a vessel of warm water. Notice that the expansion of the glass causes a momentary sinking of the liquids; and that ultimately the expansions of the liquids are very different.

iii. Expansion of gases.—(*a*) Procure a well-made paper bag and tightly tie a piece of tape round the open end. Hold the bag in front of the fire and notice that the air inside gets larger and inflates the bag.

(*b*) Or, obtain a flask with a cork and tube as in Fig. 59, i. Remove the cork and tube, and, by suction, draw a little red ink into the end of the tube near the cork. Re-insert the cork, and gently warm the flask by clasping it in the hands. Notice that the air in the flask gets larger and pushes the red ink along the tube.

(*c*) Turn over and place the open end of the tube beneath the surface of some coloured water in a beaker. Warm the flask with the hand or a flame so as to expel some of the air, and let the liquid rise in the stem (Fig. 59, ii.). This arrangement constitutes an *air thermometer*.

(*d*) Fasten in an air-tight manner two bulbs or flasks together by a tube bent six times at right angles, and containing some coloured liquid in the middle bends (Fig. 60). Show that the liquid moves if one flask is warmed more than the other. This instrument is known as a *differential thermometer*.

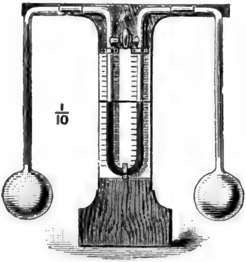


FIG. 60.—A simple form of differential thermometer.

Change of size. Expansion.—As a rule all bodies, whether solid, liquid, or gaseous, get larger when heated, and smaller when cooled.

The change of size which a body undergoes is spoken of as the amount it expands or contracts; or, heat is said to cause expansion in the body. This expansion is regarded in three ways. When dealing with solids, expansion may take place in length **(linear expansion)**, in area **(superficial expansion)**, and in volume **(cubical expansion)**. In the case of liquids and gases we have only cubical expansion. Similar terms can be used with reference to contraction.

The expansion which substances undergo when heated must be allowed for in many engineering projects. Railway lines, for instance, are usually not placed close together, but a little space is allowed between the separate rails, so that they can expand in summer without buckling. Steam pipes used for heating rooms are also not fixed firmly to the walls at both ends, but are left slightly loose or are loose-jointed, so that they can expand or contract without doing any damage. For the same reason the ends of iron bridges are not fixed to the supports upon which they rest. Iron tyres are put on carriage wheels by first heating the tyre, and while it is hot, slipping it over the wheel. As the tyre cools it contracts and clasps the wheel tightly.

The common occurrence in domestic life of the cracking of thick glasses, when boiling water is poured on them, may be explained by this expansion of solids by heating. The part of the glass with which the hot water comes in contact is heated and expands; but the effect is quite local; the heating is confined to one spot, because glass does not allow heat to pass through it readily. It is this local expansion of the glass which results in the cracking of the vessel.

It is not to be supposed, however, that substances always expand on being heated; it will be learned later that water, under certain conditions, shrinks in volume as its temperature rises. A piece of rubber too, stretched by a weight, contracts considerably on being heated. This result is, however, deceptive. Unstretched rubber acts in the usual way and expands as its temperature rises. The explanation is that hot rubber is not so easily stretched as cold. The amount of stretch due to the weight is lessened, and masks the expansion due to the heat.

Measurement and change of temperature.—Change of temperature means change in the state of hotness or coldness of a body. The change of size which takes place when a thing is heated gives a good way of measuring the change of temperature which it undergoes. Think of the experiment with the coloured water in the flask with a long tube attached to it. Suppose the coloured water in the tube rises through a certain number of inches after the water has been heated, and that when the flask is placed into some other liquid, or some more water, the coloured water is found to rise up the tube to just the same place, we should have every right to say that the second liquid is exactly as hot as the first was. This plan proves a means of measuring temperature. The flask and tube with the water have become a "temperature measurer," that is, a **thermometer**.

CHAPTER XIV.

MEASUREMENT OF TEMPERATURE.

29. TEMPERATURE AND THERMOMETERS.

i. The sense of feeling may be deceived.—Arrange three basins in a row; into the first put water as hot as the hand can bear, into the second put lukewarm water, and fill the third with cold water. Place the right hand into the cold water and the left into the hot, and after half a minute quickly put both into the lukewarm water. The left hand feels cold and the right hand warm while in the same water.

ii. Measuring temperature.—(*a*) Place the flask of water, with fitted tube used in Expt. 28, ii. (*a*), in hot water, and notice the height of the liquid in the tube. Transfer it to cold water, and observe that the liquid in the tube sinks.

(*b*) Procure, or make, a thermometer tube with a bulb at one end. With a little practice it is easy to blow a bulb upon a piece of thermometer tubing. One end of the tubing is held in a blowpipe flame and twirled round until the glass melts and runs together so as to seal up the tube. A small blob of glass is then allowed to form, and while the glass is molten the tube is taken out of the flame and blown into steadily. To introduce mercury afterwards, heat the bulb and while it is hot invert it and put the open end into mercury. As the bulb cools, mercury will rise in the tube and take the place of the air driven out of the tube by heat. Repeat the operation until the mercury fills the bulb and part of the stem.

(*c*) Place in hot water the bulb of the instrument just constructed, and make a mark at the level of the mercury in the tube. Now place the instrument in cold water, and notice that the mercury sinks in the tube. The mercury is thus seen to expand when heated and contract when cooled.

(*d*) Examine a thermometer. Notice that it is

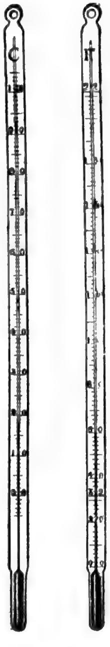


FIG. 61.—Thermometers of the form used for scientific work.

similar to the simple instrument already described, but the top is sealed up, and divisions or graduations are marked upon it, so that the height of the mercury in the tube can be seen easily. These divisions are called *degrees* (Fig. 61).

Feeling of heat and cold.—Some people feel cold at the same time that others feel warm. It is easy to understand that the sense of feeling cannot be depended upon to tell us accurately whether the air or any substance is hot or cold. Some instrument is needed which does not depend upon feeling, and cannot be deceived in the way that the senses can. Such an instrument is called a **thermometer,** and it is used to measure temperature, that is, the degree of hotness or coldness of a body.

How expansion may indicate temperature.—It has already been learned that substances usually expand when heated and contract when cooled. A flask filled with water, for instance, and having a stopper through which a glass tube passes, can be used to show the expansion produced by heat and the contraction by cold. But this flask and tube make but a very rough temperature measurer. The water does not get larger to the same amount for every equal addition of heat. Neither is it very sensitive, that is to say, it does not show very small increases in the degree of hotness or coldness, or, as the student must now learn to say, it does not record very small differences of temperature, and for a thermometer to be of value it must do this. Then, too, as every one knows, if water is made very cold it becomes ice, which, being larger than the water from which it is made, would crack the tube. For many reasons, therefore, water is not a good thing to use in a thermometer.

Choice of things to be used in a thermometer.

1. The substance used should expand a great deal for a small increase of temperature.

Gases expand most, and solids least, for a given increase of temperature. Liquids occupy a middle place. The most delicate thermometers are therefore those where a gas, such as air, is the substance which expands. But in common thermometers a liquid, either quicksilver or spirits of wine, is used. Both these liquids expand a fair amount for a given increase of temperature, and, to make this amount of expansion as obvious as possible, they are used in fine threads by making them expand into a tube with a fine very bore.

2. If a liquid is used it should not change into a solid unless cooled very much, nor into a gas unless heated very much.

It is difficult to be sure of both these things in the same thermometer. When a thermometer is required for measuring very low temperature it usually contains spirits of wine, because this liquid has to be cooled a great deal before it is solidified, that is, made into a solid. But this thermometer cannot be used for any great degree of temperature because spirits of wine is soon changed completely into a vapour when heated to only a comparatively small extent. If it is wished to measure higher temperatures a quicksilver or mercury thermometer is used, because mercury can be warmed a good deal, or, as it is better to say, raised to a high temperature, without being changed into a gas.

3. The liquid should be in a fine tube of equal bore with a comparatively large bulb at the end.

Liquids have to be contained in some sort of vessel or else they cannot be kept together. There must be a fine bore, so that the liquid may appear to expand very much for a small change of temperature. The bore must be equal all the way along, that is, the width or diameter of the inside of the tube must be the same all the way along, so that a given amount of expansion in any part of the tube shall mean the same change of temperature; and, lastly, there must be a large bulb, so that there is a large surface to take the same temperature as that of the substance the temperature of which is being measured.

Reasons why mercury is used for thermometers.—There are many reasons for selecting mercury as the liquid for an ordinary thermometer in addition to those already mentioned.

It is a liquid the level of which can be easily seen.

It does not wet the vessel in which it is contained.

It expands a considerable amount for a small increment of temperature.

It is a good conductor of heat, and consequently it very quickly assumes the temperature of the body with which it is placed in contact.

Very little heat is required to raise its temperature, and there is therefore very little loss of heat due to warming the thermometer.

Construction of a thermometer.—Having selected a suitable piece of thermometer tubing, first a bulb must be blown on one end. The glass is melted at this end and allowed to run together and so close up the bore. While the glass is still molten, air is blown down the tube from the other end, the tube being moved round, so that the bulb may be placed symmetrically with reference to it. The bore of the tube is so fine that it is impossible to pour the liquid down it; some other plan must therefore be adopted. The top of the tube may be enlarged into the shape shown at *A* in Fig. 62, or a small funnel may be connected with it, as at *B*, and the liquid with which the thermometer is to be filled poured in. Let us suppose mercury is being used. Warming the tube makes the air inside it expand, and of course some is driven out. As the tube cools the mercury is forced in, by the weight of the atmosphere, to fill the place of the expelled air. By repeating this alternate process of warming and cooling, in the circumstances described, enough mercury is soon introduced into the tube. The next step is to seal up the tube, leaving no air above the mercury; to do this the bulb is heated to a temperature slightly higher than the thermometer is expected to register. The mercury expands, and when it has reached the drawn-out part, *C*, of the tube, a blowpipe flame is directed against the glass, and the tube is thus closed up. This method of closing a tube and keeping the air out is called **hermetically** sealing it. The thermometer at this stage should be put on one side for some days at least, in order that it may assume its final size, which it does very slowly.

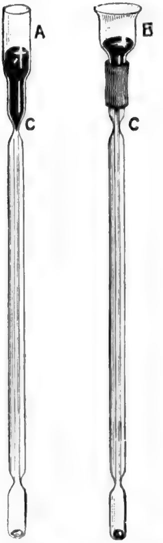


FIG. 62.—Thermometers in course of construction: (i) tube enlarged at top; (ii) tube with small funnel attached.

30. GRADUATION AND USE OF THERMOMETERS.

i. The temperature of melting ice.—(*a*) Place some pieces of clean ice in a beaker or test-tube and push a thermometer amongst them. Notice the reading of the thermometer; it will be either *no degrees* (0°) or very near it.[1] Warm the beaker or test-tube, and observe that as long as there is any ice unmelted the reading of the thermometer remains the same.

(*b*) Repeat the experiment with pieces of some other blocks of ice, and observe the important fact that the temperature of clean melting ice is the same in all your tests.

ii. Effect of adding salt to the ice.—(*a*) Add salt to the melting ice, and notice that the thermometer indicates a lower degree of temperature.

iii. The temperature of boiling water.—(*a*) Boil some distilled water in a flask, test-tube (Fig. 63), or beaker, and hold a thermometer in the boiling water. Notice the temperature. Raise the thermometer until

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[1] A Centigrade thermometer is supposed to be used. If a Fahrenheit thermometer is used the reading will be 32°.

the bulb is just out of the water and only heated by the steam. Again record the temperature. In both cases the reading is the same. It is either *one hundred degrees* (100°), or very near it, if you use a thermometer with Centigrade divisions.

(*b*) Repeat the experiment with a second lot of pure water, and note that the temperature of boiling water is again about 100°.

(*c*) Add salt to the water. Hold a thermometer in the steam of the boiling water, and notice that the temperature is the same as before, namely 100°. Push the thermometer into the water, and notice that a higher degree of temperature is indicated.

(*d*) Again place the thermometer in clean ice in a test-tube or flask. Gently heat the vessel, and notice the following changes:

(1) The mercury remains at 0° until all the ice is melted.

(2) When the ice is melted the mercury rises gradually until it reaches 100°.

(3) The mercury remains stationary at 100° until all the water is boiled away.

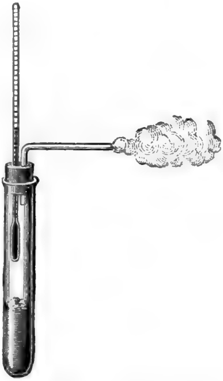

FIG. 63.—Test-tube with thermometer fitted for observing the boiling point of water.

iv. A thermometer cannot be deceived.—Arrange three basins of cold, lukewarm, and hot water side by side. Place the thermometer in the cold water and then in the lukewarm water. Notice the temperature indicated in the lukewarm water. Now place the thermometer in the hot water, and when it has been there a minute or two put it into the lukewarm water. Notice that the temperature indicated is practically the same as before. It is thus seen that, unlike the sense of feeling, a thermometer is not deceived by being made hot or cold before using it to indicate temperature.

v. The clinical thermometer.—Examine a clinical thermometer. Notice the large space between separate degrees. Observe that just above the bulb the bore of the thermometer is constricted. Hold the bulb in the hand and watch the expansion of the mercury. On exposing the thermometer to the air and allowing it to cool, notice that the mercury thread breaks at the constriction. The mercury can be shaken back into the bulb (Fig. 67).

The fixed points on a thermometer.—In the graduation of a thermometer the plan usually adopted is to choose "two fixed points" from which to number degrees of temperature. The most convenient lower fixed point to select is the temperature at which ice melts, or water freezes, for this is always the same if the ice is pure, and remains the same so long as there is any ice left unmelted. Whenever the thermometer is put into melting ice, the mercury in it always stands at the same level; or melting ice is always at the same temperature and thus may be used to give one fixed point. The "higher fixed point" chosen is that at which pure water boils at the sea-level. This stipulation must be made, for the boiling point of a liquid is altered when the pressure upon it is changed, being raised if the pressure is greater, and lowered if the pressure is less. When the water boils, the temperature of the steam is the same as that of the water, and remains so as long as there is any water left. The lower fixed temperature is referred to as the "Freezing Point of Water," and the higher as its "Boiling Point."

Marking the freezing point.—For this purpose an arrangement like that shown in Fig. 64 is very suitable. The funnel is filled with pounded ice, which before powdering has been washed carefully: or snow may, if more convenient, be used. The glass dish catches the water which is formed from the melting of the ice or snow. A hole is made in the pounded ice by thrusting in a pencil or glass tube about the size of the thermometer, and into this hole the thermometer is put and is so supported that the whole of the mercury is surrounded by the ice or snow. The arrangement is left for about ten or fifteen minutes, until it is quite certain that the tube and mercury are at the same temperature as the melting ice. When this is so the tube is raised until the mercury is just above the ice, and a fine scratch is made with a three-cornered file on the tube at the level of the mercury.

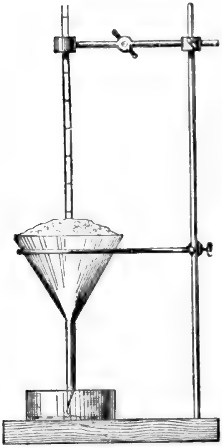

FIG. 64.—Thermometer in ice for the observation of freezing point.

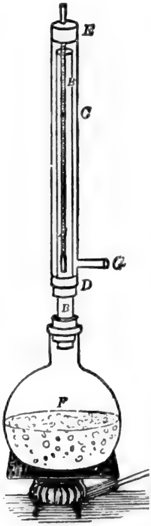

FIG. 65.—Flask fitted for the determination of boiling points.

Marking the boiling point.—On account of the condensation of vapour upon the thermometer, the method used in Expt. 30, iii. to find the boiling point is not very accurate. More exact determinations can be made by means of the apparatus shown in Fig. 65. A can or flask *F* is fitted with a cork, through which a glass or brass tube *B* passes. Surrounding this tube is a wider tube *C*, fitted upon the inner tube by means of a piece of thick india-rubber tubing *D*. At the top of the outer tube is a cork *E* having a hole in which a thermometer can be fitted. When the water in the flask is boiled, steam passes up the inner tube *B*, and down the wide tube *C*, and escapes at the outlet *G* into the open air. To use the apparatus, the top of the stem of the thermometer being graduated is pushed gently into the cork which fits in the outer tube, and adjusted so that the point reached by the mercury at the temperature of boiling water is just above the cork. The cork is then fitted into its place, the water boiled, and when steam has been coming off for about a quarter of an hour, an observation is made of the point reached by the top of the mercury in the tube. The observation is repeated after a few minutes and when two readings obtained at an interval of about ten minutes agree, the point at which the top of the liquid stands is marked upon the stem. The temperature observed is the boiling point of water under the particular conditions existing at the time and place of the experiment.

Precautions necessary in marking the fixed points.—Since experiments have shown that the mixture of common salt with ice or snow lowers the temperature of the ice or snow, great care must be taken that pure clean ice is used when the lower fixed point of a thermometer is being marked. It must also be noted that the presence of substances other than common salt similarly have an effect on the temperature.

The temperature of a boiling solution of common salt in water is higher, too, than that of the steam which is given off from it. Moreover, since it is the temperature of the steam from boiling water which alone remains constant at sea-level (the nature of the containing vessel and the presence of substances in solution affecting the temperature of the liquid), in marking the higher fixed point of a thermometer the instrument should be surrounded by the steam and not placed in the liquid. It will be seen more fully later that an increase of atmospheric pressure, represented by an increased barometric height, raises the temperature at which water boils, hence the height of the barometer must be recorded when the higher fixed point is being marked.

Thermometer scales.—Some value must now be given to the two fixed points which have been obtained as previously described, and of course they could be called anything the maker of the thermometer liked, but for the sake of comparing one man's observations and experiments with those of other people, it is most convenient to graduate all thermometers in the same way. The thermometers used in this country are graduated in two ways—(1) the Centigrade scale, (2) the Fahrenheit scale. A third scale—the Réaumur scale is extensively used in Germany.

The Centigrade scale.—Here the freezing point is called **zero or no degrees Centigrade,** written 0° C. The boiling point is called **one hundred degrees Centigrade,** and is written 100° C. The space between these two limits is divided into 100 parts, and each division called a **degree Centigrade.**

The Fahrenheit scale.—On thermometers marked in this way the freezing point is called **thirty-two degrees Fahrenheit,** written 32° F., and the boiling point **two hundred and twelve degrees Fahrenheit,** written 212° F. The space between the two limits is divided into 180 parts, and each division is called a **degree Fahrenheit.** The reason of this difference is interesting. The physicist Fahrenheit, after whom the thermometer is named, obtained a very low temperature by mixing common salt with the pounded ice when marking the lower fixed point, and he wrongly imagined that he had obtained the lowest temperature which could be reached, and called it zero.

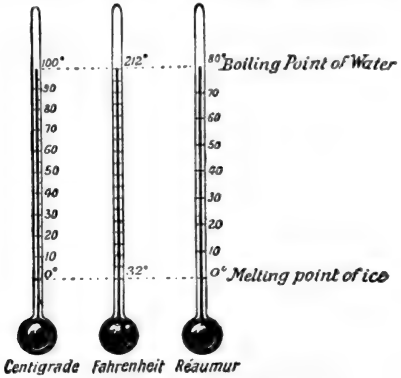

Fig. 66.—Thermometric scales.

Fig. 67.—A clinical thermometer.

The Réaumur scale.—Upon thermometers graduated according to this scale the freezing point is marked 0° and the boiling point 80°. The relation between the three scales is shown in Fig. 66.

The doctor's thermometer.—For the measurement of the temperature of the body, what is termed a **clinical thermometer** is best (Fig. 67). As the temperature of the living human body is never many degrees above or below a temperature of 98° F., a clinical thermometer is only graduated from about 95° F. to 110° F. If the bulb of such a thermometer is put into the mouth, or under the armpit, of a person in health and left there for two or three minutes, it will be found, on taking it out, to indicate a temperature from 97°·8 F. to 98°·6 F. The thread of mercury in the stem of the thermometer remains in one position, though the air is cooling the mercury while the thermometer is being read. This is because of the constriction at the top of the bulb, which causes the thread of mercury in the stem to be left behind while the mercury in the bulb contracts. To "set" the thermometer for a fresh observation, it is only necessary to jerk it slightly, when the thread of mercury will again join up to the liquid in the bulb.

CHAPTER XV.

RELATION BETWEEN THE VOLUME AND TEMPERATURE OF A GIVEN MASS OF GAS.

31. COEFFICIENT OF EXPANSION OF AIR.

i. Coefficient of expansion of a gas.—Obtain a piece of thermometer tubing of about 1 mm. bore and 20 cm. long. Suck into it a length of about 1 cm. of mercury. Seal one end of the tube and arrange that the index of mercury comes near the middle of the tube when the end has been closed and the tube is cool. Fasten the tube to a thermometer, closed end downwards (Fig. 68). There is in the tube a certain volume of air, and its volume at different temperatures can be found as with liquids. Place the combined thermometer and tube in melting ice and notice the position of the air column with reference to the thermometer scale. Repeat the operation for every 10° up to 100° C., taking care that the air column is immersed completely in each case, and giving the tube two or three taps before making an observation, in order to make sure that the mercury is not sticking to the tube. Record the observations thus :

Temperature.	Length of Air column.	Expansion for 10° C.	Average expansion for 1° C.

As the tube is cylindrical and uniform in bore, the volume of the air in it is proportional to the lengths of the air column. The average increase of volume for 1° C., expressed as a fraction of the volume at 0° C., is the *coefficient of expansion*. Find from the results the coefficient of expansion of air.

When a gas is heated in circumstances where, as in these experiments, free expansion is possible, it is said to expand *under a constant pressure.* Both at the beginning of the experiment and after the gas has been heated the pressure to which it is subjected is simply that of the atmosphere.

Measurement of expansion.—While a definite rise of temperature causes most bodies to expand, the amount of the expansion varies within wide limits. In the case of certain special alloys it is almost negligible, while gases expand more than double their volume on being heated from 0° C. to 300° C.

Having now a means at command of measuring temperatures, a more exact comparison of expansions becomes possible. The **coefficient of expansion** is the term used for the rate of expansion when the temperature is increased. In the case of solids we are usually concerned with the **linear** coefficient of expansion, while for liquids and gases the coefficient of **cubical** expansion is of importance.

The fraction of its volume at 0° C., which a gas expands when its temperature is raised through 1° C. is called its coefficient of cubical expansion.

Expansion of gases.—The expansion of gases is very much greater than that of either solids or liquids. Thus, 273 c.c. of dry air at 0° C. become 274 c.c. at 1° C. and 373 c.c. at 100° C. The coefficient of expansion is consequently $\frac{1}{273}$; and this is practically the coefficient for all gases. This law is not, however, strictly obeyed by all gases, though air and some other gases conform to it with great accuracy. The air thermometer is often used as a standard, since the expansion of air with increase of temperature is both regular and large.

The coefficient of cubical expansion of a gas can be determined by enclosing a definite quantity of air in a long narrow tube closed at one end, the gas being separated from the air by a tiny thread of mercury (Fig. 68). The length of the column of air represents the original volume of the gas. The expansion the air undergoes will push the index outwards, and these movements can be observed. If temperature observations are simultaneously taken, the expansion per unit volume per degree of temperature can be obtained.

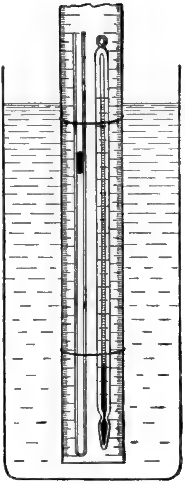
FIG. 68.—Determination of coefficient of expansion of a gas.

Normal temperature and pressure.—It should now be evident that the volume of a given mass of gas depends on the temperature at which it is measured, and also (section 27) on the barometric pressure at the time of measurement. This being the case, certain standard conditions for taking measurements of gases have been decided upon, viz. 0° C. and a pressure equal to that of a column of mercury 760 mm. in height. In practice the quantity of gas under consideration is measured under the conditions existing at the time, and by calculation the volume is found which it would occupy at standard conditions of temperature and pressure. For example, the volume of a quantity of a gas measured when the temperature is 15° C. and the barometer reads 750 mm. at normal temperature and pressure would be calculated as follows:

(273 + 15) c.c. would become 273 c.c. at 0° C. and 750 mm.;
∴ 1 c.c. „ „ $\frac{273}{288}$ c.c. at 0° C. and 750 mm.,
and (section 27) 1 c.c. „ „ $\frac{273}{288} \times \frac{750}{760}$ c.c. at 0° C. and 760 mm.

In a similar manner the volume of any quantity of gas under Normal, or Standard, conditions of Temperature and Pressure (N.T.P.) can be found.

Coefficient of expansion of gases.

	Increase of Volume at Constant Pressure.
Hydrogen,	0·00366
Air,	0·00367
Carbon dioxide,	0·00371

QUESTIONS ON CHAPTERS XIII.-XV.

66. Bodies in general expand when heated. From every-day life give three typical examples of this fact and state one remarkable exception.

67. In what respects does a gas differ from a liquid? Describe how you would show the points of difference.

68. How would you investigate the effect of increase of temperature on the volume of a quantity of water?

69. How is a Centigrade thermometer constructed and graduated?
When the temperature of a room is 15° on a Centigrade thermometer, what would be the reading of a Fahrenheit thermometer? Explain your method of calculation.

70. Having filled a thermometer, how would you proceed to graduate it so as to enable you to use it for measuring temperatures?

71. Define the coefficient of expansion of air, and describe how you would find its value.

72. How would you determine the coefficient of expansion of air?
A vessel contains 500 c.c. of air at 0° C. and a pressure of 77 cm. of mercury.
First, the air is allowed to expand until the pressure is 70 cm., no change of temperature occurring during this expansion, and then the temperature is raised to 30° C., the gas being now allowed to expand in such a way that no further change of pressure takes place. What is the final volume? (The coefficient of expansion of air $= \frac{1}{273}$.)

SECTION II.: CHEMISTRY.

CHAPTER XVI.

32. PHYSICAL AND CHEMICAL CHANGES.

i. Physical changes.—(*a*) Heat a piece of platinum wire in the flame of a Bunsen burner. The wire soon becomes red-hot. Remove the wire from the flame and it quickly resumes its former appearance.

(*b*) Magnetise a sewing-needle. Though the needle now possesses a new property, it is still easily recognised as a needle and is clearly not a new substance.

(*c*) Place a piece of ice in a test-tube, heat it gently and observe the changes. Notice that the solid ice changes, first to liquid water, and on further heating to a gas, steam. By reversing the processes the steam can be reconverted into water, and the water to ice.

ii. Chemical changes.—(*a*) Roll a piece of paper to form a pipe-light. Light the paper and allow it to burn. Observe the changes. The paper disappears and in its place is a black charred residue. During the burning, smoke is formed.

(*b*) Consider the burning of a lump of coal in an ordinary fire. Note the properties of the coal before it is placed on the fire. As burning proceeds smoke and flame are produced. Compare the ash left after the burning is complete with the original coal.

(*c*) Heat one end of a piece of magnesium ribbon in the flame of a Bunsen burner. The metal soon bursts into flame and burns brilliantly. Compare the white ash formed with the original ribbon. There is no doubt that there has been a complete change of substance.

Physical and chemical changes.—All substances are subject to two kinds of change. Hitherto only those changes which influence the properties of substances, leaving their composition unaltered, have been studied. It has been seen that a body, such as a piece of iron, may gradually increase in temperature, changing from cold iron to hot, and, becoming hotter and hotter, may change in colour, passing from a dull grey to red, and from red to almost white, becoming incandescent and emitting light rays. But, if left to itself, the iron will begin to cool, passing through the same changes in the reverse order until it reassumes precisely its former condition.

Or, again, a piece of soft iron having silk-covered copper wire wound round it several times, becomes altered if an electric current is passed through the wire. It is found, on examining the iron, that new properties have been imparted to it, that it is now able to pick up other pieces of iron, or has become magnetised. If the electric current is discontinued, the new power too disappears.

Such changes as these, and those dealt with in Expt. 32 i., where the substance or composition of the body remains unchanged, are known as **physical changes.**

On the other hand, if a piece of iron be left exposed to damp air for some hours it becomes covered with a reddish-brown powder, which a superficial examination will show is a different substance from the iron originally exposed. There is a very large number of changes of the same kind as this continually taking place around us. When a lump of coal burns, it is easy to recognise a great difference between the ash and smoke resulting from the burning and the original coal. So, too, the burning of magnesium ribbon affords an excellent example of the formation of a new substance.

Changes which result in the formation of altogether new substances with entirely new properties are called chemical changes.

It is with the nature and cause of changes of this kind that Chemistry is concerned.

In the remaining pages of this book the student will be concerned with chemical changes and his study of the simple examples here described will serve to introduce him to the important subject of Chemistry.

CHAPTER XVII.

33. OBSERVATION OF THE PROPERTIES OF SUBSTANCES.

BEFORE the changes which a body undergoes when acted on by heat, or by other substances are noted, it is essential to examine it carefully and to make an accurate record of the original properties of the body.

i. Describe the properties of any three very familiar substances, say: glass, salt and mustard.

Description of substances.—Comparing the description of any substance given by different members of a class, it is noticed that one pupil leaves out properties which another gives, and *vice versa.* If all the properties mentioned are written down it is noticed that they can be placed under definite headings, and in order to make certain of an accurate description, in every case the properties decided on must not only be recorded, but recorded in a definite order.

Colour.—One of the first things to notice about any body is its colour. Where possible the exact shade should be given. Some substances such as glass and water can be seen through and they are said to be **transparent.** (Note that although solid glass is colourless, when it is ground into a powder it appears to be white.) Other substances allow light to pass through them although they cannot be seen through and such substances are said to be **translucent.** Things which do not allow light to pass through them are said to be **opaque.**

Taste—Taste is a most important property and many substances can be recognised by it alone. *It should be distinctly understood that in no circumstances should a substance be tasted without the permission of the teacher, as it may be a deadly poison.* Many poisons resemble quite common substances in appearance. In any case only a *very* small quantity of a substance should be taken into the mouth. Substances without taste are said to be tasteless; other taste sensations are sourness or acidity, sweetness, bitterness, burning, etc.

Odour.—A very large number of substances cannot be perceived by the sense of smell and are said to be **odourless.** Other substances have odours which may be described as: sharp, irritating, pleasant or pungent. A substance like ammonia, which has a well-known odour of its own, is said to have a *characteristic odour.*

Form.—Certain substances are found in pieces having a regular shape, which is always the same for the same kind of thing; such pieces are called **crystals**, and the solids are said to be **crystalline.** The term crystalline is also applied to all transparent substances like glass. Things which are not crystalline are called **amorphous**, a word which is made up of two Greek words (*a* not and *morphe* form) meaning without shape. Mustard, flour, soot, are examples of amorphous things.

State.—Solid, liquid or gas. (See Chapter I.)

Density.—The time available will not permit the pupil to find accurately the density of each substance examined, but he should be able to give an idea of its density from observation. For example, he should have no hesitation in saying that a small piece of lead is very dense and that the density of cork is small. It will be enough in most instances to refer to water as a standard.

Solubility.—If salt or sugar is put in a glass of water and stirred, it will be noticed that the salt or sugar disappears and becomes as far as can be seen a part of the liquid itself. We say that the salt or sugar has **dissolved** or is **soluble**, and the liquid is now described as a **solution.** Substances which will not dissolve in water are said to be **insoluble** in it.

As a rule only solubility in water will be tested, but it is well to remember that many substances which are insoluble in water dissolve in other liquids. Beeswax is insoluble in water, but dissolves in turpentine.

Hardness.—If asked to say in what respects iron differs from soap a pupil would probably mention that it was harder than soap.

Now consider exactly what is meant by the property of hardness. A stone is hard, so is a piece of iron, and also a piece of wood, but they are not of the same hardness. Some things are harder than others.

It is often easy to decide which is the harder of two things. For instance, a knife is harder than a piece of wood; for a thumb- or finger-nail can often be dug into the wood, but a finger-nail cannot be dug into a steel knife. Also, wood can be cut with a knife, but not with a piece of india-rubber, because the india-rubber is softer than the wood. All things which a knife will cut or scratch are softer than the knife, and all things which it will not cut or scratch are harder than it.

In the same way, things like potatoes, some woods, chalk, bread, blotting paper, and soap can be scratched by the finger-nail, and are therefore softer than the finger-nail. Things like iron, glass, and flint cannot be scratched or cut by the finger-nail, and are therefore harder than it.

The test of hardness.—It is now possible to understand the way to find out which is the harder of two things. What has to be done is to test which will scratch or cut the other. If asked whether glass or flint is the harder, try whether the flint will scratch the glass. It does. Will the glass scratch the flint ? It will not. Which is the harder then ? The flint, of course.

In the same way, any number of different things may be arranged in the order of their hardness, by taking any one of the substances and finding which of the others it would scratch and which it would not scratch. Then another would be taken, and the same tests made, and so a list like the one below would be made. This is the method always adopted to find out if one thing is harder than another.

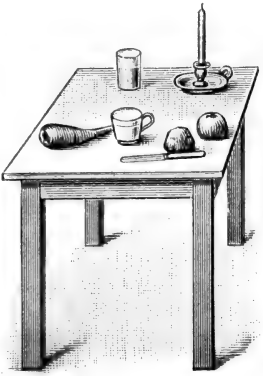

FIG. 69.—The things on the table differ in hardness as well as in other ways.

1. Diamond.
2. Rock-crystal.
3. Glass.
4. Steel.
5. Iron.
6. Copper.
7. Lead.
8. Wax.

The hardest substance is first in the list, the next hardest is second, and the softest is last. Any of the substances will scratch a substance lower in the list, and can be scratched by substances higher in the list. Diamond is seen to be the hardest substance ; it will scratch every other thing. Emery and carborundum are also very hard, and therefore used for polishing many things.

In actual practice in class it will be sufficient to use a penknife, penny, and the finger-nail as standards.

Condition.—Under this heading a number of properties may be included. If a plate of glass is dropped it breaks into fragments and it is said to be **brittle.** A biscuit is easily crumbled or powdered and is said to be **friable.** Solids such as lead which can be beaten out into sheets without breaking are called **malleable.** Copper, gold and platinum are malleable, but gold has more **malleability** than any other solid. Lead can be bent and does not spring back; its shape, after bending or twisting, remains just as it was left; it is on this account said to be **pliable.** Substances like india-rubber which do spring back to their first shape and size after being forced out of it are called **elastic.**

Other properties.—It frequently happens that a body has some special or characteristic property not included under the foregoing headings, *e.g.* salt becomes moist when exposed in damp air. It is most important to note and record all such properties, as they may be a means of distinguishing a body from others of a similar nature.

Litmus.—It has been found that many substances have the power of altering the colour of certain vegetable colouring matters. For example, a variety of cabbage which when freshly cut is a purple colour changes to a reddish colour on adding vinegar to it. This property has been found useful in classifying substances, and even although its significance will not be understood yet, it will be used in the examination of the properties of substances. The colouring matter most commonly used in the laboratory is **litmus,** which is obtained from a certain lichen, and from a kind of spurge common in South Europe. When treated with water a purple solution is obtained. **Litmus paper** is made by soaking unsized paper in the solution and drying. The paper should always be moistened with water before applying as a test.

34. EFFECT OF HEAT ON SUBSTANCES.

Effect of heat on substances.—From the very earliest times it has been customary to try the effect of heat on substances, and no wonder, for of all the experiments performed, perhaps the most interesting effects are obtained by this means. But to obtain the fullest information it is most important that great care should be taken, and the experiment closely watched. If the substance under examination is a solid, a little of it should be placed in a test-tube and heated, using a *very small flame.* Every change should be noted, and the sides of the test-tube examined from time to time, when there is no further change apparent as a result of gentle heating, the heat may be increased and careful observations again made. The properties of a residue left after heating should always be described, and if fumes are given off their action on moist litmus should be noted.

Source of heat.—For the purpose of heating substances it is desirable that we should have a flame which is not smoky. A spirit lamp can be used, but where a gas supply is available it is cheaper and more economical to use a Bunsen burner, which fulfils the above conditions. It will be noticed on reference to Fig. 70 that at the bottom of the tube *A* there is a collar *B*, which can be turned so as to uncover two holes in the tube by which air may be allowed to enter and mix with the gas before ignition. When the collar is in the position described the flame is non-luminous, and if a cold body, such as a porcelain basin, be placed in it, no soot is deposited. But when the collar is rotated so as to close the holes, an ordinary luminous flame is produced, and it is found that a cold body placed in it becomes coated with soot. The real nature of the Bunsen flame will be better understood after the question of combustion has been dealt with.

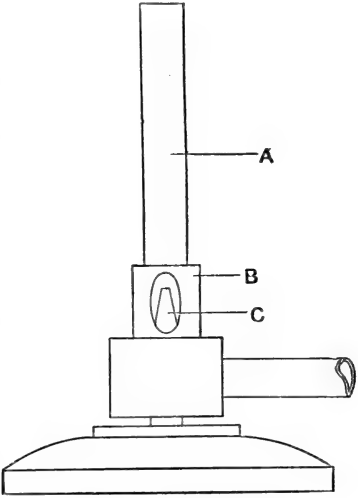

FIG. 70.—A Bunsen burner.

CHAPTER XVIII.

EXAMINATION OF COMMON SUBSTANCES.

35. EXAMINATION OF SAND.

i. Examine the sand supplied as directed in Chapter XVII. Solubility need not be tested. Why ?

ii. Form—Examine a portion of the sand with a lens.

Sand.—Sand is generally a pale yellowish or brownish, tasteless, odourless crystalline powder. From the fact that it is sometimes seen suspended in the surf by the seashore it is concluded that it is not much denser than water, in which it does not dissolve. It scratches the blade of a penknife and is therefore harder than it. On account of this property fine sand is used for polishing metals.

Examination with a lens shows that each little grain is transparent like glass ; the edges of it are smooth, due to rubbing against other particles. The fact is that each grain is really a fragment of quartz or a portion of a ' rock ' from which it has become detached.

When heated in a test-tube sand does not appear to change ; but when heated very strongly with an oxy-hydrogen flame it can be melted, and made into test-tubes, etc., which have an appearance like glass. Quartz vessels made in this way have the advantage over glass that they can be subjected to sudden changes of temperature without breaking.

Certain kinds of sand are used in the manufacture of ordinary glass. Sand is also used along with lime in the making of mortar.

36. EXAMINATION OF CHALK.

i. Examine a specimen as directed in Chapter XVII.

ii. Does chalk dissolve in water ?—When powdered chalk is shaken up with water, a " milky " liquid is produced, due to the fact that fine particles of the chalk are suspended in the liquid. If the test-tube is allowed to stand for a time, the chalk will settle to the bottom, leaving a clear liquid on top.

The suspended chalk can also be removed by pouring the milky liquid on to a piece of porous paper (filter paper) supported in a funnel. Note the method of folding the paper (Fig. 71).

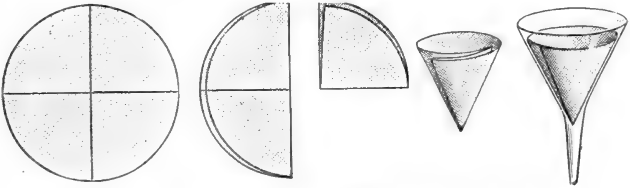

FIG. 71.—How to use a round piece of blotting-paper for filtering.

The method of pouring a liquid into a filter is shown in Fig. 72. (Never attempt to pour a liquid into a funnel without using a rod as shown.) A clear liquid passes through the filter paper, and it is evident that most of the chalk has not dissolved, but it is possible that some of it may have. Pour a little of the clear liquid into an evaporating basin and heat over a Bunsen burner on a wire gauze. The water is driven off as steam, and no powder remains, showing that chalk is insoluble in water.

A more rapid way of evaporating a little of the solution is to dip a clean stirring rod into it and hold the wet rod *above* the flame. If the rod remains clean there is no solid dissolved in the solution.

iii. Structure of chalk.—Shake up a small quantity of powdered chalk with water in a test-tube. Pour off the water containing suspended chalk, and examine a portion of the sediment from the bottom of the tube by placing a few particles on a glass slip, covering it with a suitable "cover slip" and examining under a microscope.

Note. The chalk you have examined is not the substance which is used for writing on the blackboard.

FIG. 72.—How to pour a liquid into a paper filter in a glass funnel.

Chalk.—A white, amorphous solid which is very easily powdered. It does not dissolve in pure water, but it is slightly soluble in rain-water owing to the presence of carbon dioxide in the latter (p. 170), So far as can be seen heat has no effect on it, but this action also will be more fully investigated at a later stage. When first examined chalk appears to be a perfectly uniform (homogeneous) solid, but when examined suitably under a microscope it is found to contain a number of shell like particles (Fig. 73).

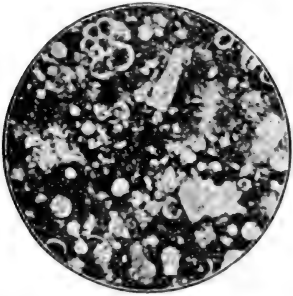

FIG. 73A.—Chalk seen under a microscope.

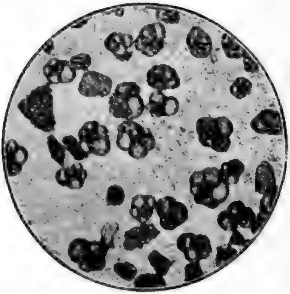

FIG. 73B.—Deposit obtained from the bottom of the Atlantic, and viewed under a microscope.

(From photographs by Mr. J. E. Barnard.)

These particles, as a matter of fact, are the remains of tiny animals such as are still to be found in sea-water. When the animal dies its shell remains on the sea bottom. The shells of multitudes of similar creatures, accumulating on the ocean floor, in the course of many ages, build up a deposit of chalk.

37. EXAMINATION OF LIME.

i. Examine as described in Chapter XVII.

ii. **Slaking of lime.**—Add a few drops of water to quicklime. Note the effect.

iii. **Solubility.**—Examine its solubility as described under chalk (p. 95).

Lime.—Lime is a white amorphous solid and is prepared by strongly heating limestone. When water is added to freshly made lime, **quicklime**, it gets very hot and steam is given off; the lime crumbles into powder and is said to be **slaked**. If left exposed to air, **lime takes up moisture out of the atmosphere** and becomes slaked. On account of this property it is said to be **hygroscopic**. It is slightly soluble in water; the solution, known as **limewater**, turns litmus blue.

Large quantities of lime are used along with sand for making mortar.

Lime is also most valuable in agriculture. Its action in getting rid of "sourness" in soil and making food available for plants will be explained later.

38. EXAMINATION OF ALUM.

i. Examine as described in Chapter XVII.

ii. Liquid formed when alum is heated.—Notice that a vapour is given off which condenses to a clear liquid at the top of the test-tube. In order to identify this liquid collect a quantity of it. Half fill a large test-tube with powdered alum; fit the test-tube with a long right-angled delivery tube, place in a horizontal position in a clamp attached to a retort stand, with the end of the delivery tube dipping into a dry test-tube (Fig. 74). Heat the alum gently and note the condensation of the vapour and collection of liquid. When the alum has become solid, cease heating and examine the liquid. It is clear like water, and has little if any taste. If it really is water it ought to have the same density and boiling and freezing points as water.

The whole class production of liquid should be collected and these determinations made.

Density.—Use the small U-tube method (p. 48).

Boiling point.—Adopt the method described for finding the boiling point of water (p. 78).

Freezing point.—Place some of the liquid in a test-tube with a thermometer and surround it with a mixture of powdered ice and salt. Stir well and notice the temperature at which the mercury remains stationary.

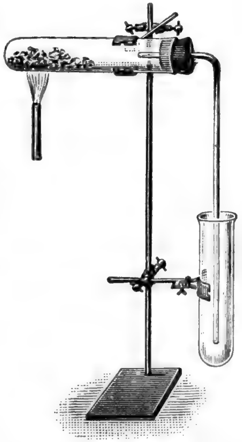

FIG. 74.—The collection of the liquid which is given off when alum is heated.

Alum.—Alum is a transparent crystalline solid, with a characteristic astringent taste. It is soluble in water, the solution turning litmus red. When heated it is noticed that alum melts (*i.e.* changes from the solid to the liquid state) and then froths up; a vapour is formed and drops of liquid are noticed on the sides of the test-tube. When a quantity of this liquid is collected and examined it is found closely to resemble water in appearance and properties, and investigation shows it to have the same density and boiling and freezing points. (Sometimes a little alum makes its way over during the heating process and causes the liquid to have a slight taste and action on litmus.) It is concluded that the liquid is water. This is certainly a fact worthy of notice. The water in the alum must exist there in some peculiar manner, since the ordinary solid form of water—ice—only exists at a low temperature, whereas in alum it exists at ordinary temperatures in a permanent form. When this water is removed from the alum the latter loses its crystalline appearance.

Water of crystallisation is the water contained in certain crystals in solid form. It seems to have something to do with their appearance and shape.

The white solid which remains after the water has been removed is said to be **anhydrous** (Greek—without water).

39. EXAMINATION OF BLUE VITRIOL.

i. Examine as described in Chapter XVII.

ii. Effect of heating.—When heated it is noticed that blue vitriol changes colour and that a liquid is given off. Allow the test-tube to cool, then add a few drops of water to the residue and note the effect.

iii. Estimation of water of crystallisation in blue vitriol.—Grind some blue vitriol in a mortar, and take about two grams of the powder in a weighed crucible with a cover; place the crucible on a pipe-clay triangle on a tripod stand (Fig. 75). Heat gently with a small flame. Note the appearance from time to time, and continue heating till the solid ceases to change colour. Allow to cool and weigh. Re-heat, cool and weigh. The weight should not have changed further; if it has, the process must be continued till a constant weight is obtained. Calculate the loss of weight for 100 grams. Add a little water to the residue. Note the change of temperature.

Fig. 75.—How to heat blue vitriol in a crucible.

Blue vitriol.—Blue vitrol is usually supplied in the form of blue transparent crystals. When powdered the substance is a paler blue. It dissolves in water formimg a blue solution.

When heated the solid becomes white, and a vapour is given off resembling steam. As the addition of a drop of water restores the colour we may assume that blue vitriol contains "water of crystallisation." When a weighed quantity of powdered blue vitriol is heated carefully in a crucible it is found that it loses weight, and when the class results are collected they show that the weight of anhydrous blue vitriol remaining is the same in every case. Assuming that the loss of weight is due to the water given off, it can be stated that blue vitriol contains a definite percentage of water of crystallisation. This is an important fact.

Blue vitriol in solution with lime or soda is used as a fungicide for spraying potato plants and fruit trees. It is also used in electroplating, and in certain electric batteries.

40. EXAMINATION OF SAL AMMONIAC.

i. Examine as described in Chapter XVII.

Sal ammoniac.—Sal ammoniac is a white crystalline powder. It has a taste rather like salt, but not so pleasant. It is very soluble in water. When heated it does not melt, but is converted into a vapour, from which the solid is re-formed higher up on the cooler parts of the test-tube. A solid which on heating changes from a solid to a vapour, without becoming liquid, is said to **sublime.** The process is called **sublimation.**

Sal ammoniac is used for charging Leclanché cells. It is also used as a flux in soldering and in the making of galvanized iron.

41. EXAMINATION OF GREEN VITRIOL.

i. Examine as described in Chapter XVII.

Green vitriol.—Green vitriol is a green crystalline solid. It has a peculiar taste rather like ink (of which it is a constituent). It dissolves in water, forming a pale green solution which soon turns yellowish. On gently heating a crystal of it becomes a whitish colour and gives off water. When strongly heated it turns to a dark red colour and sharp smelling fumes are given off. The action of heat on green vitriol will be studied more fully later, Chap. XXIII.

42. EXAMINATION OF WASHING SODA.

i. Examine as described in Chapter XVII.

ii. Effect of exposure in air.—Wash a large crystal of soda with distilled water, dry with filter paper and leave it on a watch-glass in the open laboratory. Examine next class day.

iii. Water of crystallisation.—Find the percentage of water of crystallisation in soda.

iv. Effect of strong heating—Try the effect of (*a*) strong heating and (*b*) water on the residue from iii.

Washing soda.—Washing soda is a transparent crystalline solid, with a bitter taste. It is very soluble in water, and the solution turns red litmus blue. When left exposed in air it loses its crystalline appearance and becomes coated with a white powder. The same effect can be produced by heating a crystal very gently for a short time, and the cause is the same in each case, viz. the loss of a portion of its water of crystallisation. This property of losing water of crystallisation at ordinary temperatures is called **efflorescence.** Soda melts at a low temperature, and gives off its water of crystallisation, of which it contains nearly 63 per cent. The anhydrous soda melts on very strong heating and is more difficult to dissolve than when crystalline.

FIG. 76.—Fresh crystals of soda. (From a photograph by Mr. H. E. Hadley.)

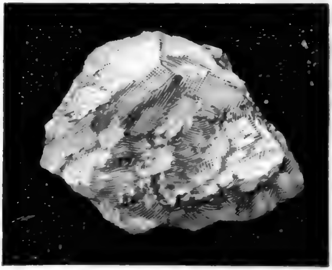

FIG. 77.—Crystals of soda after exposure to the air. (From a photograph by Mr. H. E. Hadley.)

43. EXAMINATION OF NITRE.

i. Examine as described in Chapter XVIII.

ii. Saturated solution of nitre.—Procure a supply of nitre and powder it. Put some of the powdered solid in a test-tube and half fill with water. Shake them up together for some time, and if the powder dissolves, add more and shake again. Continue this addition of the powder and the shaking until some of the powder remains undissolved however much it is shaken. The solution is now *saturated*, that is, it contains as much of the solid as it will hold.

Now warm the cold saturated solution. The powder which before remained on the bottom of the test-tube dissolves. Continue to add more nitre, and notice that a great deal must be added before a hot saturated solution is obtained.

Place the test-tube on one side to cool. As cooling proceeds some of the nitre separates out in clear, well-formed crystals, because as the solution cools it cannot dissolve as much nitre as before.

Nitre.—Nitre is a transparent crystalline solid. If supplied in the form of powder it is found on close examination to consist of small crystals. When heated it melts to a clear liquid; it melts, but does not give off water.

Nitre is used in the preparation of explosives and fireworks. It is also used as a preservative in the pickling of meat.

Saturated solutions.—When any given amount of water has dissolved as much of a solid as it can be made to, without warming, it is said to be **saturated.** But though cold water, for instance, may be saturated with, say, nitre, it can if warmed be made to dissolve more nitre, and the more the solution is heated the greater the quantity of nitre which can be made to dissolve. As cold water will not dissolve so much nitre as hot, it can be seen why a hot saturated solution deposits nitre on cooling.

Most other solids behave in a similar manner to nitre, but there are some exceptions. Lime is less soluble in hot water than in cold.

44. THE SOLUBILITY OF NITRE IN WATER AT DIFFERENT TEMPERATURES.

Solubility may be defined as the mass of a given substance which dissolves in 100 grams of a given liquid at a given temperature.

i. The solubility of nitre in water at a temperature of 60° C.—Heat a beaker of water to about 65° C. and regulate the flame so as to maintain it at this temperature. Take less than 10 c.c. of water in a test-tube, and add *finely powdered* nitre till saturated. Place the test-tube in the beaker of hot water, and continue to add nitre till some of it remains undissolved. To make sure of this it will be necessary momentarily to remove the test-tube from the hot water and shake it, replacing it immediately.

When the saturated solution has been prepared, the flame should be removed and the thermometer placed in the test-tube. When the temperature falls to 60° C. the clear liquid should be decanted into a weighed basin, and covered with a clock glass. Care must be taken not to transfer any of the undissolved solid. The basin, contents and glass should be weighed and then placed on a water bath (Fig. 78), evaporated to complete dryness, cooled and weighed. Re-heat and weigh till a constant weight is obtained. Record your results in the following form :

FIG. 78.—The shallow basin has solution in it and is kept hot by the steam rising from the water in the glass beaker.

Wt. of basin, glass and solution	59·68 gm.	
Wt. of basin and glass	45·79 gm.	
Wt. of solution		13·89 gm.
Wt. of basin and nitre	52·99 gm.	
Wt. of nitre		7·20 gm.
Wt. of water		6·69 gm.

Since 6.69 gm. of water dissolve - - 7·2 gm. nitre.

100 gm. of water will dissolve - - $\frac{7\cdot2 \times 100}{6\cdot69}$ gm.

The solubility of nitre in water at 60° C. is 108 gm. - - = 108 gm.

The solubility at other temperatures should be found by different members of the class, the results collected and a graph obtained. It is customary to mark the horizontal (OY) line in degrees. If time permits a similar set of results should be obtained for common salt, and plotted on the same page.

Solubility.—From experiment it is found that as a general rule the solubility of substances increases with a rise of temperature; but each substance is not affected to the same extent. The solubility curves for nitre and salt show that, while the solubility of nitre increases very rapidly with rise of temperature, as the steepness of the curve shows, the curve of salt is not greatly affected, being nearly a horizontal line.

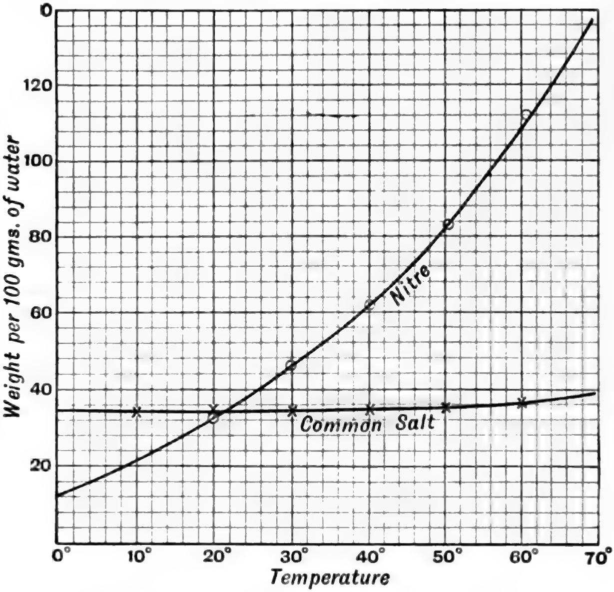

FIG. 79.—Solubility curves of common salt and nitre.

45. CRYSTALLISATION.

i. (*a*) **Effect of cooling a hot saturated solution.**—Make a hot saturated solution of alum, and divide it into two portions. Quickly cool one under a tap, and place the other aside till next class day. Make observations.

(*b*) Repeat the last experiment using (*a*) nitre, (*b*) blue vitriol.

How crystals can be made.—*Warm* water when saturated with any soluble substance often contains more of the solid dissolved than an equal quantity of a *cold* saturated solution. The consequence of this is, that if a warm saturated solution is allowed to get cold, the water can no longer keep all the substance in solution, and it separates out in the solid state, which, in these circumstances, always takes a crystalline character. The crystals of alum, formed in this way, generally have eight sides, or the shape of the crystal is the same as the solid called the **octahedron**, shown in Fig. 80. But in some circumstances the crystals only have six sides, or are cubes.

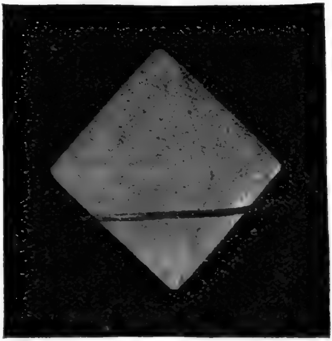

FIG. 80.—An eight-sided crystal of alum. (From a photograph by Mr. H. E. Hadley.)

It is found that when a solution is cooled rapidly, that the crystals resulting are very small, but that when allowed to cool slowly larger crystals are formed. The best crystals are formed when a solution just saturated at the temperature of the air, or very slightly above it, is allowed to evaporate, in a place free from dust. By selecting a small perfect crystal of alum, and suspending it by means of a hair in the solution from which it was obtained (all the other crystals being removed) a crystal similar to that shown in Fig. 80 may be grown. It is important that wide differences of temperature should be avoided during the growth of the crystal.

46. EXAMINATION OF WOOD.

i. Examine as described in Chapter XVII.

ii. Dry distillation of wood.—As a result of heating wood in a test-tube, it is noticed that both a liquid and a gas are given off. Arrangements must be made to collect and examine all the products given off. For this purpose the apparatus shown in Fig. 81 may be used.

If necessary, in order to secure sufficient of the products for examination, the test-tube may be recharged with wood, without interfering with the other apparatus.

Each of the final products must be examined separately.

Solid.—Examine and identify if possible, using any test you think suitable.

Liquid.—Note that there are two liquids, one like tar, and the other lighter and clear in colour. Record the odour, taste and action on litmus. Does the liquid at all resemble any liquid with which you are familiar ?

Gas.—Slip a greased plate under the mouth of the gas jar and remove from the pneumatic trough. Note the odour; add a little lime water and shake thoroughly, and again note odour. Apply a lighted taper.

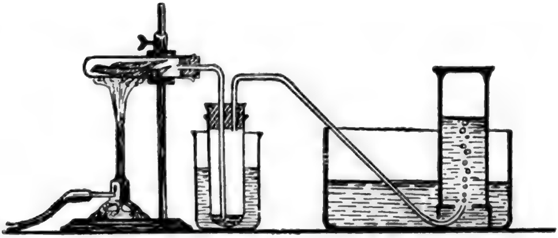

FIG. 81.—Apparatus for dry distillation.

Dry distillation of wood.—The appearance and many of the properties of wood differ according to the variety under examination. When pieces of wood are heated in a closed vessel and the products—solid, liquid, and gaseous—are collected separately, it is found that each kind of wood gives much the same products.

A **black solid** remains in the test-tube or retort, which is easily recognised as *charcoal*, by its appearance and the way it burns in air when heated.

The **liquid** which collects in the cooled receiver contains a little tar and a lighter liquid, which can be shown to be water with something else present, which has a sharp odour and bitter taste and turns litmus red. As a matter of fact it is a mixture of different liquids. One of them—acetic acid—may be recognised from its resemblance to vinegar.

The **gas** is found, after shaking up with lime water, to be colourless and nearly odourless. When a lighted taper is plunged into a jar of the gas it is extinguished, but the gas burns with a blue flame.

47. EXAMINATION OF SULPHUR.

i. Examine as described in Chapter XVII.

ii. Effects of heat upon sulphur.—(*a*) Put some finely powdered sulphur into a large test-tube, using sufficient to fill the tube to a height of about 1½ inches, and heat carefully with a small Bunsen burner flame. Occasionally take the tube out of the flame and shake it. When the sulphur has all melted notice that an amber-coloured liquid has been formed. Pour a little of the liquid into a beaker of water. Observe that a lump of yellow sulphur is again formed, which when broken reveals a crystalline structure.

(*b*) Continue to heat the remainder of the liquid sulphur obtained in (*a*) until the liquid boils. Carefully observe the changes in colour of the liquid. Pour a little of the boiling liquid into cold water. Examine the cooled sulphur; it is *plastic* and not unlike india-rubber.

(*c*) Notice that a yellow deposit has been formed on the cold, upper part of the test-tube in which the sulphur was heated. This is the result of the condensation of sulphur vapour. The deposit is known as *flowers of sulphur.*

iii. Plastic variety of sulphur.—After examining the plastic solid which results from suddenly cooling boiling sulphur, weigh a lump of it and place it on one side for a few days. After this interval examine it again and weigh it. There is no alteration in weight, but the lump has changed again to the ordinary kind of sulphur.

iv. Crystalline varieties of sulphur.—(*a*) Dissolve some powdered roll-sulphur in carbon bisulphide. The solution must be made entirely by shaking, for *the carbon bisulphide must on no account be heated.* When the sulphur has all dissolved, pour the solution into an evaporating basin, and put it into a draught cupboard free from dust to evaporate slowly. Examine the basin after an hour or so. Observe the crystals which have formed. Take the largest and most perfect and sketch its form. This kind of crystalline sulphur is known as the *octahedral variety* (Fig. 82).

(*b*) Place some powdered roll-sulphur in a clean, dry, evaporating basin and heat gently on a piece of iron wire gauze. When it has all melted remove the flame and allow it to cool. As soon as a solid film has formed on the top of the liquid, pierce two holes in it and quickly pour out the remaining liquid sulphur through one of the holes. Remove the film of solid sulphur and examine the yellow needle-shaped crystals on the sides of the basin. This kind of crystalline sulphur is known as the *prismatic variety* (Fig. 83).

Examine the crystals after a few days. Observe they are now opaque. The prismatic sulphur has changed back again to ordinary sulphur.

How sulphur is found in nature.—Sulphur is found naturally both alone—that is, in an uncombined condition—and also united with other substances in the form of chemical compounds. The uncombined or **native sulphur** is rarely pure. It is found most abundantly in the neighbourhood of volcanoes—as, for example, in Sicily—from which island a large part of the **brimstone**, as sulphur is commonly called, of commerce is obtained. Before being placed on the market the sulphur is purified, or separated from the earthy impurities by distillation, in suitable retorts connected with large cool chambers, in which the sulphur vapour, driven off by heating, is condensed. In the first stages of the first process the sulphur condenses as "flowers of sulphur," but later, when the condensing chambers have become heated, as a liquid on the floor. This liquid is cast into the familiar "roll-sulphur" or brimstone which is well known to most people.

In recent years the centre of the world's sulphur supply has been transferred from Sicily to the United States, where large underground deposits are melted by superheated steam and forced to the surface in a liquid state, being afterwards led to large bins where the sulphur is allowed to cool.

Varieties of sulphur.—Sulphur is one of the few simple substances which exist in several forms. When such substances have more than one modification, all of them with the same chemical composition but possessed of different physical properties, such as density, colour, crystalline form, and so on, it constitutes an instance of what is called **allotropy**, and the different varieties of the substance are called **allotropic** forms. Sulphur, oxygen, carbon and phosphorus all have allotropic forms. Sulphur has four allotropic forms, though it is only necessary here to mention three of them. These are **octahedral, prismatic** and **plastic** sulphur. Though the properties of these varieties of sulphur are so different, yet all the varieties are composed entirely of sulphur.

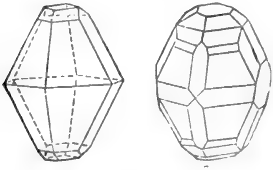

FIG. 82.—Crystals of sulphur.

Octahedral sulphur.—Ordinary roll-sulphur, or brimstone, is composed of tiny crystals of this variety of sulphur compactly massed together. This can be seen by breaking a roll of sulphur in two and examining the broken ends, when crystals will be distinctly visible in the centre of the roll. But much larger crystals are obtained by dissolving powdered roll-sulphur in carbon bisulphide and allowing the solution to evaporate slowly into the air, when fairly large, perfectly formed octahedra of sulphur will be obtained (Fig. 82). This kind of sulphur is the most **stable form**; the other varieties gradually change into **octahedral** sulphur if left to themselves exposed to the air.

Prismatic sulphur.—The second crystalline variety of sulphur is called **prismatic** sulphur. It is obtained in the form of clear, needle-shaped crystals by carefully melting powdered roll-sulphur in and evaporating basin and allowing the liquid obtained to cool slowly. As soon as a film of a solid sulphur has formed on the liquid sulphur, two holes are pierced, and the remaining liquid rapidly poured through one of them. If the film be removed and the inside of the basin examined, a number of clear, needle-shaped crystals of prismatic sulphur are seen (Fig. 83). But when the basin is examined again after an interval of a few days the crystals are no longer clear; they have become opaque owing to the transformation of each needle into a number of minute crystals of octahedral sulphur, which, as has been remarked, is the stable form of sulphur.

Plastic sulphur.—When boiling sulphur, which may be obtained by melting powdered roll-sulphur in a large test-tube, is cooled suddenly by pouring it into cold water, it undergoes a remarkable change. If a piece of the sulphur, solidified in this manner, is taken out of the water and examined, it is found to be like caoutchouc ; it can be pulled about like chewing gum, and it is quite as elastic. This springy material is **plastic** sulphur. But if plastic sulphur be left to itself for a day or two it gradually changes back into octahedral sulphur, another reason for regarding the octahedral as the stable form of the element. In this process of reconversion there is no change of weight.

FIG. 83.—Crystals of prismatic sulphur.

Effects of heat upon sulphur.—Sulphur undergoes a series of changes as it is heated. To follow the changes satisfactorily the heating must be very gradual. When powdered roll-sulphur is heated in a large test-tube it first melts, at about 114° C., into an amber-coloured liquid, which when poured into cold water solidifies into ordinary yellow sulphur. In continuing to heat the melted sulphur above 114° C., however, it gradually gets darker and darker in colour, becoming thicker and thicker in consistency, until at about 250° C. it is so viscid that the tube containing it can be inverted and the liquid will not flow. But if the temperature be still further raised, the thick liquid becomes mobile again, and by and by, at 440° C., it boils, changing into a dark orange-red vapour. The vapour, by sudden cooling, can be changed into a yellow solid, known as "flowers of sulphur." If the boiling sulphur be poured into cold water it is converted into plastic sulphur.

48. EXAMINATION OF COMMON SALT.

i. Examine as described in Chapter XVII.

ii. Make a solution of salt in water, and evaporate it by gently heating in a basin. When all the water has disappeared examine the residue. Note the shape of the small crystals, using a lens.

Properties of common salt.—Common salt crystallises in six-sided solids, or **cubes** (Fig. 84). When the crystallisation is brought about by evaporating a solution of salt, the crystals are very small. Some natural crystals, known as **rock salt**, are, however, of a considerable size. There is no water of crystallisation in the crystals, and consequently when they are heated no steam is given off. The crackling noticed when salt crystals are heated in a tube is spoken of as **decrepitation,** and is due to the breaking up of the crystals into smaller pieces.

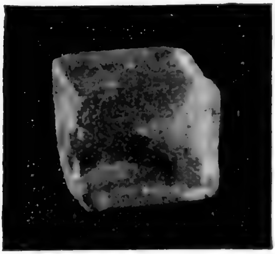

FIG. 84.—A crystal of salt.

When exposed in damp air, salt becomes moist, on which account it is said to be **deliquescent.**

Common salt dissolves in water, and the amount dissolved by cold water is almost as great as by warm water. The solution has no action on litmus papers, and is said to be **neutral.**

Rock salt is found in the earth in some countries as layers of varying thickness. The largest amount is found in the salt mines of Austria, but considerable quantities are found in Carrickfergus, Co. Antrim, and in Cheshire. Common salt is also present in large quantities in sea water, but is mixed with other substances.

Salt is used largely to prevent the decomposition of meat, and enormous quantities are employed in the manufacture of sodium carbonate.

49. EXAMINATION OF LEAD.

i. Describe its physical properties.

ii. Effect of heating lead strongly in air.—Heat a fire-clay crucible to redness, and then drop small pieces of lead foil in it one at a time. After adding the lead, stir with a piece of iron wire and continue heating for at least ten minutes, stirring all the time. Allow to cool, and observe any changes.

Lead.—When freshly scraped with a knife lead has a rather bluish colour. The bright metal, however, soon tarnishes in air. It tarnishes more rapidly in damp air. It is about 11·5 times as dense as water. It is very malleable. On account of the fact that it melts at 326° C. (a low melting point for a metal) it can easily be cast in moulds. When heated in air it becomes covered with a scum—sometimes called "dross"—of a yellow colour (litharge). If the heating is continued long enough all the metal is changed into the powder. If this yellow powder is heated for a long time at the temperature at which lead melts, it slowly changes into a red powder known as **red lead.** It is not an easy matter to carry out the latter operation in an elementary class, but traces of the formation of the red substance may be noticed during the heating of lead.

The metal is made into sheets and pipes, and is extensively employed by the plumber. When melted with other metals it forms **alloys** such as pewter, solder and type metal.

50. EXAMINATION OF COPPER.

i. Describe its physical properties.

ii. Effect of heat on copper.—By means of crucible tongs, hold a piece of copper foil in a Bunsen flame and note the changes. Allow the copper to cool, and examine the surface of the metal.

Copper.—Copper is a reddish coloured metal, with a characteristic lustre. It is hard, but not so hard as iron. It is about nine times as dense as water. It is very malleable and can be rolled into very thin sheets.

When heated in air, a wave of colours spreads over the metal, which after a time becomes covered with a black scale, or powder.

Copper is a good conductor of heat and electricity. Copper wire is largely used for conducting electric current from one place to another.

Brass is an alloy of copper and zinc. **Gun metal** an alloy of copper and tin.

51. EXAMINATION OF TIN.

i. Describe its physical properties.

ii. "Cry of tin."—Bend a piece of tin while holding it near the ear. Notice the crackling sound.

iii. Change of weight in heating.—Weigh a crucible, and place it on a pipe-clay triangle over a Bunsen flame. Heat till the bottom of the crucible is red hot. Weigh about a gram of tin foil (such as is used for wrapping chocolates) and cut into pieces of less than a square cm. Add these one at a time and note that they appear to burn. When the whole lot has been added, heat the crucible strongly for some minutes. Allow to cool and weigh. Note the appearance of the residue.

Tin.—Tin is a brittle white metal, about seven times as dense as water. It is malleable and in sheets is familiar as "silver paper." The noise which is heard when tin is bent is due to its crystalline structure. The individual crystals rubbing against one another during bending cause the characteristic sound known as the "cry of tin."

On account of the fact that it does not tarnish easily, tin is used for coating other metals, such as iron. When heated it increases in weight and changes into a greyish powder.

52. EXAMINATION OF PLATINUM.

i. Describe the physical properties of the piece of platinum foil supplied, but be careful not to injure it in any way, as it is more costly than gold.

ii. **Effect of heat.**—Hold the platinum foil in the Bunsen flame and make observations. Allow it to cool. Has the metal changed in any way ?

Platinum.—Platinum is a white metal rather like silver in appearance. It has a density of 21·5, or nearly twice that of lead. When heated in a Bunsen flame it does not melt, and on cooling is found to be quite unchanged. This property makes it most valuable for making certain laboratory apparatus, such as crucibles which have to withstand high temperatures.

Its rate of expansion when heated is the same as that of glass, and hence it can be fused into glass vessels without fear of their cracking on cooling. As platinum is a good conductor of electricity, this property is used in the making of electric light bulbs.

53. EXAMINATION OF MAGNESIUM.

i. Describe the physical properties of the piece of magnesium ribbon supplied.

ii. **Effect of heat.**—Hold the magnesium ribbon by means of a tongs in the flame. Make observations.

iii. **Change of weight when magnesium is burnt.**—Weigh a crucible and its lid with a piece of magnesium, which, folded lightly, is placed in the crucible. Heat it strongly in a burner, taking care to let no fumes escape (Fig. 85). To do this, keep on the lid, and only raise it a little when the flame is removed. The magnesium is seen to burn brightly in places, but if care is taken, no fumes are lost. When finished, the mass should be in the form of a white powder. Allow to cool, and weigh the crucible with the lid and

FIG. 85.—How to heat magnesium powder or ribbon in a crucible.

powder. Subtract the weight of the crucible and lid to find the weight of the powder. The powder will almost certainly be found to weigh more than the original piece of metal. Calculate the percentage increase in weight.

Magnesium.—Magnesium is a bright white metal, which remains bright in dry air, but soon tarnishes in moist air. Its density is less than twice that of water (light for a metal).

Magnesium burns in air with an intensely white light, and on this account is used in photography as a source of light. When a weighed quantity of it is burnt carefully, so that none of the fumes or powder are lost, it is found that there is a decided increase of weight. When the class results are collected it is found that the increase is the same in every case, allowing for errors in experiment. Careful experiments show that the increase is equal to about 66 per cent. on the weight of the magnesium. This is a noteworthy fact and requires explanation. This process of a metal changing into a powder by heating in air used to be called **calcination**, and the powder was called a **calx**. This name will be adopted until a fuller study of this question has been made.

QUESTIONS ON CHAPTERS XVI.-XVIII.

73. How are chemical changes in matter distinguished from physical changes ?

State, giving your reasons, whether the following changes are chemical or not :

(*a*) When magnesium wire is heated in a Bunsen flame.
(*b*) When platinum wire is similarly treated.
(*c*) When iron rusts.
(*d*) When water boils.

74. You are given a substance to examine. State exactly what physical properties you would look for. Illustrate your answer by giving examples of substances possessing the various properties you mention.

75. You are given some sulphur, chalk and bluestone, and asked to determine :

1st, whether they are soluble in water ;
2nd, whether crystalline or amorphous ;
3rd, the effect of heat on them.

Explain how you would proceed in each case, and, if you have worked such experiments, state the conclusions you have arrived at.

76. Describe all you have observed, and state the nature of the changes that take place when the following are respectively heated in a hard glass tube : (*a*) mercuric oxide, (*b*) green vitriol and (*c*) washing-soda crystals.

77. Explain what is meant by the terms efflorescence, decrepitation, sublimation. Give examples of substances which exhibit these phenomena. For what purposes are the following processes employed: evaporation, distillation, filtration, crystallisation? Give examples of experiments illustrating the use of each process.

78. State which of the following substances are (*a*) crystalline and containing water of crystallisation, (*b*) crystalline without water of crystallisation, (*c*) amorphous, (*d*) soluble in water, (*e*) insoluble in water: green vitriol, red lead, nitre, sulphur, salt, chalk, washing soda, caustic potash, lime and magnesium oxide.

79. How would you obtain pure specimens of sand, salt and sal-ammoniac from a mixture of these three substances?

80. Describe briefly how you would investigate completely the nature of sublimation, with the aid of a suitable substance.

81. Give a short account of what you have learnt on the subject of crystallisation.

82. You are required to draw a curve to show the effect of change of temperature on the solubility of a given substance; state in detail your method of obtaining it.

83. How would you ascertain whether a given solution were saturated? Describe how you would determine the amount of water required to dissolve 100 grams of nitre at 20° C.

84. On evaporating 25 grams of a clear saturated aqueous solution at 15° C., 1·85 grams of residue are obtained. In what terms would you express the solubility of the substance in water?

85. The following table gives the weights of residues obtained by evaporating given quantities of nitre solution saturated at the temperatures indicated:

Weight of Solution.	Temp.	Weight of Residue.
12·43 grams	15° C.	2·42 grams
18·78 ,,	34° C.	6·28 ,,
15·74 ,,	45° C.	6·64 ,,
15·77 ,,	52° C.	7·43 ,,
13·90 ,,	60° C.	7·23 ,,
11·88 ,,	70° C.	6·89 ,,

Draw a graph showing the solubility of nitre, and from it determine the least quantity of water required to dissolve 100 grams of nitre at 57° C.

86. What is a saturated solution? What effect has change of temperature on such a solution of alum? Has the rapidity with which such a change takes place any effect upon the result?

87. Five samples of a certain substance were strongly heated until no further change in weight occurred, with the following results :

	Weight of Substance in grams.	Weight of Residue in grams.
(1)	·500	·305
(2)	·450	·275
(3)	·630	·383
(4)	·560	·374
(5)	·750	·458

What conclusions can you draw from the above data ? Name two substances which might be used in a similar manner to support your main conclusion.

88. What do you mean by the term solubility ? The solubility of a certain salt at various temperatures is given by the following table :

Temperature in ° C.	8	16	24	32	36	40	52	60
Solubility - -	9·0	14·0	23·5	40·5	54·0	50·2	45·5	43·5

If 250 gm. of water saturated with the salt at 55° C. be allowed to cool to 20° C., what weight of salt will crystallise out ?

89. Describe in detail the method of finding the solubility of nitre at 60° C. The following table gives the solubility of nitre in 100 grams of water at various temperatures :

10° C. - - - - -	19·1 grams.
30° C. - - - - -	43·2 ,,
40° C. - - - - -	62·3 ,,
50° C. - - - - -	84·1 ,,
70° C. - - - - -	138·5 ,,

Draw a graph, and from it calculate what weight of solution at 60° C. will contain 5 lb. of nitre.

CHAPTER XIX.

DISTILLATION.

54. DISTILLATION OF WATER AND ALCOHOL.

i. How to obtain water from the air.—Put some pieces of ice into a test-tube, or other glass vessel, which is clean and dry on the outside. In a few minutes the outer surface of the vessel will become covered with moisture.

If ice is not available, put a little ether in a test-tube and make it evaporate quickly by blowing vigorously down a narrow tube on to the surface of the ether. As the air is cooled by the evaporation of the ether, the moisture in the outside air is deposited on the outside of the test-tube in the form of minute drops; in other words, the water vapour in the air is changed to a liquid.

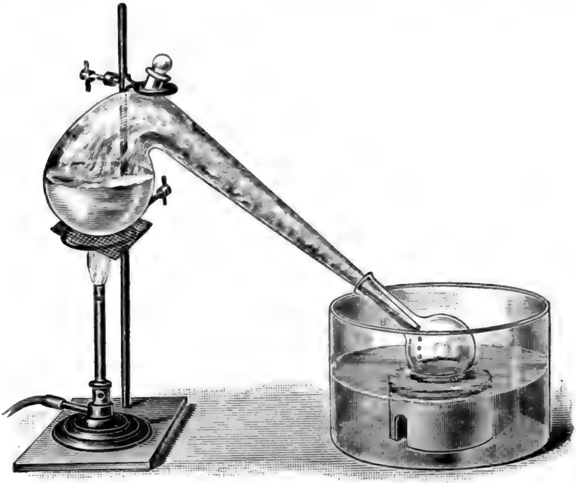

FIG. 86.—As the water boils it is converted into steam, and this steam is condensed again into water in the cooled flask.

ii. The distillation of water.—Obtain a glass retort containing some ordinary tap-water coloured with ink and in which some sand has been placed, and put it on a sand-bath, as shown in Fig. 86. Let the neck of the retort pass into the neck of a flask suitably supported over a basin of water. By means of a Bunsen burner boil the water in the retort, and keep the flask cool by continually pouring water on to it. Notice that the steam which passes over into the flask is condensed again into water, which is quite clear and tasteless.

iii. **The distillation of alcohol.**—Fit up the apparatus as shown in Fig. 88. Note that the lower tube of the condenser is to be attached to the water tap, and the upper to run into the sink. Place a mixture of alcohol and water in the flask. Heat gently, and collect as much liquid as possible that distils at the boiling point of alcohol, 78° C. Examine the resulting liquid. Will it burn ?

Condensation.—By heating a liquid the liquid is changed into vapour. This change may take place slowly and gently, as in evaporation ; or quickly and vigorously, as in boiling. But by whichever process vapour is obtained, it can, by the reverse plan of cooling it, be reconverted into liquid. This change, from the state of vapour back again to the condition of liquid, is called **condensation**, and the vapour is said to have been condensed to a liquid. Water-vapour is invisible, but when it is cooled sufficiently, visible particles of water are formed from it. This can easily be seen by boiling water in a flask fitted as in Fig. 87. Water-vapour is produced inside the flask, but it cannot be seen. When, however, the hot vapour issues from the tube into the colder air outside it is condensed, and particles of water become visible as steam. A cold plate held close to the spout of a kettle from which steam is coming will cool the steam, converting it into water, which will be seen trickling down the plate.

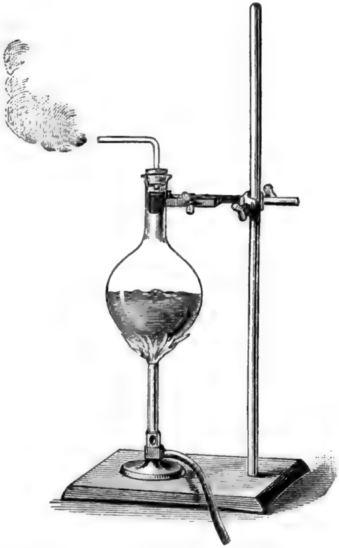

FIG. 87.—The invisible water vapour is condensed into visible steam when it comes in contact with the cooler air.

Common instances of condensation.—Most boys and girls have noticed the condensation of vapour taking place at some time or other on a cold day. For instance, if the doors and windows of a room are kept tightly closed, and there is a good fire burning, the inside of the window panes soon becomes covered with moisture, which, forming drops, trickles down the glass and collects on the window frame as liquid water. The water must evidently come from the air in the room.

What causes condensation?—The air outside a room is in winter much colder than that inside. This cools the glass of the windows very much, and consequently the air next to the cold surface itself becomes cooled and then cannot hold as much vapour as when it is warm, and some of the vapour which can no longer be held by the air is changed into water. So that condensation is caused when air containing water vapour is cooled.

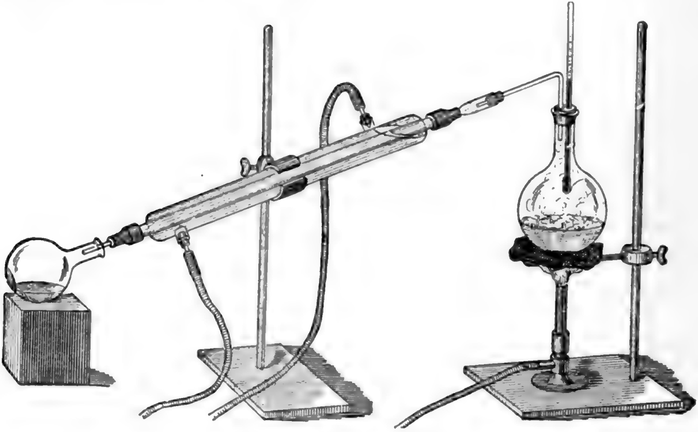

FIG. 88.—Distillation of a mixture of alcohol and water.

Distillation.—The change of liquids into vapour by heating them, and the condensation of the vapour into liquid by cooling it, is employed in an important process called **distillation.** This plan is frequently made use of for purifying water and other liquids. Perhaps the most useful application of distillation is to obtain fresh water for drinking purposes from sea water or other water not good to drink. Large ships, carrying as they often do more than a thousand people, cannot take enough fresh water on board for the needs, throughout a long voyage, of so many persons. Instead of attempting this difficult task, it is the custom to change sea water into fresh water by distillation. There is no difficulty in understanding this process after the way in which it is done on a small scale in laboratories has been described. Ordinary water may be boiled in a flask or retort, the neck of which fits into a flask continually kept cold by pouring water upon it, as in Fig. 86. As the water boils it is converted into steam, and this steam is condensed again into water in the cooled flask. If some sand and salt were first added to the water in the retort they would be left behind, and the water found in the flask would neither taste of the salt nor be coloured by the sand; it would be purified from these by distillation. So that impurities which a liquid contains are left behind in the retort, and the liquid obtained by condensing the vapour is pure.

Liebig's condenser.—Many devices have been invented for the more perfect condensation of the vapour during distillation. Most of them consist of some arrangement for cooling a tube through which the vapour passes, by surrounding it with cold water. This is the principle of the **Liebig's condenser**, Fig. 88. It can be used for the separation of a mixture of alcohol and water, advantage being taken of the difference in the boiling points of the two liquids.

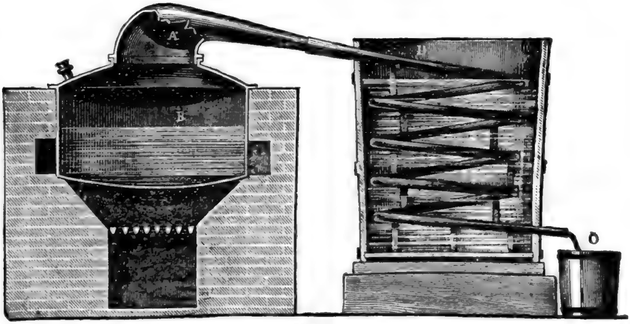

FIG. 89.—A large metal "still." The vapour of the liquid boiling in B passes through the neck A and the spiral tube in D. The spiral is surrounded with cold water, so the vapour in it is condensed, and trickles in drops of liquid into O.

Differences between distilled and ordinary water.—It has already been learnt that the water which is found naturally in rivers, lakes, and especially in the sea, contains a large number of substances in solution which it has at some time in its history dissolved out of the earth. When any natural water is heated in a distilling flask it is only pure water which is converted into vapour; the solids dissolved are left behind. When all the water has been thus changed into vapour the solids are left behind on the inside of the flask in the form of a **residue.** This residue, consisting of substances which were dissolved in water, forms the crust found inside a kettle or boiler which has been used for a long time. The great difference, then, between distilled and ordinary water is that the former contains no dissolved solids. Rain water is really water which has been naturally distilled from seas and other large quantities of water on the earth.

CHAPTER XX.

CONSTITUTION OF AIR.

55. EFFECT OF HEATED METALS ON AIR.

i. Action of heated tin on a confined volume of air.—(*a*) Get a long hard glass combustion tube, closed at one end, and in it place a long strip of tin foil. Close with a rubber stopper, and place in a clamp on a retort stand. *Take care not to close the jaws of the clamp too tightly.* Why? Heat gently at first, but strongly afterwards. Turn the tube from time to time so as to expose a fresh surface of molten tin. Continue heating till no further action is observed. Allow to cool. Remove the cork under water and note the height to which the water rises in the tube. Transfer the remaining gas to a small gas jar. Try the effect of plunging a lighted taper into this gas.

(*b*) Repeat this experiment using " reduced " iron or " reduced " copper.

Air changes in properties when metals are heated in it.—It has been seen that certain metals increase in weight when heated, and change into calces (Section 53). It was assumed that this increase must come from the air. It is now found that when metals are heated in a confined volume of air, that although the air does not undergo any visible change, yet it is changed in a very remarkable manner. First a portion of it has disappeared, as is shown by the water which enters the tube, when the cork is removed under water, and the remaining gas extinguishes a lighted taper. Hence it is clear that the change which takes place in certain metals when heated in air is accompanied by a change in the air itself. It is reasonable to suppose that the gas which disappears is concerned in the formation of the calx of the metal which is heated in it. During calcination the tin takes out of the air that part of it which helps burning, and the portion of the air which has disappeared has combined with the tin, forming a calx. These results may be summarised. **Tin on calcining gains in weight, taking some material from the air. This material is the part of the air which causes substances to burn in it, and might be called the "active part." The remaining and greater portion, which does not assist burning or calcination, might be called the "inactive part."**

56. CHANGES WHICH TAKE PLACE DURING BURNING.

It has been seen that there is a portion of the air, and only a portion, which allows of the burning of a taper in it. It is important that the nature of the changes which take place, both with regard to the substance burnt and the air in which it burns, should be thoroughly understood. For this purpose we must use a substance which is easily ignited. **Phosphorus** is a suitable substance, as it takes fire at a temperature of 30° C. As this is below body heat, it is very dangerous to handle it with the fingers.

CAUTION.—Phosphorus must never be taken in the hand, it must be cut under water and removed by means of a tongs.

i. White powder produced when phosphorus burns.—Carefully dry a wide glass cylinder and a small crucible. Cut off (under water) a piece of phosphorus about half as big as a pea, and dry it between blotting paper or filter paper. Using a pair of tongs, place the phosphorus in the crucible, touch it with a hot wire, and quickly put the cylinder over it, as in Fig. 90.

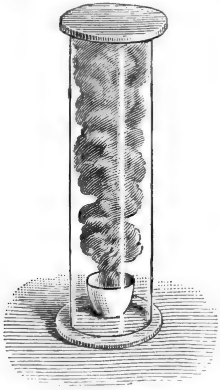

FIG. 90.—Arrangement for collecting the white powder produced by burning phosphorus.

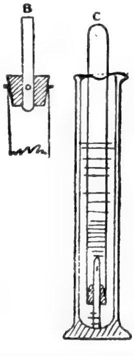

FIG. 91.—Method of opening "side hole" plug, under water.

A white powder is deposited upon the sides of the cylinder. When the phosphorus has ceased to burn, lift up the cylinder and pour a little water into it. The white powder dissolves with a slight hissing noise. Add a little blue litmus to the water; the colour is changed to red.

Note that the powder is deliquescent.

ii. Proportion of air removed by burning.—Take a long test-tube, and in it place a piece of phosphorus which has been carefully dried with filter paper. Close with a one-hole rubber stopper, fitted with a piece of glass tubing sealed at one end and having a small hole blown in one side (Fig. 91).

The small hole is closed by the rubber stopper. Heat the tube gently, and when the phosphorus melts give a sharp flick, so as to cause the molten phosphorus to run down the tube. When cool invert the tube and push it down in a cylinder of water, thereby uncovering the hole in the side of the plug and admitting water which rises in the tube. Raise the latter till the water in it is at the same level as that in the cylinder. Measure the length of the tube and of the water which has entered it.

iii. Volume of air used up when phosphorus burns.—*Demonstration.*[1] Cut a slit in a large cork so that the handle at the top of a crucible lid will fit into it. Place the lid upon the cork and float on water contained in a pneumatic trough. **Dry most carefully** with filter paper a piece of phosphorus about the size of a large pea, and place it on the floating lid. Cover with a bell jar from which the stopper has been removed. Close the jar with a cork fitted with a short right-angled glass tube, sealed at the end, and containing a little sand. Note the height of the water in the bell jar. Heat the end of the tube containing sand, and then give the jar a smart tap on the side opposite to the tube. Hot sand falls on the phosphorus and ignites it. At first the water in the jar falls. Why ? Then it slowly rises. Note the height to which the water rises, and compare it with the amount of air present before the burning of the phosphorus. Note that when the bell jar is removed the phosphorus again takes fire.

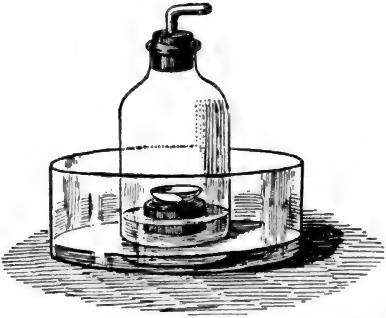

FIG. 29.—Burning of phosphorus in a confined volume of air.

iv. Increase of mass when phosphorus burns.—Procure a hard glass tube, having the shape of BA in Fig. 93, and loosely pack asbestos fibre into its drawn-out end. When this is done determine the mass of the tube by weighing. Place a small piece of dry phosphorus, the mass of which is about one-fifth of a gram, in the tube, and after putting it in weigh the tube and its contents again, so that the exact mass of the phosphorus may be known. Attach the narrow end A of the tube to an aspirator. As the water runs out of the bottle a current of air is drawn through the tube, and, in order to dry the air, it is first made to pass through the test-tube C, containing strong sulphuric acid, to absorb the moisture from the air which passes through the tube. Slightly warm the phosphorus ; it soon catches fire. When this happens, remove the burner and allow the phosphorus to

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[1] To be performed by the teacher.

continue to burn. The fumes which are formed are stopped by the asbestos fibre. A quantity of red deposit is also found (this is really another form of phosphorus), but may be got rid of by strongly heating it.

When the apparatus is cool, disconnect the tube *BA* from the test-tube and aspirator, and again weigh. It will be found that a decided increase of mass has occurred.

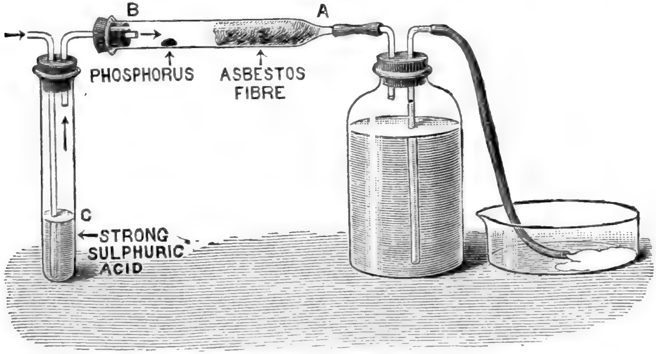

FIG. 93.—Experiment to show that the fumes produced by burning phosphorus in a current of air are heavier than the phosphorus used.

Different kinds of phosphorus.—There are two kinds of phosphorus, one called **yellow** phosphorus, and the other **red**, or **amorphous**, phosphorus. Yellow phosphorus catches on fire very easily; the warmth of the hand is quite enough to inflame it. For these reasons it is always kept under water. It is generally bought in the form of sticks, which, when freshly manufactured, are of a yellow waxy colour. This phosphorus can easily be cut with a knife, but the cutting should always be done under water. However small the piece, it must never be touched with the bare fingers but always lifted by small tongs or forceps. If this precaution is not taken, the warmth of the fingers may cause the phosphorus to catch on fire, and, as it is difficult to shake it off when once alight, the burn which it causes is very severe and dreadfully painful. In all experiments with yellow phosphorus these warnings must be borne in mind. The red or powder form of phosphorus is not so inflammable as the yellow kind, but it must be used with care.

Phosphorus readily burns in the air.—It is only necessary to touch a piece of dry phosphorus with a hot wire to make it catch on fire and burn. It burns with a dazzling bright flame, and at the same time dense clouds of white fumes are formed, which spread throughout the room. These facts are noticed until all the phosphorus has disappeared.

What happens when phosphorus burns in this way ? Is the change anything like that when tin calcines ? Does the phosphorus gain or lose in mass ? These and several other questions present themselves, and they must be answered in this lesson.

Change produced in air by burning phosphorus.—To decide whether phosphorus in burning causes the same change in air as tin does when calcining, it is best to burn some phosphorus in an enclosed amount of air. One way to do this is to place a little phosphorus on a cork or basin which floats on the surface of water, under a bell jar, or a stoppered bottle having no bottom. After the experiment is over and the fumes have disappeared, the water is seen to have risen in the jar, indicating that there is less gas in the jar than before the phosphorus was burnt in it.

From what has been learnt already it can be understood at once that phosphorus in burning takes out the active part of the air and leaves the inactive part behind. So far, then, the changes which occur when phosphorus burns are very like those when tin is calcined. Some differences will be studied a little later.

The fraction of the air which disappears as a result of the burning of the phosphorus in a stoppered jar, can be measured easily enough after the jar has been raised a little, so that its mouth is still under water, but it no longer rests on the bottom of the basin. One-fifth of the air is taken out of it by the phosphorus in burning.

Another way is to burn a little phosphorus in a closed tube, then, by opening the tube under water, the volume of water which takes the place of the air used can be found, and its proportion to the whole volume of the tube can be determined.

That the gas left behind is really the inactive part, which is called **nitrogen**, can be proved by quickly pulling out the stopper of a jar in which phosphorus has been burnt, and introducing a lighted taper. The flame is at once extinguished.

Phosphorus slowly takes out the active part of the air without being lighted.—When a piece of clean phosphorus is exposed to an enclosed quantity of air over water, the rapid changes described in the last paragraph take place slowly. The only difference in the two cases is the rate at which the active part of the air is taken out. Burning phosphorus combines with the active part very quickly; ordinary phosphorus but slowly. Still, given time enough, ordinary phosphorus will remove all the active part of air, and at the end of the experiment it will be found that again one-fifth of the air has disappeared.

Gain of mass when phosphorus burns.—If a piece of phosphorus is put into a closed vessel of air, and the whole is weighed before and after the phosphorus has been burnt, one would expect the mass to be the same in both cases. To begin with, there would be the flask, the air in it and the phosphorus; after the phosphorus had burnt there would be the flask, the white powder and the inactive part of air. As phosphorus uses up the active part of air in burning, and as, after burning, the flask only contains the white powder and the inactive part of air, evidently the white powder weighs more than the phosphorus (see 65 i.).

To determine exactly the increase in mass when phosphorus burns, a piece of phosphorus is weighed and placed in a tube through which a current of dry air is caused to pass. The phosphorus is set alight by warming it, and the white fumes are prevented from passing out of the tube by some asbestos fibre placed in it. If the mass of the tube and asbestos fibre is determined to begin with, and then another weighing is made when the phosphorus has been consumed, it is found that the tube with the asbestos fibre increases in mass, and that the white powder which causes the increase weighs more than the piece of phosphorus used. If the mass of the phosphorus to begin with is one-fifth of a gram, the mass of the powder produced would be nearly one-half a gram.

Properties of the substance formed when phosphorus combines with the active part of the air.—It has as yet only been noticed that the substance which is formed when phosphorus unites or joins with the active part of the air is a white snow-like powder; this corresponds to the calx formed when tin is heated in air. With a little care the quick disappearance of the white material which forms the fumes of burning phosphorus can be prevented. All that need be done is to burn a piece of dry phosphorus in a dry vessel. In these circumstances the white fumes settle down on the inside of the vessel in the form of a snow-like solid. But the white powder has so great an attraction for water that as soon as the vessel is opened it extracts the moisture from the air, and, first becoming moist, is quickly replaced by drops of liquid. If water is put into a test-tube or other vessel in which dry phosphorus has been burned, the white powder rapidly dissolves with a hissing noise, like that noticed when water comes in contact with hot iron. When a piece of blue litmus paper is dipped into a solution of the white powder produced by burning phosphorus it is immediately reddened.

QUESTIONS ON CHAPTERS XIX. AND XX.

90. If you were given water to which sand and nitre had been added, describe how you would proceed, and make sketches of the apparatus you would employ, in order to prepare pure specimens of each.

91. A careful observation of the effect of heat on some substances has led to important evidence as to the nature of their composition. Give two examples to illustrate this, recording the changes that may be observed, and stating any conclusions you would draw from them.

92. What is a crystal ? Describe exactly how you would proceed in order to prepare *large* crystals of copper sulphate from an impure specimen of the salt.

93. Being given a liquid, what experiments would you perform to ascertain whether it contains water, whether it is water, or whether it is a solution of salt in water ?

94. When substances burn they are said to have removed something from the air ; how has this been proved ? Illustrate your answer by reference to burning phosphorus or a burning candle.

95. How would you show that air consists chiefly of two gases, and how would you proceed to determine the proportion in which they occur ?

96. When phosphorus is burned in a closed vessel it uses up only a portion of the air. Why is this ?
Mention the properties of the gas remaining.

97. Describe what you have observed when (*a*) a candle, (*b*) a piece of phosphorus, burns in an enclosed volume of air. Sketch the apparatus used. What are the properties of the gas left in each case ?

CHAPTER XXI.

STUDY OF THE ACTIVE PART OF THE AIR.

57. SEARCH FOR THE ACTIVE PART OF THE AIR.

ALTHOUGH the properties of the active part of the air have been noted, when mixed with four times its volume of the inactive part, the exact properties of the former gas itself are not known. It is, however, safe to predict that substances which burn in ordinary air will burn still better in the unmixed active part. As the calces are formed as a result of the active air becoming united with the metal, it is possible that they may be made to yield it up again, and so provide a supply of the gas for examination.

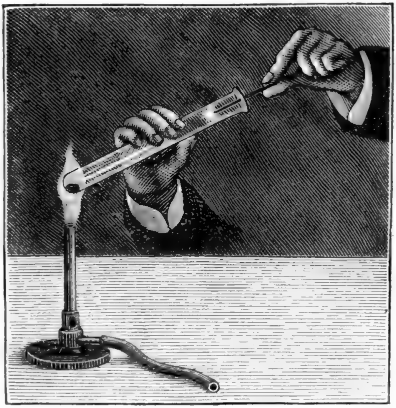

FIG. 94.—Method of testing with a glowing splinter.

i. Gas produced by heating red lead.—Place a little red lead in a hard glass tube, and strongly heat the tube as in Fig. 94. Notice that the red lead undergoes a change of colour. Into the tube insert a glowing splinter. Observe that the splinter is re-kindled. Why is this ?

ii. Change produced by heating mercury calx.—Repeat the preceding experiment with some red oxide of mercury, and notice the formation of the silvery, mirror-like deposit of mercury, or quicksilver, round the cold upper part of the tube. Insert a glowing splinter of wood and watch it re-kindle.

Where to look for the active part of air.—Since, in suitable circumstances, iron, copper and lead take the active part out of the air, and unite with it to form fresh substances, it should not be difficult to make these, or similar substances, give up the part of the air which they take up, and so procure the active constituent of air by itself in a pure form. But a little more thought suggests that probably some of these substances would do better than others. It is quite certain that some are formed more easily than others. Will those which are most easily formed be the best from which to get the active part? No. The reason is this. When a chemical change takes place easily it generally means that the substances taking part in the change have, as it were, a great liking for one another, and when they combine together they form a compound which it is difficult to separate into its parts again. The easiest way to set to work is, therefore, first to find some substance which only combines with the active part slowly and with difficulty, for the compound such a substance forms with the active part will most likely be a weak one, and easily broken up again.

The compound which lead forms with the active part of the air.—When lead is heated in contact with the air, a yellow powder which is much darker in colour when hot is formed. If the heating is continued long enough, all the metal is changed into powder. The change takes place fairly easily, so that from previous reasoning it may be concluded that it is probably difficult to get the active part of the air again from this powder. This is so. But it is found that when some of the yellow powder is heated for a long time at the temperature at which lead melts, it slowly takes up still more of the active part of the air, and changes in colour, becoming red. The first powder obtained, which is yellow, is in some states called **litharge**; the second red powder is known as **red lead.** It is easy to get the second lot of the active part of the air again from the red lead. A third powder of a black colour, and having less of the active part of air than either litharge or red lead, is also formed when lead is heated in air.

How the active part of air is obtained from red lead.—When red lead is heated it changes in colour, and if the heat has not been great it regains its original red colour when allowed to cool. But if strongly heated the red lead gives up some of the active part of the air which it contains, and is reconverted into litharge. The amount of the active part of the air which it thus gives up on being heated is the second quantity referred to in the last paragraph, which is taken up slowly when the heating of lead is continued for a long time. If red lead is strongly heated in a tube, as in Fig. 94, and a glowing splinter of wood is pushed down the tube, the splinter bursts into flame and continues to burn brightly. The active part of air has been obtained alone, and supports burning strongly.

Other ways of obtaining the active part of air.—Quicksilver, or mercury, when strongly heated in the air slowly combines with the active part, and gradually becomes converted into a bright red powder which is the calx of mercury. If some of this calx of mercury is heated in a hard glass tube, as in Fig. 94, it soon changes in colour; and as the heating is continued it is noticed that a mirror-like deposit is formed round the top, cold part of the tube. When this deposit is rubbed with a penholder or pencil, it runs together and forms little drops of quicksilver. Moreover, if a glowing splinter of wood is introduced into the tube it bursts into flame, showing that the active part of the air is being driven out of the red mercury rust. This change is just the reverse of what takes place when mercury itself is heated. The active part of the air, with which hot mercury slowly combines, is driven out of the red mercury rust when that is strongly heated. But red mercury calx is expensive, and it is too costly a plan to heat it to obtain a quantity of the active part of air sufficient for the study of its properties. Several other and cheaper substances easily give up the active part of air when heated.

The active part of the air is called oxygen.—As it will be more convenient in the future to speak of the active part of the air by the name chemists use for it, it may be stated here that it is always called **oxygen**, but the meaning of this name will be better understood later.

58. PREPARATION OF OXYGEN.

i. Oxygen from potassium chlorate.—Place a little potassium chlorate or chlorate of potash (which is the same thing) in a test-tube, and heat it as in Fig. 94. Observe that the powder crackles, melts and gives off a gas. Test by a glowing splinter of wood, and see that the gas behaves like oxygen, the active part of the air.

ii. Preparation of a small quantity of oxygen.—Powder some crystals of potassium chlorate, and mix the powder with a little manganese dioxide (sometimes called pyrolusite). Heat some of the mixture in a test-tube as in the last experiment. Observe by putting in a glowing splinter that oxygen is given off. Notice that in this case there is no melting, and the gas comes off more readily.

iii. Preparation and collection of oxygen.—Into a hard glass tube, closed at one end, fit an india-rubber stopper, with one hole in it, through which a tube, bent as in Fig. 95, is

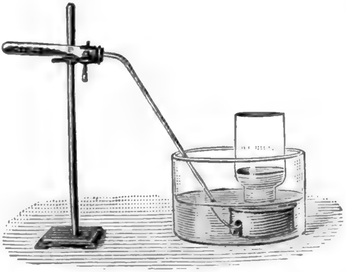

FIG. 95.—Preparation of oxygen by heating a mixture of potassium chlorate and manganese dioxide. The gas is being collected over water.

passed. The other end of this tube, called the delivery tube, dips under water in a trough. Mix together four parts of potassium chlorate and one part of manganese dioxide, as in the previous experiment, and place the mixture in the tube. Support the tube and delivery tube as shown in the illustration. Fill several jars with water, and invert them in the trough. Gently warm the tube, and place one of the jars of water over the end of the delivery tube. As the oxygen is driven off, it displaces the water and gradually fills the bottle. When the jar is full of oxygen, cover its mouth with a greased glass plate, and lift it out of the trough. In this way fill five or six jars with oxygen.

Caution.—*Be careful not to take away the burner from under the hard glass tube before removing the delivery tube from the trough.*

Preparation of oxygen from potassium chlorate.—As the quantity of oxygen obtained by heating red oxide of mercury is comparatively small, and the oxide of mercury is expensive, a more convenient source of the gas is the white crystalline powder called potassium chlorate.

If this white crystalline compound is heated, in the same way as the red oxide of mercury, it melts and gives off bubbles of oxygen, and after all the oxygen has been given off a white substance like table salt is left behind.

By heating, the potassium chlorate is broken up into two things, a gas and a white substance like common salt, which is called *potassium chloride*.

POTASSIUM CHLORATE	gives when heated	POTASSIUM CHLORIDE	and OXYGEN.

Use of oxygen mixture.—By adopting a slightly different method, oxygen can be obtained more readily and easily, for it has been found that by mixing the potassium chlorate with certain other substances, as, for instance, a black compound, manganese dioxide, the oxygen from the chlorate comes off more easily and at a lower temperature. This mixture may be called *oxygen mixture*. If after all the oxygen has been driven out of the mixture the residue left behind is boiled up with water in a flask and the turbid liquid filtered, the manganese dioxide remains unchanged on the filter paper.

59. PROPERTIES OF OXYGEN.

The jars of oxygen prepared as described in Expt. 38 iii. are required.

i. Physical properties of oxygen.—Take one of the jars of oxygen (one of those collected last should be chosen). Notice everything you can about the contents of the jar. The gas in the jar may be slightly cloudy, and have a sharp odour, but if left standing over water for a time, both these disappear. Try the taste by breathing some of the gas. See if the gas has any effect on moist litmus paper.

ii. The burning of a candle in oxygen.—Attach a piece of stout wire to a wax taper and having lighted the taper plunge it into another of the jars of oxygen. Notice that it is *not extinguished,* but continues to burn and with a *larger* and *brighter* flame.

iii. The burning of charcoal in oxygen.—Into another jar of oxygen thrust a splint of wood red-hot at the end, or a piece of red-hot charcoal placed in a deflagrating spoon (a small up-turned iron spoon with a long handle) (Fig. 96). Note the brilliancy of the combustion. Now pour into the jar some clear lime-water (*i.e.* some of the clear liquid which is formed if lime and water be shaken together and allowed to stand. It is really a solution of lime in water). Notice that it turns milky owing to a white powder being diffused through the liquid.

iv. The burning of phosphorus in oxygen.—In another jar burn a piece of phosphorus about half the size of a pea, contained in the deflagrating spoon. Note the great brilliancy of the combustion and the dense white fumes. Add water and shake—the fumes dissolve. Into the solution put a blue litmus paper. Observe that it is turned *red.*

v. The burning of sulphur in oxygen.—Perform, with another jar, a similar experiment with sulphur. There are few fumes, but a strongly smelling gas is obtained, also soluble in water, turning blue litmus red.


FIG. 96.– Jar with deflagrating spoon in it.

vi. The burning of magnesium ribbon in oxygen.—Ignite a small piece of magnesium ribbon and hold it by means of crucible tongs in a jar of oxygen. Notice the white solid formed. Test its solubility in water, and show that, unlike the previous products, it will not turn blue litmus solution red, but will turn red litmus solution blue.

vii. The burning of sodium in oxygen.—Put a small piece of sodium in a dry deflagrating spoon, light the sodium, put it into another jar of oxygen. Observe the fumes formed. Dissolve these in water and try the effect of the solution on litmus paper. It does not turn the blue colour to red, but has exactly the reverse effect, it changes red litmus to blue. Feel the water ; it has a soapy feel.

viii. The burning of iron in oxygen.—Obtain some fine iron wire and twist a few strands of it together. Heat in a Bunsen flame, and while white hot plunge into a jar of the gas. Test the solubility of the solid formed.

Properties of oxygen.—Oxygen gas was discovered on August 1st, 1774, by Dr. Joseph Priestley, who obtained it by heating mercury calx. In accordance with the theory of combustion prevailing at the time he named it "dephlogisticated air," since substances burned better in it than in ordinary air.

When first collected it may have a slight smell like bleaching powder, due to the presence, as an impurity, of a little chlorine, but this odour disappears on standing over water for a time. It can be got rid of more quickly by passing the gas through a solution of caustic potash during the preparation.

The pure gas has no colour, no odour and no taste. It has no action on litmus, and is for this reason said to be neutral. Ordinary combustible substances burn more brightly in oxygen than in the air.

Oxygen has no effect on substances like sulphur and carbon when they are at the same temperature as the room, but if these elements are heated to the point of ignition the oxygen *combines* with them very readily, causing them to burn vigorously.

Some substances which will not burn under ordinary conditions can be made to burn in oxygen, and the case of iron affords a good example of this. Consider what the result would be if there were only oxygen in the air. As soon as iron got red-hot it would start burning. Iron could not in these circumstances be used for grates, furnaces and similar things.

Oxygen is not very soluble in water—one hundred parts of water dissolve three parts of this gas. That the amount of oxygen dissolved by water is very small is seen by the fact that oxygen prepared for experiment is usually collected over water. But though the amount is small it is of great importance in the economy of nature, for it is due to this dissolved oxygen that water animals are able to breathe. Oxygen can, however, be readily dissolved by some liquids, such as a solution of pyrogallol in caustic potash.

Though oxygen exists in a gaseous condition under ordinary conditions of temperature and pressure, yet it can, by lowering the temperature and very much increasing the pressure, be made to assume the liquid, and even the solid, state.

Oxygen is indispensable to life. It is the constituent of the atmosphere which is used up in the processes of combustion, decay and fermentation.

Formation of oxides.—In all the cases of burning studied experimentally new substances have been formed, differing greatly from the original substance and from the oxygen in which it was burned.

To such new substances the name **compound** is given. To the compounds formed by the burning of substances in oxygen the general name **oxide** is given.

Classification of oxides.—At an earlier stage it was mentioned that the action of substances on litmus was of assistance in classifying them. In the case of the oxides tabulated above it is noticed that their action varies. The oxides of magnesium and sodium turn reddened litmus blue, while those of carbon, phosphorus and sulphur turn it red. Iron oxide has no action, but it does not dissolve in water. Notice that the first two substances named are metals. Solutions of the oxides of carbon, phosphorus and sulphur have a sour taste, and it is to this fact that oxygen owes its name, which means acid-producer. It has been found that the following statements are true of other oxides besides those examined.

Metallic oxides when soluble in water turn reddened litmus blue.

Non-metallic oxides when soluble in water turn litmus red.

OXIDES.

Oxide.	State.	Solubility in water.	Action of Litmus.
Carbon	Gas	Slightly soluble	Turns litmus claret colour.
Sulphur	Gas	Very soluble	Turns litmus a bright red.
Phosphorus	Solid	Very soluble	Turns litmus a bright red.
Magnesium	Solid	Slightly soluble	Turns litmus blue.
Sodium	Solid	Very soluble	Turns litmus blue.
Iron	Solid	Insoluble	No effect.

Combustion.—At all times the phenomenon of burning has attracted attention and excited speculation as to its nature. For long the idea was held that something of a material nature—some sort of fiery matter—was the cause; but no one was successful in obtaining a specimen of such a body.

Early in the seventeenth century Jean Rey showed that the increase in weight of metals when they were calcined was due to air. A little later Robert Hooke proved that while charcoal might be heated very strongly without burning if air was excluded, it takes fire if brought into the presence of air at the same temperature.

John Mayow (1674) burnt a candle in a closed vessel over water and noticed the diminution of the volume of the air, and the difference of the character of the remaining air after the candle was extinguished. He concluded that air contains two kinds of particles, one of which is withdrawn when a candle is burnt.

In spite of these discoveries, which were leading up to the truth, another and wrong theory was accepted as true for nearly 100 years. It is difficult to understand why this happened, but it must be remembered that at that time the balance had not come into general use in chemical studies.

The phlogistic theory.—This theory was first stated by Becher about the middle of the seventeenth century, but not much notice was taken of it till it was restated by Stahl some years later. According to it, all combustible substances are composed of two parts; one of these being the same in all of them, and the other different for each. The name phlogiston was given to the common substance, and *burning was supposed to consist in or be caused by the escape of this substance from the body.* Substances which after burning left a bulky ash were supposed to contain very little of the phlogiston, and those which burnt away leaving little or no ash were supposed to be rich in phlogiston. There were many points which were difficult to explain in accordance with this theory. For example; a lighted candle placed under a bell jar is extinguished long before the material of the candle is used up. The explanation offered was that a given volume of air can only take up a limited amount of phlogiston, when this limit is reached the air is said to be phlogisticated, and no further burning can take place in it, hence the candle goes out. Again, certain substances were known to increase in weight when calcined. To this Stahl replied that phlogiston is "a principle of levity" the presence of which in a substance causes it to weigh less than when without it.

The true theory of combustion.—The discovery of oxygen led Lavoisier to put forward the true theory, viz. that **burning consists in the combination of substances with oxygen.** This theory meets all difficulties, but for a time there were many believers in the old theory, and Priestley maintained its truth till his death, although his discovery of oxygen was the cause of its overthrow.

60. THE INACTIVE PART OF AIR.

i. The inactive part of air.—Repeat the experiment 56 i. but as soon as the phosphorus becomes ignited place the jar mouth downwards over water in a pneumatic trough, and satisfy yourself that the gas left behind (*a*) extinguishes a flame; (*b*) has no action on a litmus paper; (*c*) does not turn lime-water milky.

ii. Air can be obtained again by mixing oxygen with the inactive part of air.—Again burn phosphorus in an enclosed amount of air. Place the delivery tube from an oxygen apparatus, similar to that used in the experiment 58 iii., and by heating the hard glass tube bubble oxygen into the jar until it is again full of gas. Cover the mouth of the jar with a ground glass plate, lift out the jar, and test the gas it contains with a burning taper: it behaves just like air.

Nitrogen.—The gas which is left in a jar of air after a metal has been calcined in it, or phosphorus has burnt in it, or in which any one of the instances of burning which have now been studied has occurred, will no longer allow a candle or taper to burn in it. This is one reason why it is called the inactive part of air. The name by which this gas is known to chemists is **nitrogen.**

Properties of nitrogen.—The gas nitrogen is very inert, since it can be made only with difficulty to combine with any other substance. It does not burn, nor will it allow other substances to burn in it; or, as usually expressed, **it does not support combustion.** A mouse dies if put into this gas.

If the negative nature of the properties of nitrogen is borne in mind and is considered in connection with the very active powers of oxygen, it will be understood that its presence in the air serves the purpose of diluting the oxygen, weakening its powers, and making combustion much less intense than it would otherwise be.

If nitrogen obtained from air be heated with either magnesium or lithium, it is found that nearly 1 per cent. of it remains unabsorbed. This residue is another substance present in the atmosphere to the extent indicated, and called **argon.** Argon is also very inert, more so even than nitrogen. Owing to this, even until the year 1894, its presence in the air had been completely overlooked, although, nearly a century before, the eminent chemist Cavendish had unknowingly obtained some, regarding it as an impurity which he had overlooked—an example of the importance of giving attention to the minutest details in scientific investigations. After the discovery of argon, four other new gases—helium, neon, krypton, and xenon—were found to be permanent constituents of air, but they occur in such minute quantities that their presence need only be mentioned here.

Chemical composition of air.—What may be termed the fundamental gases in air are oxygen and nitrogen. Argon and the four other new gases may for our purposes be regarded as part of the nitrogen. Carbon dioxide and water vapour are practically always present, and various other gases or vapours frequently occur in small quantities; but these may be regarded as impurities, and ought not to be considered as constituents of pure air. The following table shows the percentage composition of air as regards volume; that is, for instance, the number of cubic feet of the various gases present in 100 cubic feet of the atmosphere:

Oxygen, a gas which supports combustion,	21·00
Nitrogen, an inert gas,	78·03
Argon, an inert gas,	0·94
Carbon dioxide, a suffocating gas,	0·03
Water vapour,	Variable
Nitric acid, Ammonia, Ozone,	traces

These proportions are remarkably constant in ordinary air, and it is only in localities or under conditions of an exceptional character that they vary to any noteworthy degree. In the air of mines the oxygen has been found as low as 18·6 per cent., but this represents almost the lowest percentage of oxygen ever obtained from a place where human beings could live. In the midst of vegetation, or open ground, especially in the daytime, oxygen is present in the proportion of about 21 per cent., but never more.

Nitrogen not only serves to tone down the activity of oxygen as a supporter of combustion, but it is also useful to the life of plants. A few lowly plants appear to absorb nitrogen direct from the air, but the majority of them obtain it indirectly as the result of the action of bacteria existing in the soil or in their roots.

The proportion of carbon dioxide rarely exceeds 3 parts in 10,000 in pure air, and is not often less than 2·7 parts per 10,000. During the night the proportion is slightly greater than in the day. In the streets of a town the amount of carbon dioxide only exceeds the average amount of the open country by about 1 part in 10,000. In rooms, however, and badly-ventilated places, carbon dioxide is often greatly in excess, and oxygen is present in a much smaller proportion than it ought to be. Carbon dioxide is not essentially a poisonous gas, but it is often found in bad company, and when it occurs in excess the air of which it forms a part is unfit to breathe.

Air always contains a certain proportion of invisible water vapour, and when the air is cooled to a sufficient degree this vapour becomes visible in the form of mist, fog, cloud, rain, or other familiar forms of water. **Ozone** is a peculiar form of oxygen, and is usually present in the air of the open country or over the sea, but not in that of towns.

In addition to these and other gases, numerous minute solid particles are suspended in air, some of them being living germs. They are more abundant in the town than in the country.

61. QUANTITATIVE STUDY OF THE ACTION OF HEAT ON POTASSIUM CHLORATE.

i. Weight of oxygen given off from potassium chlorate when heated.—Fit up a Winchester quart bottle as shown in Fig. 97. Fill the bottle with water, and blow through the short tube, so as to cause water to syphon over through the tube *A* into the cylinder *B*. Fit the test-tube on the cork. If the apparatus is air-tight the water will cease to rise in *B*. If it does not, examine the corks, and detect the leak.

Put a little manganese dioxide in the dry test-tube and a piece of asbestos wool on the balance pan, weigh. Then add about one gram of powdered potassium chlorate, and weigh again. Put the plug of asbestos *loosely* in the top of the test-tube, and connect to the apparatus. Heat *gently* till water ceases to be displaced. Allow the test-tube to cool and then weigh. Also note the volume of water in the cylinder.

From the results calculate the weight and volume of oxygen given off from one gram of potassium chlorate.

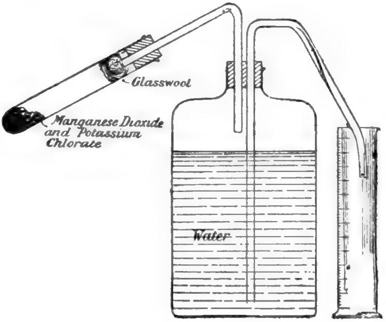

FIG. 97.—Displacement bottle.

Density of oxygen.—By weighing a quantity of potassium chlorate, before and after heating, and measuring the volume of gas given off the weight of a certain volume of oxygen is found, and from this can be calculated the weight of 1000 c.c. Careful experiments show that one litre of oxygen weighs 1·43 gm. at standard temperature and pressure (N.T.P. = 0° C. and 760 mm.). As a litre of air weighs 1·293 gm., oxygen is slightly heavier than air, and therefore nitrogen is slightly less dense than air.

QUESTIONS ON CHAPTER XXI.

98. How is oxygen usually prepared? Make a sketch of the apparatus you would employ, and describe three experiments illustrating its properties you have made with the gas.

99. Indicate four methods of preparing oxygen: point out the advantages and disadvantages of the various methods.

100. Why was oxygen gas so called? Who gave the gas this name and what experiments led him so to name it?

101. You are given some mercuric oxide and a jar filled with nitrogen standing over water; state exactly how you would proceed to make a mixture having the approximate composition of air.

102. Sketch an apparatus suitable for obtaining a number of jars of inactive air. How would you satisfy yourself that the gas obtained by any two selected methods is the same?

103. Chlorate of potash loses 39·15 per cent. of its weight on being heated. What volume of gas would you obtain by heating 3 grams of the chlorate? Sketch the apparatus you would use. No description of the experiment need be given. One litre of oxygen (N.T.P.) weighs 1·43 grams.

104. Define the term Density as applied to gases. If you were asked to find the density of the gas resulting from the action of heat on potassium chlorate, sketch the apparatus you would employ, and state how you would proceed.

105. Contrast the ancient view as to the nature of combustion with that now held. What led to the adoption of the modern view ?

CHAPTER XXII.

ELEMENTS. COMPOUNDS. LAWS.

62. ELEMENTS.

When studying the action of heat on substances, one of three observations is made.

I. Certain substances do not permanently change in appearance or weight, even after prolonged heating, *e.g.* platinum and silver.

II. Certain substances when heated increase in weight and change in properties, *e.g.* magnesium, tin and copper.

III. Other substances change in appearance and decrease in weight, *e.g.* mercury oxide and blue vitriol.

In the **first case** it is evident that the platinum and silver have remained unchanged, and that no new substance has been obtained from them.

In the **second case** an oxide has been formed. This is a new substance which differs greatly from both the metal and the oxygen which go to form it, and is more complex than its constituents.

In the **third case** new substances are obtained from the original substance, each of which may be regarded as being simpler than the substance from which it was obtained. Mercury oxide might be called a complex substance, and mercury and oxygen simple substances; unless, indeed, still more simple bodies are subsequently obtained from them.

A simple substance, or, as it is more commonly called, an element, may be defined as a distinct kind of matter from which no other kind of matter has been obtained except by addition to it.

But it must be borne in mind that as the methods which chemists adopt become more refined it is likely that some of these may be found to be wrongly regarded as elements. Up to the time of Davy (1807) the substances soda, potash and lime were regarded as elements. He found, however, that they could be split up into simpler constituents. From soda he obtained the metal sodium, and oxygen and hydrogen, and from that time, of course, soda could not be regarded as an element. Similarly, if in the future it should be found that any of the forms of matter which we call elements can be split up into simpler bodies with different properties, the element which is thus decomposed will have to be struck off the list. Of the elements known to chemists some exist in the gaseous state under ordinary conditions of temperature and pressure. The commonest of these are hydrogen, nitrogen, oxygen and argon. Two elements, bromine and mercury, are liquids. The rest are solids, and some of the commonest are given in the table, the metals being printed in small capitals.

SOME OF THE COMMONEST ELEMENTS.

Gases

Hydrogen
Nitrogen
Oxygen
Argon

Liquids

Bromine
MERCURY

Solids

ALUMINIUM
ANTIMONY
BARIUM
CALCIUM
Carbon
COPPER
GOLD
IRON
LEAD
MAGNESIUM
Phosphorus
POTASSIUM
SILVER
SODIUM
Sulphur
TIN
ZINC

63. COMPOUNDS.

1. **Mixture and compound.**—Take about one gram of powdered sulphur and mix thoroughly with the same quantity of iron filings. Divide the mixture into two portions. Place one portion in a crucible covered with a lid and heat till all the sulphur appears to be burned off. Now examine both portions. (*a*) Appearance. (*b*) Try with a magnet. (*c*) Test the solubility in carbon disulphide. (*d*) Action of hydrochloric acid.

What is a complex substance?—A definition for a simple substance has been arrived at, but when it is attempted to state clearly what is meant by a complex substance difficulties arise, owing to the fact that the same two elements may give substances differing widely in properties. It is evident, for example, that mercury oxide is not a **mixture** of mercury and oxygen, although it is known to contain both. Again, if equal weights of iron filings and sulphur are mixed as thoroughly as possible the resulting mixture has the properties which would have been anticipated. In colour it looks not so dark as iron and not so light as sulphur. It is possible to remove the iron from the mixture by means of a magnet, or the sulphur may be removed by solution in carbon disulphide, leaving the iron in the same condition as before mixing. When a portion of the mixture of iron and sulphur is heated in a covered crucible, it is found that a new substance is obtained which possesses neither the properties of iron nor of sulphur. It is not attracted by a magnet and it does not burn when heated. Nor can the iron or sulphur be recovered from it by simple means. A distinction must be made between such a *mixture* and the *new* substance which results on heating it.

A complex substance, or compound, may be defined as a new substance formed by the union of two or more elements, and from each of which it differs more or less completely in properties.

Compounds are homogeneous, *i.e.* each particle of them, no matter how small, is the same in composition and properties as every other particle. Another property of compounds is dealt with in Section 64.

64. LAW OF CONSTANT COMPOSITION.

In the course of experimental work, in which the substances used have been weighed, it has been noticed that frequently each member of the class obtained the same result or a result differing very little from the mean class result. This is evidently more than a coincidence. The following results were obtained :

Washing soda contains 63 per cent. water and 37 per cent. anhydrous soda.

Blue vitriol contains 36 per cent. water and 54 per cent. anhydrous vitriol.

Magnesium oxide contains 60 per cent. metal and 40 per cent. oxygen.

More accurate work than can be carried out in class has shown that these and other pure substances always contain their constituents in exactly the same definite proportions by weight. When facts so invariable as these are observed they are stated as a **law**. It is important to bear in mind that a Law in Science is simply a statement made as a result of observations, which have never been found to be contradicted, and it is believed to be true for other substances although they may not have been fully examined.

Law of definite proportions, or constant composition. "The proportion by weight of each element in a compound is fixed and constant."

65. LAW OF INDESTRUCTIBILITY OF MATTER.

i. The burning of phosphorus in an enclosed volume of air.—Take a 500 c.c. round-bottomed flask and in it place a small piece of phosphorus which has been carefully dried on a piece of filter paper. Close the flask with a tightly fitting rubber stopper. Warm the flask very gently, avoiding the spot where the phosphorus lies. As soon as the phosphorus is ignited discontinue heating. When the flask is cool, weigh it. Remove the stopper and weigh again.

Indestructibility of matter.—Although in many experiments there appears at first sight to be a loss of the materials employed; it is found that when care is taken to collect and preserve all the products that the total mass of the substances used remains unaltered throughout an experiment. When a small piece of phosphorus is burned in a closed flask, it is found the mass of the flask and contents is the same before and after the experiment. When air is admitted to the flask after the burning there is a decided increase in mass. During the process of burning a portion of the air is used up in forming phosphorus oxide, but the amount of mass lost by the air is exactly counterbalanced by the gain in mass of the phosphorus (Section 56, iv.). In the same manner, in every case it can be shown than an apparent loss of mass in one substance is compensated for by an increase in mass of something else.

This fact is generally stated as follows: **Matter cannot be created or destroyed, although it may be changed from one form to another**; or, from the chemical point of view, it may be put in a slightly different way although the meaning is the same. **The quantity of matter which takes part in a chemical reaction remains unaltered throughout.** This statement is known as the "Law of the Indestructibility of matter."

QUESTIONS ON CHAPTER XXII.

106. What do you understand by an element? Water was formerly regarded as an element—on what grounds? What experimental evidence can you quote to show that this view was incorrect? Classify the following substances into elements and compounds: lime, copper, mercury, litharge, salt, nitre.

107. Tabulate the properties of alum, sulphur and iron rust according to the following scheme:

(*a*) Whether element or compound.
(*b*) Effect of water.
(*d*) Possibility of obtaining the crystalline forms.
(*d*) Effect of strongly heating.

108. What do you understand by a law? Illustrate your answer by examples in both Physics and Chemistry.

109. The older chemists asserted that there were only four elements, viz. earth, air, fire and water. How would you proceed to show that the above are not elements in the chemical sense of the word?

110. What constitutes a true chemical compound ? Give examples.

111. What important Law of Chemistry enables you to predict the exact amount of oxygen you should get from, say, a gram of chlorate of potash, or the amount of carbon dioxide from a gram of chalk ?

Give an account of any experiments and class results which led to this generalisation.

112. It is stated that when a chemical change takes place, matter is neither created nor destroyed. Explain this statement. How can you reconcile it with the results obtained when copper and chalk are strongly heated in open crucibles ?

113. What do you understand by the Indestructibility of Matter ? How do you reconcile your views with the effects observed when (*a*) magnesium, (*b*) chalk, is strongly heated ?

CHAPTER XXIII.

ACIDS.

66. SULPHURIC ACID.

i. Action of heat on green vitriol.—Fit a hard glass test-tube with a long right-angled delivery tube. Half fill the test-tube with powdered green vitriol. Fix in a horizontal position in a clamp attached to a retort stand. Let the end of the delivery tube lead into a test-tube.

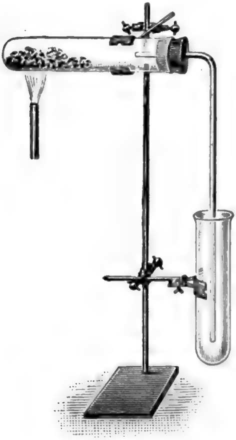

FIG. 97*a*.—Heating green vitriol.

Gentle heating. Heat very gently till no further change is observed. It will be found that the green vitriol turns a dirty white colour, and that a clear liquid collects in the receiving test-tube. Remove and identify the liquid.

Strong heating. Put a second test-tube under the delivery tube and continue heating at as high a temperature as possible. It is noticed that white fumes with a choking odour are given off, and that they condense to an oily-looking liquid in the receiving tube. Examine the solid remaining in the tube. Is it soluble in water?

Sulphuric acid.—When green vitriol is heated it is found that at first a colourless liquid is obtained in the receiving vessel, and if a sufficient quantity of this liquid is collected it is found to have the same boiling point and the same density as water, and to have all the other properties of water, which it is. On strongly heating, fumes are given off which have a very sharp odour and condense to form an oily liquid, which on account of its appearance and origin used to be known as "oil of vitriol." It is better known as **sulphuric acid.** The acid of commerce is not prepared in this way, as it would be too expensive. It is got from sulphur dioxide. Very large quantities of sulphuric acid are used in manufacturing processes. The chief use of sulphuric acid in Ireland is for the manufacture of artificial fertilizers.

67. PROPERTIES OF SULPHURIC ACID.

i. Density.—Take a test-tube about one-third full of distilled water, and carefully trickle about 1 c.c. of strong sulphuric acid down the side. Note the effect. Does the acid sink ?

ii. Solubility.—Shake up the liquid in the last test. Does the acid dissolve in water ? Feel the outside of the test-tube. Is there any change in temperature ?

iii Taste.—Take about half a cubic centimetre of the dilute solution prepared in the last experiment and further dilute it. Observe the taste.

iv. Action of dilute sulphuric acid on substances.—Try the effect of the dilute acid on the following substances and record your observations: (*a*) litmus, (*b*) chalk, (*c*) lime, (*d*) washing soda, (*e*) and the metals zinc, iron, copper, magnesium. If a gas is given off, test the gas with moist litmus and apply a flame to it.

v. Action of strong sulphuric acid on substances.—Try the effect of strong sulphuric acid on each of the following : (*a*) wood, (*b*) salt, (*c*) nitre, (*d*) sal ammoniac, (*e*) copper. If no action is observed, heat gently and carefully.

Properties of sulphuric acid.—Familiarly known as **oil of vitriol,** sulphuric acid is one of the most important and useful of chemical compounds. It is a heavy, oily liquid, which, when strongly heated, boils at 335° C., and gives off a quantity of choking, pungent white fumes. It mixes with water in all proportions, and produces during the solution so much heat that the temperature may rise above 100° C. —the boiling point of pure water—so that care has to be taken when sulphuric acid and water are mixed. Sulphuric acid absorbs moisture very readily, and, owing to this affinity for water, it chars organic matter, such as sugar and wood. When the strong acid is brought into contact with sugar it abstracts water from the sugar, and a black mass of carbon remains. Like most other strong acids, it burns the skin and destroys cloth, so that care must be always taken in using it.

This property of taking up water so greedily makes sulphuric acid a valuable agent for drying gases. The dilute acid has a sour taste and when added to washing soda effervescence takes place and a gas is given off. The dilute acid dissolves the metals zinc, iron and magnesium, and at the same time an inflammable gas is given off. No action is observed with copper. The strong acid acts on salt and on sal ammoniac. In each case clouds of fumes are given off which immediately turn litmus paper pink. Copper dissolves in boiling sulphuric acid, giving off a gas with an odour of burning sulphur.

68. HYDROCHLORIC ACID.

i. Preparation of hydrochloric acid gas.—Fit up the apparatus shown in Fig. 98. Remove the india-rubber stopper of the flask and place in it a small quantity of rock salt in small pieces, or some thoroughly dried common salt. Pour some strong sulphuric acid into the wash-bottle shown in the middle of the illustration. Re-insert the india-rubber stopper into the flask and pour down the acid funnel enough of a mixture [1] of strong sulphuric acid and water to cover the salt in the flask. Gently warm the flask. Collect jars of the gas which is evolved (after it has bubbled through the strong sulphuric acid in the wash-bottle and so become freed from water vapour) in gas jars by downward displacement in the way the illustration makes clear. When each gas cylinder is full, which you can tell by holding a blue litmus paper just below the top of the outside of the cylinder until it is turned red, cover it with a ground-glass plate, with the ground side underneath. As the cylinders are filled set them on one side for examination as presently described. Collect four jars of gas in this way.

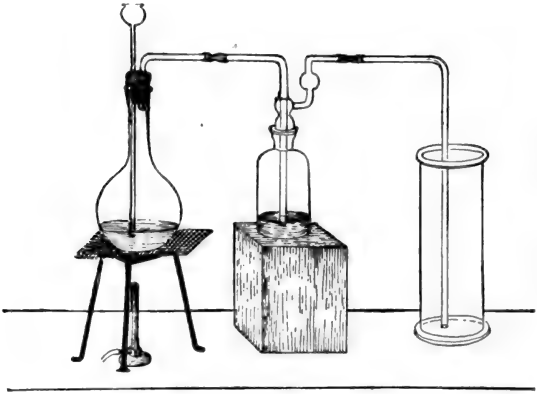

FIG. 98.—Preparation and collection of hydrochloric acid gas.

ii. Properties of hydrochloric acid gas.—(*a*) Raise the glass plate from the first jar and plunge a lighted taper into the gas. The flame is extinguished and the gas does not burn. Quickly replace the glass plate.

(*b*) Into the same jar drop a piece of moistened blue litmus paper and replace the glass plate. The paper is turned red, showing the gas has acid properties.

(*c*) Observe the fumes which the gas forms with the air when the glass plate is removed from a cylinder full of the gas. This is due to the very strong power of absorbing moisture possessed by hydrochloric acid gas.

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[1] One part of acid and one of water are convenient proportions. Be careful gradually to pour the acid into the water and not the water into the acid when mixing them, keeping the mixture well stirred throughout the process.

(*d*) Firmly pressing the glass plate, invert a cylinder full of the gas and place it upside down in a basin of water. Remove the plate when the mouth of the jar is under water, and notice that water rushes up and completely fills the jar. If the water does not completely fill the jar, it shows that the air in the jar was not altogether displaced by the gas when you should have filled it.

(*e*) Place a piece of blue litmus paper in an empty gas jar, and tilt a jar of the gas over it. Note the effect.

iv. To make a solution of hydrochloric acid.—Modify the apparatus in Fig. 98 by removing the drying bottle, and attach the delivery tube direct to the flask. To the end of the delivery tube attach, by means of a short piece of rubber tubing, a funnel. Place this dipping into distilled water in a beaker, as shown in Fig. 99. Care must be taken that the funnel is so arranged that, if the water is sucked back, it will not be able to fill it.

If there is still enough salt and sulphuric acid in the flask, again warm it gently and allow the evolved gas to bubble into the water. Notice that it is completely dissolved. The solution of hydrochloric acid gas formed in this way is the "hydrochloric acid" of commerce.

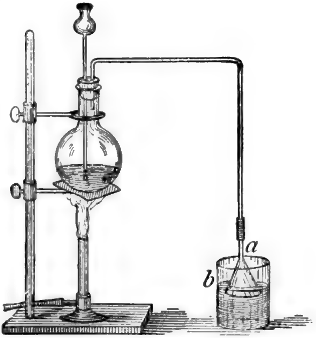

FIG. 99.—Making a solution of hydrochloric acid.

Hydrochloric acid gas.—When common salt is heated with strong sulphuric acid a gas is given off which forms steamy fumes in the air. The gas readily dissolves in water, and the solution constitutes the hydrochloric acid of commerce. The solution was, because of its preparation from salt, originally known as "Spirits of Salt." Another common name given to it is "Muriatic Acid." The "salt gas" itself is called hydrochloric acid gas. It is colourless, will not allow things to burn in it, nor will it burn itself. As is seen by its action on blue litmus paper, it is strongly acid. It is heavier than air, and can consequently be collected by downward displacement.

69. PROPERTIES OF HYDROCHLORIC ACID SOLUTION.

i. Taste.—Make a very dilute solution of hydrochloric acid and taste it.

ii. Effect of heat on the solution.—Take about one c.c. of concentrated hydrochloric acid in a test-tube, and boil it. Try to expel the gas. Do you succeed? Repeat, using a very dilute solution.

iii. Action of the solution on substances.—Try the action of the solution on each of the following: (*a*) litmus, (*b*) chalk, (*c*) lime, (*d*) wood, and on the metals zinc, iron, copper and magnesium.

Properties of hydrochloric acid solution.—When a strong solution of hydrochloric acid gas is heated, the gas is at first given off, but it is found that further heating does not drive off the whole of the acid, which continues to be given off till the last drop of liquid has evaporated. On the other hand, a very dilute solution, on being heated, becomes more concentrated up to a certain point, beyond which no further concentration can be effected.

The solution has a sour taste. It causes effervescence with washing soda, a gas being given off. Chalk dissolves in it with effervescence, leaving a clear solution. It has no action on copper, but gives off an inflammable gas with zinc, iron and magnesium, which dissolve in it.

70. NITRIC ACID.

i. **Preparation of nitric acid.**—Into a stoppered retort, such as is used in the distillation of water, place 30 or 40 grams of small crystals of potassium nitrate (also known as *nitre*). Using a funnel, carefully introduce enough strong sulphuric acid to cover the nitre. Replace the stopper. Place the retort on a stand as shown in Fig. 86, and insert its neck in that of a flask which is continually kept cool by water, just as in the distillation of water. Gently heat the retort. Brown fumes are given off in abundance, and soon drops of a light yellow liquid are seen to fall into the receiving flask. When enough nitric acid has distilled over, remove the laboratory burner, and while the materials in the retort are still liquid, pour them, after removing the stopper, from the retort into an evaporating dish.

ii. Repeat the experiments performed with sulphuric acid, using nitric acid instead.

Nitric acid.—One of the most important compounds with which chemists are familiar is an acid which has long been known under the name of **aqua fortis.** It can be prepared by synthesis from its constituents. It is generally prepared by distilling a nitrate with strong sulphuric acid. Either potassium nitrate, which is more familiarly known as "saltpetre" or "nitre," or sodium nitrate, also called "Chili saltpetre," is generally employed. The latter salt is the cheaper, and in addition yields a larger amount of nitric acid for a given expenditure of sulphuric acid, so that it is more commonly employed than ordinary saltpetre in the manufacture of nitric acid.

Pure nitric acid is colourless. When heated it gives off brown fumes. This happens slowly when the pure acid is placed in sunlight, a fact which accounts for the brown fumes above the liquid in bottles of nitric acid which have been kept for some time.

When diluted it has a sour taste, and like the other acids examined, it causes washing soda to effervesce, giving off a gas, and it turns litmus red. Its action on chalk and lime is similar to that with hydrochloric acid. Its action on most metals is violent, a red gas being given off. With copper a blue solution is left, or green if excess of acid is used. It is to be noted that none of the metals caused an inflammable gas to be given off with nitric acid.

QUESTIONS ON CHAPTER XXIII.

114. What are the chief characteristics of acids ? Describe how you would prepare nitric acid.

115. Being given an acid, which is either hydrochloric, nitric or sulphuric, how would you determine which it is ?

116. Tabulate the effects of acting on copper and zinc with the three common laboratory acids, dilute and strong (with heat if no effect is produced in the cold). Tabulate also the properties of not more than two of the gaseous products.

117. Sketch the apparatus you would employ if you wished to fill a flask with dry hydrochloric acid gas, and describe briefly how you would conduct the experiment. How would you show experimentally that the gas contains hydrogen ?

118. Weighed quantities of water and concentrated sulphuric acid are placed, side by side, in an enclosed volume of air and allowed to remain there for several days. What changes would you find on examination at the end of that period ? What conclusions would you draw from the results observed ? Mention some other substances which would afford similar results.

119. Set down in tabular form observations you have made in the laboratory which would enable you to distinguish between the metals lead, zinc, copper, tin, iron, by using dilute and concentrated hydrochloric acid, cold and hot.

120. Describe the action of dilute and strong nitric acid on copper, zinc and iron. Tabulate your results.

CHAPTER XXIV.

ALKALIES.

71. CAUSTIC SODA.

DURING the previous work a number of substances have been found which turn reddened litmus blue, *e.g.* solutions of washing soda, lime, magnesium oxide, sodium oxide. Some of these must be examined more fully.

i. **Sodium oxide.**—Burn a piece of sodium metal in a jar of oxygen (p. 129). Shake up with a little water to dissolve the sodium oxide formed. Note the soapy feeling of the solution. Try its action on litmus. Evaporate the solution to dryness.

ii. **Sodium and water.**—Into a basin containing distilled water put, one at a time, a few pieces of sodium about the size of a small pea. Note the action, soapy feeling and effect on litmus. Evaporate, just to dryness, but no further. The white solid obtained is called caustic soda.

iii. **Effect of air on caustic soda.**—Leave a stick of caustic soda exposed in the laboratory till the end of the lesson and make observations.

iv. **Effect of caustic soda on other substances.**—Examine the effect of a solution of caustic soda on each of the following: (*a*) a solution of green vitriol, (*b*) a solution of blue vitriol, (*c*) a piece of aluminium metal.

Caustic soda.—When a small piece of the metal sodium is added to water in a basin it remains on the surface and races about over it till it is entirely dissolved. The solution formed feels soapy when rubbed between the fingers, and turns red litmus blue. On evaporation a white solid is obtained, which is caustic soda or sodium hydroxide. Caustic soda is commonly supplied in the form of white sticks. It is very deliquescent and easily soluble in water, giving out heat. It has a corrosive action on most organic substances. It can be made to combine with certain oils and fats, forming **soap.** When a solution of caustic soda is allowed to act on aluminium an inflammable gas is given off.

72. CAUSTIC POTASH.

When the metal potassium is added to water, an action similar to that with sodium is noticed; but in this case the metal takes fire and burns with a violet coloured flame. The resulting solution is indistinguishable from caustic soda by its action on other substances, they are so much alike. The white solid obtained on evaporation is called **caustic potash** or **potassium hydroxide**. The compounds may be distinguished from one another by the colour which they impart to a Bunsen flame. Potash gives the flame a violet colour which is visible through a piece of blue glass. Soda colours the flame a bright yellow, but the colour disappears when viewed through blue glass.

73. CALCIUM HYDROXIDE.

i. Calcium oxide.—Take a piece of calcium metal about the size of a large pea, and hold it by means of a piece of stout copper wire rolled round it. Place some clean dry sand in the bottom of a gas jar and fill it with oxygen by displacement of the air. Heat the calcium to redness and quickly plunge it into the gas. It burns brilliantly and destroys the piece of wire rolled round it. When cool, shake up the contents of the jars with distilled water, and filter.

ii. Effect of calcium hydroxide on substances.—Examine the action of the solution obtained in i. on (*a*) a solution of copper sulphate, (*b*) a solution of green vitriol, (*c*) a solution of washing soda, (*d*) litmus.

Calcium hydroxide.—Calcium is a silver white metal, which tarnishes when exposed in air, but not nearly so quickly as sodium. It burns brilliantly in oxygen, giving out great heat. The calcium oxide formed is a white solid, which gets hot when water is added to it, forming calcium hydroxide. It is slightly soluble in cold water and less soluble in hot. The solution turns litmus blue. It has not a soapy feel, but its action on most substances resembles that of caustic soda.

74. AMMONIA.

i. Preparation of ammonia.—(*a*) Place some ammonia solution in a flask and boil, allowing the gas that is evolved to pass over quick-lime or solid caustic potash in order to dry it. Collect several jars of the gas as shown in Fig. 100. Observe the gas has the same smell as the liquid and acts similarly on red litmus.

ii. Properties of ammonia.—(*a*) In one jar place a lighted taper.

(*b*) Place a second jar in water and note the rapid absorption of the gas and the rise of the water in the jar.

(*c*) Dip a rod in hydrochloric acid and hold it over a jar of the gas. Observe the formation of white fumes.

(*d*) Shake up a jar of the gas with a little water. Examine the solution. See that it is identical with the "ammonia liquid," from which the gas was obtained, and that it loses its odour on boiling, the gas being evolved.

(*e*) Heat an ammonium salt with caustic soda or lime, or merely well mix the two in a mortar with the addition of a little water. Observe by the odour that ammonia is produced.

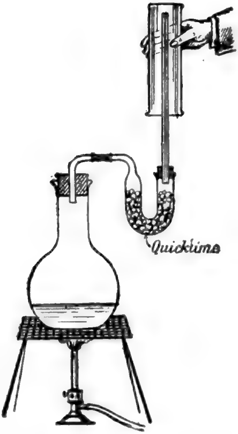

FIG. 100.—Preparation of ammonia.

Properties of ammonia.—The liquid commonly known in the laboratory as ammonia has a sharp pungent smell and colours red litmus blue. If this liquid be heated, a gas is given off which may be dried by passing over quick-lime, and collected by upward displacement, as illustrated in Fig. 100. It is found that the "ammonia liquid" is merely a solution of a gas, viz. ammonia, which is very soluble in water, turns red litmus blue, does not support combustion and is apparently not combustible. In an atmosphere of oxygen, however, the gas readily burns, forming water.

CHAPTER XXV.

HYDROGEN.

DURING the study of the properties of acids, it was found that when certain metals acted on dilute sulphuric and hydrochloric acids an inflammable gas was produced. For a reason which will presently appear, this gas is called **hydrogen.**

75. PREPARATION AND PROPERTIES OF HYDROGEN.

i. Preparation of hydrogen.—Select a flask and fit it up as is shown in Fig. 101. Be very careful that the stopper and the tubes respectively fit very closely. Into the flask put enough granulated zinc to cover the bottom.

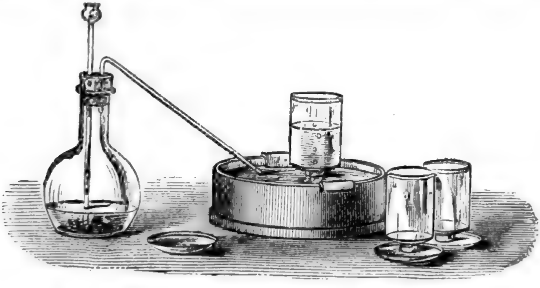

FIG. 101.—Zinc turns hydrogen out of dilute sulphuric acid.

Pour some water upon the zinc. Arrange the delivery tube in the trough as when making oxygen. Pour a little sulphuric acid down the thistle-headed acid funnel, and be quite sure that the end of the funnel dips beneath the liquid in the flask. Do not collect jars of the gas until you are sure pure hydrogen is being given off, which you can find out in this way. Fill a test-tube with water and invert it over the end of the delivery tube. When it is full of gas, still holding it upside down, take it to a flame (which should not be near the flask in use; notice that there is a slight explosion. Continue this until the hydrogen burns quietly down the test-tube. When this happens proceed to fill one or two jars. When the jars have been filled, it is better not to remove them from the water until you want to use them. Collect also a soda-water bottle half full of the gas.

Caution.—*Be careful not to bring a light near the thistle funnel or tube delivering the gas, even when the action in the flask seems to have ceased, or a dangerous explosion may occur.*

Be careful also that none of the acid used gets upon your fingers or clothing.

ii. The liquid left in the flask.—Filter off the liquid in the flask from the undissolved zinc (sufficient zinc should be used to leave a quantity still undissolved : if all has disappeared add more and wait till the action ceases. Partially evaporate the liquid and allow it to crystallise. A quantity of clear colourless crystals is formed. Examine the crystals and sketch the most perfect. Heat some of the crystals in a tube and observe that they melt, give off water (which can be collected and proved to be water), and leave a white powder.

iii. Hydrogen burns but extinguishes a flame.—Test one jar of the gas by means of a lighted match or taper. Observe that the gas burns at the mouth of the jar and that the taper is extinguished when thrust into the tube ; on being taken out, the taper again becomes alight on passing through the flame of the burning hydrogen (Fig. 102).

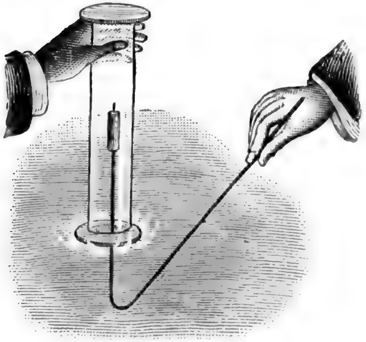

FIG. 102.—The hydrogen burns at the mouth of the jar, but the candle is extinguished when inside the jar of hydrogen.

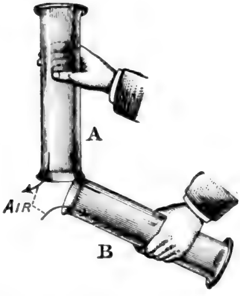

FIG. 103.—Hydrogen is lighter than air and can be poured upwards.

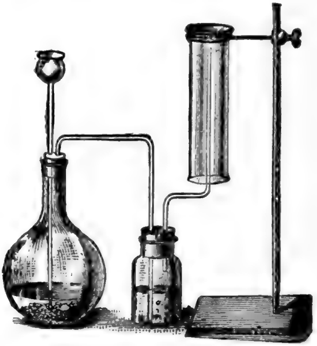

FIG. 104.—Hydrogen being lighter than air can be collected by upward displacement.

iv. Hydrogen is lighter than air.—Take a full jar of the gas and hold it mouth upwards below a second smaller jar held mouth downwards, as shown in Fig. 103. On testing with a lighted taper observe that the gas has left the lower jar and filled the upper. Many experiments, as the filling of balloons or soap bubbles, may also be performed to demonstrate the extremely low density of hydrogen.

v. Hydrogen forms an explosive mixture with air.—Wrap your hand in a duster and with it hold the soda-water bottle (Expt. 75 i.). Take it out of the water so that the water runs out, and the bottle is now filled with a mixture of hydrogen and air. Apply a light and you will not fail to observe that an explosion results.

vi. The flame of burning hydrogen.—Fit a right-angled tube drawn out to a point to a hydrogen generator, as shown in Fig. 105. After pouring a little sulphuric acid down the thistle funnel, collect a test-tube of the gas issuing from the straight tube, and hold the mouth of the test-tube near a flame, *which must be a few feet away from the generator.* The gas will at first go off with a pop, or burn with a squeaking noise, but after two or three trials it will burn quietly, with a blue flame. *When you can carry this flame of burning hydrogen to the apparatus from which the gas is being produced, do so, and use it to ignite the gas escaping from the pointed tube.* If you remember always to do this, there can be no danger, for when you are able to carry a flame of hydrogen in a test-tube of the gas, for a distance of two or three feet to the generating apparatus, you may be sure that the hydrogen issuing from the apparatus is not mixed with air.

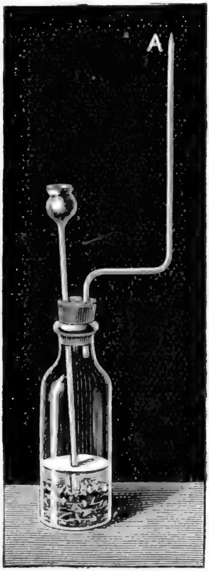

FIG. 105.—Arrangement for obtaining a flame of hydrogen.

Light the hydrogen in this way.

Observe that it burns with a pale blue flame, which after a time becomes yellow. This coloration is due to the glass becoming hot, and some of the substances in it being burnt in the flame of hydrogen.

vii. Hydrogen can be collected by upward displacement.—Substitute for the delivery tube in Expt. 75 i. a tube bent twice at right angles, as shown in Fig. 104, and arrange a jar on a retort-stand in an inverted position, Place a test-tube over the upright tube, and allow it to stay there for a minute, and test the gas as described in Expt. 75 i. When it is pure, substitute the inverted jar, and after a few minutes lift it off the stand and apply a light, first taking the precaution to wrap a duster round the jar and to hold it away from your face. The jar will be found to contain hydrogen.

Preparation of hydrogen.—The most convenient method of preparing hydrogen in the laboratory is by the action of zinc on dilute sulphuric acid (one part acid to four parts water). The method is described in the experimental work. Owing to the slight solubility of hydrogen in water, it can be collected in the same way as oxygen over the pneumatic trough. If, when the chemical action in the flask has completely stopped, the liquid is filtered from the still undissolved zinc, as previously explained, and then partially evaporated in a basin and afterwards allowed to crystallise, a quantity of clear colourless crystals is formed. These crystals melt if heated in a tube, give off water, and leave a white powder. They consist of a compound formed from the zinc and part of the sulphuric acid and known as **zinc sulphate** (Fig. 106).

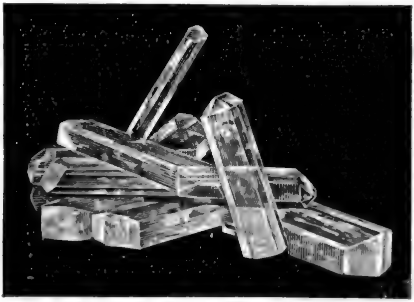

FIG. 106.—Crystals of zinc sulphate.

We may therefore state that *sulphuric acid and zinc form hydrogen and zinc sulphate.* Or, the same fact may be expressed in another way:

Sulphuric Acid when acted upon with Zinc gives Zinc Sulphate and Hydrogen.

Properties of hydrogen.—Having now a means of obtaining hydrogen in considerable quantity, it can be observed that it is a colourless, odourless gas, considerably lighter than air; it burns, but does not support combustion, and it forms a highly explosive mixture when mixed with air. It is now necessary to obtain and examine the compound which is produced by the burning of hydrogen—that is, the oxide of hydrogen.

76. EXAMINATION OF THE LIQUID WHICH IS FORMED WHEN HYDROGEN BURNS.

i. Liquid formed by burning hydrogen.—(*a*) Arrange a flask as before for the production of hydrogen. Pass the gas through a tube containing chloride of calcium in order to dry it thoroughly. Allow it to burn under a retort which is kept cool by a stream of water flowing in at the tubule and out at the end of the neck (Fig. 107). Observe the formation, on the outside of the retort, of a clear liquid which collects and drops into a beaker placed to receive it. By this means sufficient of the liquid can be obtained to identify it, especially if several students add together the liquids formed in their experiments.

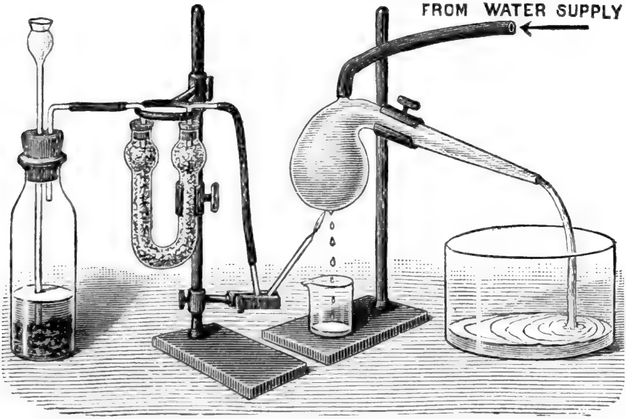

FIG. 107.—The water formed when hydrogen burns in the air can be collected and examined.

(*b*) Take the density, freezing point (a mixture of sodium sulphate and hydrochloric acid forms a convenient freezing mixture), and boiling point of the liquid formed by burning hydrogen. You will find these are 1, 0° C., and 100° C. respectively, and these results are sufficient to enable it to be stated that the liquid is identical with pure water.

ii. Analysis of water.—This may be done by means of an electric battery for generating the electric current, and a *voltameter*. The latter is most simply made by closing the bottom of a funnel by means of a tightly fitting cork through which pass two platinum wires with small plates of platinum attached to the ends remaining in the funnel (Fig. 108). Over

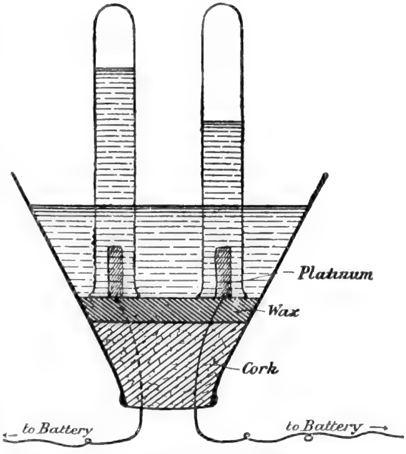

FIG. 108.—A voltameter in which water can be analysed by the electric current.

these plates are supported two glass test-tubes of equal capacity, and the tubes and part of the funnel are filled with water to which has been added a little sulphuric acid, as otherwise the liquid offers great resistance to the passage of the electric current. The wires from a battery of three or four cells are connected with the ends of the platinum wires, and as soon as the connection with the battery is complete, provided there is clean metal at every junction, bubbles of gas are seen to rise from each platinum plate, and to ascend into the tube and displace the contained liquid. After the experiment has gone on for half an hour, the gases may be tested and their volumes measured. It will be found that the volume of one gas is double that of the other, and that the gas of which there is the larger amount is *hydrogen*, whilst the other is *oxygen*.

Formation of water by burning hydrogen.—When a jet of burning hydrogen is brought into contact with a cold surface, such as a cold glass, the product of combustion, the oxide of hydrogen, is condensed. If after a sufficient quantity of the liquid has been collected it is examined, it is found (*a*) to have a density of 1, (*b*) to freeze at 0° C., and (*c*) to boil at 100° C. These are the physical characteristics of water and of no other substance, so we are justified in stating that this liquid, formed when hydrogen burns, is water.

From these experiments the evidence is conclusive, that hydrogen and oxygen are the sole constituents of pure water, and therefore water is an oxide of hydrogen. The reason for giving the inflammable gas the name hydrogen, which comes from Greek words meaning " water producer," is now apparent.

77. COMPOSITION OF WATER BY WEIGHT.

i. Action of hydrogen on heated copper oxide.—Arrange an apparatus like that shown in Fig. 109, in which a flask *A* for the making of hydrogen is connected with the bottle *B* containing strong sulphuric acid. The passage of the hydrogen through the strong acid completely dries the gas.

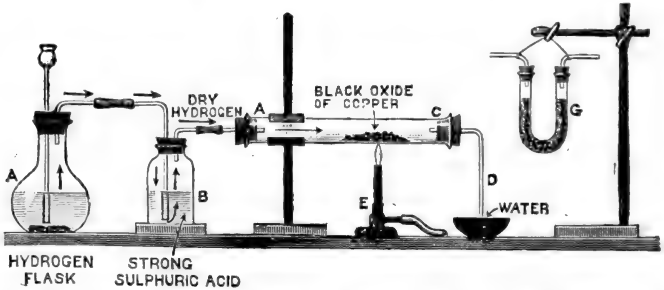

FIG. 109.—When hydrogen is passed over heated oxide of copper it extracts the oxygen, with which it forms water, and leaves copper behind.

A small amount of the black oxide of copper is placed in the hard glass tube *AC*, which is about 1·5 centimetres in diameter, and fitted with the tube *D* in the manner shown. When you are sure that pure dry hydrogen is escaping from the open end of *D*, heat the oxide of copper in the tube by means of the burner *E*. In a few minutes moisture will be seen to collect in the tube *D*, and presently to drop into the dish put to collect it.

Examine the residue in the hard glass tube, and note its change to a reddish colour; this is due to the presence of copper.

Rearrange the apparatus, using the **U**-tube *G* in the place of the tube *D*. The **U**-tube *G* containing lumps of calcium chloride is employed to collect the water formed.

ii. Composition of water.—Put some copper oxide into the tube *AC*, and carefully weigh the tube. Similarly the weight of the **U**-tube *G* and its contents must also be determined. As in the last experiment, when you are sure pure dry hydrogen is escaping from the open end of *G*, heat the copper oxide by means of the laboratory burner. Be sure that all the water formed is collected by the **U**-tube. If any condenses at the end of the hard glass tube *AC*, drive it over by heating the tube at this place.

Allow the tube *AC* to cool. Disconnect it at *A* and *C*, and again determine its weight and that of its contents. Notice the diminution in weight. Also weigh the **U**-tube again, and observe its increases in weight.

Relative weights of the constituents of water.—To find the composition of water by weight—that is, the weights of oxygen and hydrogen which combine to form water—it should be noticed that only the weights of two out of the three substances concerned are required, *i.e.* the weights of hydrogen and water (or of oxygen and water) are known, the weight of the oxygen (or hydrogen) is readily calculated. The experiment is done by finding the weights of the oxygen and water, and for this it is best to use, not oxygen itself, but some oxide which readily gives up its oxygen to the hydrogen, so that by weighing the oxide before and after the experiment we can ascertain the weight of oxygen which it has lost. The oxide used for this purpose is usually oxide of copper, a black powder, which it has been seen is obtained when air is passed over red-hot copper. Pure dry hydrogen is passed over the heated oxide, and it combines with the oxygen of the oxide to form water and leaves the copper behind.

By weighing the water produced and subtracting from it the weight of the oxygen used, the weight of the hydrogen can be found. If the experiment is carefully performed it is found that **water is formed of eight-ninths its weight of oxygen with one-ninth its weight of hydrogen.**

This experiment, also taken in conjunction with other experiments upon the volumes of the gases, proves further that any volume of oxygen is 16 times heavier than the same volume of hydrogen.

QUESTIONS ON CHAPTER XXV.

121. Describe carefully, illustrating your answer with a sketch, how you would prepare hydrogen. What precautions would you consider necessary to obtain a jar of it in a pure state ?

122. How would you obtain hydrogen (*a*) from water, (*b*) from dilute acid ? How can it be shown that, when this gas burns, water is formed ? State clearly the properties that distinguish water from other liquids.

123. Sketch an apparatus by means of which you could burn a jet of dry hydrogen. State how you would proceed to collect the product of combustion with as little loss as possible and what experiments you would make to test the purity of this product.

124. Make a careful sketch of the apparatus you would use to show that when hydrogen burns in air a liquid is formed. How is this liquid identified ?

125. On placing a flask containing cold water over a gas or spirit flame a dew-like deposit is formed ; how would you proceed to collect sufficient of the liquid for the purposes of examination, and what experiments would you perform in order to identify this liquid ?

126. What is the effect of passing dry hydrogen over heated oxide of copper ? Sketch the apparatus you would employ if asked to make the experiment.

If water contains 88·89 per cent. of oxygen, what weight and volume of hydrogen would be required to produce a litre of water ?

CHAPTER XXVI.

SALTS.

78. PREPARATION OF SALTS.

It has been found that certain substances—acids—turn litmus red while others turn reddened litmus blue. It will be interesting to investigate the effect of adding these two classes of substances to one another in the presence of litmus.

i. Preparation of salts.—(*a*) *Sodium chloride.*—Make a solution of sodium hydrate, as in Expt. 71 ii., or by dissolving caustic soda in water. To a portion of the solution in an evaporating basin add dilute hydrochloric acid, drop by drop, until the solution has no effect upon either a red or blue litmus paper. The solution is then said to be *neutral.* Gently evaporate the solution on a sand-bath until a dry white residue is left. Then, by tasting the solid, satisfy yourself that it is common salt, or sodium chloride.

(*b*) *Magnesium sulphate.*—Take about 10 c.c. of dilute sulphuric acid in a basin, and warm gently. Then add, in the presence of litmus, small quantities of magnesium oxide, till the litmus is changed in colour. Evaporate to half the bulk and place on one side. After a time crystals are formed which may be recognised as Epsom salts.

Acids.—Hydrochloric, sulphuric and nitric acids have now been studied, and it is necessary here to consider what characters these three compounds, so dissimilar in many of their properties, have in common, which leads chemists to call them all **acids.** The experiments described from time to time have shown that the three compounds all have a sour taste, and that they all possess the power of turning a blue litmus paper red. These two properties are possessed by almost all acids. In addition, moreover, **all acids contain hydrogen**, which can, in suitable circumstances, be replaced by a metal. The metal may itself turn out the hydrogen by acting on the acid directly. An example of this is afforded by the preparation of hydrogen by acting on sulphuric acid with zinc. Or, the hydrogen of the acid may be replaced by the metal in a compound like caustic soda (sodium hydrate), or lime (calcium oxide). These facts may be collected together in the form of a definition of an acid. Thus:

An acid is a chemical compound with a sour taste, which usually has the power of turning blue litmus red. It always contains hydrogen, which can be replaced by a metal, either directly, or by the action of an oxide, or hydroxide, of a metal upon it.

Bases and salts.—Another class of compounds, of which sodium hydrate and lime may be taken as typical examples, all possess properties of an opposite character to those which distinguish acids. They all have the power of destroying, or **neutralising**, the properties of an acid. These compounds are called **bases.** Some, such as caustic soda and caustic potash, are **soluble in water**, and are called **alkalies.** A solution of ammonia in water behaves like the two soluble bases mentioned, and is also classed with the alkalies.

Bases are always oxides or compounds of a metal with oxygen and hydrogen, known as **hydrates**, or, better, **hydroxides.** When added to an acid, the metal of the base replaces the hydrogen of the acid forming a **salt.** This is a fact of great importance. We may write it in the form of an equation:

A BASE + AN ACID = A SALT + WATER.

This equation provides a definition of a base. It may be stated that:

A base is an oxide or a hydroxide of a metal, capable of neutralising an acid with the formation of a salt and water.

The equation, too, also furnishes a convenient definition of a salt, for, added to the knowledge previously obtained, it enables the statement to be made, that:

A salt is a chemical compound formed by acting upon an acid with a base, in this way replacing the hydrogen of the acid with the metal.

Names of salts.—Before chemistry became the systematised study which it is at present, it was the custom to name a salt after its discoverer, or a name was sometimes given indicating some characteristic property of it, or again the source from which it was obtained. The system now adopted is to name salts in such a way as to show their composition, or the metal and acid from which they are derived. Salts obtained from sulphuric acid are named **sulphates**, those from hydrochloric **chlorides**, and those from nitric acid **nitrates.** For example, the salt obtained in Expt. 78 i., which was recognised as *common salt*, is called **sodium chloride.** The salt produced by the action of sulphuric acid on magnesium oxide (*Epsom salts*) is called **magnesium sulphate.**

The following table gives the common and chemical names of a number of well-known substances. The student should consider how each could be prepared.

Common Name.	Chemical Name.
Baking soda	Sodium bicarbonate
Caustic soda	Sodium hydroxide
Chalk	Calcium carbonate
Common salt	Sodium chloride
Copperas or Green vitriol	Iron sulphate
Epsom salts	Magnesium sulphate
Glauber's salt	Sodium sulphate
Lime	Calcium oxide
Muriatic acid or Spirits of salt	Hydrochloric acid
Oil of vitriol	Sulphuric acid
Plaster of Paris	Calcium sulphate
Sal ammoniac	Ammonium chloride
Soda, or washing soda	Sodium carbonate

QUESTIONS ON CHAPTER XXVI.

127. You are given two bottles containing solutions of hydrochloric acid gas; how would you ascertain which bottle contains the more strongly acid liquid?

128. What are oxides? In what important particulars do the soluble metallic oxides differ from the non-metallic oxides? How are salts formed? Give examples.

129. Oxides of the metals in solution affect vegetable colours differently from oxides of the non-metals. These different solutions when mixed in certain proportions produce a substance affording no colour changes. What names are given to these solutions and to the product of their interaction, and why?

130. What exactly do you mean by an acid, an alkali? Why were these bodies so named? Give an example to show the effect of their interaction.

131. Explain clearly what you understand by the term neutralisation as applied to acids. What class of substances do you obtain by means of this process? Give three examples.

132. What do you understand by a salt?
You are given some dilute hydrochloric acid. State exactly how you would determine the volume of this acid, which, when neutralised with soda, will give one gram of common salt.

133. What class of substance is produced by substituting the hydrogen of an acid by a metal? Give the general properties of these products with special reference to the behaviour of their solutions on litmus.

134. Distinguish between simple and chemical solutions. Illustrate your answer by contrasting the action of water with that of hydrochloric acid on washing soda. Give other examples from your own experiments of the two kinds of solution.

CHAPTER XXVII.

CARBON AND SOME OF ITS COMPOUNDS.

79. FORMS OF CARBON.

i. Carbon is contained in organic substances.—Heat a series of organic substances, such as meat, wood, potato, egg, etc., in a crucible, and notice in all cases the production of a black residue, consisting largely of carbon. Heat more strongly and observe that it burns away, leaving an almost colourless ash.

ii. Properties of carbon.—Examine and write down the properties of as many of the following forms of carbon as you can obtain: diamond, blacklead, wood-charcoal, bone-black and soot.

iii. Charcoal is porous.—(*a*) Show that charcoal floats in cold water. In boiling water charcoal sinks after a time, and then will not float again unless thoroughly dried. This is because air is driven out of the charcoal by the warmth of the water.

(*b*) Fill a large test-tube with ammonia gas over mercury and show that it can be absorbed by introducing a small lump of charcoal. This is a striking proof of the porosity of carbon.

Forms of carbon.—Carbon is a substance which is very widely distributed in nature, being present in all living matter, and in most products resulting from vital activity.

Carbon occurs, combined with other substances, in many rock masses, being a constituent of all the minerals known as *carbonates*. Combined with oxygen as carbon dioxide, it occurs in the atmosphere and dissolved in spring waters.

Diamond.—In the pure state carbon exists in various allotropic forms. Of these the purest and the most valuable is the **diamond.** This form of carbon is crystalline and very hard, being capable of scratching all other materials. Its refractive index is very high, and on this depends its brilliancy as a gem. Diamond is proved to consist of carbon by burning it in air or oxygen, when only **carbon dioxide** results.

Blacklead or **graphite** is another form of almost pure carbon, with properties totally different from those of the diamond. It is opaque and black, and so soft that it will mark paper. It is really a crystalline form of carbon, although good crystals are not very common. It occurs naturally in mines, chiefly in California, and was formerly largely obtained from Cumberland. Besides its use for lead pencils, it is also used as a lubricant.

Amorphous varieties.—Other forms of more or less pure carbon in an **uncrystallised** or **amorphous** state are **coke**, and **gas carbon**, which result from the heating of coal; **lampblack**, which is the carbon deposited by oils, etc., burning in an insufficient supply of oxygen, and **wood charcoal**, obtained by heating wood in closed retorts or in stacks under earth.

Charcoal has the power of absorbing colouring matter, and on the latter account it is used for decolorising solutions coloured by organic matter. **Animal charcoal** is really a misleading term, as the quantity of carbon present is usually only about 10 or 12 per cent., the remaining being chiefly bone ash.

Both animal and wood charcoal are very porous substances, and they have the power of absorbing gases to a large extent. Wood charcoal is used considerably on the Continent for heating purposes. Both kinds are useful in destroying noxious vapours.

Coal contains large quantities of carbon, especially the harder or anthracite coals, where the quantity may reach 94 per cent., being, however, only about 65 in brown coal or lignite.

Whenever any of the kinds of carbon burn freely in a good supply of air or oxygen, carbon dioxide is the compound formed, thus affording evidence that the three varieties are only allotropic forms (see p. 106) of the simple substance carbon.

80. CARBON DIOXIDE PRODUCED BY BURNING AND BREATHING.

i. When carbon is burnt carbon dioxide is formed.—(*a*) Heat strongly a piece of charcoal in a closed hard glass test-tube and show that without air it does not burn.

(*b*) Suspend a piece of glowing charcoal in a jar containing lime-water, Shake up and show that the lime-water is turned milky. Carbon dioxide can always be distinguished by this action upon lime-water, for it is the only common, colourless, inodorous gas which turns lime-water milky.

ii. Carbon dioxide is produced by a burning candle.—(*a*) Burn a candle or taper in a clean dry glass jar (Fig. 110). After the flame has been extinguished, withdraw the taper. Pour a little freshly made lime-water into the jar and shake it up. Notice the milkiness of the lime-water.

(*b*) Cut a long thin chip of wood, hold it in the flame of a laboratory burner until it burns brightly, then thrust it into a cylinder, the bottom of which is covered with lime-water to the depth of about an inch. When the stick ceases to burn, withdraw it and shake the lime-water.

iii. Carbon dioxide produced by breathing.—(*a*) Blow through a piece of glass tube into some clear, freshly made lime-water contained in a beaker. Milkiness is at first produced, but if the blowing is continued long enough it disappears.

FIG. 110.—Burning a candle in air in a glass jar.

FIG. 111.—The air breathed out from the lungs contains carbon dioxide, and will turn clear lime-water milky.

Fill a jar with water and invert it in a basin of water. Blow air from your lungs into the jar by means of a tube. When the jar is full of air place a glass plate under it and lift it out of the water. Show that the air will extinguish a lighted taper.

(*b*) Repeat the two preceding experiments by blowing air from a bellows instead of from the lungs. Notice that this unbreathed air has not the same effects upon a lighted taper or lime-water as breathed air.

iv. The air contains carbon dioxide.—Pour some clear lime-water into a watch-glass, or some other shallow vessel. Leave it exposed to the atmosphere for a little while. Notice the thin white scum formed on the top. The carbon dioxide in the air has turned the top layer of liquid milky.

v. Breathing changes the character of air.—Fit two bottles with corks and tubes as shown in Fig. 112. See that the corks are air-tight. Put some clear lime-water into each bottle. Place the tube *C*, or an india-rubber tube leading from it, in your mouth. When you suck at the tube, air is drawn in through the glass tube which

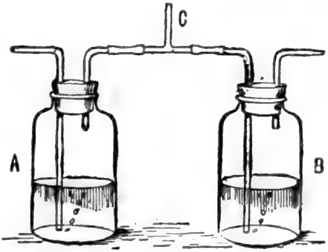

FIG. 112.—Air drawn into *A* does not turn the lime-water milky; but when blown by *C* through *B* from the lungs the lime-water in *B* is turned milky.

dips into the lime-water in the bottle *A*. When, however, you blow instead of sucking, your breath passes out through the tube which dips into the lime-water in the bottle *B*. Notice that the lime-water in *A* remains clear, but that in *B* is rendered milky by the air you breathe out. You thus see that fresh air has little effect upon lime-water, but breathed air quickly turns clear lime-water milky.

vi. Oxygen from plants.—Take a bunch of fresh watercress, or water weeds, and put it into a beaker or glass jar very nearly filled with water saturated with carbon dioxide. Cover the plants with a funnel nearly as wide as the jar, as shown in Fig. 113. Fill a test-tube with water and invert over the funnel. If properly managed there should at first be no gas in the test-tube. Place the jar in bright sunlight for an hour or two and then examine it. You will notice bubbles of a gas have collected at the top of the tube. Test the gas with a glowing splinter of wood. It is found to be oxygen.

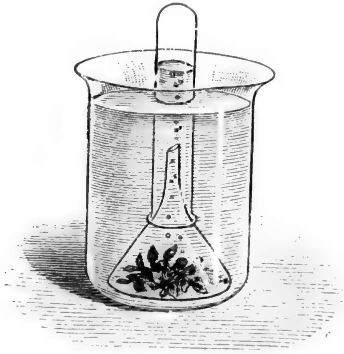

FIG. 113.—Green plants in bright sunlight can decompose carbon dioxide. They keep the carbon for themselves and liberate the oxygen.

vii. Plants in sunlight and in darkness.—Repeat the whole experiment, but instead of putting the bottle in bright sunlight place it in the dark. Observe that in such circumstances no bubbles of oxygen are formed.

viii. Carbon in plants.—Take some green portions of a plant (leaves will do) and heat them on a piece of tin plate over a laboratory burner. Note that they become charred, showing the presence of carbon in them.

Production of carbon dioxide by burning.—When things such as candles, oil, gas and wood are burnt, either in the air or in pure oxygen, a gas is produced which has the power of turning lime-water milky. All these substances contain, in one form or another, a constituent called carbon. As has been seen in previous chapters, the gas produced when these substances burn is carbon dioxide, that is, the gas obtained by burning carbon in air or oxygen. In fact, whenever a substance rich in carbon burns in a plentiful supply of air or oxygen, this carbon dioxide is produced. Knowing how many fires there are in houses, furnaces, engines and so on, it is not difficult to understand that at every hour of the day very large quantities of carbon dioxide are formed, which escape, sooner or later, into the air.

Carbon dioxide is given off in breathing.—If a person blows with the mouth into clear lime-water, the lime-water is turned milky.

This is another important fact. It is clear that carbon dioxide escapes from our mouths in breathing; and so it does from every animal. Not only, then, do all cases of ordinary burning result in the addition of carbon dioxide to the air, but also every act of breathing. It does not matter how small an animal is, all the time it is alive it is continually adding to the atmosphere a certain amount of the colourless, odourless gas which puts out flames and turns lime-water milky.

Purifying action of plants.—That there is always a certain amount of carbon dioxide in the air can be proved by exposing fresh lime-water in a shallow vessel. Very soon the lime-water becomes covered with a thin white layer of chalk, which is formed by the combination of the carbon dioxide in the air with the lime in the lime-water. One reason why there is never very much carbon dioxide in the air out of doors is because there is an agency continuously at work getting rid of this gas. This agency is the green parts of plants which occur everywhere.

When fresh watercress is put into a bottle completely full of water containing carbon dioxide in solution, and the bottle is inverted in a basin of water without allowing air to get into the bottle, it is found that, when the bottle and its contents are exposed to bright sunlight, bubbles of gas collect at the top of the bottle. These bubbles, when tested, are found to be pure oxygen. If, however, the bottle with the cress in it is kept in the dark no bubbles of oxygen collect. Or if a bottle of water in which carbon dioxide is dissolved be put in sunlight, without any watercress, no oxygen collects in the top of the bottle.

In other words, two things are necessary for the formation of the bubbles of oxygen collected from the green plant as described. They are (1) the vegetation, (2) the sunlight. The same conditions have been found to hold true always, thus proving that **green plants in the presence of bright sunlight have the power of turning oxygen out of carbon dioxide.** They keep the carbon for themselves, and it helps them to grow. If the experiment were carefully performed it would be found that the watercress had increased in weight after being exposed in these circumstances to bright sunlight for some time.

81. PRODUCTION OF CARBON DIOXIDE FROM CHALK BY THE ACTION OF AN ACID.

i. Preparation of carbon dioxide.—Into a flask or bottle, fitted like that in Fig. 114, place some chalk, or small pieces of marble. Place the delivery tube in a glass cylinder or a jar with a wide mouth. A disc of cardboard, through which the delivery tube passes, should rest on the top of the jar. Pour dilute hydrochloric acid down the funnel. During the effervescence a gas is given off and collects in the jar. When a burning taper is extinguished immediately it enters the jar, take out the delivery tube and put it into another jar. Cover the first jar of gas with a glass plate. In the same way collect several jars of the gas.

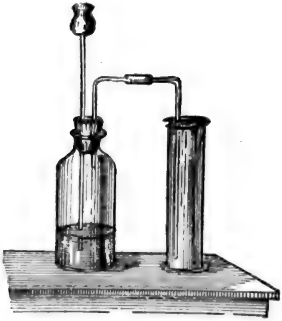

FIG. 114.—Apparatus for the preparation and collection of carbon dioxide.

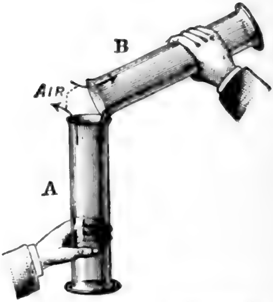

FIG. 115.—Carbon dioxide gas is heavier than air, and can therefore be poured from *B* into *A*, like a liquid.

ii. Properties of carbon dioxide.—(*a*) Notice that the gas is (1) invisible and without taste or smell; (2) extinguishes a lighted taper; (3) must be heavier than air or it could not be collected in the way described.

(*b*) Pour the gas from one jar (*B*) into another (*A*), as shown in the diagram (Fig. 115), and test both jars by a lighted taper. It will be seen that the lower jar contains the carbon dioxide.

iii. Acid solution formed by carbon dioxide.—Pour a little water made blue with litmus into a jar of the gas and shake it up. Some of the gas dissolves, and the colour of the solution turns red. Boil the solution; the carbon dioxide is driven off, and the blue colour is regained.

iv. Action of carbon dioxide on lime-water.—Pass the gas from the delivery tube through some lime-water. Observe that a milkiness is produced, owing to the production of a white powder or *precipitate*, which disappears after a short time.

Boil the solution thus obtained, and notice that the milkiness again appears.

Filter the milky solution, and so obtain the white powder on a filter paper. Add a few drops of dilute hydrochloric acid to the powder. Notice the effervescence. Test the gas which is given off; it puts out a flame.

v. Burning of magnesium in carbon dioxide.—Hold a piece of magnesium ribbon in a crucible tongs, ignite and place in a jar of the gas (take care not to allow the burning metal to touch the sides of the jar). It burns with a crackling noise, and black specks appear on the sides of the jar. A white powder, which may be identified as magnesium oxide, is also formed. In order to dissolve the oxide add a little dilute hydrochloric acid to the jar, and shake up. The contents of jars from the class should be collected in one, filtered, and the black powder remaining on the paper washed with distilled water. The filter paper containing the black powder should now be placed in an oven and allowed to dry.

The powder looks like carbon. Heat a deflagrating spoon till it is red hot, place a portion of the powder on it and quickly put it into a jar of oxygen. Identify the gas formed.

Preparation of quantities of carbon dioxide.—Though carbon dioxide is produced naturally by burning and breathing, and it also escapes from the earth in some regions, there are more convenient ways of obtaining the gas. Experiments show that when an acid is added to chalk, marble, or limestone, a gas is given off which puts out flames, turns clear lime-water milky, and possesses all the properties of carbon dioxide—it is, in fact, carbon dioxide. The best way to prepare bottles or jars of the gas is to place pieces of chalk or marble about the size of peas into a bottle fitted like that in Fig. 114. Dilute hydrochloric acid is poured down the thistle funnel, and when it comes into contact with the marble, the gas is given off. Enough acid is poured in to cover the lower end of the funnel, so the gas cannot escape up the funnel; the carbon dioxide passes through the other tube in the cork. The gas given off is heavier than air, and can therefore be collected as shown in Fig. 114. As the gas accumulates in the jar, the air is pushed out at the top. After several bottles or jars have been filled, the properties of the gas can easily be examined.

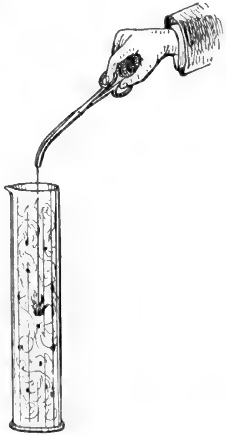
FIG. 116.—Burning of magnesium in carbon dioxide.

Properties of carbon dioxide.—An examination of the gas shows that it is colourless and has no odour. As it is heavier than air it can be poured downwards just like a liquid (Fig. 116).

Carbon dioxide is slightly soluble in water, and the solution which is thus formed turns a blue litmus paper red, just as acids do. For this reason the solution of carbon dioxide in water is often called **carbonic acid**, and the carbon dioxide itself is sometimes spoken of as **carbonic acid gas.** The gas has the property of extinguishing the flame of a taper or match, and is consequently called a non-supporter of combustion.

Further evidence of the identity of the gas obtained from chalk with that produced by burning carbon, is furnished by the burning of magnesium in the gas. The magnesium burns with a crackling noise, as if the gas was being torn asunder; and it is. Magnesium oxide is formed while the sides of the tube are bespattered with black specks. When the oxide is dissolved away with hydrochloric acid, and the black residue collected on a filter, washed with water and dried, it is found that it has the properties of carbon.

Action of carbon dioxide on lime-water.—If carbon dioxide is passed into lime-water a milkiness is seen, but if the passage of the gas is continued the milkiness by and by disappears. If the clear solution which results after the disappearance of the white powder or precipitate is boiled, the milkiness again makes its appearance. The reason of this is that the white substance of which the precipitate is formed dissolves in water which has become saturated with carbon dioxide. When the clear solution, which appears after the solution of the powder, is boiled, the carbon dioxide is driven out of it, and the liquid again becomes pure water. The precipitate reappears, because it will not dissolve in water.

The chemical change when carbon dioxide is passed into lime-water.—What are the facts taught by the experiments described? When an acid is added to the white powder formed when carbon dioxide is passed into lime-water a brisk effervescence is noticed, and the colourless, odourless gas given off is found to put out a flame. But this is just what happens when the acid is dropped upon chalk, and putting the facts together they suggest that the white powder is really chalk, so that carbon dioxide gas combines with the lime in the lime-water to form chalk. We can write:

$$\text{Carbon Dioxide} \ \begin{matrix}\text{unites}\\ \text{with}\end{matrix} \ \text{Lime} \ \begin{matrix}\text{to}\\ \text{form}\end{matrix} \ \text{Chalk.}$$

From this statement it will also be seen that chalk consists of lime and carbon dioxide. Further evidence of the truth of this will be learnt later.

Uses of carbon oxide.—The fact that carbon dioxide is a non-supporter of combustion is made use of in many forms of chemical fire extinguishers, which generally contain solutions of the gas under pressure, or a means of producing large quantities of it by the action of acid upon a carbonate solution. As burning cannot take place in this gas, the flames are therefore extinguished when the gas reaches them.

The solubility of carbon dioxide in water is increased by pressure, and the sparkling nature of the various aerated waters, like soda-water, is due to the carbon dioxide with which they have been charged at high pressures, which escapes when the pressure is reduced to that of the atmosphere by opening the bottle. As the pressure is very great, the bottles used as fire extinguishers have to be very thick.

Carbon dioxide is also produced during fermentation, the "rising" of bread being due to the escape of the gas which is generated by the fermentation, under the influence of the yeast, of the saccharine matters formed from the starch.

Carbon dioxide will not support life, and is sometimes used for suffocating stray dogs and cats.

By cold and pressure carbon dioxide may be liquefied and also solidified, forming a soft white substance, which when mixed with ether forms a powerful freezing mixture, the temperature sinking to about $-100°$ C.

Occurrence of carbon dioxide.—Carbon dioxide has been already stated to exist in the atmosphere, and to be produced by the oxidation of animal and vegetable tissues; also that under the influence of sunlight it is reconverted by the green parts of plants into its constituents, of which the carbon is used by the plant in the formation of new tissue. The gas is found in many natural gaseous emanations, and is frequently present to a large extent in the gases of caves and underground passages, where, owing to its high density, it tends to accumulate if formed by fermentation or other natural processes. In expired air, carbon dioxide is present to the extent of about 4·7 per cent. Although such air is not again respirable, this is probably due to the diminution of the oxygen, for it is doubtful whether carbon dioxide has any direct poisonous effect. The proportion of carbon dioxide may be increased even to 20 per cent. without immediate serious effects if only the quantity of oxygen be increased simultaneously.

82. PRODUCTION OF CARBON DIOXIDE BY HEATING CHALK.

i. Change produced by heating chalk.—Place a little powdered chalk (*not* blackboard chalk) on a piece of platinum foil, and heat it strongly for some minutes in the flame of a laboratory burner. If platinum foil is not at hand, heat a lump of chalk on a piece of coarse wire gauze for some time. After heating, shake the powder on to a damp red litmus paper. Observe that the red litmus paper is in places changed to a blue colour.

ii. Action of lime upon litmus.—Test some wet lime with litmus paper. Observe that it changes the colour of red litmus to blue.

iii. Powder obtained by heating chalk.—Shake up in ordinary tap-water some of the powder obtained in Expt. 82 i. by heating chalk on platinum foil. Filter, or allow the powder to settle; then taste and notice the peculiar taste of lime-water.

iv. Loss of weight when chalk is heated.—Weigh a crucible; add a gram of finely powdered chalk. Place it in a muffle furnace for about an hour at a bright red heat, allow to cool and weigh. Heat the crucible again for a further twenty minutes, cool and weigh. The weight ought to be the same as in last weighing.

v. Add a little water to the residue in the crucible, and note the effect.

vi. **Gas given off when chalk is heated.**—Place some *very* finely powdered chalk in a hard glass test-tube fitted as shown in Fig. 117, and heat as strongly as possible with a blow-pipe flame, even if the test-tube is destroyed. A gas is given off which collects in the receiving tube. Test with a *little* lime-water and with litmus.

The chalk will not be entirely decomposed before the hard glass tube is destroyed, but sufficient gas is obtained for the purpose of identification.

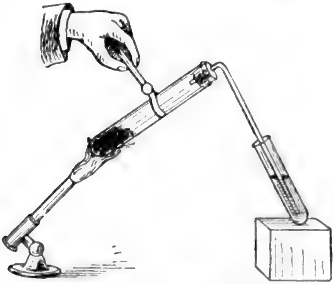

FIG. 117.—Carbon dioxide produced by heating chalk.

Chalk undergoes a change when heated.—It is easy to prove by putting some powdered chalk upon a piece of moist red litmus paper that this substance is unable to change the colour of the paper. If, however, some powdered chalk be strongly heated on a piece of platinum foil in a laboratory burner and then placed on a piece of moist red litmus paper, the red colour is changed to blue. The chalk undergoes some change when heated, or it would not acquire this new property. The same chemical change as occurs when a little powdered limestone, or some chalk, is heated on platinum foil takes place on a large scale in the limekiln. In other words, when chalk and limestone are strongly heated they are changed into quick-lime. The change is brought about by driving carbon dioxide out of them.

CHALK when heated splits up into LIME and CARBON DIOXIDE.

Changes in mass when chalk is heated.—That chalk does undergo a change when heated has already been learnt. But by studying the changes in mass which take place when chalk is strongly heated several important conclusions are arrived at. To obtain satisfactory numerical results it is desirable to heat a fair amount of powdered chalk in a muffle furnace, though a very small quantity of chalk can be completely converted into quicklime when heated in a good laboratory burner. After heating the chalk in a crucible in a muffle furnace for about an hour the crucible is taken out, allowed to cool again, and then reweighed. This plan is continued until there is no further loss of mass on reheating. A simple proportion sum will now show how much 100 grams of chalk would lose if heated strongly, seeing that a given mass of chalk loses the amount noticed.

Percentage loss of mass of chalk when strongly heated.—Whenever chalk or limestone—no matter from what locality—is strongly heated in a kiln or in a laboratory furnace, it loses about 44 per cent. of its mass, as shown in the above experiment. That is to say, 100 grams, or pounds, or tons, of chalk lose 44 grams, or pounds, or tons, in mass when strongly heated. But it has been learnt before, that this loss in mass is due to the escape of a gas which puts out flames and is called carbon dioxide, so that when 100 grams of chalk are strongly heated, 44 grams of the gas, carbon dioxide, are driven out and 56 grams of quicklime are left behind.

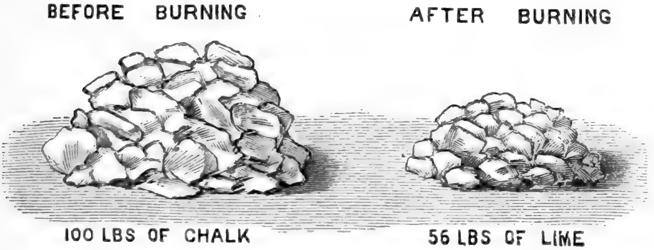

FIG. 118.—When 100 lbs. of chalk are heated, 56 lbs. of lime are produced.

83. QUANTITATIVE STUDY OF THE ACTION OF ACIDS ON CHALK.

i. Composition of chalk.—Place a few small pieces of chalk in a test-tube. Add dilute hydrochloric acid until the effervescence, due to the production of carbon dioxide, ceases (Expt. 81 i.). Filter the solution remaining in the test-tube and evaporate it. Notice the substance left; it is not chalk, but *calcium chloride.*

ii. Solution of lime in acids.—Dissolve some lime in hydrochloric acid, and evaporate the solution to dryness. Note the formation of a white solid, which rapidly absorbs moisture from the air and liquefies. It has been previously used in experiments under the name of *calcium chloride.*

iii. Loss of mass when an acid acts on chalk.—Into a flask *A* (Fig. 119) fitted with india-rubber stopper and tubes, as shown, pour some dilute hydrochloric acid, and weigh the flask with the contained acid. Weigh a small test-tube *B* (i) empty, (ii) with some chalk in it. Place this in the flask, taking care not to allow the acid to reach the inside of the tube. Replace the cork and shake the flask so as to allow the acid and chalk to come in contact with one another. The gas (carbon dioxide) formed escapes through the tube *C*, in which cotton-wool is packed to prevent the escape of moisture. When the action is at an end draw a current of air through the flask by sucking the tube *C*, in order to displace the gas still left inside. Weigh the flask again, and subtract the mass so obtained from that of the flask + mass of tube and chalk. The resulting number is evidently the mass of the gas evolved. Calculate the loss of mass for 100 grams of chalk.

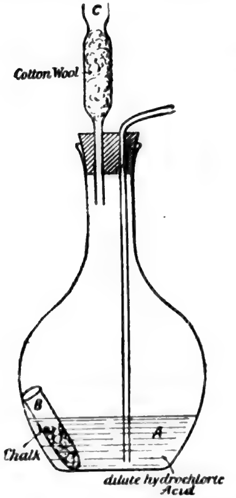

FIG. 119.—When the chalk in *B* is acted upon by the acid in *A* it loses 44 per cent. of its mass.

If the experiment is repeated several times, or done independently by different students, the results obtained may be added together and divided by the number of experiments, in order to obtain an average result. It will be found that 100 grams of chalk lose about 44 grams of gas when acted upon by an acid.

Changes produced by adding acid to chalk.—When hydrochloric acid is added to a little chalk a brisk effervescence occurs and a colourless, odourless gas which turns lime-water milky is given off.

If hydrochloric acid is poured upon chalk, and if the solution remaining, after the effervescence of the chalk with the acid has ceased, is filtered and evaporated, a new substance known as **calcium chloride** is obtained. So that, from chalk and hydrochloric acid, it is possible to produce carbon dioxide, calcium chloride and water (which is driven off by evaporation).

Substances which, like chalk, evolve carbon dioxide when acted upon by an acid are known as **carbonates**, and numerous carbonates exist, all possessing similar characteristics. Many of these on heating also give off carbon dioxide, the residue being known to be an oxide of a metal, so that they consist of carbon dioxide and a metallic oxide.

This leads to the idea that lime is also the oxide of a metal, and this view is now known to be correct, the metal being named **calcium.** Lime, therefore, is **calcium oxide,** and chalk **calcium carbonate.**

A test for chalk and limestone.—The geologist, or man of science who examines and studies rocks, makes use of this experiment to discover whether rocks are limestones or not. Some limestones are black and not easily recognised by the eye. But if, when a little acid is dropped upon them, they fizz like the chalk in a test-tube does when acid is poured on it, the geologist can be fairly sure, even if they are black, that they are limestones all the same.

Changes in mass when acids act upon chalk.—If a certain mass of chalk or limestone is weighed, and an acid is then allowed to act upon it until no more gas is given off, it is found that the chalk always loses mass in a given proportion. When 100 grams of chalk are treated with enough acid to expel all the gas which puts out a flame, the loss in mass is always 44 grams. Now, as none of this gas is given off when acids come into contact with quicklime (the substance which is left when chalk is strongly heated), the conclusion which must be come to is that quicklime contains none of it, and that strong heat acts in the same way as the acid in expelling the gas. The loss of 44 grams in mass therefore shows that 100 grams of chalk contain 44 grams of the gas which puts out flames, and called carbon dioxide; and the experiment of strongly heating chalk proves that the remaining 56 grams are quicklime, or that

100 grams	are	56 grams		44 grams
of	made up	of	and	of
Chalk	of	Quicklime		Carbon Dioxide.

Action of acids and of heat on chalk compared.—These results are of such great importance that it is worth while to look at them in another way to try to understand them thoroughly. It has been learnt that 100 grams of chalk when strongly heated lose in mass and give 56 grams of quicklime. Also, when 100 grams of chalk are acted upon by acids there is a loss of mass too, and 44 grams of an invisible gas, which puts out flames, are given off, and this gas turns blue litmus red.

But by adding 56 and 44 together the result is exactly 100, and it looks as if the quicklime and carbon dioxide together make up chalk. As chalk can actually be made from the combination of these two compounds, there need be no room for doubt, and it may be written as a fact, that

Chalk	is made up of	Quicklime	and	Carbon Dioxide.

QUESTIONS ON CHAPTER XXVII.

135. Describe how you would prepare carbonic acid gas by two different methods. In what respects does this gas differ from nitrogen?

136. Give a brief summary of the conclusions arrived at from your study of chalk.

137. Chalk when strongly heated affords two constituents; state clearly the properties of these constituents; do you regard them as elements? if not, state the evidence in favour of their being compound substances.

138. How is chalk distinguished from lime? Name the elements present in chalk. If you are provided with these elements how would you prepare a sample of pure chalk?

139. A stream of carbonic acid gas is passed through lime-water; what is the first effect observed, and what is the effect of the continued action of the gas? How do you account for these different effects?

140. Water is carefully added to weighed specimens of quicklime and chalk; what will be observed? The residues are now quickly dried at 100° C. and weighed again; account for any differences which may be observed.

141. Dilute nitric acid is (*a*) poured into each of three porcelain dishes, the first containing calcium, the second lime, and the third chalk. The contents are then (*b*) evaporated to dryness and (*c*) strongly heated. Describe what is observed at each stage, and mention the products formed.

142. 1·5 grams of a mixture consisting of chalk and lime when treated with excess of hydrochloric acid afforded 200 c.c. of gas at 15° C. and 74 cm. pressure. What is the percentage of chalk in the mixture?
(A litre of carbonic acid gas at N.T.P. weighs 1·98 grams.)

143. You are given washing soda, potassium chlorate, hydrochloric acid and metallic calcium. Describe how with these materials you would make a specimen of chalk. Given that calcium on being heated increases in weight by 40 per cent., calculate what weight of chalk could be made from one gram of calcium.

GENERAL QUESTIONS.

144. Sketch the apparatus you would employ, and describe fully how you would determine which of two given metals is capable of liberating the greater volume of hydrogen from an acid.

145. State exactly how you would proceed to collect the following gases: hydrogen, ammonia, carbon dioxide and hydrochloric acid.

146. For what purpose are the following substances employed in the laboratory: (*a*) calcium chloride, (*b*) litmus, and (*c*) lime-water? Give examples of their use.

147. Mention any three substances whose composition alters when left exposed to the atmosphere at ordinary temperature. Explain the cause of the alteration in each case and the experimental evidence on which your explanation is based.

148. What are oxides? Give examples of metallic and non-metallic oxides. What oxides enter into the composition of chalk? How would you experimentally determine the proportions in which they are present in a sample of chalk?

149. Oxygen may be prepared by heating potassium chlorate alone, or by heating a mixture of potassium chlorate and manganese dioxide. Why is the latter process usually employed?
If you were required to ascertain the source of the oxygen prepared by the second process, how would you carry out the investigation?

150. On heating water in an open vessel it gradually disappears. Mercuric oxide and carbon, when strongly heated in open crucibles, also disappear. Has matter been destroyed in any of these cases? Briefly outline the experiments you would make to ascertain the nature of the change in each case.

151. When a candle is burned in air enclosed over mercury, it soon goes out, the vessel being bedewed with moisture and the mercury rising slightly; if lime-water be substituted for mercury it rises to a greater extent and turns milky. Give reasons for all the phenomena here stated.

152. How do you account for the fact that sea-water has such a pronounced salty taste, whereas the water from the rivers which flow into seas has a scarcely perceptible taste? How would you estimate the amount of soluble matter present in one litre of sea-water?

153. Give a detailed account of any investigations you have made as to the effect of heat on green vitriol, sal ammoniac and red oxide of mercury respectively.

154. You have placed before you four jars containing respectively hydrogen, carbon dioxide, oxygen and nitrogen. What experiments would enable you to say which was which?

155. Explain how it is that phosphorus readily oxidises in air which consists of about 23 per cent. oxygen, while it remains unchanged in water which consists of about 89 per cent. oxygen. Explain also why it is impossible to keep metallic sodium in water.

156. Carefully weighed quantities of the following substances are left exposed to the air and weighed from time to time: iron filings, caustic potash, oil of vitriol. State whether you would expect any change in weight, and, if so, account fully for such change.

157. Describe exactly what happens when each of the following bodies is heated and then allowed to cool: sulphur, ammonia solution, chalk, nitre.

Compare the physical and chemical properties of the residues with those of the original substances.

158. Describe fully the effect of strongly heating (*a*) red oxide of mercury, (*b*) green vitriol. Show from your knowledge of the history of chemistry that special interest attaches to these experiments.

159. Show that the introduction of the balance had an important influence on the progress of chemistry as a science.

160. State clearly your views on combustion. How would you investigate the conditions which limit the combustion of a candle in an enclosed volume of air?

ANSWERS TO PROBLEMS

p. 10. **3.** 28·06 ft. **6.** 25·06 in.
p. 21. **7.** 3776 sq. ft., 7 % **10.** 255·43 sq. cm., 232·26 c.c.
16. 284·4 c.c **17.** 1·5 in., 2·9 in.
p. 36. **21.** 5·12 in. beneath table top at centre. **23.** 0·722 in.
p. 43. **24.** 6·88 ft. from end. **26.** 30, 24, 20, 16, 15 cm., 24 gm.
30. 18 cm. from the centre.
p. 58. **35.** Sp. gr. $8{\cdot}\dot{3}$. **38.** Sp. gr. 0·65. **42.** 3 lbs., 3·44 lbs.
44. Sp. gr. 4·19. **46.** Sp. gr. 0·705. **47.** 8·36 ; 1.
48. 0·6 in.
p. 69. **53.** 1036 gm. per sq. cm. **59.** 27·2 in. **61.** 2·72 in., 4·72 in.
p. 86. **69.** 59° C. **72.** 610·4 c.c.
p. 112. **84.** 100 gm. water dissolve 7·99 gm. solid at 15° C.
85. 100 gm. **88.** 66·75 gm. **89.** 9·63 lbs.
p. 135. **103.** 821·3 c.c.
p. 158. **126.** 111·1 gm. or 1240 litres.
p. 176 **142.** 55·3 % **143.** 2·5 gm. chalk.

INDEX

(The numbers refer to the pages.)

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