

AN INTRODUCTION
TO
CHEMISTRY AND PHYSICS

VOLUME I

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BY

W. H. PERKIN, JR., PH.D., F.R.S.


AND

BEVAN LEAN, D.Sc., B.A. (LOND.)

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PREFACE

TO THE FIRST EDITION

THE assured position that Chemistry now holds in English education is shown by its inclusion as a compulsory or at least an optional subject in almost all examinations preliminary to entrance on the learned professions, and by the fact that many of the best schools now possess well-equipped laboratories and give regular instruction in the elements of both Theoretical and Practical Chemistry.

In directing the attention of our nation to the study of Chemistry as a means of scientific education, Sir Henry Roscoe has led the way. His text-books have given to generations of students their introduction to a knowledge of the science. His Primer of Chemistry and List of Experiments prepared for the Science and Art Department contain the subject-matter which every schoolboy has till recent times been expected to know, every examiner has asked for, and few text-books have dared to disregard. The output of admirable text-books on these lines has during recent years been so abundant, that the issue of another would be unjustifiable, if it differed from them merely in detail and in unimportant features.

But while the knowledge of the actual facts of Chemistry is worth much, there has been a growing feeling that the methods of gaining knowledge are often of more educational value than the knowledge itself, and that a lad who starts life with his mind merely stored with facts is not so well equipped as one who has

also *learnt how to learn*, not from books alone, but by direct contact with realities. In a school curriculum Science is chiefly valuable as a means of culture, and because there are faculties of the human intellect, namely, correct observation, sound reasoning from observation and experiment, and imagination in seeing into the invisible causes of visible phenomena, which can alone find their development in a carefully considered course in Physical Science. We are, therefore, even more anxious to insist on the methods by which knowledge has been obtained than on the actual facts of Chemistry. Sound methods of observation, moreover, remain valid, and may continually enlarge and correct our ideas in studying the further development of the science. At the same time the student must be sure of his facts before he begins to theorise. "Why is it," Charles II. is reported to have asked, "that when you put a dead fish into a bowl full of water it runs over, but if the fish is alive it does not run over?"

Some time ago Professor H. E. Armstrong, a leader in the movement which has sought to place the teaching of Chemistry upon what is regarded by many as a more rational basis, appended a detailed scheme for a school course on Chemistry to the Report upon Methods of Teaching Chemistry, which was presented to the British Association at Newcastle.¹

Very recently, the Incorporated Association of Headmasters has issued a syllabus of a course of instruction in Elementary Science, including Physics and Chemistry, drawn up by a Committee on Science Teaching, of which Mr. C. M. Stuart, of Catford, was chairman, and Professor Armstrong a prominent member. To quote from this syllabus, "This course is intended for all boys and girls commencing the study of Science. It represents in the opinion of the Committee a suitable commencement for those who continue the subject, and indicates the manner in which it may be made of true educational value to those who do not pursue it further. . . . While the main object of the course should be to train students to solve simple

¹ See *Report of the British Association*, 1889. A subsequent paper elaborating certain features in this scheme may be found in *Nature*, vol. xliii. p. 593.

problems by experiment—to work accurately, and with a clearly defined purpose, and to reason from observation—the instruction given should eventually lead them to comprehend the nature of air, water, 'fire,' earth and food."

This syllabus of the Association of Headmasters reached us after we had nearly completed our work on the present book, and it was a matter of satisfaction to us to find that almost the whole of the syllabus and the detailed experiments for the course on Elementary Chemistry were dealt with in our manuscript.

While the aim of the course described in the present book, and the object of the courses referred to above have alike been the cultivation of a spirit of inquiry and the prosecution of Science as among the best means of culture, there is, nevertheless, one difference between the lines on which they have been worked out.

The principle which has been our general guide in the selection and treatment of subject-matter may be expressed in words used in another connection by Professor H. A. Miers, in a recent inaugural lecture¹ at Oxford, in which he applied the saying of the Darwinists that the development of the individual is an epitome of the evolution of the race.

"The order," he said, "in which a subject can best be unfolded before a student's mind is very satisfactorily marked out by the historical development of the subject: a profitable course of teaching is suggested by the history of a science, and *the order in which problems have presented themselves to successive generations is the order in which they may be most naturally presented to the individual.*" This is, we believe, particularly applicable to the study of Chemistry on the part of beginners. At the same time the progress of Science has often of course been like a paper-chase with many false scents, and it is not necessary for the logical application of the above principle to follow all these blind alleys.

A glance at the Table of Contents will indicate the course which we have been led to adopt in our attempt to work out an

¹ *Nature*, vol. liv. p. 208

Introduction to the Study of Chemistry. We have referred at the outset to Alchemy and to some of the errors which were current till the seventeenth century, showing the readiness with which errors arose unless checked by well-devised experiment and careful measurement. In this way the student is led to see the *necessity of practical work and of exact measurement at a very early stage*. Measurements are then made of length, of mass, of the volume of liquids, of temperature, of density, of the pressure of the air, and of heat.

Then follows the practice of important chemical operations, such as solution, filtration, evaporation, and crystallisation, and the student's powers of observation and description are exercised by the examination of certain important raw materials, found native in the crust of the earth. He next learns how to prepare the most important acids and alkalies from his raw materials, after the manner of Geber, Glauber and other alchemists.

Having thus acquired in some measure quickness of eye in observation, accuracy of expression in description, delicacy of hand in manipulation, and a passing acquaintance with the apparatus and chemicals in the laboratory, the student is ready to set to work upon a series of researches upon the Action of Acids on Chalk, Fire and Air, the Rusting of Iron, the Discovery of Oxygen, the Action of Acids on Metals, the Composition of Water, and other subjects, in which he is led to work so far as may be along the lines which were successfully pursued by Black, by the Oxford chemists of the seventeenth century, by Priestley, Scheele, Lavoisier, Cavendish and by other famous investigators. Many of the experiments introduced have been suggested by their researches, and frequent quotations of interest have been made from their accounts of their own experiments.

The student will thus become acquainted with several Gases, and is therefore ready to investigate some of their principal properties. Researches follow which have a direct bearing on the law of Definite Proportions; for example, a complete research is made upon Chalk, following the lines of Black's classical investigation, and another upon the quantitative composition of Water. * * * * *

A chapter on the Law of Definite Proportions closes the book.

We hold then that not only the *order*, but also the *method*, by which Chemistry can best be unfolded before a student's mind is suggested by its historical development. It is a necessary corollary that Chemistry cannot be learnt in the lecture-room alone, and that the *experiments should be for the most part quantitative in character*. Qualitative analysis has hitherto been largely used as the practical exercise in the science. It has some merits. We think, however, that practical chemistry should not consist, to such an extent as is usually the case, of "qualitative analysis," and that the right place for such a course is after and not before a simple quantitative course. While practical work should be introduced at a very early stage, the lecture-room has of course a necessary place, and the student should not start an experiment until he thoroughly understands what the object of the experiment is, what he means to do, and how he means to do it.

There are somewhat rigid limitations of time and circumstance which must be considered in selecting experiments suitable for a school laboratory. The experiments must not require more than two pairs of hands; the manipulation must not be so difficult that accurate results cannot be obtained; the apparatus must be cheap enough to allow of all the boys doing the same experiment at once, and the whole experiment must not take more than 1 or $1\frac{1}{2}$ hours, including the time required for putting the apparatus away. These conditions exclude many important experiments; for instance, the determination of the composition of water by passing hydrogen over copper oxide may be done by the teacher as a lecture-table experiment, but it must generally be omitted by individual members of the class. Weighings should generally be conducted to the nearest centigram, because a boy's errors of manipulation are apt to exceed 1 centigram, and therefore milligram weighings will only give a fictitious accuracy and waste time. One per cent. accuracy is the standard that may be aimed at. We have dwelt at some length upon the details which are necessary to exactness in the

results of experiments. It is a great advantage to have several boys working at the same time at an experiment, for example, on the weight of a litre of air or on the percentage of water in a crystal; results should afterwards be compared and discussed. A great deal of time is saved by letting boys work in pairs, moreover, a spirit of co-operation is cultivated, and they learn at least as much as in doing the experiment alone, perhaps more, because they teach one another. * * * *

We would add that experience shows that a student's mathematics should be ready before he touches the apparatus of the laboratory. And this applies not only here. If any would-be chemists have not yet mastered the elements of arithmetic, decimals, the unitary method, percentages and proportional parts, we recommend them to close this book and go back to their ciphering. There can be no sound knowledge of Physics or of Chemistry without mathematical backbone. There is nothing more distracting to teacher and to student than to find that laboratory results cannot be worked out for want of adequate mathematical knowledge.

It may perhaps be objected that we have abandoned the principle we laid down as our guide in relegating Black's research on Chalk to so late a position. Our reply is that the aforesaid principle while invaluable as a general guide, must be applied with discrimination, and other considerations must not be lost sight of. For instance, it is desirable that a study of "every-day" phenomena of great interest, such as Fire and Air, should precede what is after all of a more special character. Moreover, the research on Chalk is the most complete in itself of any entered upon, and one which makes too great a demand upon the powers of hand and of mind to be introduced in its true chronological place immediately after the discovery of Fixed Air.

The value of this book would be little in practice, whatever its general merit, if it came straight from the writing-desk, and had not been employed in actual class work. Not only have many of the chapters been worked through by elementary students in the laboratories of the Owens College, but we have been most fortunate in securing the co-operation of Hugh

Richardson, M.A., Science Master of Sedbergh School, and of S. H. Davies, M.Sc., Head of the Chemical Department of the Battersea Polytechnic, London. Using typed copies of our MSS., they have taught through almost the whole in their classes. We have amended our text wherever their trial of it showed alteration to be required, and we are convinced that the practicability and success of our scheme in actual class work has thus been secured and verified. Mr. Richardson also made at our request the first draft of the chapters on the Properties of Gases, and he has from the first helped us with his experience in the teaching of Science in schools.

We gladly thank also Mr. S. M. Walford, of the Hyde and Oldham Technical Schools; Mr. Foster, of the Nelson Technical School, and Mr. S. E. Brown, of the Friends' School, Ackworth, for testing some chapters in their own classes.

Our colleagues, Mr. G. J. Fowler, Dr. A. Harden, and Dr. W. A. Bone, have also given us kind assistance in certain details, while Mr. P. J. Hartog has helped us with criticism upon our chapter on the Law of Definite Proportions.¹

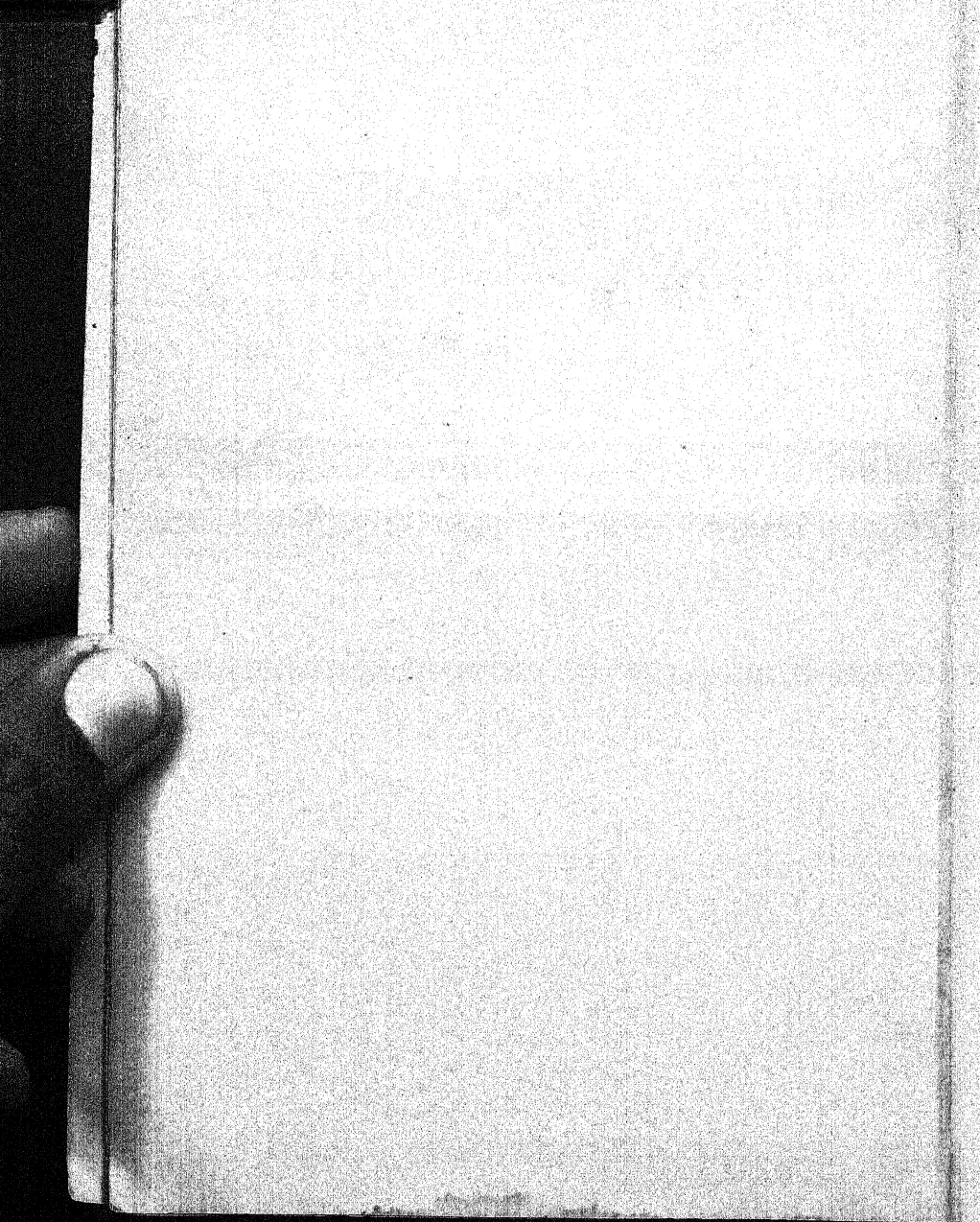
While we have spared no trouble in a task which has been no light one, we cannot hope that we have avoided all error either in judgment or in points of detail, and suggestions and corrections offered by fellow teachers will be freely welcomed.

We may add that it is our desire to include in any subsequent edition chapters, which are already in hand, on the Discovery of the Metals and on Fuels and Food-Substances.

W. H. P.
B. L.

OWENS COLLEGE,
July, 1896.

¹ See *Nature*, vol. 1, p. 149, and *British Association Report*, 1894, p. 518.



PREFACE

TO THE SECOND EDITION

WE have been encouraged by the sympathetic reception which has been accorded to the first edition of our little book to submit it to a thorough revision.

The marked trend of elementary science teaching during recent years along the lines which we in common with many others have been advocating and practising is seen for instance in the present syllabus of Elementary Experimental Science of the Association of Headmasters, and in those of the Oxford and Cambridge Local Examinations, the Joint Scholarship Board, and the Scholarship Examinations of the London County Council.

These all show an increasing recognition of the vital necessity of keeping alive in both boys and girls that power of "making knowledge"¹ through their own experiences which, though possessed by every young child, is so commonly stifled and lost as years advance. To form a habit of inquiry, and to train the intelligence, and not the accumulation of facts, is now generally admitted to be the first aim of an elementary course in science. It is desired, in a word, *to teach boys and girls to think for themselves and to learn from their own observations.*

It has sometimes been alleged as an objection to the method which our book illustrates, that sooner or later some fact

¹ Professor J. G. Macgregor, *Nature*, September, 1899.

must be assumed which a boy has not yet proved. That is no doubt the case, but his position is also that of the original investigator of to-day. John Dalton refused to accept the results obtained by others, and to his loss. The scientific method does not forbid such aid. Still, in an introductory course it is desirable that not many facts should be borrowed in the few typical researches pursued, in order that a very habit of research may be formed; afterwards the student may safely bring facts established by others to bear upon his own inquiries. Nor should it be forgotten that the scientific method includes deductive methods as well as inductive. Theories and working hypotheses inspire research, and without ideas progress stops. It is for lack of them that so few boys make any profitable use of a laboratory if left alone without the help of books or teachers. We have endeavoured to illustrate the importance of ideas by the use we have made, for instance, of Lavoisier's Theory of Combustion, when investigating the burning of Carbon and the explosion of Inflammable Air and Oxygen.

In this connection it may be added that the comparison of results and the discussion of them by the teacher is of the utmost importance. Instead of each boy repeating an experiment many times as the original investigator may do, he should be led to compare his own result with those obtained by his comrades, and to calculate the *mean* result. The comparison of results keeps alive a spirit of rivalry and a keen interest, and the mean result of quantitative experiments is usually surprisingly accurate, for example, the percentage of oxygen in the air or of carbonic acid in chalk can be determined with an error of less than one per cent. by means of the simplest apparatus.

There have been ominous warnings from Prof. G. H. Darwin¹ and from Prof. Karl Pearson² against the growing luxury of our laboratories and the costliness of apparatus; and this not because money is grudged for the cause of scientific education,

¹ Preface to his book on *The Tides*.

² *Lecture on Science and National Life*

but because these material things cloud our eyes from seeing the really essential thing, the encouragement of the spirit of research. To this end we are content to work in rooms with simple fittings, and we have tried to avoid specialised apparatus, and to use combinations of rubber stoppers and flasks and tubes. This simplicity of apparatus results in great economy of expenditure and allows of a sufficient supply of balances and other essential instruments of general research. Moreover, as Prof. Ostwald has very recently pointed out, the simplest apparatus possible should be used that the attention may be fixed on the chief features of an experiment.

In this Second Edition the book is now divided into two volumes. The original volume covered two years' full work, and there is a distinct disadvantage in leaving a book long unfinished in a boy's hands. He begins to think he is stale of it. A new book is always a stimulus.

At the beginning of Volume I there are several chapters on Physical Measurement. Formal instruction, in our opinion, in the use of the rule, the balance and the burette is essential to progress. In some schools several hours are devoted to Science every week, and a more complete course in Physics than that we have introduced can be pursued. But in many schools other subjects will continue to make such demands upon the curriculum that the course of Physics, which should precede any study of chemistry, must not be too detailed. It is quite possible, of course, to spend a couple of terms or more upon measurements of the densities of all manner of materials, but the teacher will do well to recall the warning of Prof. J. J. Thomson, lest in teaching his boys to measure every physical quantity, he should deprive them of all wish to measure any of them. Grammar is not literature, nor is science measurement, though skill in the use of instruments of precision is essential. In collecting these physical chapters we have put together so much as seems to us necessary for progress in the chemical laboratory, but we rely on individual teachers to use them with discretion, selecting or postponing according to the circumstances of the class and the

school, and it must always be kept in mind that a boy, whilst learning to use his tools, will need to be inspired with a desire to employ them.

Volume II is purely chemical in character, and is intended to lead the schoolboy of 15 or 16 years of age along the paths of inquiry, which engaged the attention of the great chemists of the eighteenth century.

These two volumes contain the bulk of what was in our first edition. To this we have added some new chapters. The growing disposition to use Graphic Methods of representation and calculation, justify us in devoting a whole chapter to this subject, to which before we only referred incidentally. Many exercises are added to show the use of squared paper both for the records of results and as a means of arriving at new or unsuspected facts and relations. A new chapter on the Discovery of the Common Metals introduces simple blow-pipe work, minute smelting operations, which will give training in clean and neat manipulations and valuable practice in the examination of unfamiliar minerals. In another chapter on Fuels and Food Stuff's an examination of a variety of organic substances leads to a recognition of the fundamental importance of Carbon. Our treatment is in keeping with the educational maxim of taking familiar objects as texts, proceeding from the known to the unknown, and following the lines of the historical development of knowledge.

Three somewhat difficult chapters,¹ which are of less general interest, have been omitted, and are held over for a Third Volume, which we have in preparation. In our opinion they will find their proper setting in a Third Year's Course in which the Atomic Theory may arise for discussion.

The increasing number of Public Examination papers which recognise a change in the methods of teaching, and their great value both to boys and masters in directing attention to problems of interest, and in stimulating the very necessary

¹ Dalton's Law of Mixed Gases, The Combining Properties of Acids and Alkalies, and The Equivalent Weights of the Metals.

work of revision, have encouraged us to add, at the end of Volume II, some papers which have been set by Public Bodies. We think also that they are important as evidences of the growing harmony of view between teachers and examiners. Some private examination papers which have been set in schools using our book have been appended to Volume I.

It has been the fashion lately with some to decry any use of text-books. It is very noteworthy that young boys (æ. 12) learn very little from a chapter in a text-book, even after a lesson on the same subject, while it is equally remarkable how an older class (æ. 16) may realise, with fidelity of detail, experiments they have never seen, and make rapid progress with comparatively little help from a master. These two volumes are not intended to take the place of a teacher or of a boy's own note-book. The impressions which a boy's mind receives as he works in the laboratory must be expressed then and there in his own note-book, without the use of a text-book as a crib. But when a lengthy inquiry has been completed and he comes to write up his laboratory notes, some help and some additional information will often be needed, and then it is a part of a wise education to encourage the use of a text-book. We know that many teachers have found our book more useful to themselves than to their classes; an able boy can perhaps do without a text-book entirely, but his duller comrade often needs a helping hand in the revision of his work. It need scarcely be pointed out that there are experiments for which detailed and frequent instructions are necessary, and in such cases the necessary information may be obtained by a boy directly from a text-book, and its use will remove a burden from the teacher. And further, it is a great convenience to have mathematical examples in type. A few of them should be worked in class, and afterwards be set for revision without help. But enough of detail; every teacher must work in his own way to do his best work.

In conclusion, we have much pleasure in stating that Mr. Hugh Richardson, of the Friends' School, York, made for us

the first draft of the chapter on Graphic Representation, and we have throughout had the assistance of his experience. We gladly also thank Mr. S. H. Davies, late Head of the Chemical Department of the Battersea Polytechnic, for notes which we used in writing the new chapters on the Discovery of the Metals and on Fuels and Food Stuffs.

W. H. P.

B. L.

August, 1901.

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CHAPTER I

THE BIRTH OF CHEMISTRY

The Earliest Book on Chemistry.—Chemistry is one of the most modern of sciences, and one which has been placed upon a sure foundation only in the course of the last century. Yet the actual beginning and the dawning knowledge of Chemistry can be traced back to Egypt many centuries ago.

The oldest chemical record¹ of which we have any knowledge is a papyrus found at Thebes in the wrappings of the embalmed body of a goldsmith, who had lived about the time of the second or third century after Christ. From this note-book of the goldsmith, which is now in the library of the University of Leyden, we find that he was acquainted not only with gold and silver, but also with copper, lead, tin, zinc, and arsenic. Recipes are given for preparing imitation gold and silver, and for increasing the weight of gold and silver by the addition of inferior metals such as lead without altering the apparent character of the former.

The notes of this fraudulent workman, which have been preserved in perfect condition in an Egyptian mummy, and have escaped the accidents of fifteen hundred years, show that at that time the knowledge of the metals possessed by some of the Egyptians was very considerable. Not only could the author counterfeit true gold or silver, the "noble" metals as they were called, by mixing inferior or "base" metals, but

¹ See Berthelot, *Les Origines de l'Alchimie*.

tests are given by which the imitation could be detected and distinguished from the pure and genuine metal.

For instance, it is stated that gold should keep its colour when melted. If it becomes whitish it contains silver, if it blackens it contains lead, and if it becomes rough or hard when cold it contains copper or tin. So also silver should remain white and shining when fused. If it is blackened it must be falsified with lead, and if it is rendered hard and yellow, copper is present.

Alchemy : The Transmutation of Metals.—The imitation of the noble metals by the process of mixing the base metals in various proportions, as the writer of the papyrus described in the fourth century, was accepted in the next century as an actual change or *transmutation* of the base metals into the noble metals. Extraordinary as it may seem to us, for more than a thousand years such transmutation was believed in not only by the ignorant and unlearned but even by philosophers.

Indeed, it seems probable that the very word Chemistry is derived from the Greek *Chemeia* or *χημεία*, *transmutation*, since the transmutation of the metals was the chief object of the Egyptians.

About the year 640 the Arabians overran Egypt, and becoming acquainted there with Chemistry, they prefixed to the word for it the Arabic Article *al*, so that with them the science was spoken of as *Alchemy*. The Arabians spread through Northern Africa into Spain, and founded many Universities to which students flocked from all parts of Europe.

From the Egyptians the Arabian Alchemists had acquired the firm belief in the possibility of the change of one kind of material into another, and of the transmutation of the base into the noble metals. Their most earnest endeavours were given to the search for the *philosopher's stone*, a substance of such extraordinary virtue that according to Roger Bacon one part of it was believed to have the power to convert 1,000,000 parts of base metal into pure gold. The possibility of the existence of such a substance was accepted by the most learned; for instance, Van Helmont, a distinguished chemist, a physician and an honest man, recorded how in 1615 he obtained one quarter of a grain of the philosopher's stone and converted eight ounces of mercury into gold. Helvetius, again, was an opponent of the Alchemists until 1666, when he managed to obtain a little piece of the philosopher's stone, with which he said he

converted lead into a yellow metal, afterwards tested by the Master of the Mint, and pronounced to be gold.

In some cases we can trace the fraudulent means by which the apparent transmutation was effected. A certain Duke of Wurtemberg kept an Alchemist in his service, and following his directions, the Duke was accustomed to place the necessary ingredients in a crucible, and then, all having left the room, the door was locked, and the next morning, behold, pure gold was found in the crucible. It is recorded, however, that the gold was obtained, not by the miraculous agency of any philosopher's stone, but by the hand of the Alchemist's son, who had been concealed in a cupboard within the laboratory.

Remarkable as the belief in the transmutation of the metals appears to us, many facts and experiments may be mentioned which at first sight seem to confirm such a belief. The following experiments will illustrate this point:—

EXPT. 1. To obtain Lead from Galena.—Galena is a

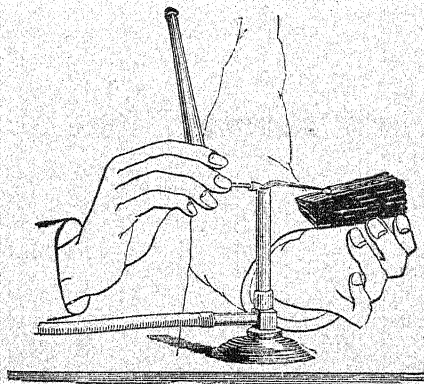


FIG. 1.

metallic-looking mineral. Grind a small portion to a fine powder in a mortar. Make a small cavity in a lump of charcoal, and place in it a little of the powder (as much as a small shot). Turn a Bunsen flame so as to form a small luminous flame. Place the nozzle of the mouth-blowpipe just within the

flame, and force a gentle current of air into it. Direct the tip of the flame upon the galena on the charcoal. Heat till white fumes are evolved, and a bright mirror-like surface of molten metal appears, which will perhaps gather into a bead. Allow it to solidify, and then transfer the bead to a mortar and crush it. It is flattened out into a little plate of metal. It is soft, marks paper, and has all the characteristics of lead.

The production of lead from galena was familiar to the Alchemists. Moreover, since by suitable treatment they found how to extract a small amount of silver from a large amount of lead, they supposed that an actual change of galena into lead and then of lead into silver was effected. It does not seem ever to have occurred to the Alchemists that the galena at the outset perhaps contained lead and silver. Their error, for it is now known to be such, might have been avoided if they had paid any attention to the weights of the substances employed and obtained in their experiments.

EXPT. 2. To convert Zinc apparently into Copper.—Pour some solution of blue vitriol (copper sulphate) into a test-tube till it is one-third full. Cut some strips of zinc narrow enough to enter the tube, drop in one strip, and notice the change of appearance. Boil, and observe that the zinc disappears, while a red and black deposit of copper is formed, as though the zinc had been transmuted into copper. Pour the solution into another test-tube, add another strip of zinc, and see whether any more copper is deposited on boiling.

EXPT. 3. To convert Copper apparently into Silver.—Clean a strip of copper thoroughly by means of sand-paper. Place it for a moment in a solution of sulphate of mercury.¹ On removing the copper and rubbing it with a flannel, it acquires a bright white silvery surface, as though the copper had been transmuted into silver. Heat the strip in a flame, and the bright silvery appearance is destroyed. The copper therefore has not at any rate been entirely changed into another metal, although, at first sight, such would appear to have been the case.

These experiments show that first impressions and ideas as to phenomena may be very far from the truth, and should in every

¹ Nitrate of mercury may be used in place of the sulphate, though the result is not quite so good.

case be severely tested and supported by further experiment before they are finally accepted.

Conclusions must be tested by Experiment and by Measurement.—Another illustration of the way in which ideas may change is found in the opinions which have been held as to whether dead matter of itself can ever become alive.

The Roman poet Vergil¹ records the notion as being widely prevalent in Egypt, Persia, and the East, that bees could be generated from the putrid carcasses of oxen. He specially recommends a two-year-old bullock for the purpose, and prescribes the mode of killing it by suffocation and beating, then the pounding and mashing of it, hide and all, and laying it on thyme and cassia in a small shady chamber, with access of air at fit intervals, and so forth. Vergil does not tell us that he had seen it done, nor does he hint that the occasional selection by a swarm of bees of the hollow of a beast's skeleton as a convenient home may have originated the idea on which he enlarges with so much minuteness.

Our forefathers certainly believed that dead matter could beget life, and would have instanced vinegar, which breeds vinegar-eels. Yet this great question of spontaneous generation has since been answered by experiments, which show that all life is derived from pre-existing life—“*omne vivum e vivo.*”

The laborious and life-long studies of the Alchemists did much to increase the knowledge of chemical operations, of minerals and metals, and of chemical agents, but the true method of scientific investigation was not theirs. They were too ready to be satisfied with a shallow explanation of what they saw. They were not accustomed to weigh the materials with which they began an experiment, nor to determine the weight of anything left at the end. In fact, measurement was neglected by them. The result was that they were led to conclusions which were to a large extent erroneous. Nature does not reveal her secrets to the careless, nor to impatient inquirers. And until the chemist made the balance his constant servant, and was willing to test all his conclusions by rigorous experiment, there was and could be no sure foundation for the theories and laws which he endeavoured to construct.

¹ Vergil, *Georgics IV.*, 281-314.

CHAPTER II

THE METRIC SYSTEM. THE MEASUREMENT OF LENGTH AND OF AREA

Standards of Measurement.—It has already been insisted upon that measurement is the first and most important step in the progress of any science, and it will be observed that in the statement of any measurement there is the mention firstly of a number, and secondly of a thing of the same kind as the quantity to be measured, which is referred to as a *standard* or *unit*. For instance, if it is desired to state the sun's distance from the earth, the distance of the moon from the earth may be selected as a standard or unit, and it may be stated that the sun's distance is 400 moon's distances. Similarly, the moon's distance may be stated in terms of the earth's radius, as about 60 earth's radii.

But whilst such statements convey definite meanings to the mind, since in either case the magnitude of the unit chosen is comparable with the quantity to be measured, yet it would be very inconvenient in scientific work to state quantities in terms of standards not universally recognised, for it would be extremely troublesome to compare the statement of one man with that of another if they employed different standards. It is on this account that a statement of quantities in yards, gallons, or pounds is almost meaningless to a Frenchman or German. On the other hand, the "Metric System" of measurement, to which reference is made below, is understood by scientific men of all nationalities, and indeed it is almost exclusively employed in scientific investigations.

There is another objection to the English system for purposes of measurement in the fact that there is no simple relation between the different standards of length, volume, and weight. There is consequently much trouble in making calculations with quantities expressed by means of them. The Metric System, on the other hand, is a *decimal* system, and the different standards are related in a simple way to one another.

We shall now refer to the standards of Length, Area, Volume, and Weight in the Metric System, which are used in chemical investigations, and explain the nomenclature used in connection with this system.

1. STANDARD OF LENGTH.—The **Metre**¹ is the length at 0° C. of a certain bar of platinum kept in Paris.

In 1790, immediately after the French Revolution, a commission of savants nominated by the French Academy was appointed to prepare a new system of units. This commission decided in favour of a decimal system, with the unit of length, from which all other units should be derived, itself connected with the size of the earth. This happy idea put the unit beyond reach of change from temperature or from other causes.

The commission added to itself Lavoisier and other eminent men, and had a meridian arc from Dunkirk to Barcelona carefully measured, as a basis for calculating the size of the earth. The metre was intended to be $\frac{1}{10,000,000}$ of the meridian quadrant of the earth's circumference drawn through Paris, (Fig. 2). This would make the circumference of the earth 40,000,000 metres. A platinum bar of the required size was constructed. Since then a small error has been found in the calculations, so that this platinum metre is not exactly what it was intended to be. But this

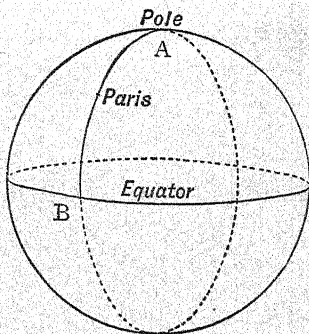


FIG. 2.

¹ Greek, *metron*, a measure.

metre bar is still taken as the ultimate standard of appeal, and careful copies have been made and distributed among the European nations.

Examine a Metre Rule. Notice firstly that there are a great number of short lines upon it. Moreover, groups of these are marked off by longer lines, which are numbered, and there are seen to be 10 divisions in each group. There are 100 of these groups, and therefore there are 1000 of the smallest divisions in the whole metre.

The width of one of the smallest divisions is called **1 millimetre** (Latin, *mille*, a thousand).

The width of a group of 10 millimetres is contained 100 times in the metre; hence this length is called **1 centimetre** (*centum*, a hundred).

The width of ten centimetres is contained 10 times in the metre; hence this length is named **1 decimetre** (*decem*, ten). Fig. 3 represents 1 decimetre; the long lines mark off centimetres, and the smallest divisions mark off millimetres.

Other names with Greek prefixes are given to lengths 10, 100, and 1000 times the metre. The most important of these is **1 kilometre**, which equals 1000 metres.

TO BE LEARNT BY HEART

1 millimetre	equals	$\frac{1}{1000}$	of a metre	; it is denoted by	mm.
1 centimetre	"	$\frac{1}{100}$	"	"	cm.
1 decimetre	"	$\frac{1}{10}$	"	"	dm.
1 metre				"	m.
1 dekametre	"	10	metres	"	Dm.
1 hectometre	"	100	"	"	Hm.
1 kilometre	"	1000	"	"	Km.

FIG. 3.

The student should try to remember roughly these metric measures by reference to real objects of known size. Thus:—

Kilometre equals distance covered by 10 minutes' quick walking.
Hectometre " " " length of football ground.

Dekametre	equals	distance	covered	by	half	length	of	cricket	pitch.						
Metre	"	"	"	"	"	"	"	width	of	door.					
Decimetre	"	"	"	"	"	"	"	width	of	hand.					
Centimetre	"	"	"	"	"	"	"	width	of	finger-nail	or	dice.			
Millimetre	"	"	"	"	"	"	"	nib	of	J	pen,	or	thickness	of	sixpence.

THE MEASUREMENT OF LENGTH

Obtain a small box-wood rule¹ marked with a centimetre scale along one edge, and an inch scale divided into tenths along the other. Obtain a sharp-pointed hard pencil.

EXERCISES²

1. Draw a line 3 inches long. Measure its length in millimetres, to the nearest whole millimetre.
2. Draw a line 2·5 inches long. Measure its length to the nearest millimetre.
3. Draw a line 4·3 inches long. Measure its length in centimetres and millimetres. Then state the length in centimetres and decimal.
4. Measure the distance between the bottom and top lines in your note-book in centimetres and decimal.³
5. Measure the length of your note-book in decimetres, centimetres, and millimetres. Then state the length in decimetres and decimal.
6. Measure the breadth of your bench in decimetres and decimal.
7. Measure the breadth of the laboratory in metres and decimal.

The Estimation of Hundredths of an Inch and of a Centimetre.—Observe an inch scale divided into tenths. It is easy to make measurements by means of it to the nearest $\frac{1}{10}$ inch. But we can be more exact. We may imagine each $\frac{1}{10}$ inch to be divided into 10 smaller but invisible divisions of $\frac{1}{100}$ inch each.

¹ They are provided by J. Raybone and Sons, Whitmore Street, Birmingham, or by J. Thornton, 109, Deansgate, Manchester.

² The eye should in each case be placed immediately above that part of the scale which is being read.

³ If a length is an exact number of cm., e.g. 13 cm., it should be stated as 13·0 cm.

Fig. 4 represents the edge of an inch scale.

$$\begin{aligned} AB &= 0.7 \text{ inch to the nearest } \frac{1}{10} \text{ inch} \\ &= 0.7 + \text{inch, i.e. between } 0.7 \text{ and } 0.8 \text{ inch} \\ &= 0.73 \text{ inch, estimating hundredths of an inch.} \end{aligned}$$

Similarly, a centimetre scale divided into millimetres can be read to the hundredth of a centimetre.

EXERCISES

NOTE.—In the following exercises read to 2 places of decimals.

1. Measure the distance between the bottom and top lines in your note-book in inches.

2. Number the lines in your note-book 1, 2, 3, 4, &c. Measure the distance between the tenth line and the first line in inches.

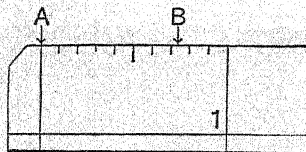


FIG. 4.

3. Measure the distance between the seventh line and the first line in inches.

4. Measure the distance of the first line from each of the following lines in inches. Tabulate your results.

5. Measure the distance between the tenth line and the first line in centimetres.

6. Measure the distance between the seventh line and the first line in centimetres.

7. Place a 10-to-the-inch scale against a scale divided into centimetres and millimetres. Read off in inches and decimals the length of 1 cm., 2 cm., 3 cm., 10 cm., estimating to the nearest hundredth of an inch. For instance,

$$1 \text{ cm.} = .4 \text{ inch.}$$

$$2 \text{ cm.} = .79 \text{ ,,}$$

8. Place two scales together as before. Read off in millimetres (and by estimation tenths of a millimetre) the lengths of 1, 2, 6 ins.

9. Measure the diameter of a penny in centimetres.

10. Take a strip of paper about 15 cm. long. Wrap it carefully round the edge of a penny until it overlaps, and then prick a small hole through both thicknesses of paper with a pin. Measure in centimetres the distance between the two pin-pricks on the paper. This

length is equal to the circumference of the penny. Then divide the length of the circumference by that of the diameter. State thus :—

$$\begin{array}{rcl} \text{Circumference} & = & \text{cm.} \\ \text{Diameter} & = & \text{cm.} \\ \therefore \frac{\text{Circumference}}{\text{Diameter}} & = & \text{--- cm.} \\ \text{,,} & = & \text{cm.} \end{array}$$

11. Measure in the same way the circumference of

- (a) a copper cylinder,
(b) a boxwood ball.

12. Repeat No. 11, but wrap the paper *three* times round the ball and cylinder.

13. Draw a circle of 5 cm. radius. Find the length of its circumference in cm. by means of a piece of cotton. Tie a knot near one end of the cotton. Mark any point on the circumference, place the knot upon it, and with finger and thumb gradually trace the cotton round to the mark again. Then pick up the cotton, and lay it straight along a metre scale. Lastly divide the length of the circumference by that of the diameter. State as in No. 10.

14. Measure the length of a road on an Ordnance Map by means of a piece of cotton, and then by reference to the scale of the map calculate the length of the road in miles.

EXAMPLES I

Metric Prefixes. Length

- How many centimetres make $\frac{1}{2}$ metre?
- How many millimetres make 1 decimetre?
- How many metres make 10 kilometres?
- How many millimetres make 1 kilometre?
- Express 1 centimetre as a decimal of 1 decimetre.
- Express 1 millimetre in kilometres.
- Express 45·6 cm. in dm.
- Bring 2·34 km. to cm.
- Subtract 1 mm. from 1 cm.; answer in cm.
- How many cm. are there in 123·456 metres?
- The diameter of a halfpenny is 25 mm. How many halfpennies in a row would stretch 1 metre?
- How many metre strides shall I take in running 5 kilometres?
- If a sixpence is 1 mm. thick, find the height in cm. of a pile of sixpences worth £1.

14. If a cricket pitch is 2 dekametres long, how many kilometres does a boy run in making a century?
15. A fishing-rod has 3 joints and a total length of 3.6 metres. Find the average length of each joint in cm.
16. How many hours will it take to walk 24 kilometres at the rate of 1 hectometre per minute?
17. How many pins, each 25 mm. long, can be made from 1 kilometre of wire?
18. Four pieces of 1 decimetre each are cut away from a metre of glass tubing. Into how many lengths of 15 cm. long can the remainder be divided?

2. STANDARD OF AREA.—Draw a square ABCD measuring 1 decimetre along each side (Fig. 5). This is called **1 square decimetre**. Mark off centimetres along each side. Through the points of division rule lines parallel to the sides of the square.

The square decimetre is then divided into smaller squares, each **1 sq. centimetre**. How many sq. cm. does it contain? There are 10 rows, each containing 10 sq. cm., or 100 sq. cm. altogether. The sq. cm. AEFG has been similarly divided into sq. mm. Thus,

$$1 \text{ sq. dm.} = 100 \text{ sq. cm.}$$

$$1 \text{ sq. cm.} = 100 \text{ sq. mm.}$$

Observe that the area 1 sq. cm. is roughly represented by the area of a finger-nail.

A square millimetre, a square centimetre, a square decimetre, a square metre, or a square kilometre may in different cases be used as the unit of area. Thus the area of a country may be suitably expressed in square kilometres, the area of a playground in square metres, and the area of the wing of a bee in square millimetres.

The Measurement of Rectangular Areas.—Obtain some paper divided by straight lines into sq. mm.

(1) With a sharp-pointed pencil draw upon it a sq. centimetre. It contains 100 sq. mm.

(2) Draw upon it a rectangle 8 cm. long and 34 mm. broad. Find its area by *counting* the sq. mm. which it contains. It is

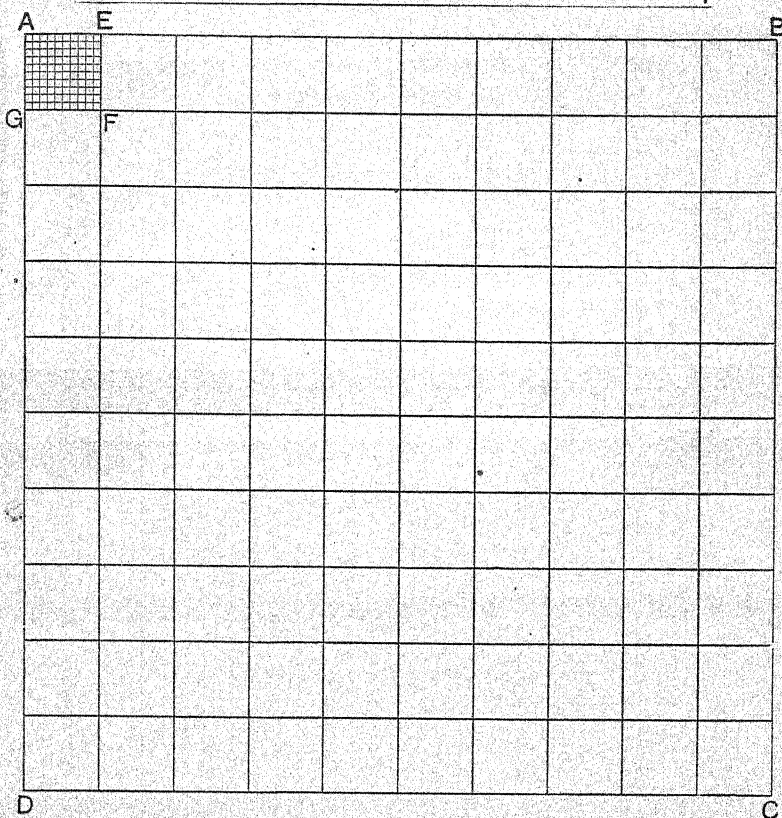


FIG. 5.

seen to contain 24 blocks of 100 sq. mm., and also a long strip containing 320 sq. mm. Thus :

$$\begin{aligned} \text{Area} &= 2400 + 320 \text{ sq. mm.} \\ &= 2720 \text{ sq. mm.} \end{aligned}$$

Observe that we can also calculate the area by considering that it contains 34 rows of sq. millimetres, with 80 sq. millimetres in each row, or

$$\begin{aligned} \text{Area} &= 34 \times 80 \text{ sq. mm.} \\ &= 2720 \text{ sq. mm.} \end{aligned}$$

(3) Draw a rectangle 6.9 cm. long and 5.3 cm. broad. What is its area?

$$\begin{aligned}\text{Area} &= 69 \times 53 \text{ sq. mm.} \\ &= 3657 \text{ sq. mm.} \\ &= 36.57 \text{ sq. cm.}\end{aligned}$$

These examples will serve to show that in the case of any rectangular surface

$$\text{Area} = \text{length} \times \text{breadth.}$$

Care must be taken to express both length and breadth in terms of the *same* unit.

EXERCISES

1. Obtain some paper divided by straight lines into square millimetres. Draw upon it a rectangle 5.5 cm. long and 24 mm. broad. Find its area

- (a) in sq. mm.
- (b) in sq. cm.

2. Draw a square inch upon the same paper. How many sq. mm. does it contain?

3. Obtain some paper divided by straight lines into square inches and hundredths of a square inch. Draw upon it a rectangle 2.2 inches long and 0.8 inch broad. Find its area

- (a) in hundredths of a square inch
- (b) in square inches.

4. On the same paper draw a rectangle 2.8 inches long and 1.7 inches broad. Find its area in square inches.

5. Open your note-book at the middle. Measure the length and breadth of the inside sheet in cm. (to 1 place of decimals). Calculate its area in sq. cm. and decimal.

6. Find the area of (a) a sheet of paper, (b) a postcard, in square inches and decimal.

7. Find the area of a visiting card in sq. cm. and decimal.

8. Trace the outline of a nasturtium leaf upon some paper divided into sq. mm. Calculate its approximate area in sq. mm.

EXAMPLES II

Metric Square Measure. Area

1. How many sq. dm. are there in 1 square metre?
2. How many sq. cm. are there in 1 square metre?
3. Multiply 3 dm. by 2 cm.; answer in sq. cm.

4. Multiply 80 mm. by 30 mm ; answer in sq. cm.
5. If 1 inch equals 25 mm., how many sq. mm. equal 1 square inch ?
6. Find in sq. cm. the area of a postcard 12.5 cm. long and 80 mm. broad.
7. How many pieces of zinc each 25 cm. by 0.8 cm. can be cut from a sheet 16 dm. by 8 cm. ?
8. How many postage stamps each 24 mm. by 20 mm. are required to paper a wall 4 metres by 3 metres ?
9. If the pressure of the air is $2\frac{1}{2}$ lbs. per sq. cm., what is that in cwts. per square metre ?
10. If 1 dm. equals 4 in. long, express 1 square metre in square yards.

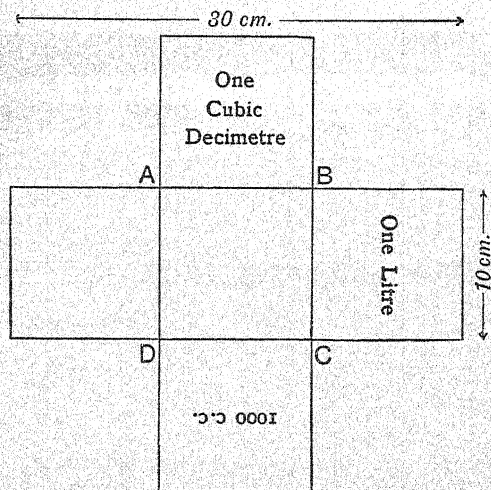


FIG. 6.

3. STANDARD OF CUBIC CONTENTS OR VOLUME.

—The unit of volume most used in laboratory work is the volume of a cube each edge of which measures 1 cm. long. This is called 1 cubic centimetre : it is denoted by 1 c.c.

EXERCISE. To cut out a Cubic Centimetre.—Cut out of soap a cube, each edge of which shall measure 1 cm

Make a mental note of its size. It is about that of a die. The capacity of a small thimble is also about that of 1 c.c.

EXERCISE. To make a Cubic Decimetre.—Cut out a stiff cardboard cross, 30 cm. by 10 cm. (Fig. 6.) Print inscriptions on the faces as indicated in the figure. Cut the card half-way through along each of the lines AB, BC, CD, DA. Bend back the four squares and stick their edges together with gummed paper both inside and outside. Let the gum dry. Then warm the hollow cube, and paint it over with melted paraffin.

Similarly make out of cardboard a cubic centimetre.

How many cubic centimetres could be packed into the cubic decimetre? Imagine first that the bottom of the cubic

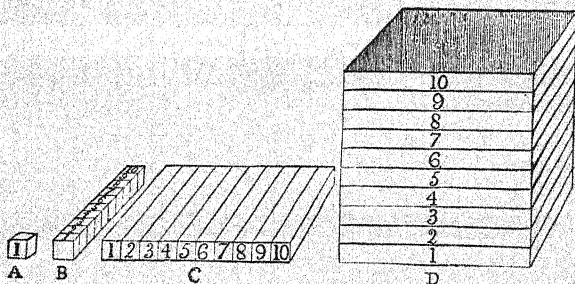


FIG. 7.—Ten cubes like *A* would make a row like *B*. Ten rows like *B* would make a layer like *C*, and ten such layers would go into the box *D*.

decimetre were paved with cubic centimetres, how many would cover the bottom? Clearly 10 rows of cubic centimetres with 10 cubic centimetres in each row, or 100 altogether, would be necessary. And ten of these layers piled one on another would be required to fill the cubic decimetre (see Fig. 7). Thus $10 \times 10 \times 10$ or 1000 cubic centimetres equal 1 cubic decimetre.

The capacity of 1 cubic decimetre is called 1 litre. On the Continent milk and oil and other liquids are measured in litres.

Take a large flask which will contain just 1 litre of a liquid, fill it with water, and find how many times you can fill a tumbler from it. It can be filled four times, so the capacity of a tumbler is about $\frac{1}{4}$ of a litre.

Fill a litre flask with water from a pint measure. It can contain about $1\frac{3}{4}$ pints.

The multiples and decimals of a litre are named like those of the metre; for instance, a *decilitre* means $\frac{1}{10}$ of a litre, and remembering that 1 litre = 1 cubic decimetre = 1000 c.c., we may construct the following table:—

1 litre	=	1000 c.c.
1 decilitre	= $\frac{1}{10}$ litre =	100 c.c.
1 centilitre	= $\frac{1}{100}$ „ =	10 c.c.
1 millilitre	= $\frac{1}{1000}$ „ =	1 c.c.

We have thus two names for $\frac{1}{1000}$ litre; it may be called either 1 millilitre or 1 cubic centimetre. The latter is more often used.

It has already been shown that the volume or capacity of any rectangular solid is obtained by multiplying the area of its base by its height. Thus, if a biscuit-box is 5 cm. broad, 8 cm. long, and 6 cm. high, its capacity is $5 \times 8 \times 6$ or 240 cubic centimetres. Of course, before multiplying breadth \times length \times height, the numbers expressing the lengths of these must all be in terms of the *same* unit.

EXERCISES

Find the volume of the rectangular block of wood provided—

- (a) in cubic inches,
(b) in cubic centimetres.

Measure inches to 2 places of decimals and centimetres to 1 place of decimals.

State thus:—

Length	=	inches
Breadth	=	„
Height	=	„
\therefore Volume	=	$\times \times$ cubic inches
	=	cubic inches

EXAMPLES III¹

Metric Cubic and Capacity Measure

1. How many c.c. are there in 1 cu. dm. ?
2. A cube measures 30 cm. each way ; what is its volume in c.c. ?
3. How many cu. mm. are there in 1 cubic inch, if 25 mm. equal 1 inch long ?
4. How many cu. mm. are there in 1 c.c. ?
5. How many c.c. are there in one cubic metre ?
6. How many c.c. are there in 3 litres ?
7. How many c.c. are there in 2 decilitres ?
8. Express 1234 c.c. in litres.
9. How many c.c. are there in 56·78 litres ?
10. Express in c.c. 1 litre + 2 decilitres.
11. The lead of a pencil is 2 mm. broad, 2 mm. thick, and 150 mm. long, what is its volume in cu. mm. ?
12. Find in c.c. the contents of a biscuit-box 15 cm. high, 12 cm. broad, and 240 mm. long.
13. Find in litres the contents of a box, 8 cm. deep, 25 cm. long, and 10 cm. broad.
14. How many times can a dekalitre bucket be filled from a trough 2 metres long, 50 dm. broad, and 40 cm. deep.
15. 600 cu. ft. of air each is allowed in a dormitory for 20 people. How many litres is that for all ? (1 ft. long = 30 cm.)
16. How many days will a hectolitre cask of water last for 7 people, allowing each 3 pints per day (1 litre = 1 $\frac{3}{4}$ pints).

4. STANDARD OF WEIGHT.—The unit or standard of weight in the Metric System is the **gram**. One gram (gm.) is defined as the weight of 1 cubic centimetre of pure water.²

A penny weighs about 10 grams.

It has already been stated that 1 gram is the weight of 1 c.c. of water. There is thus a very simple relation between the measures of the volumes and weights of given quantities of water. For instance, 1000 c.c. of water should weigh 1000 grams. Test this as follows:—Take a litre (1000 c.c.) flask, and counterpoise it on a strong balance; then add 1000 grams to

¹ *To the Teacher.*—All the questions in this set and in the succeeding sets of Examples in this chapter should not be worked on first studying the subject. Some of the questions should be set in the course of the study of the next few chapters.

²Strictly speaking at 4° C., the temperature of maximum density of water.

the counterpoise, and fill the flask with water; the balance is still maintained.

There is no such simple relation between the English measures of volume and weight.

Multiples and decimals of 1 gram are named *as before*. There is therefore no difficulty in learning the following table, whereas Avoirdupois Weight is not made at all easier by having previously learnt Long Measure :—

1 kilogram	=	1000 gm.
1 hectogram	=	100 gm.
1 dekagram	=	10 gm.
1 gram		
1 decigram	=	·1 gm.
1 centigram	=	·01 gm.
1 milligram	=	·001 gm.

EXAMPLES IV

Metric System. Weight

- How many decigrams are there in 1 gram?
- How many milligrams are there in 1 centigram?
- Express 7891 grams in kilograms.
- What decimal of a gram are 2 decigrams?
- What decimal of a gram are 50 milligrams?
- Express 6·54 kilograms in grams.
- What decimal of a gram are 100 milligrams?
- Express 3 centigrams + 4 milligrams in milligrams.
- Express 2 decigrams + 5 centigrams in milligrams.
- Add 3 gm. + 2 dgm. + 50 mgm. Express in grams.
- Add 25 gm. + 200 mgm. + 2 mgm. Express in grams.
- Add 1 Kgm. + 2 Dgm. + 3 gm. + 4 dgm.
- A flask weighs 20 gm. when empty, and 45 gm. when full of water. What is its capacity?
- A flask which will hold 80 c.c. of water weighs 107 grams when full of water. How many grams will it weigh when empty?
- A litre flask weighs 150 gm. when empty. How many grams will it weigh when full of water?
- A biscuit-tin measures 16 cm. by 25 cm. by 20 cm. How many grams of water will it hold?
- Mercury is $13\frac{1}{2}$ times as heavy as water. How many grams of mercury will fill a decilitre flask?

18. Sulphuric acid is 1·8 times as heavy as water. How many c.c. of sulphuric acid will weigh 99 grams?

19. If 5 cm. of wire weigh 25 milligrams, how many grams will 10 metres of the wire weigh?

20. If 1 square centimetre of metallic foil weighs 1 centigram, how many grams will 1 square decimetre weigh?

The Relations between English and Metric Units :—

Some rough equivalents are worth remembering :—

1 foot	equals	30 cm.
1 inch	„	25 mm.
1 decimetre	„	4 in.
1 kilometre	„	$\frac{5}{8}$ mile.
1 ounce	„	28 grams.
1 litre	„	1 $\frac{3}{4}$ pints.
1 kilogram	„	2 $\frac{1}{2}$ lbs.

More exact numbers are useful for reference :—

1 metre	equals	39·37 inches.
1 foot	„	30·5 cm.
1 gram	„	15·43 grains.
1 grain	„	·06489 grams.
1 kilogram	„	2·20 lbs.
1 litre	„	1·76 pints.
1 litre	„	61·03 cu. ins.
1 cu. in.	„	16·38 c.c.

Note.—If 1 gram equals 15·43 grains, then 1 grain equals $\frac{1}{15·43}$ gram, or ·06489 gram.

(The numbers 15·43 and ·06489 are called *reciprocals*.)

Some Disadvantages of the Metric System.—The advantages of the Metric System over any existing system for the purposes of scientific investigation are great. It is understood by men of all nationalities, and the decimal relations which exist between the units are of great convenience in calculations.

It is not, however, urged here that the Metric System is the best for all the ordinary transactions of life. Indeed, as Herbert

Spencer has recently pointed out, there is some ground for the belief that a decimal system is not the most convenient for the purposes of petty trade. Quoting from a letter which appeared in the *Times*, he gives the following instance: "In England, if an old village woman buys a measured quantity involving fractions of an article at a price reckoned, say, in shillings, pence, and halfpence, the shopman or girl never seems to find any difficulty in working quickly in his or her head the sum which has to be paid, and the old woman manages to see that she is charged correctly. But, in France, with the decimal system, it seems always to be necessary to work the simplest sum out on a piece of paper. In fact, it is not difficult to see that it is pretty easy to work out mentally the value of, say, a yard and three-quarters of riband at $3\frac{1}{4}$ a yard, while it is very difficult to work out mentally the value of, say, 1 m. 75 centimetres of riband at 35 centimes a metre."

Moreover, in different trades different units are found to be convenient, and though the Government of France enforced as long ago as 1839 the observance of the metric system, it is remarkable that its use is even to-day not general in France. "Precious stones are to-day bought and sold in carats; fire-wood in cordes; milk in pintes; gravel in toises; grain, potatoes, and charcoal in boisseaux; sugar and tea among the poor people is dealt with in livres, demilivres, &c."

EXAMPLES V

Metric System. Miscellaneous

(a)

1. How many pieces of glass rod, each 5 cm. long, can be cut from a piece 1 metre long?
2. How many times can 10 c.c. of water be obtained from a half-litre flask?
3. If 1 metre of wire weighs 4 decigrams, how many milligrams will 5 cm. weigh?
4. A flask weighs 20 grams when empty, and 120 grams when full of water. What will the flask weigh when full of mercury, of which 1 c.c. weighs 13.5 grams?
5. A train is travelling 100 kilometres per hour. What is that in centimetres per second?

6. If a metal decigram weight measures 1 sq. cm., find in grams the weight of a piece of metal of the same thickness, measuring 4 mm. by $2\frac{1}{2}$ mm.

(b)

7. Add, expressing the answer in metres :—
 1 Km. + 2 mm. + 3 cm. + 4 Hm. + 5 Dm. + 6 dm.
8. Find in sq. cm. the area of a post-card measuring 1.25 dm. by 80 mm.
9. Find in cu. dm. the contents of a box measuring 150 mm. high, 12 cm. broad, and 2.4 dm. long.
10. How many grams of soda are contained in 10 c.c. of a solution, 1 litre of which contains $\frac{1}{25}$ of a kilogram of soda?
11. If 1 metre of wire weighs 120 centigrams, what will be the length in centimetres of 60 milligrams of the same wire?
12. Find to 2 places of decimals the number of centimetres in 1 foot long, assuming that 1 metre = 39.37 inches.

(c)

13. Add together $\frac{1}{2}$ litre, 20 c.c., 200 millilitres, and 12 decilitres, and express the sum in c.c.
14. If a bath is 17.1 metres long, 5.1 metres wide, and 1.5 metres deep, how many times would a bucket, containing 8.1 litres, have to be filled before the bath was *half* empty?
15. If I deal out to each of 40 boys 3 pieces of wire, each piece being $\frac{1}{2}$ metre long, and if a metre of wire weighs 4 decigrams, what weight of wire have I dealt out?
16. If the diameter of a sixpenny piece is 2 cm. long, and it is 1 mm. thick, calculate what would be the value of a hectometre of them placed in a *row*.
17. Calculate how high the same number of coins would reach, expressed in metres, if they were placed in a *pile*.
18. If a bicyclist rides 10.5 miles in 42.5 minutes, and if 8 kilometres = 5 miles, how many kilometres an hour does he travel on the average?

(d)

19. Add 2 kilograms, 400 grams, 200 milligrams, 5 decigrams, 1 centigram. Answer in grams.
20. If a square centimetre of platinum foil weighs 1 centigram, how many grams will a square metre weigh?

21. How many times can a decilitre flask be filled from a pneumatic trough 10 cm. broad, 8 cm. deep, and 25 cm. long?

22. A solution contains 1 ounce of salt in 1 pint of water. How many grams of salt per litre of water would make a solution of the same strength?

23. If a 100 c.c. flask weighs 120 grams when full of water, what will it weigh when full of mercury, which is $13\frac{1}{2}$ times as heavy as water?

24. If 30 cm. = 1 linear foot, and 1 cubic foot of water = 1000 ounces, find the weight of 1 litre of water in ounces.

CHAPTER III

THE MEASUREMENT OF MASS—THE BALANCE

THE mass or quantity of matter in a substance is measured by comparing it with standard masses. These standard masses are usually called "weights." Hence, the measurement of mass is commonly spoken of as "weighing." The measurement of mass is usually carried out with an apparatus called a *balance*.

EXERCISE.—To make a Simple Balance.¹—Take a box-wood scale, and drill a hole near one edge at the division

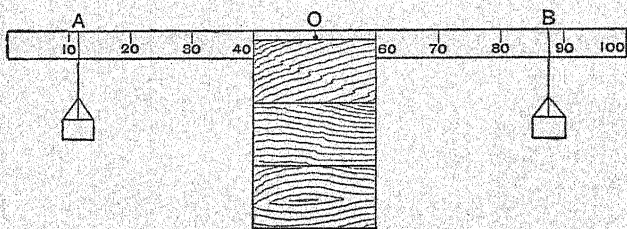


FIG. 8.

marked 50, and fix a steel knitting-needle tightly through the hole. Place two piles of wooden blocks, with their faces parallel to one another, and about 1 inch apart. Let the knitting-needle rest on the upper surfaces of the blocks with the scale between them. Does the scale balance horizontally? If not, wind some tin-foil round one end until it does. Take two pill-boxes without lids and arrange to suspend them by thread from

¹ Adapted from *Weighing and Measuring*, by E. Edser.

the scale. Hang them at equal distances from O, the middle of the scale, and on opposite sides of it. Does the scale still balance horizontally? If not, add sufficient tin-foil to one or other pill-box (Fig. 8).

Remove the two boxes and put 20 gm. into each. Hang one box 40 cm. from O on one arm. Find where the other box must be hung that the scale may balance horizontally: it has to be placed 40 cm. from O on the other arm. Now hang the first box in a new position: again, the second box must be placed at an equal distance from O to maintain equilibrium. It is clear that *equal weights will only balance one another when at equal distances from the point of support.*

Again, hang the two boxes 40 cm. from O, and on opposite sides of O. Place 30 gm. in one box, and add gram weights to the other until there is equilibrium. Is it not necessary to add an equal weight, viz. 30 gm.? Repeat with some other weight in the first box. It becomes clear that *two substances must have equal weights if they balance one another when at equal distances from the point of support.*

Suppose we desire to find the weight of some copper turnings. Place them in one box at A. Place the other box at B, such that the length BO = the length AO. Add standard masses or "weights" to the box until there is equilibrium. Then we may conclude that the weight of the copper is measured by the amount of the standard weights in the box.

Of all the instruments which are at the present day in the service of the chemist the balance is the most important. Today the balance is to the chemist as his very right hand, and without it the rapid progress of the last century would have been impossible. The use of the balance has placed chemistry among the exact sciences, and so far as may be chemical problems are now put to the test in the laboratory, and not referred for an answer to the logician in his study. How different our opinion of the balance is from that of Jean Rey, a shrewd physician, who said in 1630:

"I affirm that the examination of weights which is made by the Balance differs greatly from that which is made by the reason. The latter is only employed by the judicious, the former can be practised by the veriest clown. The latter is always exact; the former is seldom without deception."

The balance is a delicate instrument, and Jean Rey's clown cannot be allowed to play with it, or it will soon be injured.

The Student's Balance.—We shall now describe the balance to be used in almost all the experiments and problems described in this book.

It consists of a beam A B (Fig. 9) which can be raised from or lowered on to a knife edge at C by the handle D. At its two

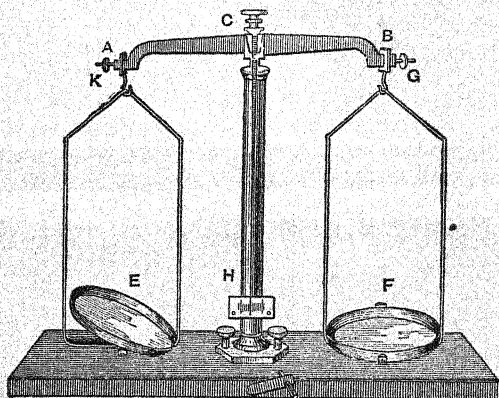


FIG. 9.

ends are suspended pans E and F. A needle H is attached rigidly at right angles to the beam, and its point is in front of a graduated scale. There are screw nuts at K and G by means of which the lengths of the arms can be slightly altered. The nuts K and G should be so adjusted that when the pans are empty the needle swings over the same number of degrees to the right and the left of the zero point at the centre of the scale.¹

Remove the pans from a balance. Examine the means of

¹ Several firms sell balances of this type. For laboratory work it is recommended to obtain them fitted with glass cases, whereby the balance is protected from the disturbing influence of air currents, and the still more dangerous effects of corrosive chemical fumes.

A Rough Balance.—It is often only desired to know quickly the approximate weights of substances. A very convenient balance is Salter's Letter Balance, No. 11, which will carry 1000 grams. (G. Salter and Co., West Bromwich. 5s. 6d.)

Another convenient balance is one built on what is known as the French type.

support, including the knife edges. Make a drawing of the way in which the pan is supported.

The Weights.—In a set of weights from 1 centigram up to 50 grams the following individual weights are usually found :—

BRASS	{	50	20	20	10
		5	2	2	1
FLAT WHITE	{	0.5	0.2	0.2	0.1
METAL	{	0.05	0.02	0.02	0.01

Notice that there are just enough weights to add up to any sum from .01 to 111.10 grams.

For example, if a substance weighing 67.89 grams were weighed, the weights would be made up as follows :—

$$\left\{ \begin{array}{l} + 50 + 10 \\ \quad + 5 + 2 \\ \quad \quad + .5 + .2 + .1 \\ \quad \quad \quad + .05 + .02 + .02 \end{array} \right.$$

The gram weight and those greater than the gram are usually made of brass.

The Smaller Weights, which are always used in accurate weighing, are differently marked by different makers. They are made of aluminium, nickel, or platinum.

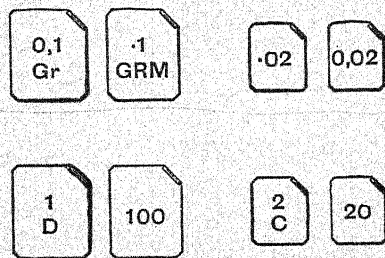


FIG. 10.

The smaller weights shown above in Fig. 10 represent :—

1 gram	.02 gram	.01 gram
= 1 decigram	= 2 centigrams	= 1 centigram
= 100 milligrams	= 20 milligrams	= 10 milligrams.

* **The Rider.**—The .01 gram or centigram weight is often a small bent wire called a *rider*, because it can ride astride the

balance beam. When placed at different points along the beam, its effective weight changes, and makes it possible to weigh to milligrams or even to tenths of a milligram. The beam of an accurate balance is divided into ten equal spaces. When the rider is put on the outermost of these, it is worth its full weight, '01 gram. On the middle division it acts as '005 gram, and on the innermost division as '001 gram.

EXERCISE.—Remove all the flat weights from the box. Arrange them on the squares drawn in Fig. 11. Ask your

.5	.2	.1
.05	.02	.01

FIG. 11.

teacher whether you are right. Rearrange the weights in their box. They should be in order of magnitude. Never force a weight into a hole too small for it. Give each weight a separate space if you can. The bent-up corner should be at the upper right-hand corner. Ask your teacher whether this has been properly done.

The Method of Weighing.—The following points should be observed :—

1. Always turn the handle D gently, and right over from one side to the other.
2. See that the pointer is swinging equally upon both sides of the scale, when the pans are empty. [Why?]
3. Stop the swinging of the balance whenever the object or the weights are added to or removed from the pans. [Why?]
4. Always place the object to be weighed in the left-hand pan, and the weights in the right-hand pan. [Why?]
5. The position of the observer should be opposite the centre of the scale. [Why?]

6. First add the largest weight necessary, and then try the smaller ones in *descending order of magnitude, omitting none*, but removing those that are too heavy before substituting lighter ones, and arranging the weights on the pan in order of magnitude, with the larger in the centre.

7. When the object is exactly balanced, write down the weights missing from their places in the box, and also the weights on the balance pan, and see that the two agree.

8. Never weigh anything which has been heated, until it is cold.

EXERCISES

Weighing

1. Find whether the balance swings true, the pointer vibrating to equal distances on each side of the middle point of the scale. Probably it does not. Take a square decimetre of paper. Tear it in two. Letter one half A. Tear the other half in two. Letter one part B. Go on bisecting; letter the weights so made A, B, C, . . . H. Use these weights to counterpoise the balance until it does swing true.

2. If the error of the balance is less than a centigram, use a centigram rider, and without any other weight find what weight would adjust the balance.

3. Pick out weights amounting to 87.65, arranging them neatly on the balance pan, the largest in the centre.

4. Add up the weights on the balance pan.

5. Add up the spaces in the box.

6. Take a set of English *grain* weights, 600 grains to 10 grains, and find the weight of each in grams to the nearest decigram *next below the true weight*. Write answers thus—11.9 +.

7. Repeat exercise 6 to the nearest centigram *next below the true weight*.

8. Cut a piece of lead to weigh exactly 10 grams.

9. Cut a piece of lead foil to weigh exactly 1 gram.

10. Weigh newly minted French (or English) silver coins.

11. Weigh the glass stopper¹ provided.

12. Weigh the copper cylinder¹ provided.

[Further exercises in weighing will be found in the succeeding chapters. Excellent exercises will be afforded by determinations of relative density.]

¹ The teacher is recommended to keep a stock of glass stoppers and copper cylinders of ascertained weight.

CHAPTER IV

THE MEASUREMENT OF THE VOLUME OF A LIQUID

The Units of Volume.—These are the litre for large volumes, and the cubic centimetre for small volumes. The litre is equal to 1000 cubic centimetres.

It is very important to remember that 1 c.c. of water weighs 1 gram.¹ Consequently, we may regard any given number of grams of water as occupying a volume expressed by the same number of c.c. For example, 1 litre or 1000 c.c. of water will weigh 1000 grams. The convenience of this relation will be learnt from the following experiment.

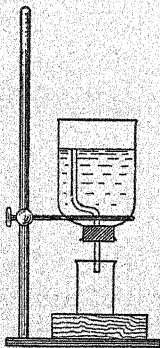


FIG. 12.

EXERCISE 1. To find the Volume of a Pebble.—We shall know its volume if we can find out how much water it can displace. Take a small bell jar with an open neck. Fit into the narrow end a rubber stopper, through which is passed a narrow glass tube, bent as shown in Fig. 12. Fix it in a clamp. Place a tumbler underneath, and pour water into the jar until it is full up to the mouth of the tube. Tie a thread

round the pebble. Weigh a beaker. Place it below the lower end of the tube. Carefully lower the pebble into the water

¹ This is strictly true only at 4° C. The average temperature of a room is about 15° C. Hence, since water expands slightly between 4° C. and 15° C., 1000 c.c. of water will at 15° C. weigh slightly less than 1000 grams.

until it is quite immersed. The water overflows, and is caught in the beaker. Weigh again. Example—

Weight of beaker... ..	8.53 gm.
Weight of beaker + water which overflowed... ..	29.97 "
∴ Weight of water which overflowed	21.44 "
and Volume " " "	21.44 c.c.
Hence, volume of pebble	<u>21.44 c.c.</u>

It is now evident that we can readily find the volume of a flask by first weighing it when empty, and secondly when filled with water. For the increase in weight expressed in grams will measure its volume expressed in cubic centimetres.

EXERCISE 2. To find the Volume of a Flask up to a mark on the neck. Gum a label on the neck, and make a horizontal mark on the label. Weigh the flask. Fill it with water up to the neck. If the top surface of the water in the neck of the flask be observed, it is seen to be curved (Fig. 13). This curved surface is called the **meniscus**. If there is a horizontal mark on the neck of a flask, it is said to be full up to the mark when the *lowest* point of the meniscus is on the level of the mark. Carefully bring the meniscus to the mark. Enter results thus :—

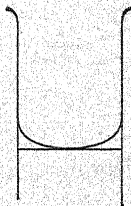


FIG. 13.

Weight of flask	24.21 gm.
Weight of flask filled with water... ..	106.42 "
Hence, weight of water in flask	82.21 "
∴ Volume of flask	<u>82.21 c.c.</u>

Measuring Flasks.—In carrying out experimental work it is found convenient to have flasks with which to measure out a definite volume of liquid, for instance, 100, 250, 500, 1000 c.c.

EXERCISE 3. To mark a Flask to contain 100 c.c.—Take a dry 3-ounce flask with a narrow neck. Weigh empty. Then add water till there is an increase in weight of 100 grams. Allow the flask to stand until any water on the side of the neck has drained down. With a sharp file moistened with water

scratch a horizontal line on the neck to mark the lowest point of the curved surface bounding the water and the air.

The measuring flasks most often used are litre, $\frac{1}{2}$ -litre, $\frac{1}{4}$ -litre, and 100 c.c. flasks (Fig. 14). They are usually fitted with ground-glass stoppers, and when filled up to the circular line etched upon their necks contain at 15° C. the volumes which are marked upon them.

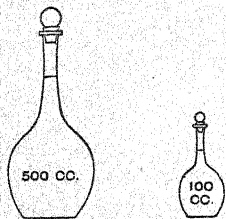


FIG. 14.

Measuring Jars.—These are tall cylindrical vessels, graduated in cubic centimetres, and holding 100, 200, 250, 500, or 1000 c.c. (Fig. 15).

The student will understand how to read the level of a liquid in a measuring jar after studying the use of a burette.

Pipettes.—Flasks and measuring jars measure the liquids they *contain*, but it is extremely useful to have vessels which will measure out, or *deliver*, accurately a small definite volume of liquid.

A pipette is a glass tube open at both ends, upon which a cylindrical bulb has been blown (Fig. 16). There is a circular mark on the stem. When filled with liquid up to the mark, the pipette will deliver exactly that volume which is marked upon it.

EXERCISE 4. To use a Pipette.—Place the pointed end of the pipette in water. Suck up the water till above the mark. Quickly close the upper end with the moist thumb. Pressing gently, rotate the pipette so that the water very slowly flows out until the meniscus reaches the mark. Then press tightly. The flow ceases. Transfer the pipette to a beaker, remove the thumb, and let the liquid flow out. As soon as the pipette appears to be empty, let it drain for five seconds, with the tip just below the surface of the water in the beaker. The drop of water which is still in the pipette is *left*, and not blown out.

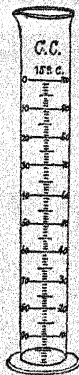


FIG. 15.

EXERCISES

Use of Pipette

1. Suck up water in a pipette. Set to the mark.
2. Take a 10 c.c. pipette and a small beaker (2 oz.). Find how many 10 c.c. of water can be added to the beaker.
3. Take a 10 c.c. pipette and a 100 c.c. flask. Find the level of the water after you have added 10 c.c. of water to the flask 10 times.
4. Repeat No. 3, using a 100 c.c. measuring jar in place of the flask.
5. Weigh an empty beaker. Measure into it 5 quantities of water of 10 c.c. each by means of a pipette. Weigh again. What is the gain in weight?

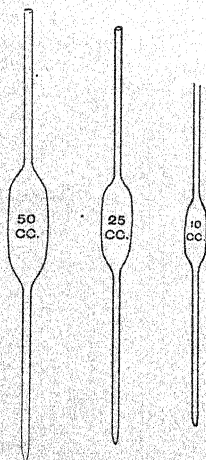


FIG. 16.

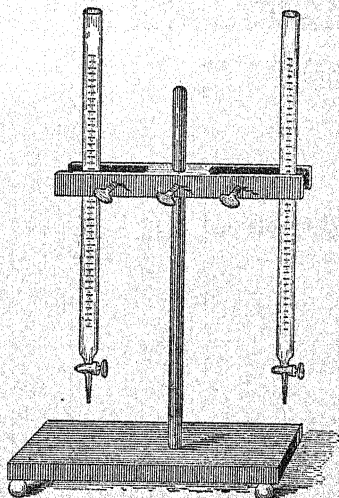


FIG. 17.

Burettes.—A burette is a glass tube with a glass tap at the lower end, used to deliver any desired volume of liquid (Fig. 17). The tube is graduated in cubic centimetres and tenths of a cubic centimetre. In place of a glass tap a burette may be fitted with a glass jet attached by an india-rubber tube, on which is

placed a spring clip; such a burette is more suitable for school use.

EXERCISE 5. To prepare a Burette for Use.—Fix the burette upright in a clamp. Place a spring clip upon the rubber tube. Pour water into the tube through a small funnel, until its level is above the top of the scale. Remove the funnel. Open the clip, and let some water run out, so that no air is left in the burette below the clip. The burette is now ready to deliver a measured volume of water.

Examine the divisions and figures upon the burette. Draw the portion from 41 to 43 four times the actual size.

Burette-Reading.—The surface of a liquid in a burette, like that in a flask, is curved, and is called the *meniscus*; it is always

the *lowest point* of the meniscus which is measured. Note the following points:—

(1) *The eye must be on the same level as the curve, otherwise the curve is seen displaced above or below the true position (see Fig. 17 A).* The level position may be obtained thus. Stand a few paces back from the

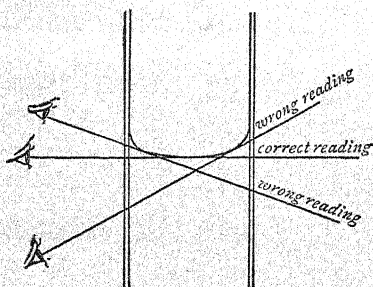


FIG. 17 A.

burette, and decide upon some object behind the burette, which appears to be on the same level as the meniscus. Then go up to the burette, and place the eye so that the meniscus is seen exactly on a level with the object selected.

(2) The extreme edge, or *lowest position*, of the meniscus is, however, sometimes hard to see, but it can be made clear. On looking into an aquarium from below the surface-line gold-fish are seen reflected above the surface. This phenomenon of reflection from the inside of a surface is called *internal reflection*, and it is this which may be made use of to render the meniscus distinct. If the background is not bright, the meniscus is made clear by holding a white object *just behind and just below the surface*. The edge of a card held behind the burette

tube answers admirably. At night, however, when the background is dark, a gas flame may be put on the bench below to illuminate the meniscus.

(3) The graduations are decimal, each fifth and tenth line being longer than the rest. Notice the figures opposite the divisions. Find whether they count upwards or downwards. Find how many c.c. or what decimal of 1 c.c. each scale-division stands for.

EXAMPLES VI

Burette-Reading

1. Read the position of the liquid in the figure of a tube shown on next page (Fig. 18), taking the two long lines W and X to represent respectively :—

A	B	C	D	E	F	G	H
30	40	7	800	120	130	250	250
and 20	50	6	900	100	150	200	300

For instance, the two lines W and X represent 30 c.c. and 20 c.c. respectively in question A, 40 c.c. and 50 c.c. in question B, and so on.

2. Sometimes the lowest point of the meniscus is not exactly opposite to any of the scale-divisions; tenths or other fractions of a division should then be *estimated* (see page 9).

Read the position of the liquid shown in Fig. 19, estimating the position of the curve in tenths of a scale-division, taking the two long lines Y and Z to represent respectively :—

J	K	L	M	O	P	Q	R
1	14	10	30	40	700	800	0
and 2	13	11	20	50	600	900	1000

EXERCISES

Burettes, Measuring Jars, and Pipettes

(a) Burettes

- (1) Pour water into a burette. Set it exactly at 10 c.c.
- (2) Pour some water out of the burette. Read the new position of the liquid.

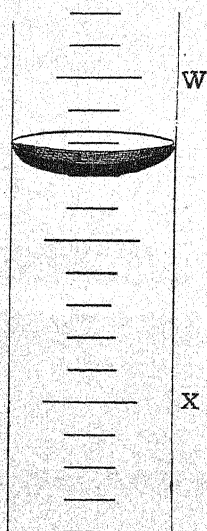


FIG. 18.

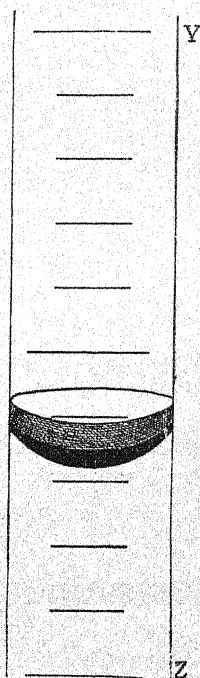


FIG. 19.

(3) Choose a 100 *grain* pipette and a 50 c.c. burette, graduated to tenths of 1 c.c. Set the liquid in a burette to 50. Measure 100 grains of water in the pipette, and run it into the burette. Read the level now. Run in another 100 grains, read again, and so on until you read the top of the burette. Take the difference between each two consecutive readings. These differences ought to be equal, for each should be the number of c.c. occupied by 100 grains of water. If the differences are not nearly equal, find out where you have made a mistake. Repeat the exercise with still greater care, estimating tenths of a division.

State your results thus:—

EXPT. 1. Measuring to .1 c.c. EXPT. 2. Estimating .01 c.c.

Differences.		Differences.	
0.	50	50	
	} . . 6.3		} . . 6.49
1.	43.7	43.51	} . . 6.46
	} . . 6.5		} . . 6.49
2.	37.2	37.05	} . . 6.46
	} . . 6.4		} . . 6.48
3.	30.8	30.56	} . . 6.46
	} . . 6.5		} . . 6.48
4.	24.3	24.10	} . . 6.48
	} . . 6.5		} . . 6.48
5.	17.8	17.62	} . . 6.45
	} . . 6.4		
6.	11.4	11.14	
7. 4.69		

4. Pour water into a burette. Set it exactly at 0 c.c. Weigh an empty beaker. Run into it exactly 5 c.c. Weigh. Run in another 5 c.c., weigh again, and so on until you have run out 50 c.c. Take the difference between each two consecutive weighings. These differences should be equal if the burette is correctly graduated.

5. You are provided with a small bottle. Find its volume by means of a burette.

(b) *Measuring Jars*

6. Make a drawing, twice the real size, of the portion of a measuring jar from 60 to 80. Show the graduations and the figures.

7. Add water to a measuring jar until the level of its surface is exactly at 74 c.c.

8. Read the level of the water in the measuring jar provided.

9. You are provided with a measuring jar and some marbles. Devise a method of finding their volume. Find also the average volume of 1 marble.

(c) *Pipettes* (see also p. 33)

10. Find the volume of a 100 grain pipette as follows: Weigh an empty beaker. Fill the pipette with water, and run the water into the

beaker. Repeat three times. Again weigh. Calculate the weight in grams of 100 grains of water.

11. Recalibrate the pipette provided, thus:—Gum a strip of paper $1\frac{1}{2}$ inch long upon the stem of a pipette. Find the volume up to the bottom, and also to the top edge of the paper, proceeding as in No. 10. Calculate where the mark should be placed upon the stem, and mark it on the paper.

CHAPTER V

THE MEASUREMENT OF TEMPERATURE

Heat and Temperature.—Every one can attach some meaning to the word *Heat*. The sensation of heat is familiar to all. And yet we shall often make mistakes in judging whether a substance is hot or cold by our sensations. For the same room may at the same time appear hot to one man and cold to another. Indeed, the two hands may judge differently of the same substance; test this by placing one hand for a few seconds in hot water and the other in cold, and then place both together in lukewarm water: one hand feels cold and the other warm. It is clear, then, that the sensations of heat and cold in our bodies do not necessarily tell us anything of the actual state of a substance with regard to heat or cold.

Nevertheless there is no room for doubt that different states of a body with regard to heat and cold can exist, and we shall proceed to consider how these different states or conditions can be measured. This we can do without stopping to find out exactly what heat itself is.

Temperature.—If a poker be placed in a fire, the end outside the fire quickly becomes warm, and we understand that heat has passed from the hot coals into the poker. If the red-hot poker be placed in a bucket of cold water, the poker will be rapidly cooled, and we say that heat has passed from the poker into the water. Many instances are met with every day of heat passing from one body to another; such bodies are said to be at different *temperatures*, and the heat passes from the body which is said to have the higher temperature to that which is said to have the lower temperature.

A difference in temperature between two bodies may be compared to a difference in level between the water in two cisterns. When two bodies are so placed that heat can flow between them, the flow of heat will always take place from the body which has the higher temperature to that which has the lower temperature, just as water will always flow from a higher level to a lower. Indeed, temperature may be defined as *that condition of a body upon which depends the flow of heat to or from other bodies*. It is important also to observe that the passage of heat from a hot to a cold body will continue until they reach a common temperature, just as water is said to find its own level.

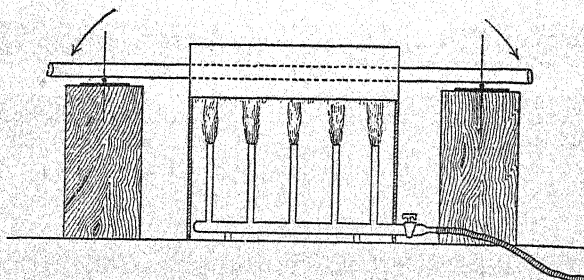


FIG. 20.

Preliminary Idea of a Thermometer.—When heat is applied to any body it may produce many effects, but none is more noticeable than change in volume. In by far the greater number of cases a body expands or increases in volume when it is heated.

EXPT. 1. To show that Iron expands when heated.¹—Support an iron gas pipe at its two ends upon two needles, resting upon horizontal glass plates, and let the central portion of the pipe lie within (but not touching) a gas furnace (Fig. 20). Fix to each needle a light straw to serve as an index. Now light the furnace. In a short time each needle begins to roll outwards turning the straw. This movement must be caused by the lengthening (expansion) of the bar. Note that the hotter the pipe becomes the greater is the expansion, and the consequent

¹ Lecture Table Experiment.

movement of the straws. Allow it to cool. Note how the needles roll back.

EXPT. 2. To show that Water expands when heated.

—Take a glass flask. Select a rubber stopper, with one hole, which fits it closely. Pass a narrow glass tube, 50 cm. in length, through the cork, so that the end of the tube is flush with the lower surface of the stopper (Fig. 21). Fill the flask quite full with water coloured with indigo. Push the stopper with the tube into the neck of the flask so that all air is expelled, and the liquid rises a short way up the tube. Mark the level of the liquid with a piece of gummed label. Immerse the whole of the flask in a vessel containing

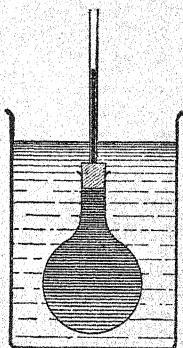


FIG. 21.

hot water. What do you observe? The liquid in the tube first

falls slightly; then it rises steadily. Why? The fall at first may have been due to a contraction of the water, or to an expansion of the glass flask. The heat had hardly time to reach the water, so probably the fall was due to the flask expanding. The second change—the slow steady rise—is due to the gradual heating and expansion of the water within the flask.

EXPT. 3. To show that Air expands when heated.

—Close a glass flask with a rubber stopper through which is passed a long glass tube. In-

vert the flask and support it so that the open end of the tube dips under coloured water contained in a beaker (Fig. 22).

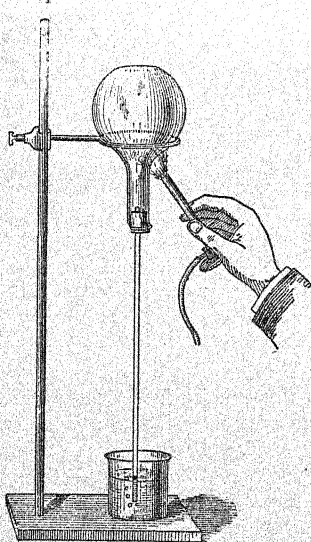


FIG. 22.

Warm the flask first with the hand. A few bubbles escape. Then warm the flask with a flame. The enclosed air bubbles out much faster. Remove the source of heat. Water rises up the tube, showing that the air contracts on cooling.

In these three experiments we have illustrations of the expansion by heat of a solid, of a liquid, and of a gas. In each case the expansion indicates the flow of heat into them from sources which are at higher temperatures. Hence, if we find that a substance is expanding in volume, we may conclude that heat is probably passing into it from something else which must be at a higher temperature.



FIG. 23.

The simple instrument described in Experiment 2 may be used to compare the temperatures on different days, or in different parts of a house, provided in each case the instrument is left sufficiently long for the passage of heat to continue until the instrument and the room have acquired a common temperature. If the glass tube be graduated—that is to say, divided into equal divisions and numbered—temperatures may be stated in terms of this scale. Such an apparatus is a simple instrument for recording and measuring temperatures—in other words, it is a rough **thermometer**. It may be remarked that

Galileo in the year 1597 was probably the first to employ a thermometer.

The Mercurial Thermometer

EXERCISE.—1. Examine a mercurial thermometer. Draw the bulb and a portion of the stem four times the actual size.

2. Try the effect of holding the bulb in the hand, and of placing it in the mouth.

The ordinary mercurial thermometer is really just such an instrument as was used in Experiment 2. Mercury, however, is used as the liquid instead of water. The principal advantages of mercury over water are, that it does not *freeze* until a very low temperature, and secondly, that it does not *boil* until a very high temperature. The mercurial thermometer consists of a very fine thick-walled glass tube, which ends in a round or cylindrical bulb (Fig. 23).

The Fixed Points on a Thermometer.—In order that a thermometer may be of practical use, its stem must be

graduated. And, further, all thermometers should be marked so as to show the points at which the mercury would stand were all exposed alike to one or two particular temperatures. Otherwise, it would be impossible to compare the readings of one thermometer with those of another.

In graduating the thermometer it is customary in the first place to mark upon its stem the points at which the mercury stands when it is heated to two standard temperatures. It was Newton who suggested that for these **fixed points**, as they

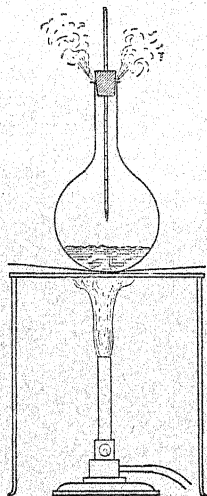


FIG. 24.

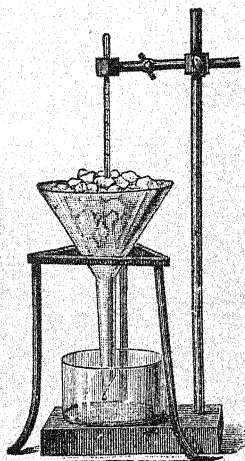


FIG. 25.

are called, the melting temperature of ice, which is found to be very constant, and the temperature of steam arising from boiling water, which also, as will be shown later, is constant so long as the pressure of the air remains the same, should be employed.

EXPT. 4. To mark the Boiling Point on a Thermometer.—Take a large flask containing some water, and fit it with a cork. File two grooves down the side of the cork, and pierce it with a hole through which a thermometer can be thrust, so that its bulb is just above the surface of the water (Fig. 24).

Boil the water vigorously. The thermometer is now surrounded by steam. Observe carefully the point to which the mercury rises. This point is marked one hundred degrees or 100° on a Centigrade and 212° on a Fahrenheit thermometer.

EXPT. 5. To mark the Freezing Point on a Thermometer.—Fill a funnel with pounded ice. Support it on a tripod. Plunge the thermometer into the mass (Fig. 25). Observe the point at which the mercury becomes stationary. This is marked 0° on a Centigrade and 32° on a Fahrenheit thermometer. Is your thermometer marked correctly?

Scales of Temperature.—The difference in temperature between these two fixed points is large. It is customary to subdivide the interval between them according to one or other of three scales, named the Fahrenheit, Centigrade, and Réaumur scales; we need only consider the first two.

On the Fahrenheit scale the freezing point is marked 32° and the boiling point 212° , while on the Centigrade scale they are marked 0° and 100° respectively (Fig. 26). The interval between these points is therefore divided into 180 equal divisions or degrees on the Fahrenheit and into 100° upon the Centigrade scale.

Hence,	180	Fahrenheit degrees	=	100	Centigrade degrees.		
or	9	" "	=	5	" "		
Therefore;	1	" "	=	$\frac{5}{9}$	" "		
and	$\frac{9}{5}$	" "	=	1	" "		

The position marked 0° on a scale is often called the *zero* of the scale. On the Fahrenheit scale zero is 32° below freezing point.

Degrees below 0° are marked by the *negative* sign: thus -10° C. means 10° below 0° C.

To Read the Thermometer.—Although the interval between the fixed points on a Centigrade thermometer is divided into only 100 degrees, it is nevertheless possible by means of the eye to mentally subdivide each degree into tenths, and therefore to read any given temperature to the nearest *tenth* of a degree.

Fig. 27 shows a piece of a thermometer stem. The graduations stand for single degrees. The end of the mercurial column lies between 11° and 12° . Subdividing the eleventh degree by the eye, the reading is $11^{\circ}.4$.

Also observe the following points :—

1. The eye should be at right angles to the scale at the point reached by the mercury (compare p. 34).
2. The thermometer should be read while its bulb is *in* the substance of which the temperature is being taken.
3. The figures count upwards.

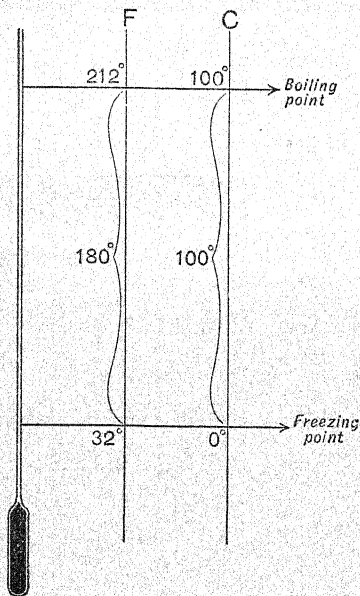


FIG. 26.

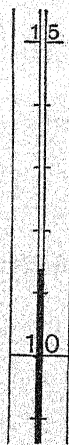


FIG. 27.

EXERCISES

1. Fill a tumbler with cold water from the tap. Place a C. thermometer in the water. Read the temperature. Change the water and read again.
2. Take some cold water in a beaker, and place a C. thermometer in it. Heat the water over a small flame. When the temperature reaches 20°C . dip your fingers into the water, and describe your impression. Continue to heat the water, and when the temperature is 40°C . and 60°C . try it in a similar way.

3. Read the thermometers placed about the room and up the wall.
4. What is the lowest temperature that can be read with the thermometer provided? What is the highest temperature?
5. Read the position of the arrow on a clinical thermometer.

Conversion of Thermometric Scales.—

(1) Find the Centigrade reading corresponding to 80° F.

80° F. is $(80^{\circ} - 32^{\circ})$ or 48° F. above the freezing point.

48° F. is $48^{\circ} \times \frac{5}{9}$ C. or $26^{\circ}\cdot6$ C. above the freezing point, which is 0° on the C. scale.

Hence, the required reading is $26^{\circ}\cdot6$ C.

(2) Find the Centigrade reading corresponding to 20° F.

20° F. is $(32^{\circ} - 20^{\circ})$ or 12° F. below the freezing point.

12° F. is $12^{\circ} \times \frac{5}{9}$ C. or $6^{\circ}\cdot6$ C. below the freezing point, which is 0° on the C. scale.

Hence, the required reading is $-6^{\circ}\cdot6$ C.

(3) Find the Fahrenheit reading corresponding to 15° C.

15° C. is $15^{\circ} \times \frac{9}{5}$ F. or 27° F. above the freezing point, which is 32° on the Fahrenheit scale.

Hence the required reading is $(27^{\circ} + 32^{\circ})$ F. or 59° F.

***Mechanical Interpolation.**—The conversion of thermometer scales is only one instance of the mathematical process called *interpolation*. Other instances occur in making out railway time-tables, in arranging prices and sizes of goods, and in the use and preparation of logarithmic, astronomical, and other tables (see Chapter XVIII.).

EXAMPLES VII

Interpolation—Conversion of Thermometric Scales

(i) * INTERPOLATION

1. In a certain street the lamp-posts and the houses do not count from the same point. The third lamp-post is opposite No. 21, and the seventh lamp-post is opposite No. 45. What number is opposite the tenth lamp-post?

2. On a certain railway journey by express K, L, and M are passed at 12.3, 12.21, and 12.27. A slower train passes K and M at 1.4 and 1.40. At what time does it pass L?

3. If the sun sets at 6.10 P.M. on September 16, and at 5.54 on September 23, find the time of sunset on September 21.

4. A solution containing 40 per cent. sulphuric acid has the density 1.306, and a 50 per cent. solution has the density 1.398. What is the density of a 46 per cent. solution?

(ii) CONVERSION OF THERMOMETRIC SCALES

What Fahrenheit temperatures are equal to

5. 25° C. 6. 5° C. 7. 80° C. ?

What Centigrade temperatures are equal to

8. 50° F. 9. 95° F. 10. 122° F. ?

Express on the Fahrenheit scale—

11. 20° C. 12. -5° C. 13. $4^{\circ}.5$ C.

Express on the Centigrade scale—

14. 40° F. 15. 14° F. 16. -4° F.

17. On the Réaumur scale the freezing point is 0° and the boiling point 80° . What Réaumur temperature is equal to 59° F. ?

18. What Fahrenheit temperature is equal to 16° R. ?

EXERCISES

Thermometry

1. Fill a beaker with cold water from the tap. Place a F. and a C. thermometer in the water. Read the lowest temperatures they show, as nearly as possible at the same time. Change the water, and again read. Repeat four more times. Enter thus :—

F.	$68^{\circ}.9$						
C.	$15^{\circ}.0$						

Convert the last C. reading to the F. scale. Does the number agree with the last F. reading?

2. Warm the water in the beaker, and when the C. thermometer reads about 40° , remove the flame, stir, and read both thermometers accurately, and as nearly as possible at the same time. Convert the C. reading to the F. scale. Does the result agree with the F. reading?

3. Heat the water again, until the C. thermometer reads about 75° , and then read both thermometers. Convert the F. reading to the C. scale. Does the result agree with the C. reading?

4. Now boil the water, and place the C. thermometer in it. Arrange two columns in your note-book headed **Time** and **Temperature**, thus—

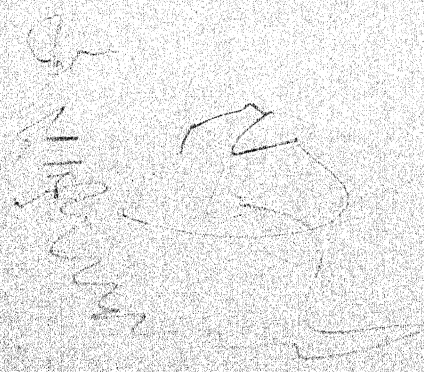
TIME.			TEMPERATURE.
3 hr.	20 m.	0 sec.	100°·2 C.
3	20	30	
3	21	0	
3	21	30	

Look at the second-hand of your watch, and when it is at the minute read the thermometer and at once extinguish the flame, but leave the thermometer in the water. Put down the temperature of the boiling water opposite to the first time reading. Then, at intervals of half-minutes, take the readings of the thermometer to the first decimal place. Take 16 readings. Plot the results on squared paper. (See Chapter XVIII.).

5. Half fill a flask with water. Add about 10 grams of common salt. Heat the flask until the water boils. Hold the bulb of a C. thermometer in the steam *above* the boiling liquid. What is the highest temperature it records?

Then place the bulb of the thermometer *in* the boiling liquid. Is the reading the same?

6. Pound some ice in a mortar. Add a small handful of common salt. Mix thoroughly. Read the lowest temperature reached.



CHAPTER VI

THE MEASUREMENT OF THE RELATIVE DENSITIES OF SOLIDS AND LIQUIDS

I. THE RELATIVE DENSITIES OF LIQUIDS

No one can fail to be aware that there are many different kinds of liquids. At the mention of mercury, water, oil, or methylated spirits, we may also perhaps admit that they are not all equally heavy. But how can we actually compare the weight of one liquid with that of another? Obviously by weighing in turn one and the same volume of each liquid. We know that 100 c.c. for instance of water weigh 100 grams. Suppose we find the weight of 100 c.c. of other liquids.

We can readily measure out equal volumes of different liquids by means of a pipette.

EXPT. 1. To find the Weight of 100 c.c. of Methylated Spirits.—Take a small dry beaker. Weigh. Measure into it by means of a pipette 20 c.c. of methylated spirits. Weigh. State thus:—

Weight of beaker	= 15.41 gm.
Weight of beaker + 20 c.c. spirits	= 31.45 „
∴ Weight of 20 c.c. spirits	= 16.04 „
∴ Weight of 100 cc. spirits... ..	= 80.20 „

EXERCISES.

Find the weight of 100 c.c. of—

- | | |
|------------------------|------------------------|
| (1) milk. | (4) hydrochloric acid. |
| (2) oil of turpentine. | (5) sulphuric acid. |
| (3) chloroform. | (6) sea water. |

Compare your results with the following numbers, which were obtained in the same way :—

Methylated spirits ...	80.2 gm.	Milk	102.9 gm.
Oil of turpentine ...	86.3 "	Hydrochloric acid ...	114.1 "
Benzene	88.2 "	Chloroform	147.1 "
Sea-water	102.5 "	Sulphuric acid	178.2 "

It is evident that equal volumes of these liquids contain different amounts of matter. This is usually expressed by saying that some liquids are **denser** than others, or that they possess different **densities**.¹

Now, it is convenient to compare the densities of different liquids with that of some one liquid, and water has been chosen as the standard liquid. We can at once obtain from the results of our experiments numbers which will express the **density** of any liquid **relative** to that of water.

Definition of Relative Density.—

$$\text{Relative density} \quad \text{or} \quad \text{Specific gravity} = \frac{\text{Weight of a substance}}{\text{Weight of an equal volume of water}}$$

For instance :—

$$\begin{aligned} \text{Relative density of milk} &= \frac{\text{weight of 100 c.c. milk}}{\text{weight of 100 c.c. water}} \\ &= \frac{102.9}{100} \\ &= \mathbf{1.029.} \end{aligned}$$

In this way we find the following relative densities, or specific gravities, as they are often termed :—

Methylated spirits	0.802	Milk	1.029
Oil of turpentine	0.863	Hydrochloric acid	1.141
Benzene	0.882	Chloroform	1.471
Water	1.000	Sulphuric acid	1.782
Sea-water	1.025		

¹ **Definition of Density.**—*The density of a substance is defined as the mass of unit volume.*

Hence, for example, from the results above—

$$\begin{aligned} \text{Density of methylated spirits} &= 0.802. \\ \text{Density of sulphuric acid} &= 1.782. \end{aligned}$$

Observe the meaning that these numbers bear. The statement that the relative density of milk, for example, is 1.029 means that volume for volume, milk is 1.029 times as heavy as water, and that 1 c.c. of milk therefore weighs 1.029 grams.

It is important to observe that determinations with different samples of any one liquid—for example, oil of turpentine—give very nearly the same numbers. This fact leads us to suspect that any definite liquid has a relative density which is a constant property of it (at a given temperature). So long as the liquid is not mixed with any other liquid, its relative density appears to remain constant, and to have a fixed value. Consequently, it is possible to distinguish between liquids which may otherwise resemble one another by a reference to their relative densities. For instance, the adulteration of milk with water, or the addition of water to alcohol, may be detected by finding the change of relative density which results from the admixture.

The Use of a Relative Density Bottle.—Determinations by the above method are rapidly made, and give excellent practice in the use of a pipette. But equal volumes of liquids can be more accurately measured in a Relative Density Bottle (Fig. 28). The glass stopper has a fine hole bored through it, and is ground to fit the neck accurately. The bottle is filled with liquid, and the stopper then lowered gently into place, so that all air bubbles and extra liquid escape through the hole. The top of the stopper and outside of the bottle are wiped dry before weighing. So accurate is this instrument that, if held in the hand for a few seconds, the liquid, expanded by the heat of the hand, will begin to exude through the hole.

These bottles are usually made to contain exactly 25 grams of water at 60° F. But this should be tested by experiment.

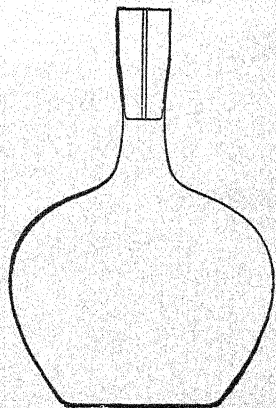


FIG. 28.

Members of a class are recommended to determine by means of this bottle the relative density of some one liquid, *e.g.* methylated spirits or a solution of salt. Results should only differ in the third place of decimals.

CALCULATIONS.—1. Find the weight of 50 c.c. of mercury : relative density of mercury = 13·5

State thus :—

$$\begin{array}{rcl} & \text{1 c.c. of mercury weighs} & 13\cdot5 \text{ gm.} \\ \therefore & 50 \text{ c.c.} & \text{,, } 13\cdot5 \times 50 \text{ gm} \\ & & \text{or } 675 \text{ gm.} \end{array}$$

2. Find the volume of 100 grams of sulphuric acid : relative density of sulphuric acid = 1·85.

$$\begin{array}{rcl} & 1\cdot85 \text{ gm. sulphuric acid occupy} & 1 \text{ c.c.} \\ \therefore & 1 & \text{,, } \text{,, } \text{,, } \text{,, } \frac{1}{1\cdot85} \text{ c.c.} \\ \therefore & 100 & \text{,, } \text{,, } \text{,, } \text{,, } \frac{100}{1\cdot85} \text{ c.c.} \\ & & \text{or } 54\cdot0 \text{ c.c.} \end{array}$$

Relative Densities of some Common Liquids.—

Ether	'73
Alcohol	'80
Benzene	'88
Turpentine	'89
Olive oil	'91
Water	1 00
Milk	1·03
Sea-water	1·026
Glycerine	1·26
Chloroform	1·47
Sulphuric acid	1·85
Mercury	13·59

EXAMPLES VIII

Relative Densities of Liquids

(a)

1. Find the weight of 100 c.c. of chloroform : relative density of chloroform = 1.47.
2. Find the weight of a litre of alcohol : relative density of alcohol = 0.80.
3. 50 c.c. of turpentine weigh 43.1 grams. Find the relative density of turpentine.
4. 60 c.c. of a sample of milk weigh 61.4 grams. Find the relative density of the milk.
5. One litre of a certain liquid weighs 1 kilogram. What is the liquid probably?
6. Find the volume of 1 gram of mercury : relative density of mercury = 13.5.
7. Find the volume of 1000 grams of olive oil : relative density of olive oil = 0.9.
8. Find the volume of 1 gram of alcohol : relative density of alcohol = 0.80.
9. Find the volume occupied by 51.3 grams of sea-water : relative density of sea-water = 1.026.
10. Find the volume occupied by 31.5 grams of glycerine : relative density of glycerine = 1.26.

(b)

[Answers to 4 significant figures.]

11. Two beakers are placed one on each pan of a balance, and shot are added until the beam swings evenly. Then 5 c.c. of mercury are placed in one beaker, and it is found that 67.5 c.c. of water have to be added to the other beaker to restore the balance. Calculate the density of mercury relative to water, and the density of water relative to mercury.
12. A pipette contains 30.86 grams of spirits of relative density 0.81. What is the volume of the pipette?
13. What weight of sulphuric acid of relative density 1.86 will the same pipette contain?
14. 20 c.c. of a liquid A of relative density 1 are added to 20 c.c. of a liquid B of relative density 0.81. What is the weight of the mixture?
15. The mixture referred to in Ex. 14 measures only 38.5 c.c., owing

to contraction having taken place. What is the relative density of the mixture?

16. A flask weighs 25 grams when empty, 125 grams when full of distilled water, and 127.6 when full of sea-water. Find the relative density of the sea-water.

17. How many c.c. of sulphuric acid (1.85) must be measured out to obtain 111 grams of it?

18. Alcohol (.8) and sulphuric acid (1.85) are to be mixed in the proportion 1 : 3 by weight. What volume of alcohol should be taken with 200 c.c. of acid?

19. A salt solution has the relative density 1.025; how many c.c. of water must be added to 1 litre of it to reduce its relative density to 1.020?

20. Find the relative density of a mixture of 100 c.c. water with 100 c.c. spirits of relative density 0.81, assuming that the liquids contract by 2 per cent. of their original volumes on mixing.

21. A 200 c.c. flask weighs 300 grams when full of water. What will it weigh when full of mercury (13.5)?

22. A bottle weighs 20 grams. Full of water at 4° C. it weighs 70 grams. When filled with water at 65° C. it only weighs 69 grams. Find the relative density of the hot water.

23. A tube 30 cm. long weighs 15 grams when empty, and 96 grams full of mercury (13.5). Find in sq. mm. the area of the cross section of the tube.

II. THE RELATIVE DENSITIES OF SOLIDS

All that has been said of liquids, as regards the differences in density which are found to exist, is of course true also of solids. Although it would appear at first sight more difficult to measure the relative densities of solids, there are several ways in which it can be carried out. Remember that it is required to compare the weight of a solid with the weight of an equal volume of water. How can we find the weight of an equal volume of water?

(A) FIRST METHOD: REGULAR SOLIDS

We can easily determine the volume of most *regular* solids, *e.g.* cubes, cylinders, or spheres, by direct measurement and a simple calculation.

EXPT. 2. To find the Relative Density of a Rectangular Oak Block.—The volume is readily calculated from the length, breadth, and height of the block. Measure each of

these accurately in cm. and decimal. Also weigh the block.
Enter measurements thus :—

$$\text{length} = 5.0 \text{ cm.}$$

$$\text{breadth} = 5.0 \text{ cm.}$$

$$\text{height} = 2.5 \text{ cm.}$$

$$\therefore \text{volume of block} = 5 \times 5 \times 2.5 \text{ c.c.} \\ = 62.5 \text{ c.c.}$$

$$\therefore \text{weight of an equal volume of water} = 62.5 \text{ gm.}$$

$$\text{Also, weight of block} = 48.16 \text{ gm.}$$

$$\text{Hence, relative density of the oak} = \frac{48.16}{62.5} \\ = 0.77.$$

EXPT. 3. To find the Relative Density of a Copper Cylinder.—Volume of a cylinder = (radius)² × $\frac{22}{7}$ × height.

Measure the diameter and height of the cylinder in cm. and decimal. Then :—

$$\text{height} = 3.78 \text{ cm.}$$

$$\text{diameter} = 1.56 \text{ cm.}$$

$$\therefore \text{radius} = 0.78 \text{ cm.}$$

$$\therefore \text{area of circular end} = (0.78)^2 \times \frac{22}{7} \text{ sq. cm.}$$

$$\therefore \text{volume of cylinder} = (0.78)^2 \times \frac{22}{7} \times 3.78 \text{ c.c.} \\ = 7.26 \text{ c.c.}$$

$$\therefore \text{weight of an equal volume of water} = 7.26 \text{ gm.}$$

Also, weight of cylinder = 64.54 gm.

$$\text{Hence, relative density of copper cylinder} = \frac{64.54}{7.26} \\ = 8.88.$$

EXERCISES

Find the relative density of—

- (1) cubes or rectangular blocks of different hard woods.
- (2) an iron or brass cylinder.
- (3) a box-wood ball or a large glass marble.

$$[\text{Volume of sphere} = \frac{4}{3} \times \frac{22}{7} \times (\text{radius})^3.]$$

How can we find the volume of a number of glass stoppers, of leaden shot, a heap of sand, or a bundle of steel screws? We shall describe three methods which may be followed in one case or another.

(B) SECOND METHOD: USE OF A MEASURING JAR

EXPT. 4. To find the Relative Density of some Glass Stoppers.—Weigh a few glass stoppers. How can their volume be found? They are too irregular for it to be found by mensuration. But we can find how much water the stoppers can displace in a measuring jar. Choose a measuring jar, the narrowest into which the stoppers will easily slide. Partly fill the jar with water, and read the level. Carefully slide in the stoppers. The level rises. Read again. Thus:—

$$\begin{aligned}
 \text{Weight of stoppers} & \dots\dots\dots = 80.5 \text{ gm.} \\
 \text{Volume of water, stoppers outside} & \dots = 100 \text{ c.c.} \\
 \text{,, ,, ,, ,, inside} & \dots\dots = 132.5 \text{ c.c.} \\
 \therefore \text{Volume of stoppers} & \dots\dots\dots = 32.5 \text{ c.c.} \\
 \therefore \text{Weight of an equal volume of water} & \dots = 32.5 \text{ gm.} \\
 \text{Hence, relative density of the glass stoppers} & = \frac{80.5}{32.5} \\
 & = 2.47.
 \end{aligned}$$

With what kinds of solids could this method not be used?

EXERCISES

Using a 100 c.c. measuring jar, find the relative densities of—

- (1) five glass marbles.
- (2) fragments of slate pencil.
- (3) lumps of marble.
- (4) a copper cylinder.
- (5) sand.
- (6) a lump of aluminium.
- (7) a cork.¹
- (8) a lump of beeswax, or of paraffin wax.

¹ Use a "sinker" which may be attached to the light solid by a piece of cotton. An old brass weight or a glass stopper will serve.

EXAMPLES IX

Relative Densities of Solids.

1. Find the relative densities of the following metals, given :—
 - (1) a litre of copper weighs 8950 grams.
 - (2) 75 c.c. of zinc weigh 540 grams.
 - (3) 676 c.c. of cork weigh 169 grams.
 - (4) 300 c.c. of lead weigh 3'42 kilograms.

2. A rectangular block of marble measures 24 cm. long, 16 cm. broad, and 10 cm. in height, and it weighs 10'368 kilograms. Find its relative density.
3. A cube of wood measures 8 cm. along each edge, and its weight is 307'2 grams. Find its relative density.
4. Solids, weighing 10, 15, 20 grams respectively, were placed successively in a graduated cylinder partly filled with water, the surface of which read 95 c.c. originally. When the first was added, the water rose to 99 c.c., when the second was added, it rose to 106 c.c., and when the third was added to 108 c.c. Find the relative densities of the solids.
5. A brass chain weighs 96 grams. When lowered into a measuring jar already containing 100 c.c. of water, it raises the water level to 112 c.c. Find its relative density.
6. A dozen shot, made from lead, of relative density 11'4, raise the level of water in a measuring jar from 98 c.c. to 122 c.c. Find the weight of each shot.
7. A shilling weighs 87 grains, and has the relative density 10'4. What would be the weight of a platinum coin (21'5) of the same size?
8. Find the value of 1 c.c. of silver (relative density 10'5), when silver is worth 2s. 6d. per oz. Reckon 28 grams = 1 oz.
9. If 1000 oz. of water fill 1 cubic foot, and cast-iron is 7'2 times as heavy as water, find the weight of an iron block 4 in. long, 3 in. broad, and 1 in. high.
10. Find the volume of 1 cm. length of platinum wire (relative density 21'5), supposing a piece 1 metre long to weigh 2'15 grams.

(C)* THIRD METHOD : FRAGMENTARY SOLIDS

The volume of a solid cannot be determined to the tenth of a c.c. by means of a measuring jar. But by means of a balance we may weigh water to a centigram, that is, to the hundredth of

a cubic centimetre. Instead of noticing the rise of level in a measuring jar, we may find the loss of weight by overflow where fragments of a solid are put in a relative density flask full of water.

EXPT. 5. To find the Relative Density of White Sand.—Weigh a dry relative density flask. Introduce some sand, so that the flask is about one-third full. Weigh again. Nearly fill the flask with water, shake gently so as to dislodge air bubbles; allow the sand to settle, and then carefully fill the flask with water, and introduce the stopper. Weigh again.

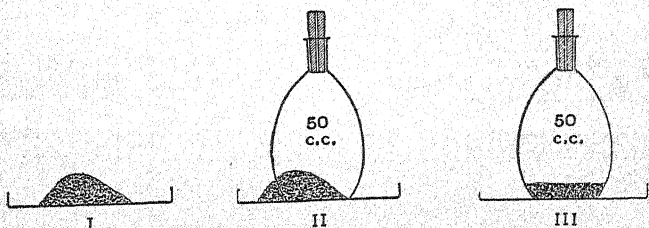


FIG. 29.

Lastly, wash the sand out of the flask, and weigh it filled with water alone.

Then calculate the weights represented in Fig. 29, I, II, and III.

Example :—

Weight of flask	= 14.80 gm.
Weight of flask + sand	= 23.84 ,,
Weight of flask + sand + water, <i>inside</i>	= 45.37 ,,
Weight of flask and water	= 39.80 ,,

Calculation :—

Weight of flask	= 14.80 ,,
Weight of flask + sand	= 23.84 ,,
∴ Weight of sand	= 9.04 ... (i.)

Again—

Weight of flask + water	= 39.80 gm.
∴ Weight of flask + water + sand, <i>outside</i>	= 39.80 + 9.04
		= 48.84 gm. (ii.)

But—

Weight of flask + water + sand, *inside* ... = 45.37 gm. (iii.)

∴ Weight of water overflowing to make room
for the sand... .. = 48.84 - 45.37 = 3.47 „

∴ Relative density of sand = $\frac{\text{Weight of sand}}{\text{Weight of overflow}}$
= $\frac{9.04}{3.47}$
= 2.60

EXERCISES

Find the relative density of—

- (1) powdered glass.
- (2) lead shot.
- (3) powdered antimony.
- (4) powdered galena.

EXAMPLES X

Relative Densities of Solids

1. Find the relative density of sand from the following weights:—
Flask, 25 grams. Flask + sand, 35 grams. Flask + sand, filled up
with water, 56 grams. Flask filled with water, 50 grams.
2. Find the relative density of a powder from the following
weights:—Flask, 25 grams. Flask filled with water, 50 grams.
Flask + powder, 40 grams. Flask + powder filled up with water,
63 grams.
3. A 100 c.c. flask filled with water to the mark weighs 125 grams.
When 57 grams of lead are added, and the water is again adjusted to the
mark, it weighs 177 grams. Find the relative density of lead.
4. A 100 c.c. flask weighs 25 grams when empty. What will it
weigh when 30 grams of powdered glass (relative density 3.0) are placed
inside, and water is filled in up to the mark?

(D) FOURTH METHOD.

Principle of Archimedes.—There is another method of
determining the relative densities of solids, which depends upon
a fact discovered about 220 B.C. by the philosopher Archimedes

of Syracuse. It is said that Hiero, King of Syracuse, sent a lump of gold to a goldsmith's to be made into a crown. When the crown was delivered he suspected that, though of the right weight, it contained nevertheless an undue proportion of silver. So Archimedes was asked to find out the truth of the matter. Turning the matter over in his mind, it is said that he went to have a bath, and noticing the rise of the water he jumped out in great excitement, and ran naked through the streets of the city,

crying, "Eureka, Eureka, I have discovered it!" But suppose we make the following experiment.

EXPT. 6. To try whether a Solid weighs the same in Water as in Air.—(1) Suspend a block of aluminium, or some other object, from a 100-gram spring balance, and read its weight in air. Then allow the block to sink below the surface of water in a tumbler, and again read the balance (Fig. 30). Is not the weight of the block in water much less than its weight in air? The water clearly supports part of the weight of the block.

Similarly, weigh both in air and in water an empty glass bottle and a block of wood. These have no weight in water, or rather the whole of their weight is supported by the water, and therefore they *float*.

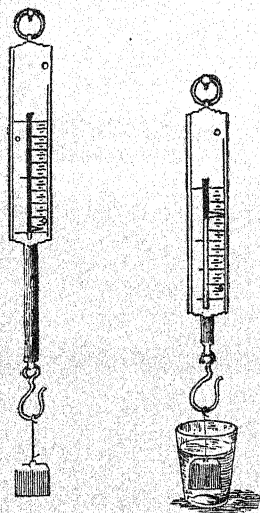


FIG. 30.

(2) Let us next make an experiment which will show a connection between the loss in weight of a solid immersed in a liquid and the volume of the solid. Hang up a pair of hand scales. Take a brass cylinder which fits exactly into a bucket. Suspend the bucket under one pan of the scales, with the cylinder hanging underneath the bucket, and within an empty beaker as shown in Fig. 31, and weigh the whole. Then add water to the beaker until the cylinder is completely immersed. There is equilibrium no longer. Pour water carefully into the bucket, so as

to restore equilibrium, taking care that the cylinder is still completely immersed and not touching the beaker. How much water have you to add? At the moment when the bucket is filled up to the brim equilibrium is restored. So the cylinder when immersed in water lost a weight which was equal to the weight of water which occupied the same volume. The experiment might be repeated, using turpentine or a salt solution or some other liquid in place of water. This experiment has verified a fact which is usually called **The Principle of Archimedes**, viz. : **A solid when immersed in a liquid loses a weight which is equal to the weight of the liquid which would occupy the same volume.**

This is of great importance. For instance, it follows that if a block of aluminium, which in air weighs 95 grams, weighs only 60 grams in water, *i.e.* the loss of weight is 35 grams, the volume of the block must be 35 c.c. We can, in fact, readily find the volume of any solid which can be weighed in water.

The Principle of Archimedes affords, therefore, a ready method for determining the relative density of any solid which can be suspended and weighed in water.

EXPT. 7. To find the Relative Density of a Glass Stopper.—Suspend a solid glass stopper by a thread above the left-hand pan of the balance (Fig. 32). Place a wooden bridge across the pan. Place a beaker on the bridge so that the stopper hangs freely within the beaker. Weigh the stopper. Pour water carefully into the beaker until the stopper is completely immersed, and remove any air bubbles which cling to the stopper with a camel's hair brush. Weigh it again. Thus in an experiment :—

Weight of glass stopper in air	=	15 gm.
Weight of glass stopper in water	=	9 „

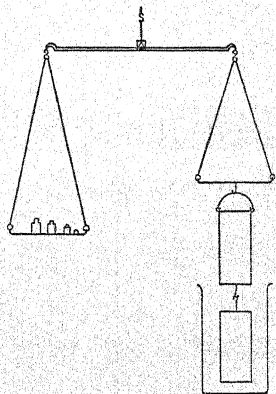


FIG. 31.

Hence, loss in weight of glass stopper = 6 gm.
 or, weight of an equal volume of water ... = 6 ,,

$$\therefore \text{Relative density of glass stopper} = \frac{15}{6} \\ = 2.5.$$

It is clear that the last *three* methods for determining the relative densities of solids cannot be applied directly in the case of one such as cork, which is lighter than water, or to another

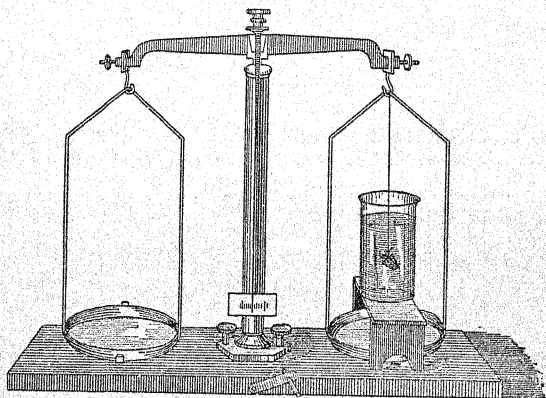


FIG. 32.

such as sugar, which will dissolve in water. The student is referred to other text-books for the special methods required in such cases.

Relative Densities of some Common Substances.—

Cork	0.25	Tin	7.3
Oak	0.84	Zinc	7.2
Ice	0.93	Iron (cast)	7.2
White sand	2.6	Brass	8.4
Marble	2.7	Copper... ..	8.9
Aluminium	2.7	Silver	10.5
Glass (crown)	2.5 to 3.1	Lead	11.4
Sulphur... ..	2.0	Gold	19.3
The Earth	5.5	Platinum	21.5
Antimony	6.7		

EXERCISES

Find the relative density of—

- (1) a solid glass stopper.
- (2) a lump of marble.
- (3) a lump of sulphur.
- (4) a copper cylinder.
- (5) a brass cylinder.
- (6) a lump of aluminium.
- (7) an iron nut, screw, or bolt, or cylinder.

EXAMPLES XI

Relative Densities of Solids

1. A gold coin weighs 12.25 grams in air, and 11.55 grams in water. Find its relative density.
2. A lump of marble weighs 20.76 grams in air, and 13.14 grams in water. Find its relative density.
3. A lump of sulphur weighs 26.1 grams in air, and 13.07 grams in water. What is its relative density?
4. The volume of a glass stopper is 5.77 c.c., and its weight in air is 15.13 grams. Find its relative density.
5. A metal cylinder weighs 64.65 grams in air, and 57.4 grams in water. Of what metal does it probably consist?
6. If a silver coin weighs 41.6 grams, and has the relative density 10.4, what will it weigh in water?
7. Find the loss of weight in water of a 56 lb. cast-iron weight (relative density 7.2).
8. If a glass stopper weighs 24 grams in air, and 16 grams in spirits of relative density .8, find the relative density of the glass.
9. A glass stopper weighs 30 grams in air, 17.5 grams in water, and 20 grams in spirits. Find the relative density of the spirits.
10. Find the relative density of a mixture of 100 c.c. lead (11.4) and 200 c.c. tin (7.3), assuming that they mix without contraction.

CHAPTER VII

THE MEASUREMENT OF THE PRESSURE OF THE AIR

The Existence of the Air.—We have no difficulty in realising the existence of solids and liquids around us. It is different with the air, and at first it is difficult to believe that there is an invisible substance surrounding us on every hand. Yet the existence of something invisible is at least suggested by the draught from an open window or door, or by waving corn and swaying branches.

EXPT. I. To show that an "Empty" Bottle contains Air.—Take an empty bottle and force it mouth downwards into

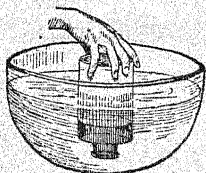


FIG. 33.



FIG. 34.

water. Notice that the water does not fill the bottle, showing that the bottle must contain something which opposes the entrance of the water. Then tilt the bottle, and notice that bubbles escape up through the water, and that only then does the water enter the bottle (Figs. 33 and 34).

Such facts help us to realise that we live in an ocean of air, which extends in all probability to a distance of at least 200 miles from the earth. This envelope of air around the earth is called the **atmosphere**.

We are aware of the pressure of the air in a strong wind, but in a room in which the air is at rest we are not conscious of any pressure.

EXPT. 2. To show that Air is exerting a Pressure even when we cannot feel it.—Bend a glass tube as in Fig. 35. Pour in some coloured water. The water stands at exactly the same height in both arms. Why? If we attach a piece of rubber-tubing to the arm A B, and suck out some of the air from that arm, what may occur? Try. The water rises in the arm A B, and falls in the other, as represented in Fig. 36. Secondly, blow in a little air. The water falls in the one arm, and rises in the other.

The rise and fall of the water in the arm A B shows that the air was originally exerting a pressure equally upon the surface of the liquid in both arms and that the pressure of the air in A B was diminished when air was sucked out, and increased when air was blown into it.



FIG. 35.

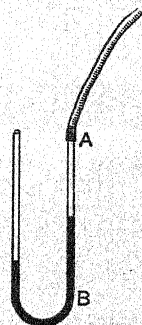


FIG. 36.

EXPT. 3.—Place one end of a *pipette* below the surface of some water. Apply the mouth at the other end, and suck gently. As the air is gradually inhaled, the water rises in the pipette.

This suggests that the air is exerting a pressure on the surface of the water outside the pipette, so that the water is forced up the tube as soon as the air is partially removed from the inside, and when the pressure there is consequently diminished.

The action of a *syringe* or that of a *suction-pump* can now be understood to depend upon the pressure which the air exerts upon the surface of a liquid.

EXPT. 4.—Take a tumbler, or, better still, a gas jar with a ground edge, and plunge it into water so that it is completely

filled. Press a piece of cardboard against the mouth. Carefully lift the vessel out of the water, mouth downwards. Note that the cardboard remains over the mouth. Why? Clearly it is held there by some pressure greater than the weight of the water within the tumbler (see Fig. 37). *

EXPT. 5.—Fit a thistle-funnel in one end of a piece of thick rubber pump-tubing. Fix the other end to an air-pump. Moisten the palm of the hand, and press the mouth of the funnel against it. Then proceed to exhaust the air from the tubing. Observe that the funnel is pressed firmly against the hand, so that some force is necessary in order to remove it.

It is clear, therefore, that the air even when quite still is exerting pressure, and the fact that no force may be felt upon the extended hand arises in the first place from the pressure being exerted above and below it equally. Then why is the hand not squeezed flat? Because there is an internal pressure of the blood in the veins and arteries and other fluids within the body, greater than that of the atmosphere.

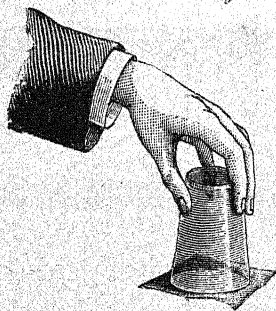


FIG. 37.

Repeat Expt. 5, applying the mouth of the funnel to the cheek. As soon as the pressure of the air within the funnel is diminished, the pressure within the cheek causes the flesh to swell outwards, while the funnel is

pressed by the air outside against the cheek.

The Pressure of the Air is due to the Fact that Air has Weight.—We are not conscious of the air having any weight, but by the following experiment we can readily show that it can be weighed.

EXPT. 6. **To show that the Air has Weight.**—Take a round-bottomed flask and fit it with a one-holed rubber stopper through which is passed a short glass tube fitted with a piece of rubber tubing and a clip, or, better, with a glass tap (Fig. 38). Suspend the flask by a string from one arm of a balance, and

weigh it. Then, leaving the weights on the pan, remove the flask, and exhaust some of the air by means of an air-syringe. Again suspend the flask on the balance, and observe that it is now lighter than the weights. Lastly open the tap: air is heard rushing into the flask, and it is then found to have regained its original weight.

It is clear therefore that the air has weight. As a matter of fact 1 litre of air weighs only 1.2 gram. But just as the weight of one eider-down quilt may be scarcely felt, yet many of them one above the other would be very oppressive, so, extending as the air does to a distance of at least 200 miles from the surface of the earth, it may be understood that the very considerable pressure of the air at the surface of the earth is due to the fact that the air has weight, slight though this be.

Torricelli, an Italian physicist, first showed in 1643 in the following way how to measure the pressure which the air is always exerting.

EXPT. 7. To measure the Pressure of the Air.—Take a thick-walled glass tube about 4 mm. internal diameter, and 90 cm. in length. Seal one end in the blow-pipe flame. Holding the open end upwards, almost fill the tube by pouring clean dry mercury¹ into it through a small funnel. Gently tap the tube so as

to dislodge air bubbles. Pass a bubble the length of the tube so as to sweep out smaller bubbles. Fill the tube completely with mercury. Close the open end firmly with the thumb, and invert the tube in a mortar containing mercury (Fig. 39). The mercury sinks in the tube, leaving a clear space A B at the top of the tube. Tilt the tube: if the mercury does not rise completely to the top, some air bubbles have not been removed; these must then be displaced by inverting the tube again, and gently tapping. Measure the length of the mercurial column above the level of the mercury in the mortar. It is about 76 cm., or 30 inches. From the way in which the tube

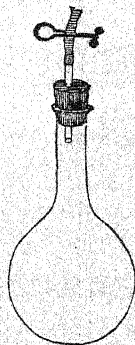


FIG. 38.

¹ On Management of Mercury, see Appendix, Vol. II.

was set up it seems impossible that the space A B can contain any air.

Let us repeat Expt. 7 in another form.

EXPT. 8.—Take a thick-walled glass tube about 90 cm. long, closed at one end, and fix a piece of thick rubber tubing about 6 inches long upon the open end. Fix a short glass tube open at both ends into the free end of the rubber tube. Resting the closed end upon the bench, carefully pour mercury into the apparatus until it reaches half way up the short tube. Sweep out air bubbles as before. Then fix the apparatus upon a board as shown in Fig. 40.

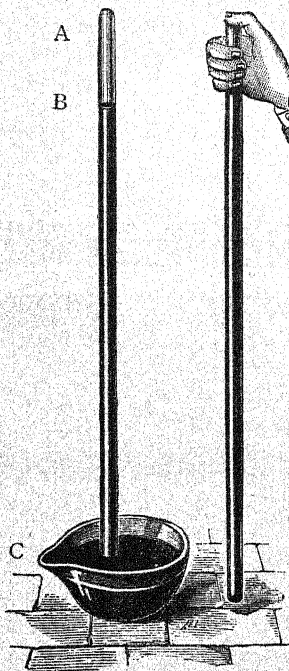


FIG. 39.

The level of the mercury in the closed arm is again about 30 inches above that in the open arm. If Expts. 7 and 8 were repeated with wider or longer tubes, the length of the mercurial column would still be found to be about 30 inches or 760 mm.

EXPT. 9. To show that A B is a Vacuum.—(1) Tilt the tube used in Expt. 7 gradually, keeping the unsealed end still below the mercury in the mortar. The mercury rises in the tube, and finally completely fills it (Fig. 41).

(2) Pour some water above the mercury in the mortar. Raise the tube carefully so that the open end C lies in the water. The mercury at once runs down, and water not only takes its place, but completely fills the tube.

It is evident, therefore, that the space A B is free from air. After the author of the experiment this space is usually called the **Torricellian vacuum**.

From this and the former experiments it appears that the *column of mercury B C must be supported by the pressure of the air on the mercury in the mortar.*

So also it appears that in Expt. 8 the column of mercury in the closed arm must be supported by the pressure of the air on the mercury in the short open arm.

EXPT. 10. To show that the Length of the Mercurial Column varies with the Pressure of the Air.—(1) Connect an exhausting air-syringe by thick rubber tubing to the short open arm of the apparatus used in Expt. 8. Exhaust some of the air from the short arm: the mercury immediately falls several inches in the other arm. Re-admit air: the mercury rises to its former level.

(2) Connect an air condenser in the same way with the open end. Gently force a little air into the short arm: the mercury is forced right up to the top of the closed arm.

Robert Boyle, who invented the air-pump, placed the whole apparatus used in Expt. 7 within the receiver of the air-pump. He found to his delight that the mercury fell further and further in the tube as the air was gradually exhausted, until it was only an inch or two above the level of the mercury outside. On re-admitting air into the receiver, he found that the mercury rose to its original level in the tube.

It is evident, therefore, that the height of the mercury in

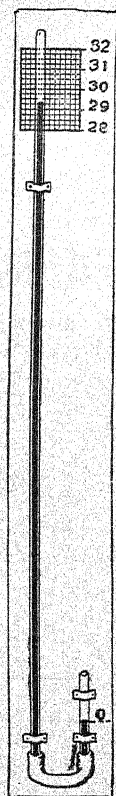


FIG. 40.

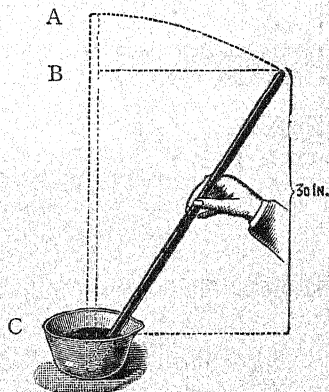


FIG. 41.

the tube above the level outside is a measure of the pressure of the air. The instrument was called by Robert Boyle a **barometer**,¹ and ever since his time it has been in use for this purpose.

Variations in the Atmospheric Pressure.—If we imagine the earth to be enclosed in a vast ocean of air, we should expect the pressure of the air to become less and less as we ascend above sea-level, and this should be made evident by our barometer. It was shown to be so by PÉRIER in 1648. He ascended the Puy-de-Dôme, in the Auvergne Mountains of France, and found that the mercury fell about 3 inches. The barometer even shows that in the course of any railway journey there are variations in the atmospheric pressure according to the height above sea-level.

Further, at any one place, for example at Manchester, the barometer shows that variations occur from time to time in the atmospheric pressure. These variations arise from the movements which are continually taking place in our atmosphere, including changes in the amount of moisture present in the air. Other things being equal, moist air is not so heavy as dry air: this may appear strange, and it will probably not be understood till later. Sometimes in a storm the height of the barometer will change by as much as 1 inch within 24 hours.

In the West Indies there are frequently very violent storms, during which the height of the barometer will vary far more rapidly than is the case in the British Isles. For instance, at St. Vincent on September 8, 1898, the barometer stood at 10 a.m. at 29.54 inches, and the wind was blowing strongly from the north-east. The wind then veered rapidly round, and by 11 a.m. a hurricane was blowing from the west, powerful enough to uproot the largest trees, and at 11.40 the barometer had fallen in 100 minutes more than 1 inch, down to 28.51 inches. The wind then died away, and there was a dead calm for three-quarters of an hour, and during this time the barometer was quite steady. At 12.25 p.m. the wind began again from the south, and by 3.0 p.m. the barometer had risen to 29.53 inches. During the day it also rained in torrents: between 9 a.m. and noon actually 4.9 inches of rain fell.

Note.—At this point, pp. 170-7 of Chapter XVIII. may be

¹ Greek: *baros*, weight, and *metron*, a measure.

taken with advantage, showing how to represent in a diagram a number of readings of a barometer.

A Weather Glass.—The weather glass shown in Fig. 42A, which is frequently seen hung up in houses, contains a barometer of the kind shown in Fig. 40. On the mercury in the short arm rests a little float, which will rise or fall as the level of the mercury alters. This float is connected by a string which passes

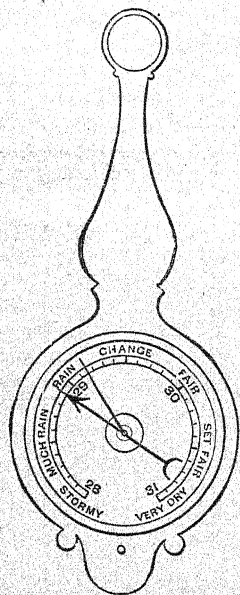


FIG. 42A.

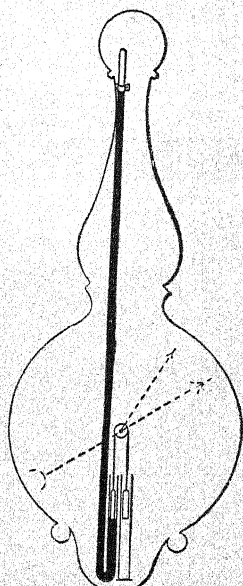


FIG. 42B.

over a wheel with a weight, which is slightly lighter than itself. The apparatus, as seen from the back, is shown in Fig. 42B. To the wheel is fixed a long needle, which is to be seen on the face of the apparatus. The motion of the float will clearly cause the wheel and also the needle to turn round. Numbers are engraved upon the face of the circle round which the needle can move, and the number to which the needle points shows the height of

the mercury in the tube. The words *rain, change, fair, etc.*, are also engraved upon the circle.

The Standard Atmospheric Pressure.—The *average* pressure, however, at any place on the sea-level is represented by about **30 inches** or **760 mm. of mercury**. It is customary, moreover, to regard the pressure which is measured by 760 mm. of mercury as the standard atmospheric pressure. In other words, the pressure exerted by a column of mercury 760 mm. in length is called a **pressure of one atmosphere**.

It is useful to remember that this pressure is about 14.7 lbs. weight upon every square inch, or 1033.6 grams weight upon every square centimetre.

A Water Barometer.—How long a column of water could be supported by the pressure of the air? We must find out how long a column of water would exert the same pressure as a column of mercury 30 inches in length.

EXPT. 11. To find out how many inches of Water will balance 1 inch of Mercury.—

Bend a long glass tube into the form of a **U**, and fix it vertically. Pour a little mercury down one arm, so as to fill the bend and rise about 2 inches up the arms; it stands at the same level in both arms (Fig. 43A). Then incline the tube, and introduce water by means of a pipette into one arm, so that the column of water is between 13 and 14 inches in length. Fix the tube vertically again. Measure CD, the length of the water column (Fig. 43B): it is, say, 13.6 inches. What balances this column of water? The short column of mercury EF *above* the level of the mercury in the other arm. Measure EF: it is 1 inch.

Thus a column of mercury 1 inch in length exerts the same pressure as that produced by a column of water 13.6 inches long.

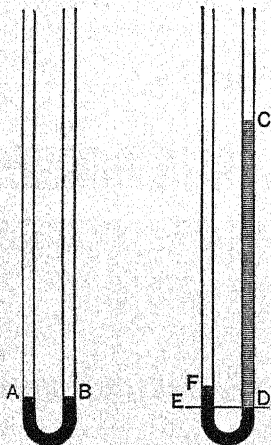


FIG. 43A.

FIG. 43B.

Hence, a column of mercury 30 inches long will exert the same pressure as that produced by a column of water $\frac{13.6 \times 30}{12}$ or 34 feet long. The tube of a water barometer must, therefore, be more than 34 feet in length.

Glycerine has occasionally been used as the liquid in a barometer: the tube must be more than 27 feet in length.

It is clear that in a water or glycerine barometer variations in height caused by the changes in the atmospheric pressure will be much greater and more evident than in a mercury barometer.

EXERCISES

1. Attach a U-tube containing water by means of a rubber tube to the coal-gas supply. What do you observe? Can the pressure of the gas be measured?

2. Find how many inches of glycerine will balance 1 inch of mercury.

CHAPTER VIII

FITTING UP APPARATUS

NOTHING is more frequently needed in any laboratory¹ than some source of heat. If a chemist had no fire or flame at his service, he would be able to do very little. The alchemists generally used some form of furnace. In quite modern times² coal gas has come into common use, and is supplied in every laboratory. Is an ordinary bright coal gas flame suitable for laboratory purposes?

Hold a piece of white cardboard for a few moments across a bright gas flame. It becomes very sooty. If, then, a glass vessel were heated in such a flame, it would quickly become coated with soot, and it would be difficult to watch the progress of what might be going on inside it. But it is frequently most important to be able to see the contents of a vessel. Consequently, a bright coal gas flame is not a satisfactory source of heat, and it would be a very great gain to have a flame which would not deposit a black curtain of soot upon any vessel heated by it. A burner which would give such a flame was devised by Robert Wilhelm Bunsen, Professor of Chemistry in the University of Heidelberg. A great teacher and a famous investigator, he died in 1899, at the ripe age of 88 years.

The Bunsen Burner.—If air be allowed to mix with coal gas before it is burnt, a blue flame hotter and smaller than the ordinary luminous gas flame results. Examine a Bunsen burner. At the bottom of the tube A (Fig. 44) is a collar B. Oval windows or holes are pierced in the tube and in the collar.

¹ See Appendix on Laboratory Fittings, and Lists of Apparatus and Chemicals

² The streets of London were first lighted with coal gas in 1812.

Either the tube or the collar can rotate so that these windows can be set opposite to each other, admitting air or alternately excluding air.

(1) Close the holes. Light the gas. The flame is yellow. Bring a piece of white cardboard into the flame; it becomes sooty.

(2) Open the holes. The flame becomes blue. Hold a piece of cardboard in it. It becomes brown and singed, but not sooty. Hold a match across the flame about one inch above the tube. The wood is burnt in two places, showing that there is an inner cooler region. Try whether you can introduce a match head into the middle of the flame without lighting it.

Is the blue flame hotter than the yellow flame? Yes. Does closing the windows make any difference to the steadiness of the flame? Yes, the yellow flame is taller, and much less steady.

(3) Light the burner with a blue flame. The tube contains a mixture of gas and air. Press gently upon the rubber gas supply tube, making the flame smaller and smaller. Suddenly the flame will shoot down the tube, and it will be seen *burning below*. Notice the curious smell. Light the burner at the top. The flame is yellow, although the holes are open. Any attempt to turn the tube will now lead to burnt fingers, since the inner flame makes the tube very hot. It is better to turn the gas out, and then light it again.

When a very small flame is required, the supply of air should be reduced as well as the coal gas.

Glass is useful to the chemist on account of (1) its trans-

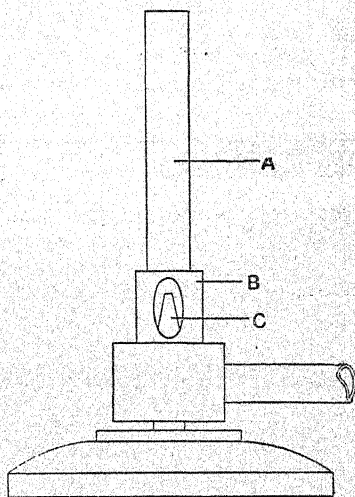


FIG. 44.

parency, (2) the ease with which it can be worked into any shape when hot and plastic, (3) its not being attacked by most chemicals. Its disadvantage is its brittleness. It is particularly liable to break on heating, because it is a bad conductor of heat, and therefore one part of the glass expands on being heated before neighbouring parts become warm.

To Illustrate the Liability of Glass to Crack when Heated.—(1) Heat one end of a thick, hard glass tube in the flame of the foot-blowpipe. The glass cracks before the fingers feel the heat.

(2) Warm another piece of glass tube very gradually. Then pour cold water on it. It cracks.

(3) Fill a test-tube one-third full of water. Hold it in the flame so that the flame plays on the glass near the water line. The tube cracks.

To Boil Water in a Glass Vessel.—(1) Try again to warm water in a test-tube, as follows. Bend a doubled strip of paper back on itself round the upper part of the tube, and hold the tube by the paper. Let the flame play on the bottom of the test-tube. Shake the tube gently so as to keep the water dancing up and down the sides of the tube, and so warm them gradually. The water boils without accident.

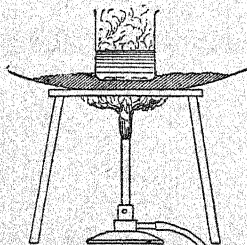


FIG. 45.

(2) Place a piece of iron *wire gauze* on a tripod-stand, or on a ring of a retort stand. Place a glass beaker or flask containing 50 c.c. water on the gauze. Put a Bunsen flame under the gauze. The gauze distributes the heat over the bottom of the vessel, and the water can be heated until it boils without accident.

(3) Place a little sand in an iron tray, and put the tray upon a tripod-stand (Fig. 45). Place a beaker containing 50 c.c. water upon the sand, and heat the tray with a Bunsen burner. The heat is distributed by the sand, so that the beaker can be heated safely until the water boils. Observe that the water heated over wire gauze boils in much less time than that heated on the sand.

The tray containing the sand is called a *sand-bath*.

To Make a Wash-Bottle

It will generally be possible to fit up for ourselves the apparatus required for our experiments. We shall require glass flasks, corks, glass tubing, and india-rubber tubing, and some knowledge of how to use them.

Suppose we learn in the first place how to make the apparatus shown in Fig. 46. This is called a *wash-bottle*. It is useful for sending a stream of water in any desired direction. By blowing at *d* a jet of water spurts from *c*, and the india-rubber connection *e* allows of the water being sent in any direction. We shall require an 8-oz. flask, a cork, about 50 cm. of glass tubing, and a small piece of rubber tubing.

The cork (wooden) selected must be just too large to fit into the mouth of the flask. It can be made to fit the neck by rolling and squeezing it under the foot. Begin gently, turning it round and round, and press harder and harder. The cork now fits well into the neck of the flask, and being somewhat elastic presses tightly against the glass all round.

An india-rubber stopper may be used instead. It should fit the mouth easily, since it will be enlarged when the glass tubes are passed through it.

To bore a Cork.—(1) A hollow brass tube sharpened at one end may be used, called a cork-borer. Choose one a little narrower than the glass tubes which are to go through the cork. At the narrow end of the cork put marks where the holes are to be made. Stand the cork on the table, narrow end up. Keep the

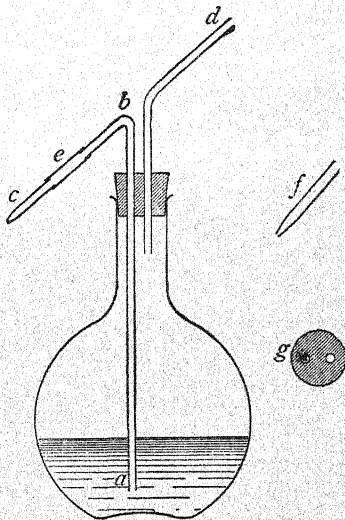


FIG. 46.

borer carefully upright. Drive it gently into the cork with a screw motion, half a turn at a time, pausing to see that it keeps upright. Along with a set of cork-borers is supplied a little solid brass rod, which can be used both to press the core of cork out of the borer and as a handle to turn the borer round. (The borers will require sharpening periodically.)

(2) Or, heat the end of a rat-tail file in a flame. Drive it gently through the cork. Enlarge the hole with the file, but take care that this is done evenly, so that the hole remains round. It should not be made so large that the tube fits loosely into the hole.

To cut Glass Tubes.—Take a sharp triangular file. Make a nick in the glass tube about 15 cm. from one end. Hold the tube in both hands. The two forefingers should be together

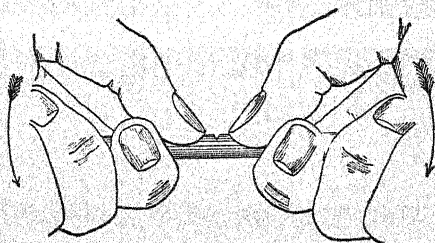


FIG. 47.

immediately underneath the nick. The two thumbs are above. Turn the wrists outwards and downwards, round the forefingers as axes, at the same time pulling the hands apart (Fig. 47). The glass cracks clean across.

To bend Glass Tubing.—Hold a piece of glass tubing lengthwise along a *flat* gas flame, as shown in Fig. 48 (fishtail or batwing, not the Bunsen flame). Keep it moving slowly round and round, and also to and fro. As soon as it feels slightly soft, hold it by one hand alone, and cease rotating it. The free end will gradually fall (Fig. 49). As soon as it is bent to the required angle, remove it from the flame. A considerable length has been evenly heated, and a smooth bend has resulted. Now cut two lengths of 15 cm. and 25 cm., and bend them at angles of 135° and 45° respectively, as shown in Fig. 46.

Mistakes to avoid in bending Glass.—(1) Hold a piece of glass tubing across the blue Bunsen flame. It becomes hot

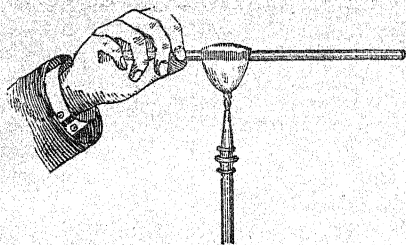


FIG. 48.

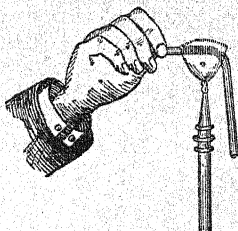


FIG. 49.

in two places. When soft, bend it. The glass is found bent in two places. (2) Hold a piece of glass tubing across an ordinary flat gas flame. The tube is heated this time in one place. Bend it. The tube is flattened on the outside of the bend and thickened on the inside. When cold, try to bend the tube. It snaps easily at the bend.

To round the Ends of Glass Tubing.—The sharp edges left at the ends of the glass tubing must be rounded, otherwise the tubes will not slide easily through corks, and are apt to cut any rubber tubes that are slipped over them. Round the ends of the tubes just made—(a) by grinding the sharp edge on a file or on a stone floor, (b) by holding the end in the edge of the blue flame, at a point about two-thirds up the flame, as shown in Fig. 53, until the glass just melts at its edges.

To draw out Glass Tubing.—

(1) Holding a piece of glass tube in both hands, heat it at one place in the *blue Bunsen* flame. Remove it when soft from the flame and pull.

The tube is drawn a little at A (Fig. 50). At B it is pulled



FIG. 50.

out much finer, because it happens to be hotter, and it has broken there because the pull was too hard. The lump C was

in the cooler central portion of the flame, and never became properly heated.

(2). Heat another tube in a *flat flame*, holding it lengthwise and turning it slowly round and round. Pull the two ends apart. There is no lump, because the tube was heated evenly,



FIG. 51.

but the finely drawn out tube has softened at once, and is bent down (Fig. 51).

(3) Heat another tube as in Expt. 2 until the tube feels limp in the fingers. *Remove it from the flame.* Draw slowly at first, then more quickly, but gently rotating each end until the hands are wide apart. Hold it rigidly for a few seconds, until the glass stiffens (Fig. 52). The glass is now drawn out perhaps a yard long or more. Cut the tube with a very sharp file gently at A and B. Notice how elastic this section A B is. It can be bent nearly to a circle. If bent too far, it snaps, and the pieces spring back straight again. Dip the end B into water, and blow at A. Bubbles arise. Dip another piece of this fine tube into an inkpot, and note the result. Then, A B is



FIG. 52.

really hollow all through. The end A C can be used for the wash-bottle jet. The very narrow tube A B is called "capillary tubing" (Lat. *capilla*, a hair).

The foot blowpipe flame is better than a flat flame for drawing out glass tubing.

To narrow or close an End of a Glass Tube.—Incline a glass tube 20 cm. long downwards at an angle of 60° , and hold the lower end in the edge of a Bunsen flame at a point about two-thirds of the height of the flame, as shown in Fig. 53. The edge becomes red-hot, fuses, and gradually contracts.

Remove the tube from the flame before it is quite sealed up. Cut the glass about 4 cm. from the end. This portion can be used for the wash-bottle jet.

Hold one end of the remainder of the tube for a longer time in the flame, and close the end entirely.

Rubber tubing sometimes becomes stiff when not in use. Warm gently and pull it about ; it soon softens. Cut off a piece an inch and a half long for future use. If a rubber tube is too wide for a glass tube, it can be made to fit by doubling it back on itself.

To fit together the Wash-Bottle.—

Moisten the end *a* of the long tube *ab* (Fig. 46). Push the long tube through the cork with a zig-zag motion. Hold the tube near the cork, and avoid holding it at the bend *b*, where it might break and cause a nasty cut in the palm of the hand. Fit *d* in the cork also. Now fit the cork into the bottle. The end *a* should come nearly to the bottom of the flask, but not so near as to risk touching it.

NOTE.—Tubes should never be pushed into corks when they are fixed in the mouths of flasks ; nor should corks be fixed in the necks of flasks when the flasks are standing on the bench ; hold the flask in the left hand.

To test whether the Flask is Air-tight.—Close the glass tube at *e* and blow at *d*. Does the air hiss out through the cork ? Again closing *e* suck at *d*. The tongue is sucked back, and held at the end of the tube. Attach the jet *c* by means of the rubber tubing. If the flask is now filled with water, it is ready for use.

To use the Wash-Bottle.—

- (1) Blow at *d*. A fine jet of water issues from *c*.
 - (2) Invert the bottle. A large stream of water issues from *d*.
- It is often well to keep the bottle filled with distilled water. If a supply of hot water is required, the bottle can be heated over a flame upon wire gauze.

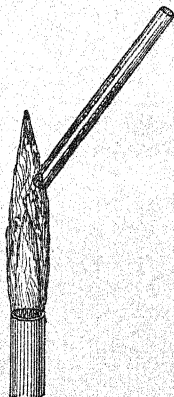


FIG. 53.

LABORATORY RULES, if not provided by a teacher, should be made by the student for himself:—

1. Each boy must use his own apparatus and work at his own place. Experiments, calculations, or notes must not be made at the bench reserved for balances.
2. Clean apparatus, and put it away in its proper place as soon as done with.
3. Bench, drawers, and cupboards must be left clean and in order. Test-tubes must be left clean, and half full of water.
4. Solids (*e.g.* matches, filter-paper, rubbish) must not be thrown into the sink, but be put into the refuse pan.
5. Never pour strong acids down the sink. Dilute them with water before throwing them away.
6. Stoppers must not be left out of their bottles on the bench, but be replaced.

EXERCISES

Working of Glass. Fitting up Apparatus

1. Cut a tube 10 cm. in length. Round the edges. Bore a cork to fit the tube tightly. Pass the tube through the cork.

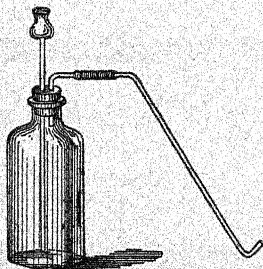


FIG. 54.

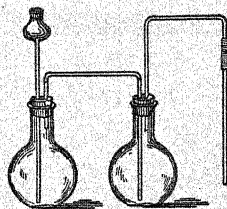


FIG. 55.

2. Prepare tubes as follows:—

a tube	15 cm.	long	bent	at	120°	in	the	middle.
"	15	"	"	"	90°	"	"	"
"	15	"	"	"	60°	"	"	"
"	40	"	"	"	into an equilateral triangle.			

3. Cut a tube 40 cm. long. Bend it at right angles near one end, and at 60° near the other end, taking care that both bends are in the same plane.

4. Cut two tubes 10 cm. long. Round their edges. Bore two holes through a cork to fit the tubes. Fix the tubes through the cork.

5. Cut a tube 25 cm. long. Draw it out near one end, and cut it

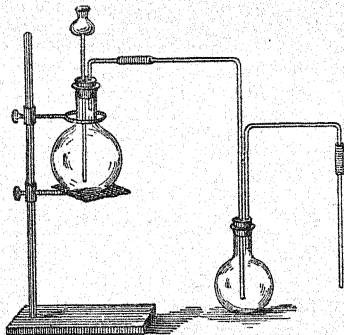


FIG. 56.

in two. The shorter pieces can serve as wash-bottle jets, the longer ones can be used as pipettes.

6. Practise pulling out glass tubes to full arms' length. Break up the capillary tube so formed into 6-inch lengths, and save them.

7. Fit up the pieces of apparatus shown in Figs. 54, 55, and 56.

CHAPTER IX

CHANGE OF STATE—LIQUEFACTION AND SOLIDIFICATION

We are acquainted with many different solids and liquids.

What is the effect of heat upon them?

Many solids are changed into liquids by heat, which solidify again on cooling.

EXPT. 1.—Place some powdered **sulphur** in a test-tube. Heat very gently in a Bunsen flame. The sulphur quickly melts, and forms a yellow liquid. The liquid becomes thick like treacle, and of a dark red colour. Then it becomes more liquid again, and begins to boil. Let it cool. The liquid quickly solidifies again, forming a solid cake of sulphur.

EXPT. 2.—Place a lump of **paraffin wax** in a test-tube, and heat it gently in a flame. The wax melts very quickly, and forms a clear liquid. Let it cool. The liquid is quickly changed into an opaque white solid mass.

EXPT. 3.—Place a few pieces of **tin** in an iron spoon. Support it on a tripod. Heat in a good Bunsen flame. The tin melts completely. Skim off the earth-like dross on the surface of the molten metal with an iron file. Pour the liquid in a thin stream into a mortar full of water. It is converted into a bright crumpled mass of the solid metal again.

EXERCISES

Describe the effect of heat and subsequent cooling upon the following substances. In each case describe the appearance of the substance before heating it.

- | | |
|-----------------------------------|------------------|
| (1) sealing wax. | (6) lead. |
| (2) shellac. | (7) zinc. |
| (3) vaseline. | (8) copper. |
| (4) bees' wax. | (9) common salt. |
| (5) material of a stearin candle. | |

A further examination of the effect of heating metals in the air and of the dross that appears upon the surface of molten metals will be made in Volume II. Chapter II.

Let us study one particular case more fully.

EXPT. 4. The Effect of Heat upon Ice.—Break some ice into small fragments, the smaller the better. Pack them into the lower half of a wide beaker. Plunge a thermometer into the mass. The temperature is 0° C. Warm the beaker on wire gauze over a very small flame, and stir the mass continually with the thermometer, occasionally removing the flame for a few seconds. Read the thermometer frequently. The ice rapidly melts. Although so much heat is added, nevertheless the temperature remains fixed at 0° to 1° C., until the whole of the ice is melted. Only then does the temperature begin to rise, and continue to do so.

It is evident that while ice is being liquefied *no rise* of temperature occurs. In other words, the change of state takes place at a definite temperature. This temperature is called the **melting point** of the ice.

Other substances behave similarly to ice, and we may now proceed to determine in a simple way the melting points of some of them.

***EXPT. 5. To find the Melting Point of Paraffin Wax.**—Take a narrow glass tube. Close one end by holding it in a flame. Introduce some small particles of paraffin wax to a depth of about 0.5 cm. Attach it to a thermometer by means of a thin slice of rubber tubing so that the wax is opposite the middle of the bulb. Support a beaker, of about 100 c.c. capacity, half-filled with water, over wire

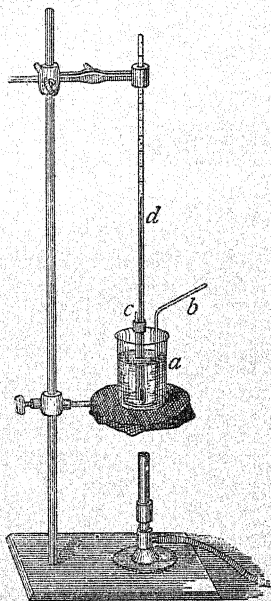


FIG. 57.

gauze on a ring of a retort-stand. Place a glass stirrer, *b*, inside, the end of which has been bent into a circle at right-angles to the stem. Support the thermometer so that its bulb is in the centre of the water (Fig. 57). Heat gently, stirring constantly. As soon as the paraffin wax melts, read the temperature. Probably you have been too late, and the reading is too high.

Allow the water to cool slightly. Take a fresh capillary tube containing paraffin wax and ascertain the *exact* melting point, removing the flame from time to time, and stirring thoroughly, so that the temperature rises very slowly. Read to the nearest degree.

EXERCISES¹

Find the melting points of (1) iodine, (2) naphthalene, (3) washing-soda, (4) Glauber's salt, (5) sulphur, (6) "hypo," (7) phenol.

We have found that different solids melt at different temperatures. Usually, moreover, the melting point is in each case very definite. Pure sulphur, for example, has a well-defined melting point. Impure substances, on the other hand, do not melt entirely at one temperature, but gradually through a range of temperature. Hence, use may be made of melting points in order to distinguish and identify substances, and to detect the admixture of foreign ingredients. The melting point is indeed one of the best defined of physical "constants." Here are a few melting points.

TABLE OF MELTING POINTS

Platinum	1770° C.	Naphthalene ...	79° C.
Gold	1062°	Paraffin, about...	54°
Copper	1054°	"Hypo"	48°
Silver	961°	Washing-soda ...	34°
Common salt ...	851°	Butter, about ...	33°
Zinc	419°	Glauber's salt ...	32°
Lead	326°	Ice	0°
Tin	233°	Mercury	-39°
Sulphur	115°		

¹ It will be found necessary to use strong sulphuric acid in place of water in the case of iodine and sulphur.

Latent Heat of Liquefaction.—We found that while ice was being melted no rise of temperature occurred, in spite of the addition of so much heat. In a similar way heat apparently disappears while any other solid is being changed into a liquid. This heat which alters the state of a solid, but does not raise its temperature, and which therefore apparently disappears, was first investigated by Dr. Black of Edinburgh about the year 1760. The question may be asked, whether, when such a liquid is reconverted into the solid state, the heat will become sensible again.

EXPT. 6. To find whether Heat is given out when a Liquid solidifies.—Take some crystals of a substance known to the photographer as “hypo,” and to the chemist as sodium thiosulphate. Put them in a perfectly clean flask. Heat gently, until the crystals are completely melted. Introduce a thermometer. Place a plug of cotton-wool in the neck of the flask. Allow to cool in a quiet place. Then remove the plug, and introduce a small crystal of sodium thiosulphate. The liquid quickly solidifies. At the same time the thermometer rises many degrees, and the flask becomes quite hot.

It is evident then that the heat, which apparently disappears when a solid is liquefied, can be recovered if the liquid be again caused to solidify. Hence the heat is said to be hidden or latent, and it is usually spoken of as **latent heat of liquefaction**.

Freezing Mixtures.—Advantage is taken of the fact that heat becomes latent in the liquefaction of any substance to produce “freezing mixtures.” For, supposing the liquefaction of a substance is induced to take place, and yet no heat is supplied from any external source, the heat which must be absorbed in the process of liquefaction will be taken out of the substance itself or out of anything in contact with it, and a considerable fall of temperature may ensue. For instance, if 1 part of common salt is mixed intimately with 4 parts of pounded ice or snow, the mass liquefies, because the freezing point of *brine* is far below that of water. Hence, at the temperature of the experiment, the mixed materials ought to be liquid. But in order to liquefy they must absorb heat. They, therefore, sacrifice their own temperature in order to get this heat, and the temperature sinks to about -22°C .

PROBLEMS

Find whether freezing mixtures can be made from—

- (1) Glauber's salt and strong hydrochloric acid.
- (2) Ice and crystallised chloride of calcium.
- (3) Salt and water.
- (4) Sal-ammoniac and water.

CHAPTER X

CHANGE OF STATE—VAPORISATION AND CONDENSATION

We have found that many solids are changed into liquids by heat. What is the effect of heat upon liquids?

Many liquids are changed into vapours by heat, which liquefy again or solidify on cooling.

EXPT. 1.—Place a few crystals of **iodine** in a test-tube. Heat the end of the tube very cautiously, holding the tube nearly level. Some of the iodine melts, but at once heavy violet vapours are given off. A dark deposit forms on the upper and cooler part of the tube. Examine it carefully. Crystals of iodine have been formed. The violet vapour of iodine when cooled has become solid iodine again. The vapour is said to be **condensed**. Finally, all the iodine disappears from the bottom of the tube

EXPT. 2.—Place a drop of **mercury** in a test-tube. Heat it carefully in a flame, holding the tube nearly level. A mirror of minute globules of mercury is soon seen upon the upper and cooler portion of the tube, although there is no visible vapour. Then the mercury boils, and the globules grow larger as the hot invisible vapour is cooled and condensed, and some of them may roll down and again be vaporised

EXPT. 3.—Heat some **water** in a flask over wire gauze. A rolling vapour is soon seen within the flask, and a dew is formed on the cool sides of the flask. Bubbles have appeared on the bottom of the flask, and some small ones may ascend through the water. The bubbles grow larger, and more and more reach the surface of the water, which is at last completely broken up, and a white cloud of "*steam*" is seen at the mouth of the flask. The water is then said to *boil*. Nothing is now seen in the

flask above the water. The white cloud consists of tiny drops of water—"water-dust" it has been called—condensed by the cold air from the invisible water vapour.

In all these three experiments liquids are found to change into vapours, and then on cooling into the same liquids again. The *state* of each substance is changed, but there is no change in the *stuff* of which each consists. The solid iodine, for instance, is changed into liquid iodine, and this into the vapour of iodine, and the iodine vapour back to the solid iodine again; the substance is all the time *iodine*; it is not changed into a different substance; only its state is altered. Such changes are called **physical changes**.

Let us study one case more fully:—

EXPT. 4. To find the Temperature of the Vapour from Boiling Water.

—Take a large flask containing some water, and fit it with a cork. File two grooves down the side of the cork, and pierce it with a hole, through which a C. thermometer can be thrust, so that its bulb is just above the surface of the water. Place the flask on wire gauze, and heat it (Fig. 58). As soon as the water boils, and steam is issuing freely from the mouth of the flask, read the temperature six times at intervals of about one quarter of a minute. Does the temperature rise? No, *it remains at about 100° C.*

Thus the change or state of water to the vapour of water occurs at a definite temperature, viz. 100° C. This is called the **boiling point** of water.

Is the same true of other liquids? Does each possess a definite boiling point?

EXPT. 5. To find the Temperature of Boiling Chloroform.¹—Take an 8 oz. flask. Support a dry test-tube within it

¹ Suggested by H. P. Highton in *Practical Quantitative Analysis* (Longmans), p. 124.

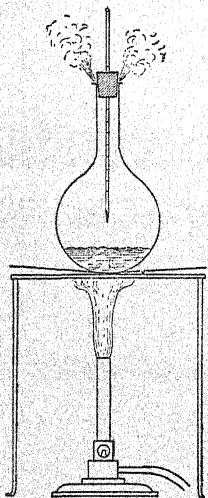


FIG. 58.

either by the neck of the flask, or if this is too large by a cork. Pour chloroform into the test-tube to a depth of about 2 cm. Add cold water to the flask, so that the test-tube just dips into the water. Place the flask on wire gauze, and fix a thermometer in a clamp, so that the bulb is within the test-tube, and just above the surface of the chloroform (Fig. 59). See that it does not touch the side of the test-tube. Place a small flame underneath, and drop two or three grains of sand into the chloroform, which will help it to boil evenly. As soon as the chloroform begins to boil vigorously, remove the flame, and read the thermometer five or six times. The heat of the water is sufficient to keep the chloroform boiling. What is the boiling point of chloroform? About 60°C .

Different liquids are thus found to boil at different temperatures. Moreover, when a liquid is pure, the boiling point is usually very definite and readily determined. Hence, it should be noted if possible in the case of every pure liquid.

TABLE OF BOILING POINTS

Sulphur	440°C .
Mercury	357°
Toluene	$110^{\circ}\cdot 3$
Water	100°
Benzene	$80^{\circ}\cdot 4\text{C}$.
Alcohol	$78^{\circ}\cdot 3$
Chloroform	$60^{\circ}\cdot 2$
Ether	35°

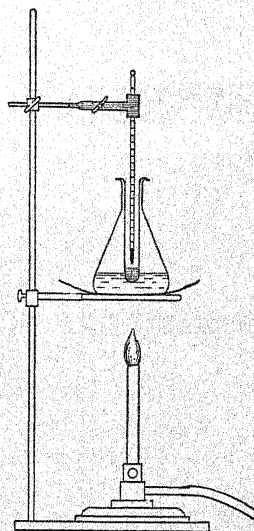


FIG. 59.

Let us repeat the vaporisation of water in another way, so as to collect the condensed water.

EXPT. 6. The Distillation of Water.—(1) Take an 8 oz. flask, *a*, and half fill it with water. Introduce a few pieces of broken glass or of clay pipe. Take a piece of glass tubing about 80 cm. long, and bend it about 10 cm. from one end, as shown. Fit the tube to the flask by means of a cork. Support the flask on wire gauze upon the ring of a retort-stand. Place

the other end of the glass tube within a flask, *b*, underneath a tap of cold water (Fig. 60). Heat the flask. As soon as the water has begun to boil, the steam passing down the tube is condensed on its cold surface, and when the glass tube becomes very hot, in the other flask.

The condensation of the steam, and the collection of the water produced, is spoken of as **distillation**. Taste the water. It is flat and insipid.

(2) The following is a better apparatus for the distillation of water. Take a flask, and fix in its neck by means of a cork a

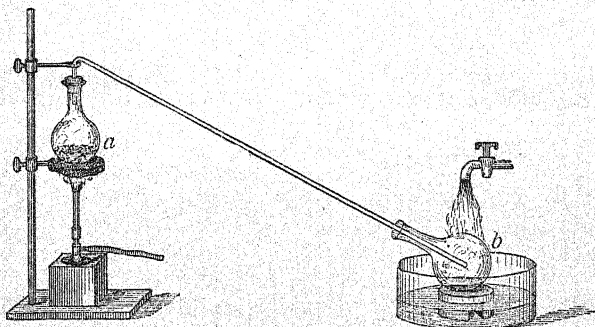


FIG. 60.

thermometer and a bent glass tube, as shown in Fig. 61. Take also a "**Liebig's condenser**," which consists of a long narrow tube, the central portion of which is surrounded by a wider tube, so that cold water can be circulated between them. Select a cork which fits the tube of the condenser. Fix the cork on the tube in the flask. Place some water and a piece of clay pipe in the flask. Then, supporting the flask upon wire gauze, connect it with the condenser. Circulate cold water through the condenser. Heat the flask. When the water has begun to boil, the vapour, passing the bulb of the thermometer and down the tube, is completely condensed on the cool surface of the tube. Collect in a flask the water thus condensed. Meanwhile read the thermometer. It stands constantly at 100° C.

Non-Volatile Liquids.—Some liquids cannot be distilled, or converted into vapours which will yield the same liquids again when cooled.

EXPT. 7. Olive Oil cannot be distilled.—Half fill a small retort with olive oil. Fit a thermometer in a tubulure, so that the bulb is in the oil. Introduce the neck of the retort into a small flask. Heat the retort over wire gauze. Observe that the temperature rises continually higher and higher. As soon as 300° C. is reached, remove the thermometer, and replace the

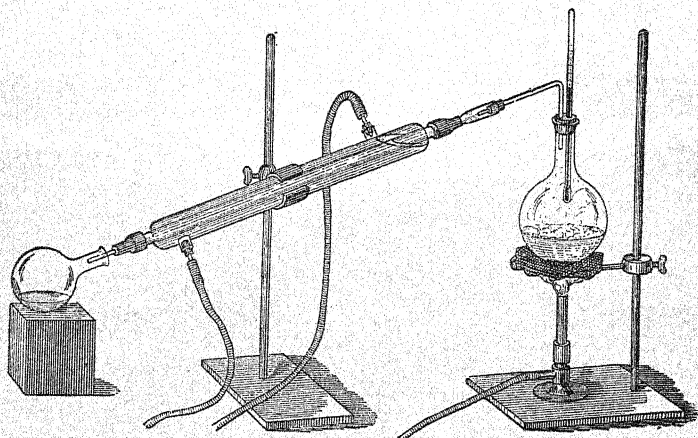


FIG. 61.

stopper of the retort. Heat further. After a time vapour is at last evolved, and being condensed in the neck, a few drops of liquid are collected in the flask. Then remove the flame. Notice the odour of the liquid in the receiver. It is very acrid, and clearly not that of unchanged olive oil. Indeed, this liquid is not olive oil at all, but another substance produced by the heat breaking up the oil.

Liquids such as olive oil, which cannot be distilled, are said to be **non-volatile**.

EXERCISES

1. Find the boiling points of—
 - (1) turpentine.
 - (2) benzene.
 - (3) pure alcohol.
 - (4) toluene.

2. Distil a mixture of 100 c.c. water, with 100 c.c. methylated spirit in the apparatus of Fig. 61. Read the temperature of the thermometer every minute. Collect the distillate in separate samples of about 25 c.c. each. Find the relative density of the first and last samples.

CHAPTER XI

CHANGE OF STATE—THE EVAPORATION OF WATER AND OF OTHER LIQUIDS

The Evaporation of Water.—We are often, perhaps, inclined to suppose that it is only when water is boiled that it is converted into vapour. Yet this is far from being true. Wet clothes become dry when hung in the open air, particularly on a windy day, and the puddles in the road after rain are quickly dried up in the sunshine. The quiet and invisible passage of a substance from the liquid to the gaseous state at all temperatures is called **evaporation**.

So large a part of the surface of the globe is covered with water, in sea, river, or lake, that an enormous amount of evaporation is always going on. It is evident, therefore, that the air must always contain a very large amount of aqueous vapour.

But is there any limit to the amount of aqueous vapour which the air may contain, and what conditions affect the rate at which evaporation proceeds? We must now search for answers to these questions.

Valuable information about the evaporation of liquids can be obtained by introducing them into the Torricellian vacuum of a barometer tube.

EXPT. 1. The Evaporation of Water in a Vacuum.—Fill two dry glass tubes, A and B, each sealed at one end, with warm dry mercury, and invert them in a trough of mercury (Fig. 62). Introduce a drop or two of water into B under the mercury by means of a pipette, C, with a curved end. The water, being lighter than the mercury, rises to the surface in

the tube. The mercury column at once falls a considerable distance. This cannot be due to the weight of the drop of water. Observe closely the top of the mercury column. The drop of water has nearly or entirely disappeared. The water has then evaporated very rapidly in the vacuum. The depression of the mercury must be due to a *pressure* which the invisible aqueous vapour is exerting, in the same way that air exerts a pressure.

Supposing we introduce some more water, will the mercury sink further? We must test this by experiment. Add a few more drops of water, so that a little water rests upon the top of the mercury column. The mercury does not sink any further after this, even though two or three drops of water are added. Measure the difference in level of the mercury in the two tubes. This difference measures the pressure which the aqueous vapour is exerting.

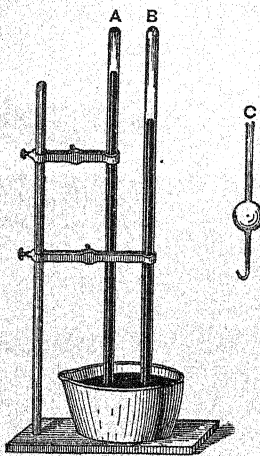


FIG. 62

From this experiment it appears that (the temperature remaining unchanged)—

(1) Water evaporates very rapidly in a vacuum.

(2) It is not possible for an unlimited amount of water to evaporate into a given space, but only a certain maximum amount. The

space is then said to be **saturated**.

(3) When water is allowed to evaporate into a given space, the aqueous vapour exerts a pressure, but this pressure cannot exceed a certain maximum.

What will be the effect of an increase in temperature upon the aqueous vapour pressure? Gently warm the upper end of the tube containing the water vapour with the hand, and then with a Bunsen burner. The mercury sinks still further. Hence:—

(4) As the temperature is gradually raised, the pressure

exerted by aqueous vapour is also increased. Can we measure the pressure corresponding to any particular temperature?

Dalton's Experiment for measuring the Pressure of Water Vapour.—John Dalton was the first to carefully study this subject, and his own account of his experiments may be quoted :—

“I take a cylindrical glass tube open at both ends, and of two inches diameter and fourteen inches in length, to each end of which a cork is adapted, perforated in the middle so as to admit the barometer tube to be pushed through and to be held fast by them ; the upper cork is fixed two or three inches below the top of the tube, and is half cut away so as to admit water, &c., to pass by, its service being merely to keep the tube steady. Things being thus circumstanced, water of any temperature may be poured into the wide tube, and thus made to surround the upper part or vacuum of the barometer, and the effect of temperature in the production of vapour within can be observed from the depression of the mercurial column.”

The experiments previously described show that water evaporates at ordinary temperatures below 100° . The pressure of such vapour formed at temperatures below 100° is often called **aqueous vapour pressure**, while the pressure exerted by the vapour arising from boiling water is commonly spoken of as *steam pressure*. The term *steam*, as commonly used, means either the vapour arising from boiling water, or else aqueous vapour out of contact of water heated to temperatures above 100° .

The following table shows the pressure of aqueous vapour for various temperatures, measured in millimetres of mercury, that

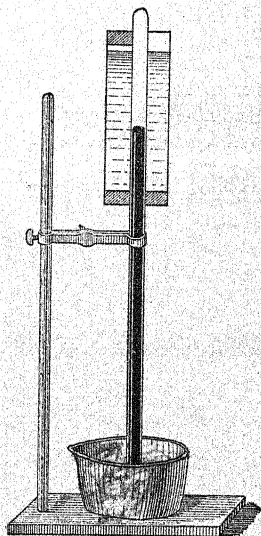


FIG. 63.

is to say, by the number of millimetres of mercury through which the aqueous vapour at each temperature would depress a column of mercury.

PRESSURE OF AQUEOUS VAPOUR

Temperature.	Pressure.	Temperature.	Pressure.
0°	4' + mm.	40°	54' + mm.
5	6' +	60	148' +
10	9' +	80	354' +
15	12' +	90	525' +
20	17' +	100	760' +

Saturated and Unsaturated Aqueous Vapour.—We have seen that if water be placed in a closed vessel the water evaporates very quickly if the vessel is vacuous, but that there is a certain maximum pressure corresponding to any given temperature which cannot be exceeded. If the vessel contains air, it is found that the only difference is that the evaporation is slower: gradually the maximum pressure corresponding to the temperature is attained, and cannot be exceeded. In either case the air is said to be **saturated** with aqueous vapour.

Briefly, then, at any definite temperature the air is capable of taking up only a fixed amount of aqueous vapour.

Suppose the temperature of air which is saturated with aqueous vapour be reduced, what will be the consequence? At the reduced temperature the maximum vapour pressure is less than at the original temperature. Hence some of the aqueous vapour will have to be parted with. What can become of it? Will it not be condensed as water or dew again?

We are really well acquainted with many instances of this. In a hot room there may be a large amount of aqueous vapour, yet not sufficient for it to be saturated. Close to the windows on a cold night, however, the temperature of the air will be greatly reduced, so that there will be more than sufficient aqueous vapour present for it to be saturated, and as we know a heavy dew will be deposited upon the cold glass. So also the dew often found in the morning on the grass, and on the surface of leaf and flower and cobweb, comes from the cooling of the air around them to such a temperature that the invisible aqueous vapour is in excess of that which can be contained in the air at this temperature, and it is therefore deposited.

EXPT. 2. To show the Deposition of Dew from the Air of the Room.—Make a freezing mixture with Glauber's salt and hydrochloric acid in a test-glass. Notice the deposition of dew on the outside of the glass.

If air containing aqueous vapour be slowly cooled, the temperature is sooner or later reached at which the air is saturated with vapour. The deposition of dew will begin at this temperature, and it is this temperature which is called the **dew-point**.

The "dew-point" is the temperature at which the aqueous vapour actually present in the air is sufficient to saturate it.

The Boiling of Water.—In the above table the number 760 mm., representing the pressure of aqueous vapour at 100° C., shows that the *pressure of aqueous vapour at 100° is equal to the normal atmospheric pressure*. Now it is exactly at this temperature that water evaporates in that rapid and violent way which we speak of as the *boiling* of water. This suggests that a liquid boils at a temperature such that the pressure of its vapour at that temperature is equal to the pressure of the atmosphere. How do experiment and experience bear this out?

If this be so, it follows that water should boil on a mountain at a lower temperature than at the sea level, for on the top of a mountain the atmospheric pressure is less than at sea level. This conclusion is borne out by Professor Tyndall's discovery that on the top of Mont Blanc water boils at 85° , or 15° below the temperature at which it boils at the sea level. Indeed, for every 600 feet above sea level the boiling point of water is found to fall about 1° F. But we do not need to ascend a mountain in order to alter the boiling point of water. At any given place the height of the barometer varies, and consequently the boiling-point of water is found to vary, slightly it is true, but perceptibly. It changes 1° C. for an alteration of 27 mm. in the height of the barometer.

EXPT. 3. Franklin's Experiment to show that Water will boil below its ordinary Boiling Point when the Pressure above it is lowered.—Fit a strong round-bottomed flask with rubber stopper, glass tube, rubber tube, and clip, as shown in Fig. 64. Introduce some water, and

support the flask over wire gauze. Open the clip, and boil the water vigorously till all the air is expelled, and its place taken by aqueous vapour. Then remove the flame, and close the clip. Allow the flask to cool until all boiling ceases. Then pour some cold water over it. The water begins to boil. The reason of this is that a large portion of the aqueous vapour in

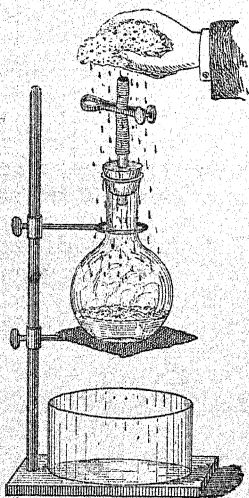


FIG. 64.

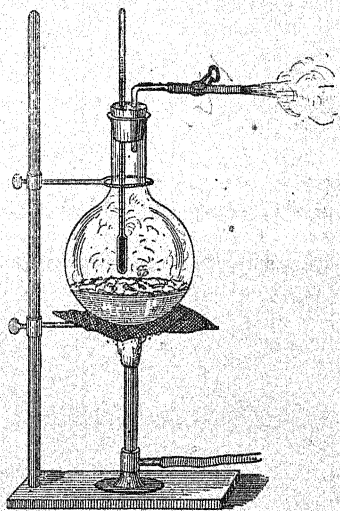


FIG. 65.

the flask is condensed through the pouring on of the cold water. Consequently the pressure on the surface of the water is so far diminished that it is less than the maximum pressure of aqueous vapour at the temperature of the water, and the water therefore boils.

PROBLEM

Place a flask containing hot water under the air pump, or connect it with a Bunsen pump. Can you make the water boil?

Then again, supposing the pressure above the surface of water is increased, should not the boiling point be raised?

EXPT. 4.¹ To show that the Boiling Point of Water can be raised above 100° C.—Pour some water into a strong flask, and close it with a rubber stopper, through which are passed a thermometer and a glass tube bent at right-angles (Fig. 65). Attach to the external end of the tube a piece of rubber tubing bearing a screw clip. Leaving the clip open, heat the flask on a flame until the water boils, and observe the temperature. Then close the clip until only a narrow exit is left for the steam. The result is that the pressure within the flask is gradually increased, and consequently the water boils at a higher temperature. *As soon as the temperature is seen to rise 2°, release the clip at once.*

Definition.—The boiling point of a liquid is that temperature at which the pressure of its vapour just overcomes the pressure of the atmosphere upon the surface of the liquid.

EVAPORATION OF OTHER LIQUIDS.

We have considered the evaporation of water in some detail. Do other liquids evaporate at temperatures below their boiling points? Can the statements made of aqueous vapour be applied to the vapours of other liquids besides water?

EXPT. 5. To compare the Evaporation of Liquids.—Place drops of ether, chloroform, alcohol, benzene, water, and olive oil upon a slate. Observe that some disappear or are evaporated before others.

EXERCISE

Repeat Expt. 1, using (1) alcohol, (2) ether, (3) chloroform, in place of water.

The mercury sinks further when alcohol is used than with water, and still further in the case of ether, the temperature being the same in each case. Try also in each case the influence of an increased temperature.

It is evident then that the vapour pressure of alcohol or ether is greater than that of water at the same temperature. Moreover, in every case the vapour pressure increases as the temperature rises. It follows from this that with alcohol or ether

¹ Lecture Table Experiment.

the temperature will probably be sooner reached at which the vapour pressure is equal to the pressure of the atmosphere—in other words, the temperature at which the liquid will boil. If reference is made to the table of boiling points, p. 91 it is found in agreement with this conjecture that ether boils at 35° , and alcohol at $78^{\circ}3$.

LATENT HEAT OF VAPORISATION.

We have found that while the vaporisation of a liquid is proceeding by boiling, although heat is being continuously added, nevertheless there is no simultaneous rise in temperature. This heat which apparently disappears but can be made sensible again is termed the **latent heat of vaporisation**.

In the process of boiling this heat is supplied from an external source.

Is heat also rendered latent when a liquid simply evaporates below the boiling point, when no heat is supplied from a flame

or fire? If heat is absorbed in such a case, it will of necessity be obtained out of the liquid itself, or from surrounding objects, and consequently their temperatures will fall. We must test this by experiment and experience.

EXPT. 6. Is Heat rendered Latent in Ordinary Evaporation?—Pour a little ether over the hand. A sensation of cold is felt.

Again, recall the sensation of cold experienced when the body while wet after bathing is exposed to the air.

EXPT. 7.—Take a small flask, fill it one-third full with ether. Place within it a test-tube containing about 1 c.c. of water. Then, by means of a glass tube connected with an air-blast

or bellows, send a rapid stream of air through the ether (Fig. 66). The ether is rapidly evaporated. Presently the water inside the tube is found frozen solid. Notice the hoar frost on the outside of the flask.

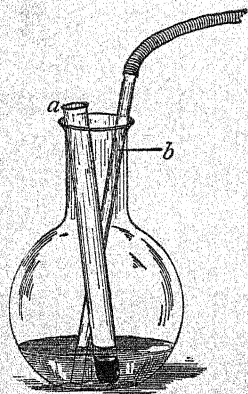


FIG. 66.

EXERCISE

Put some ether in a beaker. Wet the outside of the beaker with water. Stand it on a wooden block. Can you freeze the beaker to the block?

EXPT. 8. The Cryophorus.—This is an interesting piece of apparatus devised by the physicist Wollaston (Fig. 67). It consists of a closed glass tube, bent twice at right angles, with a bulb at each end, and contains some water and water vapour, but no air. Run all the water into the bulb B. Immerse the

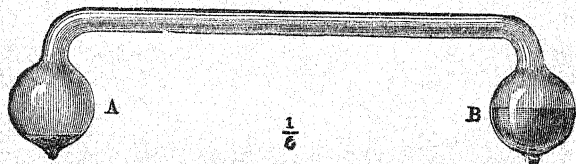


FIG. 67.

bulb A in a freezing mixture. After a time dew is seen to be deposited on the outside surface of B, and finally the water in B may even be frozen.

EXPT. 9. The Wet and Dry Bulb Thermometers.—Take a thermometer. Wrap the bulb in cotton wool. Read its temperature. Pour a few drops of methylated spirit on the cotton wool. Wave the thermometer about in the air. Read the thermometer again. The temperature has fallen several degrees. Repeat, using (1) ether, (2) water.

It is a matter of common knowledge that the air is much more moist on some days than on others. Sometimes the air appears to be laden with moisture, and everything out of doors is dank and wet. On such a day it would be useless to hang clothes out to dry, for the air is saturated or contains as much aqueous vapour as is possible at that temperature, and if a wet thermometer bulb were moved about in such air no fall of temperature would occur. It is evident, therefore, that we can judge of the un-saturation of the atmosphere by observing whether a moist thermometer sinks below the actual temperature of the air or not. *The wet and dry bulb thermometer* (Fig. 68)

is an instrument which enables us to do this. One thermometer serves to record the actual temperature of the air. The bulb of the other is surrounded by a piece of muslin, one end of which dips into a small vessel containing water, and by these means the bulb is kept moist. On a dry day the water on the wet bulb evaporates, and a lower temperature is recorded than by the dry thermometer. Moreover, the magnitude of the difference between the readings of the two thermometers gives a measure of the degree of un-saturation of the atmosphere,

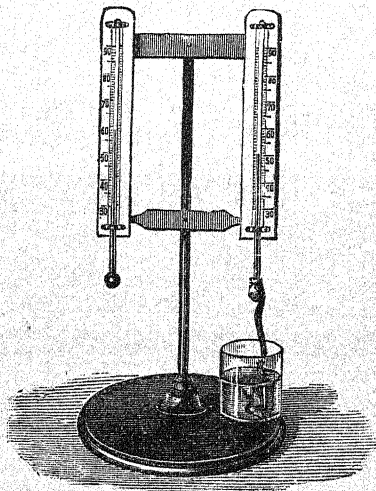


FIG. 68

for, the drier the air, the more rapid is the evaporation, and the greater the lowering of temperature. On a thoroughly damp day the two thermometers record the same temperature.

Conclusion.—It has been shown that the air always contains invisible water vapour. Sometimes the air is saturated, and the vapour is readily condensed as rain or mist or dew in the open, or as a dew on cold walls or pipes indoors. But far more often the air is not saturated, and in that case evaporation must result wherever there is water. In the open air, water is evaporated

from seas, lakes, and rivers, from the soil, and from the leaves of all trees, from the grass and all flowers, and if the heat is great, and the evaporation therefore rapid, these may fade and wither. And in a warm, closed room, which is not well ventilated, especially if it is heated by hot pipes or by a closed stove, evaporation will often take place so rapidly from growing ferns or other plants that the fronds or leaves droop and die ; evaporation of water will also take place from the surface of the skin of any occupants of the room, frequently causing exhaustion and oppression. Such a condition of a room may be prevented by not allowing the temperature of a room to exceed 60° F. by proper ventilation, and by having some water exposed in vessels in the room.

PROBLEM

Hang up a piece of salt seaweed. Compare its condition, day by day, with the readings of the wet and dry bulb thermometer.

CHAPTER XII¹

CALORIMETRY

The Unit of Heat.—We have already seen that when a solid is melted, or when a liquid is vaporised, heat becomes latent (pp. 85, 90 and 102), but we cannot endeavour to measure the *quantity* of heat which becomes latent until we have fixed upon some standard measure, that is upon a definite *unit* of heat.

The effects of heat are numerous, and any one of them which can be accurately measured might be used as a unit of heat; for instance, we might select as our unit of heat the quantity necessary to produce a definite increase of volume in unit mass of some standard substance. As a matter of fact, the effect of heat selected is that of increase of temperature, and the unit of heat is the **quantity of heat required to raise the temperature of 1 gram of water from 0° C. to 1° C.**; this unit is usually called **one calorie**. Hence, if 50 grams of water are to be heated from 0° to 1° C., 50 units of heat will be required.

Pending further inquiry, we shall assume that the quantities of heat necessary to raise the temperature of 1 gram of water 1° are equal wherever that degree may be on the scale of temperature; for instance, the quantity of heat necessary to raise the temperature of 1 gram of water from 80° to 81° may be taken as equal to that required to raise the temperature of 1 gram of water from 0° to 1°. It follows from this that if the

¹ *To Teachers.*—As the calculations involved in this chapter are usually felt to be rather difficult, the chapter may, if circumstances allow, be postponed with advantage to a later period in the course.

temperature of 1 gram of water be raised from 0° to 80° , 80 calories are required, and if the temperature of 20 grams of water be raised from 30° to 55° , 25×20 or 500 calories will be required.

It must be assumed also that, when a given weight of water, or indeed of any substance, is cooled through any range of temperature, it gives up exactly the same quantity of heat that is required to raise its temperature through the same range.

Let us next consider what occurs when we mix together two quantities of water at different temperatures.

EXPT. 1. The Effect of mixing Two Quantities of Water at Different Temperatures.—Support a small cylindrical vessel, *a*, made of sheet brass,¹ on corks inside a

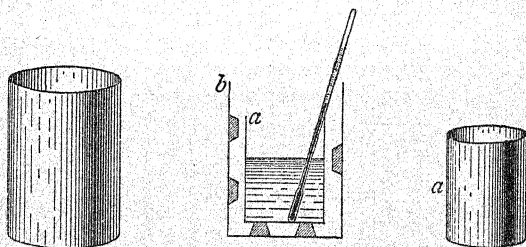


FIG. 69.

larger zinc vessel, *b* (Fig. 69). The inner vessel is called a "calorimeter" or heat-measurer. Measure into it 50 c.c. of water. Place a C. thermometer inside it. Weigh a small dry flask. Measure 50 c.c. of water into it. Place another C. thermometer inside it, and heat the flask slowly over a small flame. Meantime stir the water in the calorimeter, and read its temperature. As soon as the water in the flask is above 50° , remove the flame, stir carefully with the thermometer, quickly read its temperature, take out the thermometer, and pour the water into the calorimeter. Stir the mixture in the calorimeter, and read the highest temperature attained to one-tenth of a degree if possible. Weigh the flask, and so determine the

¹ A small glass beaker may be used if a brass vessel is not available.

quantity of water left in the flask. The following numbers were obtained in an experiment :—

Weight of warm water measured out	50 gm.
" " " left in flask	0.5 "
" " " used... ..	49.5 "
Temperature of warm water	52°·1 C.
Weight of water originally in calorimeter	50 gm.
Temperature of water in calorimeter	19°·1 C.
Temperature of mixture observed	35°·1 "

It is evident that the water in the calorimeter gained heat, while heat was lost by the warm water added to it. If there was no loss of heat, and no interchange of heat except between the two quantities of water, we can readily calculate what the common and final temperature should be, and this should agree with the final temperature observed, viz. 35°·1.

Probably the simplest and most convenient way is to calculate the quantities of heat which both quantities of water contained *above* that they would have possessed were their temperatures 0°. These were :—

50 × 19°·1 or 955 calories in the case of the water in the calorimeter.
49·5 × 52°·1 or 2578·95 calories in the case of the warm water.

Hence,

955 + 2579 or 3534 calories of heat were absorbed by 50 + 49·5 or 99·5 grams of water.

Hence, the common temperature should be $\frac{3534}{99.5}$ or 35°·5, whereas it was actually found to be 35°·1.

The difference between the observed and calculated values is probably due to a slight loss of heat from the water ; for instance, a small amount of the heat in the warm water goes to heat not only the water in the calorimeter, but also the calorimeter itself and the thermometer. (*See* Expt. 4, p. III.)

SPECIFIC HEAT

When equal quantities of heat are communicated to equal weights of water, the temperature of each is raised by the same amount. But if equal weights of *different* substances are heated until they have acquired the same temperature, will they be

found to have absorbed the same quantities of heat? In other words, have different substances the same *capacity* for heat?

EXPT. 2. To find whether Different Metals have the same Capacity for Heat.—Take cylinders of lead, tin, and copper, of the same weight, say 50 grams in weight each. Suspend them by threads in boiling water for five minutes. Take three calorimeters, and measure 50 c.c. of water at the ordinary temperature into each. Place a thermometer in each. Then remove the cylinders at the same time from the boiling water, and place one in each calorimeter. Stir carefully with the thermometer, and read in each case the highest temperature attained. Have the metals heated the water to different temperatures? Which has raised the temperature of the water the most? The copper.

It follows from this experiment that if equal quantities of heat are given to equal weights of *different* substances, the temperatures of each will be raised by different amounts.

Definition of Specific Heat.—

The quantity of heat required to raise the temperature of 1 gram of a substance from 0° C. to 1° C. is called its specific heat at 0° . Since, however, very nearly the same quantity of heat is required to raise the temperature of any substance 1° wherever that degree may be on the scale of temperature, we may state as approximately true that the

specific heat of any substance is the quantity of heat necessary to raise the temperature of 1 gram of it 1° C.

EXPT. 3. To find the Specific Heat of a Metal.—Measure into a thin brass vessel 100 c.c.² of water, and place a thermometer in it. Weigh into a hard test-tube, *a* (Fig. 70), 50 grams² of the metal. Close it with a stopper, through which is

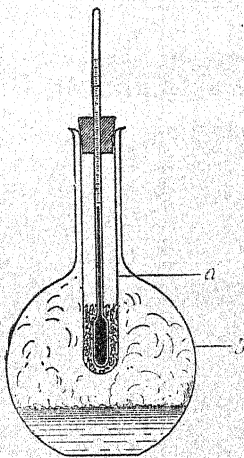


FIG. 70.

¹ E.g. granular lead, iron filings, granulated zinc, copper turnings.

² Half quantities may be used.

passed another thermometer, the bulb of which must be surrounded by the metal. Heat the metal as nearly as possible to 100°C . by suspending the test-tube for some time in a flask, *b*, containing boiling water. Read its temperature, read also the temperature of the water in the calorimeter, and then transfer the metal as quickly as possible into the water in the calorimeter, stir with the thermometer, and observe the highest temperature reached. Thus—

Weight of metal	50 gm.
„ water	100 gm.
Temperature of metal before experiment	100°C .
„ „ cold water before experiment	12°C .
„ „ water after experiment	16°C .

Assume *calories lost by metal = calories gained by water*; then,

Metal fell in temperature	$100 - 16$ or 84 degrees.
Water increased in temperature	$16 - 12$ or 4 degrees.
Water gained...	100×4 or 400 calories.
<i>i.e.</i> metal lost 400 calories.

If 50 gm. metal cooling 84° lost 400 calories,

$$1 \text{ gm. } ,, ,, 1^{\circ} \text{ lost } \frac{400}{50} \times \frac{1}{84} \text{ or } \frac{8}{84} \text{ calories.}$$

Hence,

$$\text{Specific Heat of metal} = 0.095.$$

The student should observe that (1) the metal must be thoroughly heated, (2) as the difference of temperature to be measured is very small, temperatures should be read to the nearest tenth of a degree.¹

Water Equivalent of a Calorimeter.—If warm water is added to a calorimeter, containing a thermometer, which is at the temperature of the air, the calorimeter and thermometer are heated, and on the other hand there is a loss of heat from the water, which must be allowed for in accurate determinations of specific heat. Now there must be a certain quantity of water which will absorb just as much heat as the

¹ *Note for Teacher.*—The method described above is only moderately accurate, but it is instructive, and does not take too long. The student realises at least what is meant by “specific heat,” which is the important thing. It may be noted that the “water-equivalent” is neglected, and that it is assumed that 100 c.c. = 100 grams.

calorimeter and thermometer absorb in being heated through any range of temperature. This equivalent amount of water is called the *water equivalent of the calorimeter*, and must be allowed for in accurate determinations of specific heat.

EXPT. 4. To show the Absorption of Heat by a Calorimeter.—Support a calorimeter as before on cork in an outer zinc vessel. Heat 100 c.c. water in a beaker containing a thermometer to about 35° C. As soon as this temperature is reached, remove the flame, stir well, read the temperature, at once pour the water into the calorimeter, put the thermometer in the calorimeter, and stir well, watching closely the thermometer. The temperature falls rapidly through a small range owing to the absorption of heat by the calorimeter. Read the temperature. From this point the temperature continues to fall slowly, owing to the cooling of the apparatus.

State thus :—

Temperature of water in beaker = $35^{\circ}4$ C.

„ „ „ calorimeter = $34^{\circ}5$ C.

\therefore Fall in temperature = $0^{\circ}9$ C.

For a fuller treatment of this subject the student should refer to books on Physics.

TABLE OF SPECIFIC HEATS

SOLIDS

Magnesium... ..	0.250	✓ Zinc... ..	0.096
Aluminium... ..	0.214	✓ Copper	0.095
✓ Glass	0.188	✓ Silver	0.057
✓ Sulphur	0.177	✓ Tin... ..	0.056
Diamond	0.147	✓ Gold	0.032
✓ Iron	0.114	✓ Lead	0.031

LIQUIDS

Water... ..	1.000	Oil of turpentine ...	0.462
Alcohol	0.615	Mercury... ..	0.033
Ether... ..	0.503		

It is seen at once from these numbers that the specific heat of water very greatly exceeds that of most other substances.

See p. 117 for Examples on Specific Heat.

LATENT HEAT OF LIQUEFACTION

We have already found that heat becomes latent when any solid substance is melted, and we are now in a position to measure the latent heat of liquefaction of ice. But first we must define exactly what it is we wish to measure.

Definition of Latent Heat of Liquefaction.—*The latent heat of liquefaction of any substance is the number of calories of heat absorbed by 1 gram of it in passing from the solid to the liquid state without rise of temperature.*

N.B.—This is equal to the number of calories of heat evolved when a gram of the liquid substance is reconverted to the solid state without fall of temperature.

EXPT. 5. To measure the Latent Heat of Liquefaction of Ice.—Weigh a small sheet-brass calorimeter. Add 100 c.c. of water, and weigh again. Heat it carefully to about 20° C. Then weigh out roughly on filter-paper about 10 grams of ice. Place the calorimeter on cork supports within a zinc can, and carefully read the temperature of the water in it. Dry the ice with blotting paper. Quickly introduce it into the calorimeter, stir thoroughly, and observe the lowest temperature attained. Finally, remove the thermometer, allowing the water to drain off into the calorimeter, and weigh again. The increase in weight gives the weight of ice added. In an experiment the following numbers were obtained :—

Weight of calorimeter	21.4 gm.
„ calorimeter and water... ..	119.2 „
„ calorimeter, water, and ice... ..	128.5 „
Initial temperature of water	21° 0 C.
Final „ „ „ „ „ „ „ „	12° 5 C.

Thus, $119.2 - 21.4$ or 97.8 grams of water were cooled from 21° to $12^{\circ} 5$.
 $128.5 - 119.2$ or 9.3 grams of ice were added.

This ice was first melted, and secondly heated to $12^{\circ} 5$.

Assume

heat gained by ice = heat lost by water.

Then,

heat lost by water = $97.8 \times (21 - 12.5)$
 = 831.3 calories.

Therefore,

heat gained by 9.3 grams ice = 831.3 calories.

$$\begin{aligned} \text{But the calories gained by 9.3 grams of melted} \\ \text{ice at } 0^\circ, \text{ due to } 12^\circ.5 \text{ rise of temperature...} \end{aligned} \left. \begin{array}{l} \\ \\ \end{array} \right\} = 9.3 \times 12.5 \\ = 116.25$$

$$\begin{aligned} \text{Hence, the number of calories absorbed in melt-} \\ \text{ing 9.3 grams of ice } \end{aligned} \left. \begin{array}{l} \\ \\ \end{array} \right\} = 831.3 - 116.25 \\ = 715.05$$

Therefore,

$$\begin{aligned} \text{number of calories absorbed in melting} \\ \text{1 gram of ice } \end{aligned} \left. \begin{array}{l} \\ \\ \end{array} \right\} = \frac{715.05}{9.3} \text{ or } 76.9$$

The value obtained here for the latent heat of liquefaction of ice must be slightly inaccurate, owing to a small loss of heat from the apparatus, and to the ice being slightly wet on the surface. The true value lies between 79 and 80, but in round numbers it may conveniently be taken as 80. We have arrived then at the result that in the change of 1 gram of ice at 0° into 1 gram of water at 0° no less than 80 calories of heat are rendered latent; that is to say, as much heat as could heat 80 grams of water from 0° to 1° , or heat 1 gram of water from 0° to 80° .

It is interesting to observe the values of the latent heats of liquefaction of other bodies.

TABLE OF LATENT HEATS OF LIQUEFACTION

Ice	80	Tin	14+
Sodium nitrate... ..	63	Iodine	11+
Nitre... ..	47+	Sulphur	9+
Zinc	28+	Lead	5+
Silver	21+	Phosphorus	5+

It is very noticeable that the latent heat of ice is very great as compared with those of other substances, and this is a matter of the greatest importance. If snow and ice were immediately converted into water at 0° C. with little or no absorption of heat, the sudden melting of the snow on the hills and mountains would expose the inhabitants of the valleys below them to the most terrible floods and inundations.

LATENT HEAT OF VAPORISATION

Just as heat disappears whenever a solid is liquefied, so also we have already seen that when a liquid is vaporised heat becomes latent.

Definition of Latent Heat of Vaporisation.—*The latent heat of vaporisation of any liquid at a given temperature is the number of calories of heat absorbed by 1 gram of it in*

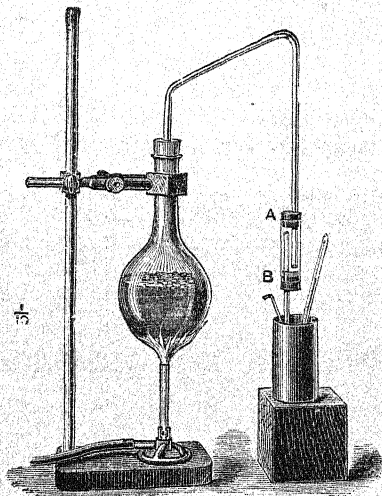


FIG. 71.

passing from the liquid to the gaseous state without rise of temperature.

When a vapour is condensed again into a liquid, the heat which was insensible or latent becomes sensible again, and the number of calories evolved is equal to the number previously absorbed.

EXPT. 6. To measure the Latent Heat of Vaporisation of Water.—(1) Place some water in a 500 c.c. flask. Close it with a stopper, through which are passed a thistle-funnel (its lower end dipping below the surface of the water, not shown in Fig. 71) and a glass tube, bent as shown. Proceed to

heat the water. Also take a wide glass tube, A B, closed by stoppers, through the lower of which a straight glass tube is passed. Connect this apparatus, which is to serve as a trap, with the flask by means of the bent glass tube. The trap serves to catch the condensed water which is carried over by the steam. When the steam issues freely from the end of the tube, measure 50 c.c. of ice-cold water into a weighed brass calorimeter or beaker, and place it so that the end of the tube reaches nearly to the bottom of the calorimeter.

The steam is condensed with a loud noise by the cold water, and the temperature rapidly rises, till finally the water in the beaker itself is raised to near the boiling point. Then remove the flame, and also the beaker, and observe that the volume of the water in the beaker is only increased by about one-fifth. If the increase in weight of the beaker is 10 grams, this shows that 10 grams of steam at 100° , in changing into 10 grams of water at 100° , set free sufficient heat to raise the temperature of 50 grams of water from 0° to 100° C. What value does this result give for the latent heat of vaporisation of water?

This experiment shows that the latent heat of vaporisation of water has a very high value, and that therefore steam should never be needlessly wasted.

(2) Try to measure accurately the latent heat of vaporisation of water. For this it will be necessary to stop the condensation of the steam, when the temperature of the water reaches about 50° , because at higher temperatures its loss of heat by cooling becomes much greater. Proceed thus:—Measure 100 c.c. of cold water into a calorimeter, and weigh. Read its temperature. Pass steam in until the temperature rises just above 50° C. Then lift the delivery tube out from the calorimeter, stir well, and at once read the temperature. Thus:

Weight of cold water	= 100 gm.
Weight of calorimeter when cold	= 141.7 ,,
" " " " warm	= 148.7 ,,
Temperature of cold water	= $13^{\circ}0$ C.
" " " " warm water	= $53^{\circ}3$ C.

Hence, 100 grams of water were warmed from 13° to $53^{\circ}3$ C., and 7 grams of steam were condensed.

Now, the steam *was first condensed, and, secondly, cooled to 53°3 C.*

Assume

$$\text{heat gained by water} = \text{heat lost by steam.}$$

Now,

$$\begin{aligned} \text{heat gained by water} &= 100 \times (53.3 - 13) \\ &= 4030 \text{ calories.} \end{aligned}$$

$$\therefore \text{heat lost by 7 gm. of steam} = 4030 \text{ calories.}$$

But the calories lost by 7 grams of *condensed* steam at 100° due to (100° - 53°3) or 46°7 fall of temperature = 7×46.7
= 326.9.

$$\therefore \text{the calories given out by the condensation of 7 grams steam into 7 grams water at } 100^\circ \dots \dots \dots = 4030 - 326.9 = 3703.1$$

$$\therefore \text{the number of calories given out by the condensation of 1 gram steam } \dots \dots \dots = \frac{3703.1}{7} = 529$$

The true value of the latent heat of vaporisation of water at 100° is found to be **537**; that is to say, in converting 1 gram of water at 100° into steam at 100°, no less than 537 units of heat are rendered latent.

TABLE OF LATENT HEATS OF VAPORISATION

Water	537	Oil of turpentine	74
Alcohol	202 +	Bromine	45 +
Ether	105 +	Iodine	24
Acetic acid	102		

It is seen from this table that the latent heat of vaporisation of water is exceptionally high, and greatly exceeds that of almost every other liquid, just as the value of its specific heat and that of its latent heat of liquefaction are higher than those of other liquids.

In view of the fact that water exists in vast quantities in the liquid form on the surface of the globe, that its vapour is always present in the air, and that ice and snow are forms in which it is often found, there can be no doubt that the above exceptional properties of water are of the highest importance in maintaining a fairly uniform temperature on the surface of the earth.

EXAMPLES XII

(a) Specific Heat

1. Calculate in each of the following cases what the common temperature should be, assuming no loss of heat:—

- i. 50 gm. water at 51°C ., added to 50 gm. water at 16°C .
- ii. 49.5 " " 50°C . " " " 15°C .
- iii. 50 " " 95°F . " " " 16°C .
- iv. 1 litre " " 99°C . " 900 c.c. " 10°C .

2. How many calories of heat are required to heat—

- i. 50 gm. copper from 0° to 100°C . (S. H. copper = 0.095)
- ii. 50 " " 10° to 95°C . " " "
- iii. 25 gm. lead from 20° to 100°C . (S. H. lead = 0.03)
- iv. 1 kilogram iron from 20° to 100°C . (S. H. iron = 0.11)

3. Find the specific heat of a silver watch chain, having given

Weight of chain	90 gm.
Temperature of hot chain	99°C .
Volume of cold water	100 c.c.
Temperature of cold water	15°C .
Temperature of mixture	19.1°C .

4. If 200 grams of mercury at 100°C . are mixed with 100 grams of water at 7°C ., and the resulting temperature is 13°C ., find the specific heat of mercury.

5. 10 grams of a certain substance are heated to 100°C . and placed in 75 grams of water at 4°C . The final temperature of the water is 10°C . Find the specific heat of the substance.

6. Find the final temperature of a mixture of 100 grams alcohol at 10°C . and 100 c.c. water at 0°C ., assuming that the S.H. of alcohol is .6, and that the only heat change is an interchange between the alcohol and the water.

7. Find the final temperature of 100 grams of water originally at 0°C ., after dropping into it 10 grams of copper nails at 100°C . (S.H. of copper = .095.)

8. Find the resulting temperature when 10 grams of water at 100°C . are poured into a copper vessel at 0°C . weighing 100 grams. (S. H. of copper = .095.)

9. What would be the rise of temperature on dropping 100 grams of lead shot at 100°C . into 100 grams of water at 0°C .? (S. H. of lead = .03.)

10. How many grams of silver at 90° if dipped into 100 grams alcohol at 10° would raise its temperature to 20° C. ? (S. H. of silver = $\cdot 057$; alcohol = $\cdot 615$.)

11. What weight of mercury at 100° if mixed with 100 grams of water at 0° would raise its temperature to 10° C. ? (S. H. of mercury = $\frac{1}{16}$.)

(b) Latent Heat

12. How much heat would be required to melt 1 kilogram of ice without rise of temperature?

13. Calculate the latent heat of liquefaction of ice from the following data :—

	(i.)	(ii.)	(iii.)
Weight of water	100 gm.	70.5 gm.	100 gm.
" " + melted ice	104.5 "	76.3 "	105.7 "
Initial temperature of water ...	$21^{\circ}.4$ C.	$21^{\circ}.5$ C.	$21^{\circ}.4$ C.
Final " " " ...	$17^{\circ}.1$ C.	$13^{\circ}.9$ C.	$15^{\circ}.9$ C.

14. Calculate the common temperature reached on mixing 10 grams of ice at 0° C. with 100 c.c. of water at 80° C.

15. How many grams of steam would be required to raise the temperature of 1 litre of water from 0° to 100° C. (L. H. of vaporisation of water = 537.)

16. Calculate the latent heat of vaporisation of water from the following data :—

	(i.)	(ii.)
Weight of cold water	100 gm.	100 gm.
Weight of calorimeter when cold	117.3 "	137.7 "
" " " warm	124.0 "	144.1 "
Temperature of cold water	$12^{\circ}.2$ C.	$13^{\circ}.8$ C.
" warm water	$50^{\circ}.4$ C.	$50^{\circ}.0$ C.

17. How much heat would be required to change 1 gram of ice at 0° C. into steam at 100° C. ?

CHAPTER XIII

SOLUTION

THE SOLVENT ACTION OF WATER.—What occurs when a solid substance is placed in water? Place a lump of sugar in a small beaker of warm water. The sugar crumbles down, melts, and disappears in the water.

The sugar is said to be **dissolved** in the water, or to be **soluble** in the water. And this sort of melting and disappearing in the water is called **solution**, and the water is said to have a **solvent action** upon the sugar. But the sugar is not lost, for the water tastes sweet. Is there any change in weight?

EXPT. 1. Is there a change in weight when sugar is dissolved in water?—Warm some water in a beaker, put in it a glass rod, and cover it with a watch-glass. Place a piece of smooth paper with a few grains of sugar on the watch-glass. Place the whole on one pan of a balance, and counterpoise it. Carefully slide the sugar into the water, and stir until it is dissolved. Is the weight altered? No.

All substances, however, do not dissolve so readily in water as sugar.

EXPT. 2. To find whether Plaster of Paris is soluble in Water.—Shake some powdered plaster of Paris in a small flask with water. The water becomes milky owing to powder *in suspension*. The plaster seems insoluble.

But does the evidence justify this conclusion? We know that the plaster *is not all dissolved*. Is it possible that some of it is dissolved? To settle this we must separate the liquid from the floating powder. This we can do by letting it stand, but that takes time. Can we not pour the milky liquid on some porous

material such as blotting-paper, which will allow the liquid part to run through while the floating solid particles are caught and retained on the paper? Try.

EXPT. 3. To filter the Milky Liquid containing Plaster of Paris.—Take a circular piece of a kind of porous paper called filter-paper. Fold it once across a diameter.

Then fold it again in two. Open the folded paper so that it forms a cone, as shown in Fig. 72. Place it in a dry funnel, the cone of which is larger than the paper cone. It springs up

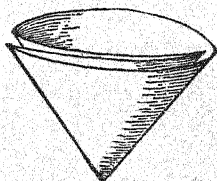


FIG. 72.

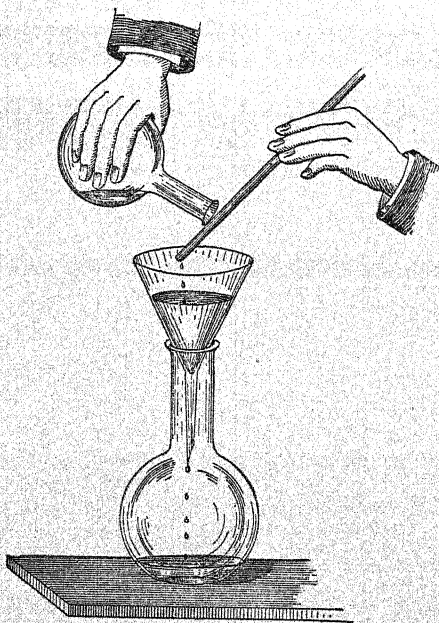


FIG. 73.

if the paper does not fit the glass closely all round; alter the folding till it fits well. Now moisten the paper with water to keep it in place. Place the funnel in the mouth of a flask, or in the ring of a retort-stand above a clean beaker.

Take the flask containing the water which has been shaken up with the plaster of Paris. Pour the contents down a glass rod into the funnel (see Fig. 73), and take care that the liquid does not rise above the top of the paper.

The plaster is retained on the paper, while a clear liquid runs through; this latter is called the **filtrate**. If the

first few drops are not quite clear, pour them on the filter again.

If the clear filtrate is evaporated, the water will be driven off. Solid plaster of Paris does not evaporate into the air. It seems probable, then, that if the clear filtrate be evaporated, solid plaster of Paris will be left behind, if the water has really dissolved any of it.

EXPT. 4. To find whether the Clear Filtrate contains Dissolved Plaster of Paris.—(1) Place a drop of the filtrate on platinum foil. Hold the foil by a pair of tongs over a flame.

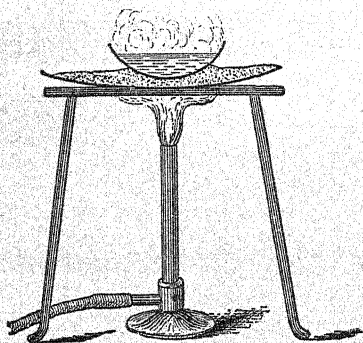


FIG. 74.

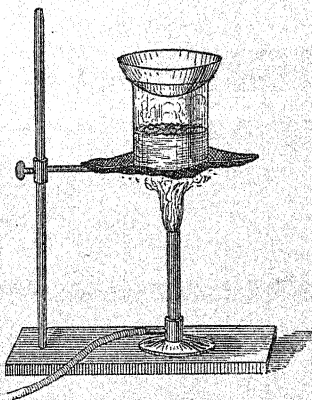


FIG. 75.

A white stain is left. It is clear, then, that plaster of Paris is slightly soluble in water.

(2) Pour some of the filtrate into a glass basin. Place the basin on a sand-bath (see Fig. 74), and heat it so that the liquid is evaporated without spurting. The volume of liquid gradually lessens, and a white deposit crusts the basin.

(3) A still surer method for avoiding loss by spurting is by evaporating the liquid on a "water-bath." The basin is in this case put on a beaker in which water is being boiled, as shown in Fig. 75.

PROBLEMS

Examine the solvent action of cold water upon powdered nitre, salt, green vitriol, garden soil, gravel, chalk, Epsom salts.

SOLIDS IN NATURAL WATERS.—Do your results not bring out the fact that several substances which occur in nature are soluble in water? Then will not water which has flowed over or through the crust of the earth be found to have dissolved something? We can readily find out whether this is so.

EXPT. 5. To find whether Spring Water contains Dissolved Solids.—Take a clean glass basin.¹ Measure 10 c.c. of clear spring water by a pipette into it. Evaporate the water by placing the basin on a sand-bath or water-bath. After the disappearance of the water, a slight solid residue is seen to be left.

PROBLEMS

Find whether rain water, river water, sea water,² do or do not leave a solid residue on evaporation.

You have probably found that while the rain water leaves little or no residue the sea water leaves very much more than the spring or river water.

Then *rain* water—*i.e.*, water which has not been in contact with the surface of the earth—is almost entirely free from dissolved solid impurity. *Spring* and *river* waters, on the other hand, are waters which have trickled through or over various strata of the earth's crust, so that it is probable that much of the solid matter they contain has been dissolved out of the rocks and soil over which they have passed. This will account, too, for *sea* water containing so much dissolved matter. For the rivers flow down into the sea, carrying with them that which they have dissolved out of the earth, and as the water of the sea evaporates the solid material remains behind, just as in the above experiment. A *mineral* water is a spring water which contains an exceptional amount of dissolved solids.

Observe the following table. Why do some river waters contain more dissolved solids than others?

¹ Or a watch-glass, 3" diameter.

² Sea water is supplied by most chemical firms.

SOLIDS IN GRAMS PER 100,000 C.C. OF NATURAL WATERS.

Rain ... average	3 grams.	Lakes :—	
Rivers :—		Loch Katrine ...	28 grams.
Neva	5.5 ,,	Thirlmere ...	27 ,,
Irwell near source	8 ,,	Mineral Waters :—	
,, below Man-		Vichy	503 ,,
chester ...	56 ,,	Eau de Seltz ...	407 ,,
Rhine	16-32 ,,	Sea Water :—	
Thames, at London,	40-45 ,,	Black Sea ...	1,770 ,,
Jordan	105 ,,	Indian Ocean ...	3,400 ,,
Nile	158 ,,	North Atlantic	
Springs and Wells :—		Ocean... ..	3,849 ,,
Average of 200		Dead Sea... ..	17,000 ,,
samples	28 ,,		

We have seen that a solid residue is left when a water containing solid matter is evaporated. If, as we have really assumed, the whole of the dissolved matter is left behind, the water which evaporates must be free from such impurity.

EXPT. 6. Can Pure Water be obtained by Distillation?—(1) Distil about 100 c.c. of sea water from a flask, as in Expt. 6, p. 91. See that the water does not boil violently. Throw away the first 10 or 20 c.c. of the distillate, which will have washed out the tube and flask, and may therefore contain some dust. Then collect some more, and find whether on evaporation 10 c.c. leave as much residue as before. Extremely little or none.

(2) Colour the rest of the sea water with a few drops of ink, and distil the coloured water. Is the distillate colourless? Yes.

We have, therefore, in the process of distillation a means of obtaining pure water from such water as sea water, which contains much dissolved solid matter, and is unfit for drinking purposes.

Rain Water is Distilled Water.—Distillation is always occurring in nature. From the wide expanse of the waters of the ocean, evaporation is continually taking place. Leaving the dissolved matter behind, the water vapour ascends, to descend again in the form of pure rain water upon the surface of the earth. The rain which falls towards the end of a downpour in the country far from the smoke of cities is the purest water in nature.

HARD AND SOFT WATERS.—The differences which are found between the amounts of solid matter contained in various natural waters are notable, and of much importance.

Waters which yield a lather with difficulty with soap, so that much soap is used up before its cleansing action can be utilised, are called **hard**.

EXPT. 7. To make a Soap Solution.¹—Cut a piece of Castille soap so that it weighs 10 grams. Cut it into thin slices, and place them in a large flask. Add 400 c.c. methylated spirits. Warm gently, and shake until the soap is dissolved. Filter the solution into a litre flask. Wash out the other flask with a little alcohol. Add spirits so as to make the volume up to 1 litre. Store the solution in a Winchester pint bottle.

EXPT. 8.² To compare the Hardness of Different Waters.—Fill a burette with the soap solution. Take a 6-oz. glass-stoppered bottle. Measure into it 50 c.c. of the laboratory tap water. Read the burette. Run into the bottle 1 c.c. of soap solution. Close with the stopper, and shake well. Place the bottle on its side on the bench. Is there a permanent lather? No! Then, add another 1 c.c. and shake, and proceed in this way until a lather remains right across the surface of the water for three minutes. Read the volume of soap solution added.

Find in the same way what volumes of the same soap solution must be added in order to produce a permanent lather with different kinds of water—*e.g.*, (*a*) rain water, (*b*) a river water, (*c*) a spring water.

To quote some results :—

50 c.c. of rain water	required	$2\frac{1}{2}$ c.c. soap solution.
„ a certain spring water	„	$3\frac{1}{2}$ „ „
„ another spring water	„	9 „ „
„ a river water	„	$11\frac{1}{2}$ „ „

Now the rain water had been found to contain very little dissolved solids, while the river water contained a considerable quantity. It might appear, therefore, that *hard* waters are waters that contain much solid matter. But the two spring waters had

¹ Lecture Table Experiment.

² This and other experiments may with advantage be performed by a few members of a class before their fellows.

been found to contain about equal quantities of dissolved solids, and therefore, since one of these was nearly three times as hard as the other, it would seem likely that the hardness of a water depends not only upon the quantity of dissolved solids, but also upon their nature. Test this :—

PROBLEM

Shake up rain water with a little of each of the following naturally occurring substances: take as much of the powdered substance as would cover a sixpence. Filter, and find the quantity of soap solution required to form a lather with 50 c.c. shaken with

(a) soda, (b) gypsum, (c) Epsom salts.

Again, some results may be quoted :—

50 c.c. rain water shaken with soda	required	2 c.c.
" " " " gypsum	"	50+c.c.
" " " " Epsom salts	"	50+c.c.

It is clear, therefore, that the presence of soda in a water will not make it hard; the spring water referred to above which required only $3\frac{1}{2}$ c.c. soap solution to form a lather as a matter of fact contained soda. On the other hand, chalk, or gypsum, or Epsom salts, will render a water hard. And it is the experience of those who live in chalk or limestone districts that the well water is very hard. (There will be a further discussion of the hardness of waters containing chalk later: see Vol. II., A Research on Chalk.)

It is evident that hard waters are ill adapted for washing purposes, and that the softness or hardness of a town's supply of water will greatly affect the annual consumption of soap in every household.

EXPT. 9. A Hard Water when boiled deposits a Solid Crust.—Take some water which you know to be very hard. Half fill a perfectly clean 8-oz. flask with the water. Boil steadily for some minutes. Then pour out the water, and examine the inner surface of the flask. A thin film of solid matter has been deposited wherever the water was in contact with the glass.

Examine the inside of a kettle in which hard water has frequently been boiled, and note the fur on the sides.

Thus it is clear that a boiler in which immense quantities of water are heated will in the course of time become coated with a thick solid crust, if the water is hard. This leads sometimes to dangerous accidents, because the iron, if protected by the incrustation from the water, becomes greatly overheated. Hence, a water which contains a considerable quantity of solid matter in solution is generally ill suited for use in boilers.

It may be added that soft waters are much better adapted than hard waters for cooking purposes, whether for making tea and bread, or boiling vegetables and meat. On the other hand, it is found that soft waters have the disadvantage of attacking metals, especially lead, much more than hard waters do.

PROBLEMS

Find the volume of soap solution required to form a permanent lather with 50 c.c. of a hard water which has been treated in each of the following ways :—

1. Heated, and then tested while still hot.
2. Boiled for a few minutes, and then filtered and cooled.
3. Distilled.
4. Shaken with a few crystals of soda, and then filtered.

Do you not find that in each case the water has been rendered softer? The use of soda crystals in the laundry can now be understood.

Make a summary of all the measurements you have made with the soap solution.

THE SOLUBILITIES OF SOLIDS IN WATER.—

Water, we have seen, can dissolve many different substances. Are all substances, however, soluble to the same extent in water? Sugar is readily soluble; plaster of Paris only slightly so. But we must make actual measurements with different substances.

Definition of Solubility.—*The solubility of a substance in water at any given temperature is measured by the weight of the substance which can be dissolved in 100 grams of water at that temperature.*

EXPT. 10. To find what Weight of Salt can be dissolved by Water.—(1) Half fill a 6-oz. glass-stoppered bottle

with cold water. Add some common salt, and shake well. If it is entirely dissolved, add more until some remains undissolved. Note that it is impossible to dissolve an unlimited amount in the water. At the temperature of the experiment the water will only dissolve a certain amount of the salt. The solution is then said to be **saturated**. Allow it to stand. If the undissolved salt does not quickly settle, filter the solution. Weigh an evaporating basin or clock-glass. Carefully suck up 10 c.c. of the clear solution into a pipette, and transfer to the basin. Evaporate the solution to dryness on a sand-bath. Weigh. Calculate the weight of salt dissolved in 100 c.c. of the solution. In an experiment,

10 c.c. solution contained 3.04 gm. salt.

(2) Notice, however, that in the last experiment we measured the weight of salt dissolved in 10 c.c. of the *solution*, not in 10 grams of water. In order to find the "solubility," we must also *weigh* 10 c.c. of the solution in a flask. Thus,

10 c.c. solution weighed 11.80 gm.

Hence,

11.80 - 3.04 or 8.76 gm. water dissolved 3.04 gm. salt.

Therefore, 100 gm. water dissolved $\frac{3.04}{8.76} \times 100$ gm. salt.

= 34.7 gm. salt.

PROBLEMS

Find what weight of each of the following substances can be dissolved by 100 grams of water, saturated at the ordinary temperature:—powdered nitre, garden soil, powdered gypsum.

Do not your results bring out the important fact that substances of common occurrence differ greatly as regards their solubility in water?

The Influence of Temperature upon Solubility.—Thus far we have only examined the solubilities of substances in water at the ordinary temperature. But is it not possible that at higher temperatures the solubility of a substance will be different from what it is at a low temperature? We must, then, study the influence of temperature.

EXPT. 11. Is Nitre more soluble in Hot than in Cold Water?—One quarter fill a test-tube with water. Add sufficient powdered nitre to saturate it at the ordinary temperature. Then heat the test-tube in a flame. The undissolved nitre dissolves. Add more and more nitre. It can also be dissolved. Hold the test-tube in a stream of cold water from a tap, and at the same time scratch the inner surface of the tube with a glass rod. As the temperature of the solution falls, some of the dissolved nitre separates out in crystals.

Then, there can be no doubt that nitre is *more soluble in hot water than in cold water*.

PROBLEMS

Find whether powdered blue vitriol, alum, bichromate of potash, chlorate of potash, common salt, gypsum, are more soluble in cold or in hot water.

All are much more soluble in hot water than in cold water, except salt, which is scarcely more soluble in hot than in cold water, and gypsum which is even less soluble in hot water than in cold.

Solubility Curves.—The results of careful experiments on the amounts of nitre or other solids which can be dissolved in 100 grams of water at different temperatures can be best expressed by a graphical method. See Chapter XVIII, p. 179.

Solutions of known Strength.—It is sometimes necessary to prepare a solution of known strength, that is, one containing a known quantity of a substance in a given volume. The method of doing this may be illustrated by an example.

EXPT. 12. To prepare a Solution containing 50 grams of Nitre per Litre.—Your object will be to make the solution of the exact strength required without needless loss of time. Grind the nitre to powder. This will help it to dissolve. Heat a little gently in a dry test-tube. If any moisture condenses on the side of the glass, this shows that the nitre is damp. In this case place the nitre in a basin over a small flame. Stir with a glass rod until dry. If the nitre is weighed on the balance pan or even on a sheet of paper, some may be lost in transferring it to the vessel in which the solution is made.

Counterpoise an empty beaker. Weigh into it 5 grams of nitre. Add about 40 c.c. of hot water. You will naturally have saved time by putting water on to boil whilst weighing; besides, any attempt to boil cold water in a beaker at the bottom of which there is a thick layer of undissolved solid is liable to crack the glass. [Why?] Shake and warm until all the nitre is dissolved. Pour the solution into a 100 c.c. measuring flask. Rinse the beaker three times with warm water into the flask. The liquid is now hot, and will contract on cooling. Run water over the outside of the flask to cool it. Dilute carefully to the mark with cold water.¹

Is there any objection to filtering a solution before it is diluted? Could you get over the objection?

If a solution has to be both cooled and diluted, which should be done first, and why?

Latent Heat of Solution.—You may have observed when some solids have been dissolved in water that the solution has appeared to become colder than the water was at first. Test this by means of a thermometer.

EXPT. 13. To find whether the Temperature is lowered when Nitre is dissolved in Water.—Take 30 grams of powdered nitre. Measure 100 c.c. of water into a tumbler, and read its temperature. Add the nitre, and stir rapidly with the thermometer. Read the lowest temperature reached. State thus:—

Temperature of water = 59° F.

„ „ solution = 45° F.

Fall in temperature = 14° F.

The tumbler feels chilled, and dew may collect on its outer surface.

How may this absorption of heat be accounted for? It will be recalled that heat becomes latent when a solid is *melted* (see page 87). And this experiment seems to show that when nitre is broken up by being *dissolved* in water, heat is also required, which can only be obtained from the liquid itself, and so its temperature is lowered. Some of the sensible heat is rendered *latent*.

¹ Ask your teacher to test the solution you have made with a "Twaddell's hydrometer," to see whether it is of the right strength.

PROBLEMS

Find whether there is any change of temperature when the following are dissolved in water. Take 30 grams of each and 100 c.c. of water:—

Sal ammoniac, potassium chloride, salt, soda crystals, ammonium nitrate, Epsom salts.

DISSOLVED AIR IN WATER.—Have you not noticed the appearance of bubbles on the sides of a beaker or flask in which water is being heated, before the water has begun to boil? These look like air-bubbles, and if they are such, the air must have been dissolved in the water. And if so, the air must be less soluble in hot water than in cold water, while the

reverse is almost always the case with solids. This suggests heating a large quantity of water till it boils, in order to see whether any air is expelled.

EXPT. 14. To find whether Air can be expelled from Water by Heat.—Fill a 16-oz. flask completely with water (preferably from a spring. Why?) Fit it with a rubber stopper, and a glass tube bent as in Fig. 76, one end of which is flush with the lower surface of the

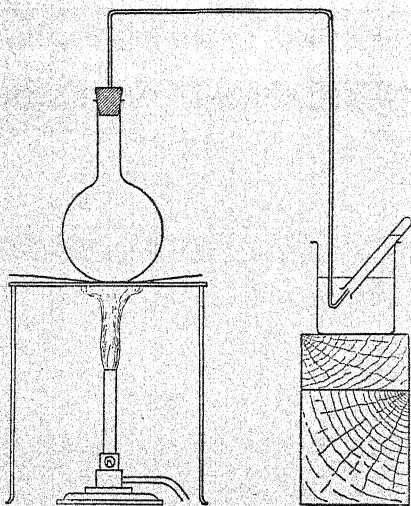


FIG. 76.

stopper. Fill the tube with water, and then push the stopper well into the neck of the flask. Support the flask on wire gauze. Let the other end of the tube dip into water in a tumbler. Fill a test tube with water and invert it above the mouth of the tube (Fig. 76). Proceed to heat the flask until the water

boils. Air is carried over into the tube. After boiling a few minutes, remove the burner.

Measure the volume of air collected when cool, and also the volume of the flask. Calculate the volume of air obtained from 100 c.c. water. A certain spring water was found to contain 3 c.c. air per 100 c.c. water. What do you find?

Taste some of the water which has been boiled in the flask. It is very flat and insipid. It seems likely that this is *due to the loss of the dissolved air*. Now drinking water is often obtained at sea from sea water by distillation; but such condensed water is very unpalatable, unless it is afterwards exposed in thin streams to the air. If such water be then heated as in the last experiment, air is given off again.

Moreover, since atmospheric air is necessary for the life of man, it seems probable that it is the *air dissolved in water which supports the life of fishes and of all aquatic animals*. This is confirmed by the fact that no fish can live in water which has been boiled.

We may, therefore, conclude that unless water is boiled or distilled it will contain dissolved air, and that the small amount of air which is found dissolved in water is really obtained from the atmosphere.

SOLVENT ACTION OF OTHER LIQUIDS.—We have up to the present studied the solvent action of water alone upon solids. But though water is by far the most important and common of liquids, there are many others in frequent use; for example, ether and methylated spirits. Will the same substances dissolve in them as in water? Are those substances which are insoluble in water also insoluble in these liquids? Try.

EXPT. 15. To try the Solvent Action of Ether on Tallow.—We know from daily experience that fats and oils do not dissolve in water. But will tallow, for instance, dissolve in ether? Fill a burette with ether.¹

Place a small lump of tallow in a dry test-tube. Add about 2 c.c. of ether. Shake, and stir with a glass rod. The tallow is dissolved.

¹ *Caution*: Take care that no flame is near, since ether vapour is very inflammable. Members of a class may conveniently obtain ether from a burette.

The experiment shows that the solvent action of ether is very different from that of water. It is probable, therefore, that other liquids will also be found which will dissolve substances though they may be insoluble in water. Carry out Exercises 9 and 10 below. Tabulate your results.

PROBLEMS AND EXERCISES

Solution

1. Find whether the substances A, B, C, D¹ are
 - (a) soluble in cold water ;
 - (b) more soluble in hot than in cold water.
2. Find whether the following are soluble in water. In each case state the evidence on which you form your conclusion :—
Sulphur, bleaching powder, pyrogallic acid, permanganate of potash, quick lime.
3. Prepare solutions in water containing 50 grams per 1000 c.c. of
Common salt ; Bichromate of potash.
4. Measure out 50 c.c. of each of the solutions made in Exercise 3, into a glass basin, and find the weights of the dissolved substances.
5. Prepare a solution of common salt containing 58.5 grams per litre.
6. Prepare a solution of silver nitrate containing 17 grams per litre.
7. From the solution made in Exercise 5 prepare a weaker solution containing 5.85 grams per litre.
8. From the solution made in Exercise 7 make another solution containing 5 grams per litre.
9. Find whether the following will dissolve in ether at the ordinary temperature :—

Paraffin Oil	Iodine.
Rosin.	Water.
Olive Oil.	Methylated spirits.
Lard.	Nitre.
Camphor.	Salt.

10. Find the solvent action of (1) methylated spirits, (2) benzene on the substances given in Exercise 9.

¹ E.g. Pb(NO₃)₂, KCl, Rochelle Salt, silica.

CHAPTER XIV

CRYSTALLISATION

CRYSTALLISATION.—In by far the larger number of cases the solubility of a substance in water is increased by increase of temperature. Consequently, in general, if hot water be saturated with a substance, some of this substance will separate out in the solid state when the solution is cooled. In many cases it will separate out in crystals.

EXPT. 1. To recrystallise Nitre.—Measure about 100 c.c. of water into a small flask and heat it. Weigh out about 60 grams of nitre. Powder the nitre and dissolve it in the hot water. Filter the hot solution into a crystallising dish, and allow it to cool gradually. When cold, pour off the solution from the crystals which have formed. Transfer the crystals to some filter-paper, and dry them by gently pressing them with other pieces of clean paper. What weight of crystals have you obtained? About 40 grams should be obtained. Select a good large crystal or bundle of crystals, and draw it twice the actual size.

Circumstances affecting Crystallisation.—Perfect crystals are hard to obtain. The conditions of their growth are shown by the following experiments :—

EXPT. 2.—Prepare a warm saturated solution of nitre in a test-tube. Divide it into two equal portions. Put one aside to cool slowly. Cool the other rapidly under the tap. Shake it, and rub the inside of the test-tube with a glass rod. A fine crystalline precipitate falls. In the first portion larger, pointed, needle-like crystals will slowly form. In fact, when a solution cools slowly, large crystals may form; shaking and rapid cooling tend to produce small crystals.

EXPT. 3.—Dissolve some crystals of Glauber's salt (sodium sulphate) in a very little hot water. Filter into a clean flask. Close the mouth of the flask with a plug of cotton-wool to exclude dust. Allow the solution to become quite cold, and avoid shaking it. It does not crystallise. Remove the plug. Let fall some tiny specks of solid Glauber's salt into the solution. Feathery crystals start from where the dust fell, and in a few seconds the flask contains a solid mass of crystals. This shows that dust particles tend to act as centres round which crystallisation may start, and that from a strong solution crystals form all massed together.

EXPT. 4. To prepare good large Crystals of Alum.—Prepare a strong hot solution of alum by dissolving 15 grams of alum in 75 c.c. of hot water. Cool rapidly to make sure that it is strong enough to deposit crystals. (1) If a great many crystals are formed, pour off the liquid with only a few of the crystals. Heat gently until the crystals are dissolved. Filter the liquid into a crystallising dish. You have now secured a solution which will be *saturated whilst warm*, and able to deposit a *small* quantity of crystals on further cooling. Place a piece of paper over the dish, and put it away in a cool quiet place. Examine in an hour's time. If many crystals have formed, the solution is still too strong. Suppose there are a few crystals. Leave them for twenty-four hours. Then pick out the six best. Transfer them with a glass rod to another dish. Pour the cold "*mother-liquor*" from which they were deposited gently over them, and put aside in a cool place. Arrange them so that they are not in contact. Try also suspending a crystal by a thread looped around it in some of the mother-liquor. After twenty-four hours examine them again. Remove any tiny crystals adhering to them. Turn them over so that they may rest on different faces. If the alum solution is able to evaporate slowly, it will go on enlarging these crystals. Notice that they are of an entirely different form from that of the crystals of nitre.

(2) If the solution is too weak to deposit crystals, it must be concentrated, either by heating it, or by leaving it to evaporate for several days into the air. If, by heating it, too much water is boiled away, the crystals which form on cooling may be all massed together. To find out when the solution has been

boiled down just enough, take out a single drop of it on a *cold* glass rod, and wait to see whether this drop crystallises. Do not let the rod get hot in the solution, for if so the drop removed would evaporate on the hot rod, and crystallise before the bulk of the solution was ready to do so.

The shape of a perfect alum crystal is that of the regular solid called the *octahedron*, bounded by eight flat sides or *faces*, each an equilateral triangle (see Fig. 77). In less perfect specimens two of these faces are often enlarged at the expense of the others.

Almost every chemical substance has a certain shape or shapes in which it can crystallise. Draw good-sized crystals

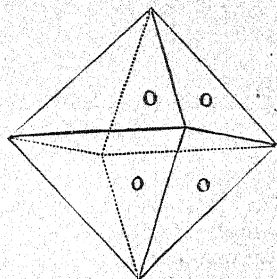


FIG. 77.—PERFECT CRYSTAL.

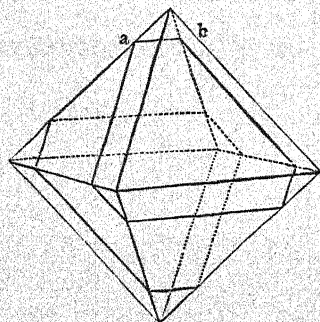


FIG. 78.

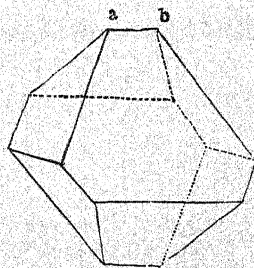


FIG. 79.

of rock-salt and of calcspar. The crystals of salt are *cubes*, those of calcspar *rhombs*. On the other hand, crystals of nitre are portions of six-sided *prisms*, with bluntly-pointed ends.

EXERCISES

1. Cut a soap model of a sugar crystal.
2. Cut a cube of soap. Mark the centres of the faces with pins. Cut away the corners to the pins so as to leave an octahedron.
3. Cut parallel slices off two opposite faces of the octahedron of soap, as indicated in Figs. 78 and 79.
4. Make a cardboard model of an octahedron.

WHAT IS A CRYSTAL?—But what is the distinguishing feature of a crystal? It is a solid of definite shape bounded by flat sides or faces which reflect the light. But does a crystal differ from a piece of cleverly cut glass? The following experiment may suggest an answer.

EXPT. 5. To show a Difference between a Crystal of Rock-Salt and a Piece of Glass.—Take a good crystal of rock-salt. Break it roughly with a hammer or in a strong mortar. Examine the fragments. They are all either little cubes or portions of cubes. Now break down a piece of glass rod in the same way. The fragments are of all shapes, and they appear to have no relation to one another.

EXERCISES

1. Crush a crystal of calcspar, and describe in words the shape of the fragments.
2. Find whether crystals of potassium ferrocyanide are equally easily cut in all directions.

The crystal of rock-salt seems then to possess definite lines along which fracture takes place, while the glass is broken as readily in one direction as in another. These lines along which the crystal can be readily broken or cut are called *lines of cleavage*. A characteristic of a crystal seems to be that it possesses a **definite structure**.

The existence of this definite structure is also impressed upon us by the fact that all the crystals of a given substance belonging to the same general form, for example crystals of calcspar, which are rhombs, have the *corresponding angles equal*.

EXPT. 6.—To show that the Corresponding Angles of Crystals of Calcspar are Equal.—Cut two slips of cardboard. Fasten two ends together by piercing a pin through them, so that they can turn about the pin as a pivot. Choose two good crystals of calcspar. Place them on the bench in the same relative position. Take up one of the crystals, and place it in the angle between the two slips of cardboard, so that each slip lies along a face of the crystal (Fig. 80). Hold the slips in this position. Now take up the other crystal, and see whether the *corresponding* angle in it will fit exactly into the angle between the slips of cardboard. It is found to do so.

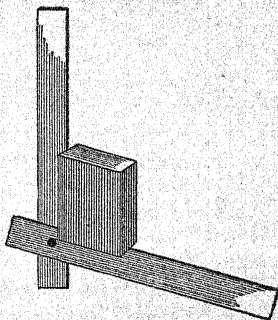


FIG. 80.

EXERCISES

1. Compare other corresponding angles on the same two crystals of calcspar.
2. Compare a third crystal of calcspar with one of those already taken.
3. Compare corresponding angles of crystals of
 - (a) rock-salt.
 - (b) alum.
 - (c) chrome-alum.

It is found that however the general shape of a crystal may be distorted *the corresponding angles remain constant.*

This is well shown in snow-crystals, which assume many fantastic shapes, but they are all built up from a simple star whose rays intersect at 60° .

The thought will readily occur that it is the *structure*, spoken of above, which determines the outward form. And it might be supposed that a cube of rock-salt consisted of a very large number of very small cubes; but this is not necessarily the case.

EXPT. 7. To build Geometrical Shapes with a Number of Spherical Balls.—(1) Take a number of five-balls or marbles, and pack them in a square so that one ball is touched by four others. To prevent the balls rolling, enclose a square by a cardboard rim fastened at the corners. Pile others in the hollows between the balls. Then add yet another layer of balls, and so on. Finally, half a regular *octahedron* is the result. Cannon balls are sometimes piled like this.

(2) Arrange the balls in a triangle, the innermost ball being touched by six others. Place balls in the hollows as before. The result is the corner of a *cube*.

A cubical crystal is, therefore, not necessarily built up with cubes.

On the other hand, it may be stated that there is every reason to believe that the forms of crystals *are* due to the symmetrical arrangement of the tiny particles which compose them.

SEPARATION AND PURIFICATION OF SUBSTANCES BY MEANS OF CRYSTALLISATION.

—We have found that substances differ greatly in solubility in water. It follows, therefore, that many mixtures can be entirely separated into their constituents by treatment with water. This will be most complete if one portion of a mixture is soluble in water while the rest is quite insoluble.

EXPT. 8. To separate a Mixture of Sand and Nitre.
—*i.e.* to obtain from it all the sand dry and free from nitre, and some of the nitre dry and free from sand.

Grind some sand and nitre together in a mortar. Place about 10 grams of the mixture in a beaker. Add some *hot* water. Boil. The nitre dissolves, while the sand remains undissolved. The finer particles of sand are floating in the liquid, but would in time settle. We could then pour off or *decant* into a second beaker a liquid which would be a solution of nitre, muddy with the finest sand, leaving in the first beaker sand wet with a solution of nitre. This would be a rough but not a complete separation. The sand must be washed, and the solution of nitre filtered.

Filter the muddy liquid. It is probably still hot, which is an advantage, since a liquid filters most quickly while hot.

(1) *To recover the Nitre from the Solution.*—If we only require *some* of the nitre, we need not wait for it all to run through the filter. Evaporate the clear nitre solution in a porcelain basin on wire gauze until it is sufficiently concentrated to deposit crystals; then pour the solution into a small beaker, and cool it by floating it in a tumbler containing cold water. Dry the crystals on filter-paper.

(2) *To clean the sand,* which was left wet with the nitre solution. Pour a little hot water on the sand in the beaker. Shake. If we wish to get all the sand, this washing must be run through the filter-paper. Wash similarly several times, transferring the sand to the paper. Some of the sand clings obstinately to the side of the beaker. Use the jet of the wash-bottle to dislodge the sand, and wash it completely on to the filter. Wait till the liquid has completely run through the paper. Gently remove the filter-paper from the funnel. Open it out flat, and place it on a pile of four filter-papers. Place the pile on wire gauze on a retort-stand ring about six inches above a small flame to dry slowly (Fig. 81).

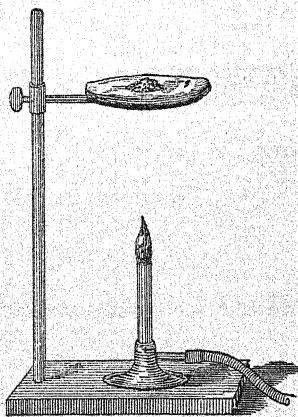


FIG. 81.

But even when both constituents of a mixture are soluble in water, it is possible to separate them, provided they are soluble in different degrees.

EXPT. 9. To separate Nitre from Common Salt—Two Substances which are both Soluble in Water.—Although both substances are soluble, they are not equally soluble. Thus 100 grams of water are able to dissolve

	at 0° C.	at 100° C.
Nitre	12 gm.	247 gm.
Salt	36 „	38 „

Salt is almost equally soluble in hot and in cold water. Nitre is less soluble than salt in cold water, but far more soluble in

hot water. What method would you suggest for separating nitre from a mixture of nitre and salt? Weigh out 30 grams of nitre and 30 grams of salt, and grind them together in a mortar. Measure 40 c.c. of water into a porcelain basin, heat it, and then slide in the mixture. Continue to heat with a small flame, and stir with a glass rod; a white granular substance refuses to dissolve (what is it?). Meantime heat a funnel fitted with a cone of filter-paper by pouring boiling water through it several times, and then at once filter the boiling solution, catching the filtrate in a small beaker. Cool the beaker by floating it in a tumbler containing cold water; long needles of nitre crystallise out. When cold, pour off the mother-liquor as completely as possible, and dry the crystals on filter-paper. These crystals may, however, still contain a small quantity of salt. Redissolve them, therefore, in the least possible quantity of boiling water. Again cool; much of the nitre again crystallises, but there is ample water to retain all the salt in solution. Dry the crystals on filter-paper. Find what weight of nitre has been recovered.

WATER OF CRYSTALLISATION

Try the following experiments:—

EXPT. 10. To observe the Effect of Heat on Crystals of Blue Vitriol.—Heat a few small crystals of blue vitriol in a small dry test-tube, holding the tube nearly horizontally. Observe that a heavy dew is deposited on the upper part of the tube, and at the same time the blue crystals become ashen grey and non-crystalline or *amorphous*. Steam also escapes from the mouth of the tube. Allow to cool: then add 4 or 5 drops of water to the residue in the tube. What do you observe?

It appears that heat drives water out of the blue crystals, and at the same time destroys their colour and crystalline form. The addition of water to the amorphous residue restores the blue colour. Can the crystalline form also be restored?

EXPT. 11. To attempt to recover Crystals by the Addition of Water to the Amorphous Residue.—Warm the test-tube containing the residue, to which a few drops of water have been added. Pour the hot deep blue solution into a watch-glass, and allow to cool. Crystals of blue vitriol are formed.

Then it seems that the crystalline form has been recovered, because water has been added. Moreover, the vapour which was given off when the crystals were heated must have been water vapour. Water which can be driven off by heat in this way from crystals is called **water of crystallisation**, and generally the crystals can be re-formed by the addition of water. We shall find that there are many solid crystalline substances, *e.g.* soda crystals, Epsom salts, which contain such water, though there are others, *e.g.* nitre, calcspar, which do not contain any water.

EXPT. 12. To find the Percentage of Water of Crystallisation in Blue Vitriol.—Weigh a small dry porcelain crucible and lid.

Add 1 to 2 grams of finely powdered blue vitriol. Weigh exactly. Place the crucible on a sand-bath with the lid on (slightly tilted), pack it round with sand (Fig. 82). Heat it for twenty minutes with the flame about three-quarters of its full height. Carefully remove the crucible from the sand-bath so as to cool. Weigh. Re-heat for ten minutes and again weigh,

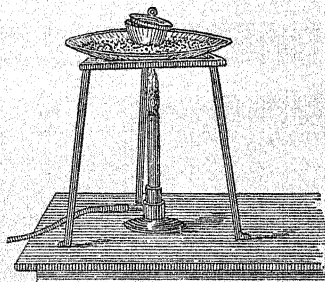


FIG. 82.

and so on until there is no further loss in weight. Record thus :—

Weight of crucible	= 12.35 gm.
Weight of crucible + blue vitriol	= 14.03 ,,
∴ Weight of blue vitriol	= 1.68 ,,
Weight of crucible + residue, heated 20 mins.	= 13.42 ,,
Weight of crucible + residue, heated 10 mins. more	= 13.42 ,,
∴ Loss in weight	= 0.61 ,,
Hence, 1.68 gm. blue vitriol lost	0.61 ,,
∴ 100 gm. blue vitriol lost	$= \frac{0.61}{1.68} \times 100$,,
		= 36.08 gm.

Do you find the same loss?

PROBLEMS AND EXERCISE

Crystallisation

1. Prepare crystals of the following substances:—Rochelle salt, chrome-alum, copper sulphate, ferrous ammonium sulphate, zinc sulphate.

2. Prepare crystals of salt by pouring a saturated solution of salt into a soup plate, and putting it where it may gradually evaporate undisturbed and free from dust.

3. Pour a dilute warm solution of sal ammoniac over a warm sheet of glass. Allow to cool. Examine the crystals which form with a microscope, and draw their shape.

4. Find what happens when crystals of chrome-alum are placed in a crystallising solution of common alum (potash alum).

5. Make a mixture of chrome-alum and potash alum solutions. Crystallise.

6. Make mixtures of the following substances, and then separate samples of the constituents by means of water:—

Chalk and common salt.

Nitre and charcoal.

Sand and sugar.

7. From the muddy liquid, containing a soluble and an insoluble substance,¹ separate the whole of the latter, clean and dry.

8. From the given powder,² containing a soluble and an insoluble substance, separate the whole of the latter, clean and dry.

9. Find the percentage of insoluble matter in the given mixture.³

10. Separate from gunpowder the part which is soluble in water.

11. Mix thoroughly 40 grams of chlorate of potash with (a) 40 grams potassium chloride, (b) 40 grams bichromate of potash. Refer to a text-book for their solubility curves. Then separate by crystallisation a sample of pure chlorate of potash from each mixture.

12. Find the percentage weight of water of crystallisation in 100 grams of the following substances: A, B, C, D, E.³

¹ E.g. KClO_3 and MnO_2 ; KMnO_4 and MgO .

² E.g. $\text{K}_2\text{Cr}_2\text{O}_7$ and CaCO_3 ; KClO_3 and MnO_2 ; NaCl and SiO_2 .

³ The following are suitable substances: $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, $\text{ZnSO}_4 \cdot 7\text{H}_2\text{O}$, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, $\text{MgSO}_4 \cdot 7\text{H}_2\text{O}$, $\text{BaCl}_2 \cdot 2\text{H}_2\text{O}$.

CHAPTER XV

RAW MATERIALS

The Earth is the Chemist's Storehouse.—The most careless observer cannot fail to conclude that the soil and rocks which form the crust of the earth are not everywhere of the same nature. The purple hue of Welsh slate, the transparency of Derbyshire spar, the sombre gray of Scotch granite, or the yellow sands of the sea-shore point to differences of material. The early alchemists understood this, and stored the different minerals and ores they found in the crust of the earth upon the shelves of their laboratories. They made experiments with them in their crucibles and alembics, and succeeded in producing many new substances.

So, too, for us the earth is a storehouse, and we have but to search it for raw material. We may obtain common salt from Northwich, chalk from Dover, iron ore from Spain, or saltpetre from India. With these and other native or raw materials we may then make all manner of experiments and hope to prepare new substances.

It will be wise, therefore, in the first place to obtain some of these materials, and proceed to examine them, noting points of resemblance and difference. We must carefully observe their appearance, forms, and their colours; their taste and the action of water upon them may also be tried, and it will perhaps be useful to try the effect of heat upon each; then when a mineral is brought to us to be named we shall often be able to identify it.

I. COMMON SALT

Common salt is always present in sea water along with other substances, and for ages men have obtained it from the waters of the oceans. If sea water be evaporated, cubical crystals of common salt separate out. These must be purified by re-crystallisation. In warm climates the evaporation is often done in great basins on the sea-shore, and the crystals of salt are raked off. "Bay salt," which may be purchased through a chemist, shows the cubical crystals.

Common salt also occurs as huge deposits of rock-salt in various strata of the earth. The chief of these in England is at Northwich in Cheshire, where there are two great beds of rock-salt, together nearly 60 feet in thickness. It is probable that this rock-salt was deposited ages ago by the evaporation of sea water. A common method of bringing the salt to the surface from these buried stores, is to bore down to them and to pour water into the boring. After lying in contact with the salt beds, the water becomes thoroughly saturated. The rich brine is then pumped up to the surface and evaporated. One hundred parts by weight of brine will yield about 23.5 parts of salt.

Expt. Is Common Salt changed by Heat?—Place a little common salt in a small dry test-tube. Heat it with a Bunsen burner. The salt is scarcely changed in appearance. If the salt be pure, no vapour is given off. Frequently common salt is moist; if so, a dew of water will be deposited on the glass. A crackling noise is also heard as the crystals break and fly into smaller crystals. If the flame is a good one, the salt may at last melt.

2. CHALK

Chalk is found right across England from Flamborough Head, along the Wolds of Yorkshire and Lincolnshire, past Newmarket and the Gog Magog Hills in Cambridgeshire, to Salisbury Plain. Thence it stretches in the North and South Downs to Dover and Beachy Head. Chalk has been shown to consist largely of the broken fragments of tiny shells or foraminifera, such as are now forming the gray ooze or soft

bed at the bottom of the Atlantic, on which the telegraph cables rest (see Fig. 83).

EXPT. To find the Tiny Shells in Chalk.—With a camel's hair brush dust the surface of a lump of chalk into a watch-glass containing some water. Carefully pour off the water from the watch-glass, and then fill it up again with water. In this way wash the dust again and again with water. The heavier particles sink to the bottom, the lighter are swept away. Examine the residue with the microscope. Real chalk must be

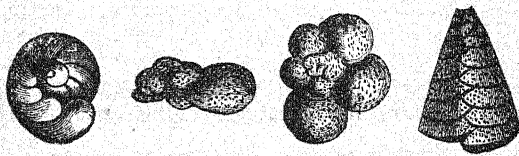


FIG. 83.

used for this experiment, not the soft "chalk" used on blackboards.

For further information the student is recommended to read Huxley's "Physiography" on this subject.

EXPT. The Effect of Heat on Chalk: Formation of Quicklime.—Heat a little powdered chalk on platinum foil in a Bunsen flame. It does not change in appearance. Remove it, and allow it to cool. Taste a very small portion. It is very caustic, and not earthy like chalk.

It has long been known that when chalk or limestone is exposed to a violent heat, it is converted into this caustic substance, called **quicklime**.

In chalk and limestone districts you may have seen the kilns in which quicklime is obtained by heating a mixture of chalk or limestone with a little coal.

Limestone is the general name given to many stones which when burnt in a kiln yield lime. To include all such stones we ought to call *chalk* and *marble* limestones. A large part of the Pennine Range consists of hard dark blue-gray stone streaked with white, called "mountain limestone." It contains numerous fossils of shells and corals, suggesting that it has once been mud at the bottom of a coral sea. It is found from Cheviot past

Crossfell and Ingleborough to the Peak. The streams which flow over it, like those in chalk districts, are hard. Their waters dissolve the rock into fantastic caverns.

Another sort of limestone called *Dolomite* or *Magnesian Limestone* extends in a narrow band from the coast of Durham past Knaresborough and Pontefract in Yorkshire, to Nottingham.

3. SODA CRYSTALS

Up to the beginning of the last century the sea-weeds cast up on the coasts of Normandy, Scotland, and Ireland were the only source of soda. The seaweeds were burnt, and their ashes—called *kelp*—were dissolved in hot water. On evaporating the liquor, crystals of soda were obtained. It is now made from common salt in very large quantities and by more complicated processes.

(1) Examine and describe some soda crystals. They are of irregular shape, white, and semi-transparent.

(2) Taste a crystal. It has a peculiar brackish "alkaline" taste.

(3) Shake a few powdered soda crystals with cold water in a test-tube. They dissolve extremely easily. Place a drop or two of the solution on the palm of the hand, and rub with a finger. It produces a soap-like sensation. Hence, the substance is called *washing-soda*, because it is found to have the same cleansing action as soap, but it is sometimes simply called *soda*.

(4) Find the effect of heat on soda crystals. Proceed exactly as in Expts. 10, 11, and 12, p. 141. Do you not find that the crystals seem to contain a good deal of water of crystallisation? How much? Nearly 63 per cent.

4. POTASH¹

When land plants are burnt, an ash is left which contains a substance very similar to soda. This used to be done in large pots, and so the ash was called *pot-ash*. The ash was then treated with hot water, and the solution so obtained was filtered from any undissolved matter. On evaporating the solution, a

¹ Potassium carbonate.

brown mass of crude potash was obtained. To purify it, it was heated strongly, and then re-treated with water.

(1) Examine and describe potash. It is a white granular non-crystalline powder.

(2) Taste a little. It has the same alkaline taste as soda.

(3) Try the action of cold water on it. It is extremely soluble, and the solution produces the same soap-like sensation.

(4) Find the effect of heat on potash in a small dry test-tube. Does it melt? Is vapour given off? No. Potash differs, in fact, chiefly from soda in not containing water of crystallisation.

5. SALTPETRE OR NITRE

In certain parts of India, the soil, after the rainy season, becomes covered with small crystals of a different salt from any referred to so far in this chapter. This "salt of the rock" was called *sal petrae* by the alchemist Geber, and is known at the present day as *saltpetre* or *nitre*. To separate this substance, the soil is scraped and treated with water, in which the nitre readily dissolves, and on the evaporation of this solution the nitre crystallises out.

(1) Examine and describe carefully the appearance of nitre. Make a drawing of a good crystal. It is a long 6-sided prism, with wedge-shaped ends.

(2) Taste a small crystal of nitre. It has a bitter cooling taste.

(3) What has been found to be the action of water on nitre?

(4) **The Effect of Heat on Nitre.**—Place about 1 gram of nitre crystals in a dry test-tube. Heat in a Bunsen flame. The crystals melt, and the liquid seems to boil, and a number of little bubbles are evolved. Allow to cool. It solidifies, forming a white opaque cake, which is crystalline unless the nitre was heated a long time.

6. SULPHUR

Sulphur or brimstone (burning-stone) has been known since very early times, for it is often present in large quantities in volcanic districts. This is particularly the case in the island of

Sicily, where it lies embedded in the rocks. Examine a specimen of native sulphur.

Sulphur, we have already found, is readily melted, and it will itself burn. In order, therefore, to separate the sulphur from impurities, it is only necessary to pile a heap of ore on the slope of a hill, and kindle the heap at the bottom; the heat from the sulphur which burns is sufficient to melt the rest, and this trickles down and is run into moulds, where it solidifies.

EXPT. To prepare "Flowers of Sulphur."—Fix a hard glass test-tube by means of a cork in the wide mouth of a lamp

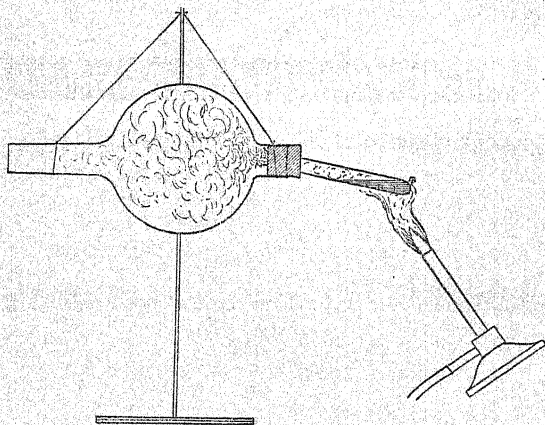


FIG. 84.

glass, as drawn (Fig. 84). Place a few lumps of sulphur in the test-tube. Sling the whole by a string from a ring of a retort-stand. Heat the sulphur gently. It readily melts, forming a mobile pale yellow liquid (at 114° C.), and then it quickly darkens in colour, and becomes as thick as treacle: with more heat it again becomes a thin liquid, but it remains very dark in colour, and presently boils; the temperature is then 448° C. The vapour passes into the globe of the lamp-glass; much of it is condensed as a yellowish powder, but as the glass walls become hot, some may collect at the bottom as a liquid, which

quickly solidifies. The fine light yellow powder is called "flowers of sulphur."

This process is carried out on the large scale by boiling sulphur in an iron retort, G, and sending the vapours into a great brick chamber, A, 200 cubic metres in capacity (Fig. 85). After a time the walls of the chamber become so hot that the vapour is not condensed as flowers of sulphur, but as a liquid, which collects on the floor. It is then run out and cast in wooden moulds in the form of sticks. This is called roll sulphur. Thus the process is a distillation, and all impurities are left behind.

Make the following simple experiments with some roll sulphur.

(1) Examine and describe a piece. Crush it gently with a pestle in a mortar; it is very brittle. Notice a fresh surface, and observe that the lump is a matted mass of crystals.

(2) Heat a few small pieces in a dry test-tube, and notice the changes it undergoes.

(3) Invert a crucible lid on a pipe-clay triangle. Place a small piece of sulphur on the lid, and heat the lid until the sulphur catches fire. Notice the blue lambent flame, and the suffocating "sulphurous" odour.

(4) Find whether sulphur is at all soluble in water. It is quite insoluble.¹

The effect of cooling molten sulphur rapidly by pouring it into water is remarkable:—

¹ Roll sulphur is, however, soluble in a liquid called carbon bisulphide. Shake some powdered roll sulphur with a little carbon bisulphide in a dry test-tube. The sulphur is dissolved. Pour the solution into a watch-glass, and let it evaporate in a fume-cupboard. Crystals of sulphur are recovered.

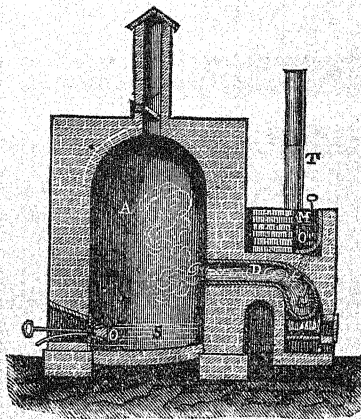


FIG. 85.

EXPT. To prepare "Plastic Sulphur."—Fill a large hard glass test-tube about one-third full with powdered roll sulphur. Fill a beaker with cold water, and invert a funnel in it. Heat the sulphur cautiously, until it is a very dark mobile liquid, and then pour it in a thin stream into the water, round and round the funnel. Lift out of the water the funnel with the sulphur, and notice that it is no longer crystalline and hard, but soft and elastic like india-rubber.

That this curious substance is still *sulphur*, though in a different form, is shown by the fact that in a few days it becomes hard and opaque, and like ordinary brittle crystalline sulphur. This soft variety of sulphur is called "**plastic sulphur.**"

7. IRON PYRITES

Iron pyrites occurs very widely in many different places. When pure it has a brass-like colour. Small quantities of it may frequently be found in lumps of coal.

Examine a piece carefully. Perfect crystals of iron pyrites take the forms of the regular solids; the cube and the octahedron are common forms.

EXPT. The Effect of Heat on Iron Pyrites.—Powder a small quantity, and heat it in a dry test-tube. A sulphurous vapour is given off, and a yellow deposit of sulphur is formed in the cooler part of the tube, while the mineral acquires a deep red colour. If the yellow deposit is heated, it melts and moves up the tube.

8. GREEN VITRIOL¹

When iron pyrites is long exposed to the air and moisture, a greenish liquor is obtained. If the liquor is evaporated, large green glass-like crystals form. These were known to the alchemists under the name of *green vitriol*.²

At the present time large quantities of green vitriol are obtained in this way from the pyrites occurring in the coal of South Lancashire.

1. Examine and describe some green vitriol.
2. Taste a small crystal of green vitriol. It has a sweet and inky taste.

¹ Sulphate of iron.

² Latin *vitrum*, glass.

3. Try the action of cold water on it. It is readily soluble.
4. **The Effect of Heat on Green Vitriol.**—Place a few crystals in a dry test-tube. Heat cautiously in a flame. It melts, boils, and a dew of water is deposited on the cool upper part of the tube. Soon it becomes dry, and a hard white mass is left. Then heat strongly. A sulphurous vapour is given off, and the mass becomes red, where it is in contact with the glass. This red substance is used as *rouge*.

9. BLUE VITRIOL

The drainage water from copper mines has frequently a blue colour. If such water is partially evaporated, and the liquid is then allowed to cool, fine blue crystals are obtained. These were known to the alchemists as *blue vitriol*.

1. Examine and describe some blue vitriol.
2. Do not taste blue vitriol. It is poisonous.
3. Try the action of cold water upon it. It is fairly soluble in cold water, and much more so in hot water.
4. What has been found to be the effect of heat upon crystals of blue vitriol? See p. 141.

10. SAL-AMMONIAC³

A remarkable salt was prepared long ago in Egypt from camel's dung. When the dung was burnt, a heavy soot was deposited, from which the salt called *sal-ammoniac* was made. In those days, just as at the present time, it was widely used as a medicine.

Later, it was prepared from the soot obtained when such animal refuse as hoofs, horns, and hair were heated.

1. Examine and describe the appearance of sal-ammoniac. It consists of minute white crystals.
2. Taste a small quantity. It is sharp and cooling.
3. Try the action of water upon it. It is very soluble.
4. **The Effect of Heat on Sal-Ammoniac.**—Place a small amount of sal-ammoniac in a dry test-tube. Heat gently. White fumes are given off. Finally the whole of the sal-ammoniac disappears from the bottom of the tube; part has escaped into the air, but much is found on the cooler part of

¹ Ammonium chloride.

the tube. Examine some of this deposit: taste it, and try the action of water upon it. It behaves in exactly the same way as the original substance.

This is an example of what is called **sublimation**. The sal-ammoniac which has been re-formed on the side of the tube by the condensation of the vapour of the original sal-ammoniac is called a **sublimate**, and the original sal-ammoniac is said to have **sublimed**.

PROBLEMS AND EXERCISES

Raw Materials

1. Describe fully a piece of slate.
2. Describe fully a lump of granite.
3. Describe fully the appearance of the substances, A, B, C, D, E, F.¹
4. Find the effect of heat on the substances, G, H, I, J, K, L.²
5. Identify the substances, M, N, O, P, Q, R.³
6. Find whether any water is contained in the substances, V, W, X, Y, Z.⁴
7. You are given a powder.² Describe
 - (1) its appearance.
 - (2) action of cold water upon it.
 - (3) effect of heat in a dry test-tube.

¹ *Note for the Teacher*—The following are suitable substances, *c.g.* HgO, HgI, Pb₃O₄, NH₄NO₃, HgS, Sb₂S₃.

² *c.g.* NaHCO₃, Pb(NO₃)₂, ZnSO₄, 7H₂O, KClO₃, camphor, iodine, cobalt chloride, Rochelle salt.

³ *c.g.* NaCl, KNO₃, CuSO₄, CaO, NH₄Cl, Na₂CO₃.

⁴ *c.g.* Na₂SO₄ · 10H₂O, Rock Salt, KNO₃, K₂Cr₂O₇, ZnSO₄ · 7H₂O.

CHAPTER XVI

THE PREPARATION OF THE COMMON ACIDS AND ALKALIES

THE Alchemists made numberless experiments with the raw or native materials which they found ready for them in the crust of the earth, and learnt how to prepare many new substances. Amongst these none have proved of greater importance and have found more frequent use than certain liquids which have strong acid or "alkaline" tastes.

We also must now learn how to prepare these liquids for future use, and how to distinguish one from another, by such simple tests as we can discover.

A. THE COMMON ACIDS

1. Oil of Vitriol¹

EXPT. 1. Preparation of Vitriolic Acid from Green Vitriol.²—Weigh out about 10 grams of green vitriol crystals into an iron dish or porcelain basin. Heat over a flame, stirring with a glass rod until all water is expelled and a brittle white mass is left. Break the mass into small lumps (but do not powder it). Introduce these into a small retort, *a* (Fig. 86). Support the retort so that it can be heated by a naked flame. Let the neck dip into a little water in a test-tube, *b*. Heat the retort at first with a small flame, and finally with the full flame for at least fifteen minutes. Remove the test-tube. Then turn the flame down gradually, so that the retort may cool slowly, but the experiment is almost certain to result in the destruction of the retort.

¹ Sulphuric acid.

² Lecture Table Experiment.

Is the water in the test-tube still merely water? Pour two or three drops into another test-tube. Taste these. The liquid has a very acid taste. This acid liquid, thus prepared from green vitriol, was called by the early chemists **vitriolic acid**. A better method of preparation has been discovered since the time of the alchemists, but this is not the place to describe it.

Transfer the residue in the retort *a* on to a sheet of white paper. Observe that those portions which were most strongly

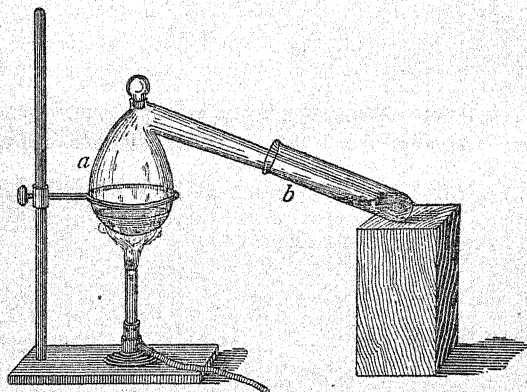


FIG. 86.

heated have acquired a deep red colour. This substance when mixed with oil is used largely as a red paint.

EXPT. 2. Vitriolic Acid when heated gives off Dense White Fumes.—Pour some of the acid prepared in Expt. 1 into a small porcelain crucible. Heat it over a small flame. It boils down to a small bulk giving off steam. Then the fumes become thick and white, and possess a pungent odour, and the remaining liquid has an oily appearance. At once remove the flame, and take the crucible to a fume-cup-board. On this account the acid is frequently called **oil of vitriol**. Another name for it, much used now, is **sulphuric acid**.

EXPT. 3. The Mixing of Oil of Vitriol and Water.—Three-quarters fill a test-tube with water. Mark the level of

the water on the side of the glass. Measure into another test-tube about one-eighth of this volume of acid. Pour some of the acid into the water in a thin stream. It sinks to the bottom; so the acid is heavier than water. Stir the water round with a glass rod. Add the rest of the acid in the same way. Grasp the test-tube in the hand. The mixture has become quite hot, and steams.

Oil of vitriol when mixed with water becomes sometimes hotter than the boiling point of water. What would happen if a little water were poured on to strong oil of vitriol? Being lighter, the water would float on the surface. The liquids would mix at the bounding surface, and become very hot, hot enough to turn some of the water to steam, and throw the acid about with explosive violence. Hence, *never pour water on to strong oil of vitriol.*

EXPT. 4. To show the Destructive Action of Oil of Vitriol.—(a) Write a word with some dilute oil of vitriol with a glass rod on a piece of *paper*. Dry the paper thoroughly by gently wafting it above a Bunsen flame. The paper is charred where the acid was traced.

(b) Place about 10 grams of lump *sugar* in a tall beaker or jar. Add about 10 c.c. of hot water. Place the beaker on a plate or dish. Then add at once about 10 c.c. of strong oil of vitriol. The mixture blackens at once, and froths up.

(c) Mix 1 or 2 c.c. of oil of vitriol with about 10 c.c. of water. Pour the mixture on a small piece of *zinc* in a beaker or basin. There is a violent effervescence. The zinc is dissolved.

Note on Sulphuric Acid.—(1) Should any acid be spilt, throw on some quicklime or powdered chalk, and mop up with an old duster. (2) Never let the acid boil; the fumes are very disagreeable.

After the discovery of oil of vitriol the alchemists tried the effect of heating it with many substances, with the result that they were led to discover other acid liquids, as will be described in the course of the following pages.

2. Spirit of Salt¹

Another acid liquid was obtained by the alchemists by heating common salt with certain substances. Glauber, who

¹ Hydrochloric Acid or Muriatic Acid.

lived at the beginning of the seventeenth century, first prepared it from common salt and oil of vitriol, and it is always made at the present time from these.

EXPT. 5. Preparation of Spirit of Salt.—Weigh out about 30 grams common salt. Introduce into a litre flask. Attach a cork fitted with a thistle-funnel, and a delivery tube bent twice at right angles. Attach an inverted funnel, *a*, to the end of the delivery tube by a small piece of rubber tubing (Fig. 87). Support the flask upon wire gauze on the ring of a retort-stand, and place the funnel within a tumbler, *b*. Add

enough water to the tumbler to cover the mouth of the funnel.

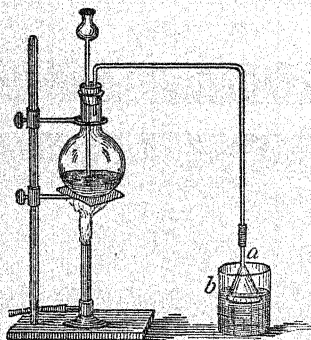


FIG. 87.

Measure out 30 c.c. strong oil of vitriol. Pour it little by little down the thistle-funnel on to the salt. Fumes are given off, and bubbles are forced through the water. As soon as all the acid has been added, heat the flask *gently*. Soon the bubbles cease to come through the water.¹ After a time the water in the tumbler begins to fume strongly. As soon as this is

the case, remove the tumbler and also the flame. Dilute the contents of the flask, and pour away.

Has anything passed into the water in the tumbler? It seems probable from the way in which it fumes. Dip a glass rod into the water, and place a very small drop upon the tongue. It is very acid, and has a different taste from oil of vitriol. Where has the acid come from? The alchemists thought that it was a spirit which the oil of vitriol had driven out of the salt. They therefore called it **spirit of salt**. Later it was called **muratic acid**, from the Latin *muria*, brine; but the fuming acid liquid is now called **hydrochloric acid**.

¹ Care must be taken that the water in the tumbler does not rush back along the tube into the flask. If it begins to do so, at once heat more strongly, or remove the tumbler.

EXPT. 6. Is Spirit of Salt Volatile?—Pour about 5 c.c. of the acid into a porcelain dish. Heat it gently over a flame. Acid vapours are given off it, and it is soon completely evaporated.

Try whether spirit of salt will attack paper and zinc.

3. Spirit of Nitre

The alchemist Geber, who lived about the twelfth century, described an acid liquid obtained by heating nitre with green vitriol. Since the acid liquid was obtained by means of nitre,

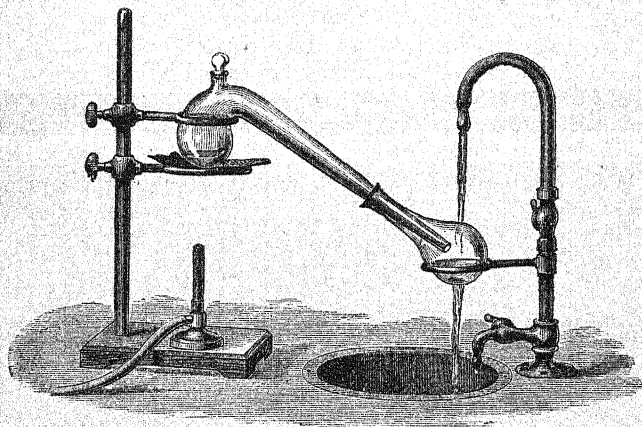


FIG. 88.

it used to be called **spirit of nitre**. It is now usually termed **nitric acid**.

Although spirit of nitre can readily be obtained by Geber's method, it is now usually prepared by heating nitre with oil of vitriol in place of green vitriol. Glauber first prepared it in this way, and he called it "spiritus nitri fumans Glauberi" or "Glauber's fuming spirit of nitre."

EXPT. 7. Preparation of Nitric Acid from Nitre and Oil of Vitriol.—Weigh out 20 grams of nitre, and measure out 30 c.c. strong oil of vitriol. Support a small retort on wire gauze, as shown in Fig. 88. Pass the neck into a small dry

flask. Support the flask under a tap or in a trough, and keep it cool with running water.

Slide the nitre into the retort, and pour the oil of vitriol in through a funnel, taking care that neither gets down the long tube of the retort. Replace the stopper of the retort. Heat gently. The nitre melts, and the liquid begins to boil. A ruddy vapour rises, and drops form and run down the tube. A yellowish liquid collects in the flask. Keep the liquid just boiling, but do not let the temperature rise unnecessarily. As soon as the liquid in the retort begins to thicken, remove the flame.

Pour the remaining liquid from the retort into an empty evaporating basin. Put it aside. On cooling it hardens into a solid mass. If this had been left in the retort, it would have been rather difficult to get it out. Wash out the retort when cold.

Can nitric acid be evaporated just as spirit of salt (hydrochloric acid) ?

EXPT. 8. **Is Nitric Acid Volatile?**—Tests for Nitric Acid.—(1) Pour about 5 c.c. of the liquid into a porcelain dish. Heat gently until the liquid is evaporated. Vapours are evolved possessing a peculiar odour quite unlike the fumes of oil of vitriol or of spirit of salt.

(2) Place a small drop of the liquid upon a finger nail, and another upon the skin of the hand. Wash them off after a few seconds. Bright yellow stains are left.

(3) Place a small piece of copper in the liquid. Red fumes are at once evolved, and the liquid becomes green or blue.

Show that oil of vitriol and spirit of salt do not attack copper in this way. On account of the corrosive action of nitric acid it used frequently to be called **aqua fortis** (*i.e.* strong water).

Try whether nitric acid will attack paper and zinc.

4. *Vinegar or Acetic Acid

Acetic acid is another acid which is frequently used, and therefore may be mentioned here, although it is not prepared from any materials which are found native in the rocks or crust of the earth. It was the only acid known to the ancients.

When beer or a weak wine such as claret is left exposed to

the air, it quickly becomes sour. This is because the alcohol has been changed into the acid called acetic acid or vinegar.

The best way to convert weak wine into vinegar is the German "quick-vinegar process." A large cask with perforated sides is filled with beech-wood shavings, which are moistened with a little vinegar. Weak wine is then poured in at the top, and it trickles slowly through the shavings, and the liquid which collects at the bottom is again and again poured down until all the alcohol is changed into vinegar. A kind of fungus called *mycoderma aceti* is found on the shavings, and it is this which in the presence of air effects the change.

EXPT. 9. Is Acetic Acid Volatile?—Heat about 5 c.c. of the acid in a basin. Observe that the vapour has a very pungent odour, which is quite different from that from any of the other acids. Little or no residue is left.

Try whether acetic acid will attack zinc, copper, or paper.

(B.) SOME COMMON ALKALIES

We must now learn a little about some substances very different in character from the acids, which the alchemists also knew how to prepare.

1. Caustic Soda

It was found long ago that if a solution of soda crystals was heated with lime the solution changed in character and acquired a very caustic nature. You will have an opportunity of trying this yourself later (see Volume II.).

If such a solution is allowed to stand till clear, and is then evaporated, a white non-crystalline solid is obtained. A piece of this solid or of the strong solution will quickly cause a blister when placed upon the skin. It also destroys vegetable fibre, e.g. wood or filter-paper. On this account it is called **caustic soda**, and must be handled with care.

Obtain a piece of solid caustic soda. Examine it, and see how readily it dissolves in water.

Dilute with a large amount of water a little of a solution of caustic soda, such as is in a bottle on your shelf. Taste it, and notice its soapy "**alkaline**" nature. Wet your fingers with the solution: they feel soapy and slippery. Next time your

hands are greasy with oil, try washing them in a dilute caustic soda solution. They will become beautifully clean. A stronger soda solution would remove the skin as well as the dirt.

2. Caustic Potash

Another caustic solution is formed if a solution of potash is heated with lime. If the clear caustic solution is evaporated, a white solid is obtained which is almost exactly similar in appearance to caustic soda. Moreover its solution in water has a very similar caustic and alkaline nature.

This substance is called *caustic potash*. It is usually sold in white sticks.

3. Volatile Spirit of Sal-Ammoniac or Ammonia

EXPT. 10. Preparation of Spirit of Sal-Ammoniac or Ammonia.—Weigh out 10 grams sal-ammoniac. Powder it.

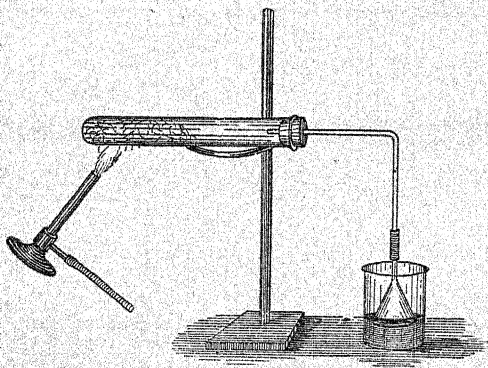


FIG. 89.

Dry it by heating it gently in a basin on a sand-bath. Allow it to cool. Meantime weigh out 15 grams quicklime, and powder it. Mix the two together. Put the mixture in a large dry test-tube. Fit the tube with a cork and a delivery tube bent once at right angles. Attach an inverted funnel to the delivery tube. Support the test-tube on the ring of a stand, so that the funnel dips into a tumbler half filled with water,

with its mouth *just below* the surface of the water, as shown in Fig. 89.

Heat cautiously by wafting the flame about under the tube. At first bubbles of air are forced up through the water, and then these cease. Heat five minutes longer. Take care that the water in the beaker is not sucked into the test-tube. Then remove the beaker, and afterwards the flame.

Observe the pungent and peculiar odour the water in the beaker has acquired. As this liquid was acquired by heating sal-ammoniac with lime, the early chemists called it the **volatile spirit of sal-ammoniac**. It is now usually called **ammonia**.

Ammonia can also be prepared by heating shavings of horn. Indeed, this method of preparation was at one time so commonly used that the pungent liquid came to be known as "spirits of hartshorn," a name which is still common for it.

EXPT. 11. Tests for Ammonia.—(1) Pour a little of the solution into a porcelain crucible. Heat over a small flame. It is quickly evaporated, and the vapours possess the characteristic odour already observed.

(2) Dip a glass rod in strong hydrochloric acid. Hold it in the vapours arising from the solution. Dense white fumes are formed.

Repeat with a rod dipped in sulphuric acid and in nitric acid.

4. Lime Water

EXPT. 12. Slaked Lime is somewhat Soluble in Water.—Half fill a test-tube with water. Add some powdered slaked lime. Shake. The liquid becomes white and milky. The lime has not all dissolved. Perhaps some of it has dissolved. Allow the test-tube to stand. The white mud, which is only in suspension, begins to settle. In time it will all settle, and leave the liquid quite clear. To save time filter the liquid. Evaporate a drop on platinum foil. A white stain is left, showing that some of the slaked lime was dissolved.

Slaked lime, then, is slightly soluble in water. The solution is called **lime water**. Observe its alkaline taste. Lime water is generally prepared by shaking quicklime with water, and allow-

ing it to settle. The shaking is repeated to strengthen the solution. The clear liquid is then decanted or syphoned off.

The acid and alkaline solutions we have been studying are commonly spoken of as *acids* and *alkalies*.

PROBLEMS AND EXERCISES

The Common Acids and Alkalies

1. Find the effect of evaporating the liquids, A, B, C, D, E,¹ in a porcelain basin.

2. You are given samples of the following pairs of liquids. Distinguish the members of each pair.

Oil of vitriol and nitric acid.

Hydrochloric acid and nitric acid.

Oil of vitriol and acetic acid.

Acetic acid and ammonia.

3. Find the effect of heating 1 part of nitre with 2 parts of green vitriol in a small retort the neck of which dips into a little water in a flask.

4. Find the effect of heating sal-ammoniac with oil of vitriol, as in Expt. 5.

5. Find the effect of heating Chile saltpetre with oil of vitriol, as in Expt. 7.

¹ *Note for Teacher.*—e.g. NH_4OH , $\text{Ca}(\text{OH})_2$, H_2SO_4 , HCl , HNO_3 , $\text{CH}_3\cdot\text{COOH}$.

CHAPTER XVII

THE ACTION OF ACIDS AND ALKALIES UPON LITMUS—THE FORMATION OF SALTS

CHEMISTS found 200 years ago that the acid and alkaline substances we have just learnt to prepare had a curious effect on certain vegetable dyes. Of these dyes *litmus* is one of those most readily acted on, and it is therefore very suitable for our experiments. It is a solid blue substance, which is extracted from certain kinds of lichen.

EXPT. 1. **To make an Infusion of Litmus.**—Weigh out about 5 grams of solid litmus. Grind it up in a mortar, and transfer into a flask. Half fill the flask with cold water, and shake well. Allow the powder to settle, and then pour off and throw away the liquor. In this way certain impurities are washed out of the litmus.

To the residue add 100 c.c. of water, and heat the flask nearly to boiling for a few minutes. Allow to settle¹ (if possible for some hours in a warm place). Decant off the intensely blue solution from the sediment, and place the solution in a bottle.

EXPT. 2. **To find the Effect of Acid and Alkaline Liquids on the Blue Litmus Solution.**—Let us write down the names of the acid and alkaline liquids we have on our shelves :—

Acid liquids.

Sulphuric acid.

Hydrochloric acid.

Nitric acid.

Acetic acid.

Alkaline liquids.

Caustic soda.

Caustic potash.

Ammonia.

Lime-water.

¹ If time does not allow of this, filter a portion of the liquor.

Try the effect of each of these liquids on the blue litmus solution in the following way :—Fill one-third of a test-tube with water ; add 5 drops of the litmus solution by means of a pipette or piece of glass tubing ; then add 5 drops of each liquid, and stir. Carefully observe any change in colour by looking down the tube. State the effect you notice in a table thus :—

Liquid taken. *	Effect on the Blue Litmus solution.
Sulphuric acid. Hydrochloric acid, etc.	Change to brick-red colour.

You have probably found that all the acid liquids turn the blue litmus solution **red**, while all the liquids possessing an alkaline instead of a sour taste leave the litmus solution **blue**, or render it still more blue. By this colour change, then, it seems that we can readily distinguish between an acid and an alkaline solution. But have no other liquids beside acid and alkaline solutions any colour effect upon litmus ? This question must be tested by experiment

PROBLEM

Find whether blue litmus solution is affected by alcohol, benzene, a solution in water of common salt, washing-soda, or of any other substance you can think of. State your observations in a table.

EXPT. 3. To make Red and Blue Litmus Papers.—Take a strong blue infusion of litmus, some very weak nitric acid, and some very weak caustic soda solution. Divide the blue litmus solution between two porcelain basins. To one portion add a few drops of the dilute acid till it is just reddened. To the other portion add one drop of the weak caustic soda solution so that it has a blue colour. Soak pieces of blotting-paper in each solution. Dry them as on p. 139. Then cut them into strips, and keep in a glass-stoppered bottle.

PROBLEM

Try the colour-effects of acid, alkaline, and other liquids upon these red and blue litmus papers, by putting drops of the liquids upon strips of each kind by means of a glass rod. State what you observe in a table thus.

Liquid.	Effect on Red Litmus.	Effect on Blue Litmus.
Sulphuric acid. Hydrochloric acid, etc.	No visible change.	Red colour.

EXPT. 4. To try the Delicacy of the Litmus Test for an Acid.—Add a cubic centimetre of strong sulphuric acid to a litre of water. Shake well. The acid is now diluted one thousand times. Put a drop on a blue litmus paper. It is reddened. Find whether the acid can still be detected when this weak solution is made 10 times and 100 times weaker still.

EXPT. 5. To prepare a Purple or Neutral Litmus Solution.—Take some blue litmus solution in a porcelain basin, very weak acid solution, very weak alkali, and a pipette. Drop the dilute acid into the blue litmus solution, and stir with a glass rod until the colour is half-way between red and blue. Probably you will overstep the mark. If so, add the dilute alkali drop by drop very carefully. If the litmus changes in colour sharply from red to blue, weaken both the acid and alkaline solutions, and try again.

You will at last succeed in getting a **purple** or **neutral** litmus solution, intermediate in colour between red and blue.

To two portions of the purple or neutral litmus solution add acid and alkali respectively. One changes to red and the other to blue. Then, if we keep a stock of this one neutral solution, we can use it to detect both acids and alkalies, instead of using both red and blue solutions or papers.

Keep the neutral litmus solution in a glass-stoppered bottle.¹

¹ If the solution loses its colour, the colour will be quickly restored on removing the stopper and shaking the solution with a little fresh air.

PROBLEMS

1. Prepare blue-purple and red-purple litmus solutions, intermediate in colour between blue and purple, and red and purple respectively.
2. Find whether lime-juice, soap, vinegar, washing-soda, bicarbonate of soda, common salt, nitre, borax, milk, sour milk, etc., have any effect on neutral litmus solution.
3. Find the effects of acids and alkalies upon solutions in alcohol or methyl orange and phenol. pht halein ; and upon infusions in water of red cabbage, logwood, and violet leaves. Fill one-third of a test-tube with water : add one drop of the solution, and then one drop of an acid or alkali by a pipette.

THE FORMATION OF SALTS

We have found that of the acids and alkalies known to us :—

1. An acid reddens blue litmus, and leaves red litmus unchanged.
2. An alkali turns red litmus blue, and leaves blue litmus unchanged.

What will be the effect of mixing an acid with an alkali in the presence of litmus ?

EXPT. 6. To find the Result of adding Hydrochloric Acid to Caustic Soda until the Liquid is Neutral.—Place about 10 c.c. of strong caustic soda solution in a basin. Add a *few* drops of litmus solution so that it is distinctly blue. Fill a pipette with dilute hydrochloric acid. Run the acid carefully drop by drop into the basin, stirring continually until the colour suddenly changes from blue to red. Then add some very dilute caustic soda solution until the liquid has a neutral or violet tint. Test also with litmus paper.

What is there now in the solution ? Is it simply a mixture of the acid and the caustic soda ? Suppose we find out what is left when the liquid is evaporated, for you will remember that hydrochloric acid, or spirit of salt as it used to be called, is volatile, so that we may expect to drive off the acid and have only the caustic soda left.

Evaporate the solution on wire gauze until the residue is dry. There will probably be some spurting, but this does not matter, as we do not need to weigh the residue in this experiment.

Do you observe the escape of any acid fumes? Does the solution still remain neutral as it evaporates?

Do you recognise the residue? Has it the appearance of caustic soda? Dissolve a little in water in a test-tube, and try whether the solution is alkaline or acid. Taste a very small portion. Do you not find it to be **common salt**?

Here, then, is an unexpected and important result. By the neutralisation of caustic soda with hydrochloric acid we do not obtain merely a mixture of the two with properties which are between those of the acid and the alkali. On the contrary, we find common salt to be produced, a substance which is very different in its properties from either of them. Is a similar result met with when other acids and alkalies are "neutralised?"

PROBLEMS

Find the products on neutralising—

1. 20 c.c. caustic potash with nitric acid.
2. 20 c.c. ammonia with hydrochloric acid.

Evaporate each solution, till the residue is dry, in a basin over a flame. Carefully examine and endeavour to identify the substances which you obtain by reference to appearance, taste, solvent action of water, and the action of heat in a dry test-tube.

Salts.—You have probably found after carrying out these problems that in each case a solid crystalline substance is obtained, which is neutral to litmus and has a salt-like taste. Other neutral mixtures of acids and alkalies might be made, and in each case you would obtain a neutral saline product. These substances are called **Salts**. The term "salt," then, as we shall use it in chemistry, applies not only to common sea-salt, but also to saltpetre and sal-ammoniac, and to many others, such as Glauber's salt and Epsom salts.

So far, then, as we can say at present a *salt is a substance, crystalline in appearance, saline in taste, neutral to litmus, which is obtained by the action of an acid upon an alkali.*

Old and New Names.—You have doubtless noticed that in the previous chapters we have sometimes given two names to the same substance—an old name, such as was used by the alchemists—and a new name adopted in more recent times. Fossils are records of the history of the progress of life on the

earth. Fossil names mark the progress of a science. For example "Spirit of salt," was the name given quite suitably by the alchemists to the volatile liquid obtained by distilling salt with oil of vitriol. The investigations of chemists during the last hundred years have however led to the abandonment of the name "spirit of salt" and the adoption of the name "hydrochloric acid." And because this latter name is the one that is probably to be found on the bottles in the laboratory, it will be more convenient to use it, although we cannot ourselves as yet attach any meaning to it, as we can do to "spirit of salt." Hence as a matter of convenience we shall adopt in many cases the more recent name, although as a matter of principle it would be better to retain the earlier name until we have found out for ourselves the reasons for employing what is really the more suggestive name.

The older names are still often used in commerce and in the household, while the more recent names find favour in the laboratory.

Names of Salts.—Salts are named after the acids from which they are formed. Thus—

Salts formed from sulphuric acid	are called	sulphates.
„	nitric acid	„ nitrates.
„	hydrochloric acid	„ chlorides.
„	acetic acid	„ acetates.

The salt formed from caustic soda and hydrochloric acid is called *chloride of soda*. *Nitrate of potash* can be made from caustic potash and nitric acid. For reasons which will appear hereafter, chloride of soda is more often called *sodium chloride*, and nitrate of potash *potassium nitrate*.

Some of these substances may be recognised in taste and appearance as already familiar under other names.

It will be convenient to have a list of these names of salts and of some other substances at hand for reference :—

TABLE OF SYNONYMS

<i>Old or Household Names</i>	<i>New Names</i>
Oil of vitriol, or vitriolic acid	Sulphuric acid
Spirit of salt, or muriatic acid	Hydrochloric acid
Spirit of nitre, or aqua fortis	Nitric acid

<i>Old or Household Names</i>	<i>New Names</i>
Vinegar	Acetic acid
Washing-soda, mild alkali, or soda crystals	Sodium carbonate
Caustic soda, or caustic alkali	Sodium hydrate
Potash	Potassium carbonate
Caustic potash	Potassium hydrate
Sal-ammoniac	Ammonium chloride
Spirit of sal-ammoniac, hartshorn, or ammonia	Ammonium hydrate
Common salt	Sodium chloride
Nitre or saltpetre	Potassium nitrate
Chile saltpetre	Sodium nitrate
Epsom salts	Magnesium sulphate
Green vitriol	Ferrous sulphate
Blue vitriol	Copper sulphate
Chalk	Calcium carbonate
Quicklime	Calcium oxide
Slaked lime or lime water	Calcium hydrate
Fixed Air	Carbonic acid, carbon dioxide
Fire-Air	Oxygen
Spoilt Air	Nitrogen
Inflammable Air	Hydrogen
Nitrous Air	Nitric oxide
Volatile Sulphurous Acid	Sulphur dioxide

PROBLEMS

1. You are provided with a stick of caustic potash and with some strong nitric acid. Prepare a well crystallised sample of nitre.
2. Try to prepare dry and clean crystals of a salt—sulphate of potash—from a stick of caustic potash and some sulphuric acid.
3. Find whether the solutions A, B, C, are acid, alkaline, or neutral.

CHAPTER XVIII

GRAPHIC REPRESENTATION¹

Areas of Countries.—It is often difficult to read quickly figures which state the sizes of things, and the figures often give no clear idea to the mind. Thus it may be stated that—

Area of British Isles	=	121,115 sq. miles.
Area of Indian Empire	=	1,700,000 „
„ „ under Native rule	=	750,000 „

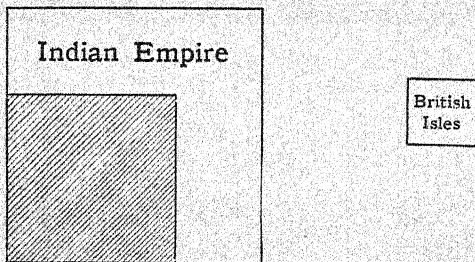


FIG. 90.

But a much clearer picture of these areas is obtained if a diagram is made showing squares drawn in proportion to these numbers as in Fig. 90.

¹ The authors call the attention of teachers to Prof. Perry's "Six Lectures at Jermyn Street in 1899" (Eyre and Spottiswoode, 6d.), from which they have obtained some valuable suggestions.

National Income.—Again, look at the following statement of the national income for the year 1900.

Excise	£33,100,000	Stamps	£7,825,000
Customs	26,252,000	Land Tax	755,000
Property and Income Tax	26,920,000	House Duty	1,720,000
Estate Duty	12,980,000	Crown Lands	500,000
Post Office	13,800,000	Suez Shares	830,000
Telegraph Service ...	3,450,000	Miscellaneous ...	2,243,000
			£130,384,000

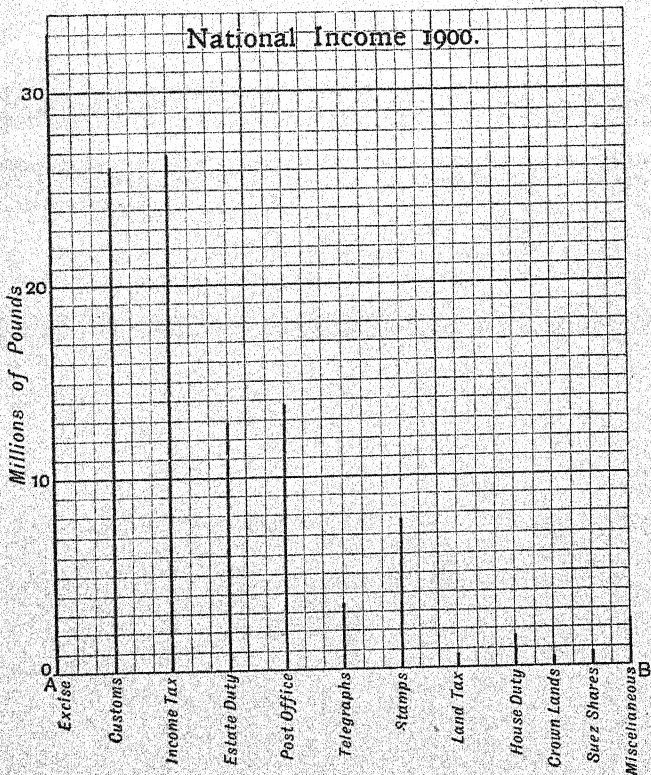


FIG. 91.

Large quantities like these are not easily grasped, but their relative sizes may be shown by drawing on paper lines proportional in length to each. The most convenient paper is what is called "squared paper." This is covered with parallel horizontal and vertical lines at equal distances, and every tenth line is more distinct than the rest. On a sheet of squared paper draw a horizontal line, AB, along one of the thick lines. Then lines drawn at right angles to AB may be taken to represent millions of pounds. If, as is convenient, the paper is ruled in inches and tenths of an inch, a line one inch long may be taken to represent 10 millions of pounds. On this scale of a line one inch long representing 10 millions of pounds, proceed to draw with a sharp hard pencil straight lines at right angles to AB, to represent in order the above items of the nation's income (see Fig. 91). Test the accuracy of your diagram by reading from it the magnitudes of the several items, and comparing your readings with the above table.

Barometric Readings.—The following numbers show the height of the barometer at York, in inches, on each day of October, 1899, at 9 a.m.

October 1	29'45 inches.	October 17	30'34 inches.
2	29'60	18	30'39
3	29'78	19	30'45
4	29'79	20	30'44
5	30'14	21	30'42
6	30'19	22	30'27
7	30'20	23	30'06
8	30'39	24	29'95
9	30'27	25	29'89
10	30'18	26	29'84
11	30'01	27	29'76
12	29'87	28	29'77
13	29'71	29	29'70
14	30'16	30	29'75
15	30'23	31	29'84
16	30'29		

These barometric heights might be represented by a row of 31 straight lines, as in Fig. 91, all standing in order upon one horizontal line, but it will be sufficient to mark the *tops* only of these lines by points. Moreover, since each height exceeds

29 inches, we need only represent the heights *above* 29 inches, and so make the diagram more compact.

Begin, therefore, by drawing two lines OX and OY at right angles to one another (see Fig. 92), along two of the thick lines upon squared inch paper. (These are often called the *axes*). Mark every tenth line along OX, 10, 20, 30, and along OY mark O as 29, and the next thick line 30, and the next 31. Proceed to mark points on the paper with a sharp-pointed hard pencil to represent the above barometric heights.

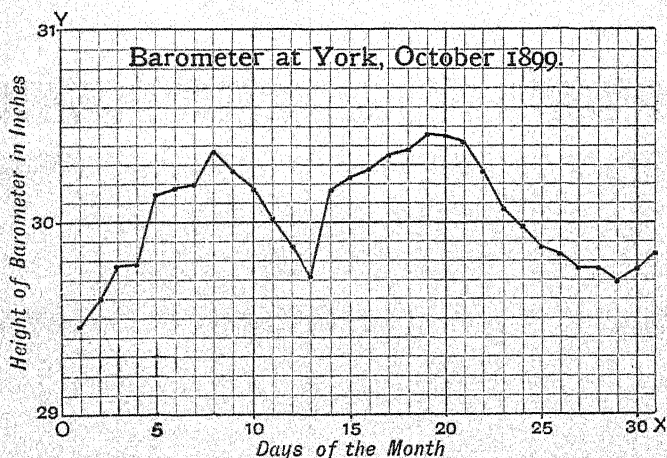


FIG. 92.

Further, since the barometer changes continuously from one height to the next, we may indicate this by joining the points by fine, short, *straight* lines, as in Fig. 92, for according to the data we do not know how the barometric height may have varied during each twenty-four hours. The irregular line so obtained is called a **barometric curve**.

Reference to a self-recording barometer would probably show that the height did not change between the successive readings quite regularly, as our curve implies. Sometimes the barometer is read more often than once a day. Thus at the office of the

Daily Chronicle in London it is read four times a day; the barometric curve consequently takes a more wavy appearance. This is shown by the thick wavy line in Fig. 93, which shows a *Daily Chronicle* weather chart for four days.

Observe that there was no reason to join the tops of the straight lines in the Revenue Diagram (Fig. 91), because the lines in it represent *different* things, not one thing which varies continuously in magnitude. So also if we were to plot the daily

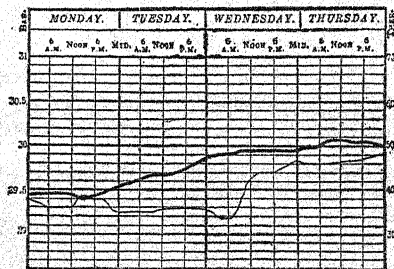


FIG. 93.¹

rainfall we should not attempt to draw a curve connecting the measurements, since the quantity of rain changes abruptly, not continuously, from one day to another.

EXERCISES

In the following exercises plot the tabulated quantities on squared paper. In all cases state on the paper what it is you have represented, and also name the different scales. In choosing the scales avoid those which would make the diagrams either too large, too long or broad, or too crowded.

¹ From Gregory and Simmons's *Elementary Physics and Chemistry*, Part I., p. 141.

1. *Annual Expenditure of Families.*

	A. Average of Twenty Artisan Families.	B. Average of Twelve Middle Class Families.
Rent	£15	£46.4
Food	48.5	101
Clothing	8.8	51.3
Fire and Light	4.8	17.2
Doctor	1.5	9.3
Education	—	16.3
Travelling and Holidays ...	13.6	41.6
Charity	—	24.3
Insurance	—	29.1

2. *Food Expenditure in a Boarding School, 1899.*

Flour	£234.6	Dried Fruits	£24.0
Meat and Fish	1011.6	Vegetables and Fresh Fruit	314.3
Milk	594.3	Coffee and Cocoa	24.4
Butter	228.3	Sugar and Syrup	106.0
Cheese	11.0	Rice	10.8
Tea	49.9	Sundries	174.0

3. *Expenditure of the British and Foreign Bible Society.*

(Founded 1804.)

1804	£692	1854	£149,000
1814	81,000	1864	169,700
1824	94,000	1874	217,400
1834	84,200	1884	231,000
1844	85,800	1894	215,000

4. *Punishment Lines given in a School in successive weeks of a Term.*

Week 1	1050 lines.	Week 8	2750 lines.
2	2190	9	2695
3	2210	10	3535
4	2000	11	2900
5	2170	12	4360
6	1845	13	4490
7	3280	14	4230

5. *Date of Appearance of Wild Flowers.*¹ (Average of 10 years.)

The numbers give the date counting from January 1; e.g., the numbers 1-31 indicate dates in January.

Hazel 32	Marsh Marigold 77	Purple Orchis ... 124
Coltsfoot 52	Cowslip 81	Hawthorn... .. 135
Primrose 54	Cuckoo-flower.. 95	Red Clover ... 137
Lesser Celandine 60	Bluebell 113	Forget-me-not ... 148
Sweet Violet ... 71	Crab 116	White Clover ... 155
Anemone 76	Upright Buttercup 116	Wild Rose ... 170

6. *Barometric Readings at York, January 1884.*

Jan 1 30·44 ins.	Jan. 11 29·81 ins.	Jan. 21 30·30 ins.
2 30·16	12 30·32	22 29·98
3 29·89	13 30·39	23 29·63
4 30·09	14 30·30	24 29·62
5 29·98	15 30·49	25 29·41
6 29·61	16 30·60	26 29·09
7 29·83	17 30·54	27 28·67
8 30·04	18 30·54	28 29·25
9 30·08	19 30·48	29 29·68
10 30·17	20 30·28	30 29·54
		31 29·78

7. *Barometric Readings at York, December 1886.*

Dec. 1 29·77 ins.	Dec. 11 29·31 ins.	Dec. 21 30·35 ins.
2 29·86	12 29·08	22 29·80
3 30·02	13 29·69	23 29·62
4 29·64	14 29·59	24 29·47
5 29·96	15 29·18	25 29·78
6 29·62	16 29·33	26 29·81
7 29·21	17 29·64	27 29·66
8 28·34	18 29·58	28 29·69
9 28·18	19 29·74	29 29·78
10 29·08	20 30·04	30 30·30
		31 30·56

8. *Maximum and Minimum Temperatures.*

Observed at Manchester, during parts of September and October, 1895.
Plot on the same paper.

	Maximum.	Minimum.
Sept. 20	64°·7 F.	44°·4 F.
21	65°·2	43°·0
22	70°·3	44°·0
23	75°·8	43°·7
24	81°·2	56°·4
25	79°·4	61°·6
26	78°·5	54°·5

¹ At Ackworth, Yorkshire. From *The Natural History Journal*, 1888, p. 116.

			<i>Maximum.</i>			<i>Minimum.</i>
Sept.	27	...	81°5	58°1
	28	...	83°9	55°3
	29	...	81°4	52°5
	30	...	76°9	52°2
Oct.	1	...	71°9	52°4
	2	...	52°0	47°3
	3	...	58°8	36°1

9. *Sunshine and Rain.*

Observed at Manchester, during parts of September and October, 1895.

			<i>Sunshine.</i>			<i>Rain.</i>
			Hrs.	Mins.		Ins.
Sept.	20	...	4	38	...	—
	21	...	7	14	...	—
	22	...	8	48	...	—
	23	...	5	56	...	—
	24	...	6	15	...	0°041
	25	...	4	14	...	0°017
	26	...	4	30	...	0°032
	27	...	7	40	...	—
	28	...	7	56	...	—
	29	...	8	32	...	—
	30	...	6	32	...	—
Oct.	1	...	1	24	...	—
	2	...	1	8	...	0°792
	3	...	0	43	...	0°226

10. *Average Heights and Weights of Children.¹*

Age last birthday.	Boys.		GIRLS.	
	Height.	Weight.	Height.	Weight.
5 years	41'74 ins.	41'20 lbs.	41'47 ins.	39'82 lbs.
6	44'10	45'14	43'66	43'81
7	46'21	49'47	45'94	48'02
8	48'16	54'43	48'07	52'93
9	50'09	59'97	49'61	57'52
10	52'21	66'62	51'78	64'09
11	54'01	72'39	53'79	70'26
12	55'78	79'82	57'16	81'35
13	58'17	88'26	58'75	91'18
14	61'08	99'26	60'32	100'32
15	62'96	110'84	61'39	108'42
16	65'58	123'67	61'72	112'97
17	66'29	128'72	61'99	115'84
18	66'76	132'71	62'01	115'80

¹ From Annual Report of State Board of Health, Massachusetts, 1877.

Pressure of Water Vapour.—On page 98 a table was given of the pressure of water vapour measured in millimetres

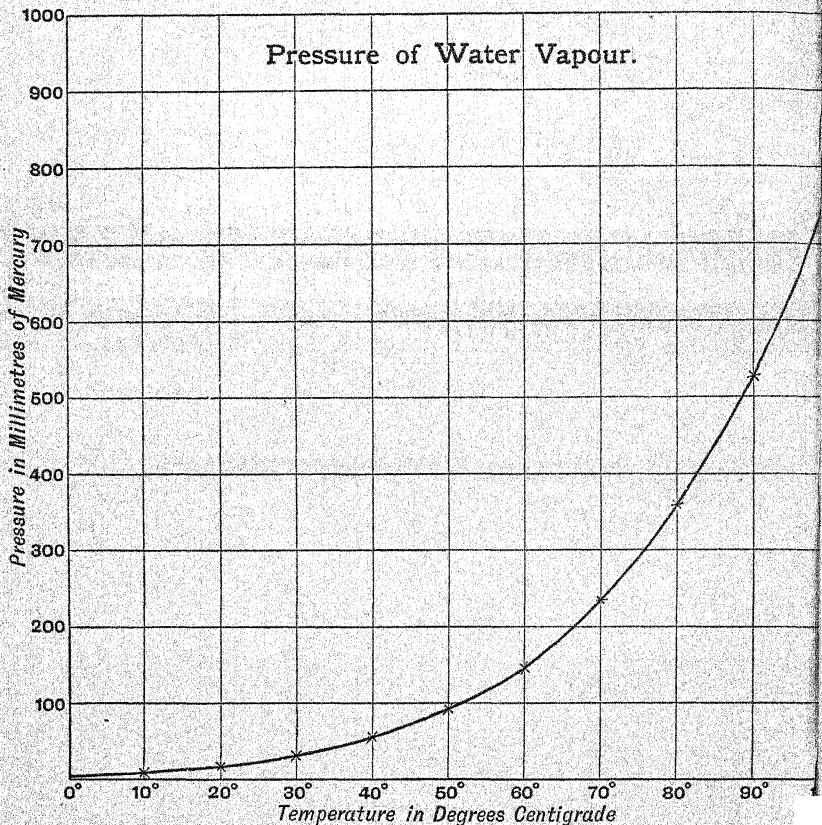


FIG. 94.

of mercury at various temperatures. This may be repeated here and somewhat extended.

Temperature	Pressure	Temperature	Pressure
0° C. ...	4' + mm.	60° C. ...	148' + mm.
10° ,, ...	9' + ,,	70° ,, ...	233' + ,,
20° ,, ...	17' + ,,	80° ,, ...	354' + ,,
30° ,, ...	31' + ,,	90° ,, ...	525' + ,,
40° ,, ...	54' + ,,	100° ,, ...	760' + ,,
50° ,, ...	91' + ,,	105° ,, ...	906' + ,,
		110° ,, ...	1075' + ,,

Proceed to mark these measurements by means of small crosses upon squared millimetre paper, marking degrees Centigrade off along a horizontal axis, and mm. of mercury up from it along a vertical axis (see Fig. 94). Observe that a curved line may be drawn sweeping through the points marked by these crosses. Endeavour to draw the curve.¹ Any fault in the curve can be more readily detected if the paper is held in such a way that the eye can look *along* it, close to the paper. Observe (see Fig. 94)—

(1) That the *slope* of the curve becomes steeper and steeper as the temperature rises : this shows very clearly that the pressure of water vapour increases much faster as the temperature rises.

(2) That the curve enables us to find the pressure of water vapour at temperatures lying between those at which actual measurements have been made : *e.g.* the curve shows that the pressure of water vapour at 85° is measured by 430 mm. of mercury. Values obtained in this way are said to be obtained by **interpolation**. Thus the careful plotting of a few accurate measurements enables us to obtain many other values

Solubilities of Solids.—It has been found in Chapter XIII. that the solubility of a solid in water usually changes with change of temperature. This is seen in the following statement of the quantities of nitre, salt, and chlorate of potash, which can be dissolved in 100 grams of water at certain different temperatures :—

¹ *Note on Drawing a Curve through a Number of Points.*

(i.) Put needles in at the points marked : bend a strip of whalebone or a straight edge round them, and then rule the line with a pencil.

(ii.) Draw the curve freehand, keeping the wrist on the *inside* of the curve whilst drawing, so that the natural motion of the hand round the wrist as centre may help in forming the curve. Sketch the line with a pencil.

	Nitre.	Salt.	Chlorate of Potash.
0° C.	13 gm.	35.5 gm.	3 gm.
10°	21	35.8	4
20°	31	36.1	6
30°	45	36.4	9
40°	64	36.6	13
50°	86	36.9	18
55°	100		
60°		37.2	24
70°		37.5	32
80°		37.8	40
90°		38.1	49
100°		38.4	60

The meaning of these numbers, each one the result of a careful experiment, can be best seen if they are plotted on squared paper.

Along a horizontal axis on a sheet of squared millimetre paper, mark off *temperatures* from 0° to 100° C., and mark off *solubilities* (i.e. grams of solid dissolved in 100 grams of water), from 0 to 100 grams up a vertical axis. Proceed to mark the measurements of the solubility of *nitre* by crosses upon the paper. Observe that a curved line might be drawn sweeping through the points so marked. Endeavour to draw the curve. This line is called the **solubility curve of nitre**. Similarly proceed to draw on the same sheet of paper the solubility curves for *salt* and *chlorate of potash* (see Fig. 95).

A solubility diagram is not only a record of experiments: it also supplies the key to many problems. This will be understood after working through the following examples:—

EXAMPLES XIII

[To be answered by reference to the diagram you have made.]

1. How many grams of nitre will dissolve in 100 grams of water at 4°, at 35°, at 45°?
2. At what temperatures will 100 grams of water be just saturated by 5 grams of potassium chlorate, by 19 grams, by 55 grams?
3. How much more chlorate of potash will dissolve in 100 grams of water (i.) at 60° than at 15°, (ii.) at 80° than at 55°?

4. At what temperature are salt and chlorate of potash equally soluble in water?

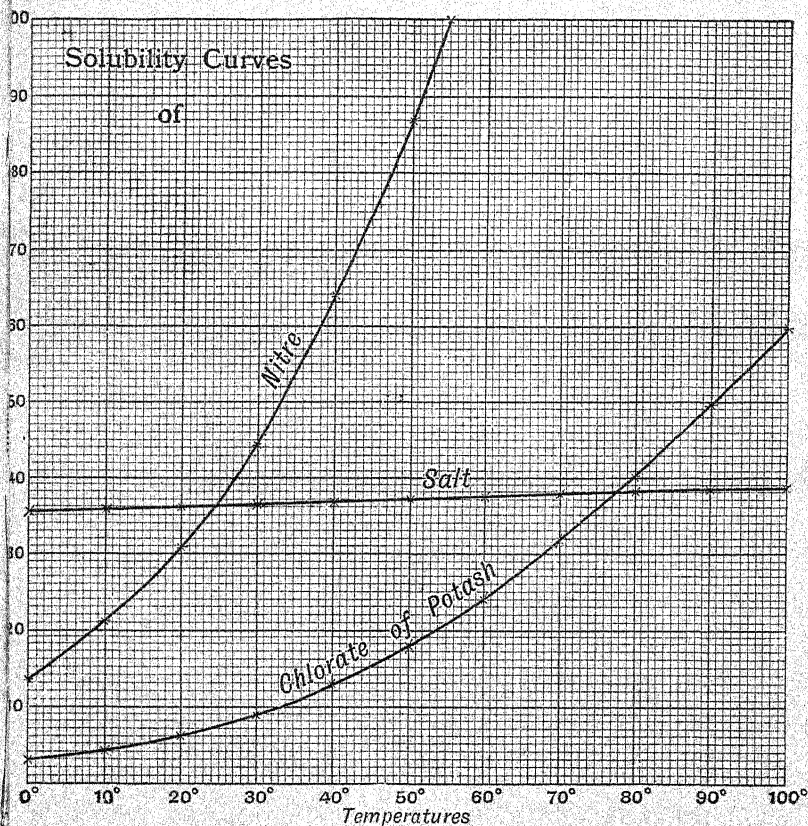


FIG. 95.

5. Which of the three substances is the most soluble (i.) at 50°, (ii.) at 10°?

6. By how many grams per degree does the solubility of chlorate of potash increase between 80° and 100°?

7. At what temperature will 100 grams of water just dissolve (i.) 25 grams of nitre, (ii.) 25 grams of potassium chlorate?
8. How much water will just dissolve (i.) 90 grams of nitre at 30° , (ii.) 20 grams of chlorate of potash at 80° ?
9. How many grams of nitre will crystallise out when a solution of nitre in 100 grams of water, saturated at 35° , cools to 10° ?
10. How could I make a solution containing 25 grams of chlorate of potash per 100 grams of water, using a thermometer but without a balance?

Volume of 1 gram of Water at Different Temperatures.—It is a fact of far-reaching importance that water has

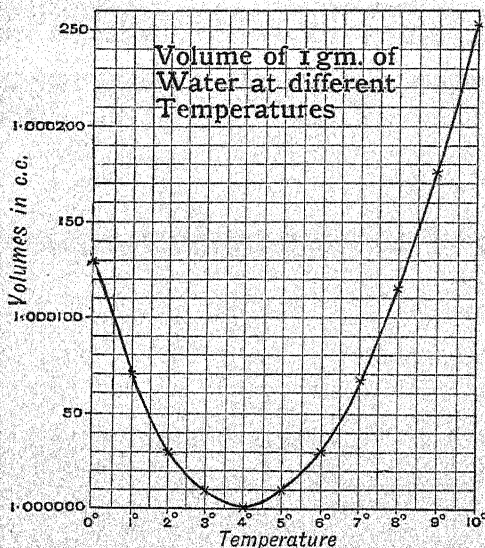


FIG. 96.

at 4°C . a greater density than at 0°C . This behaviour of water, as it approaches its freezing-point, is an exception to the almost universal fact that liquids contract on cooling. Water at 4°C . if heated expands, as we should expect. Now 1 gram is by definition the weight of 1 c.c. of water at 4°C ., and therefore

1 gram of water below or above 4° C. occupies more than 1 c.c. The following table shows the volume of 1 gram of water at various temperatures :—

0° C.	1'000129 c.c.	6° C.	1'000030 c.c.
1°	1'000072	7°	1'000067
2°	1'000031	8°	1'000114
3°	1'000009	9°	1'000176
4°	1'000000	10°	1'000253
5°	1'000010		

These volumes differ very slightly in actual magnitude, only in the last three places of decimals ; but nevertheless, on plotting them, such scales may be chosen that the diagram shall be of a reasonable size, and so that the differences between the volumes shall be evident. It will be necessary to magnify the scale representing the volumes. Use squared inch paper ; mark temperatures along a horizontal axis, and mark volumes along a vertical axis, letting each 0'1 inch represent 0'000010 c.c. Draw a curve through the points (see Fig. 96).

Times of Sunrise and Sunset.—An almanac predicts the following times of sunrise and sunset for a part of September :—

	A.			B.		
	<i>Sunrise.</i>			<i>Sunset.</i>		
Sept. 16	5.38 a.m.	6.12 p.m.
17	5.40	6.9
18	5.41	6.7
19	5.43	6.5
20	5.44	6.2
21	5.46	6.0
22	5.48	5'58
23	5.49	5'55
24	5.51	5.53
25	5.53	5.51
26	5.54	5.48
27	5.56	5.46
28	5.57	5.44
29	5.59	5.42
30	6.1	5.39

Plot the predicted times of *sunrise* on successive days upon squared paper, marking days of September, starting with the

15th along a horizontal axis, and times starting with 5.30 vertically up from it. (Express the minutes as decimals of an hour.) Join the points marked, and observe that the joining line is not a regular curve (see Fig. 97, AA). Why is this? The times tabulated are not quite accurate; they are only stated to the nearest whole minute. And, indeed, it will frequently be

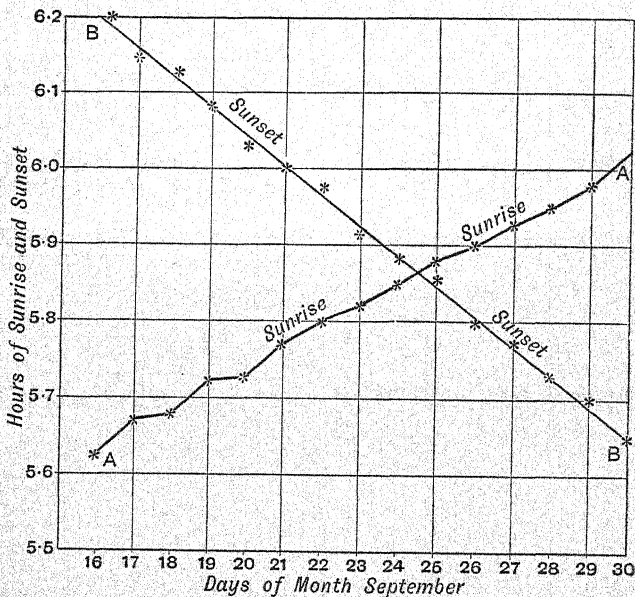


FIG. 97.

found that a series of *not quite accurate* statements will give an irregular curve such as this.

So, also, laboratory experiments involving measurements are certain to contain some errors, and if the observations are plotted, the joining line will be somewhat irregular.

In such cases, if we are confident that the points ought to be connected by a smooth line, we may *draw evenly among the points a smooth line*. Proceed now to plot the predicted times of *sunset*; observe that a straight line may be drawn passing

evenly among the points (see Fig. 97, BB), *i.e.* so that there are as many points on one side of the line as on the other. This line represents the most probable connection between the times of sunset, and even enables us to see which of the tabulated times are incorrect, either too early or too late.

Price Lists.—The price of any article frequently varies from time to time. A record of the price, which can be read at a glance, is afforded if the prices are plotted. Thus Fig. 98 is a record of the price of india-rubber from 1877 to 1898. It shows

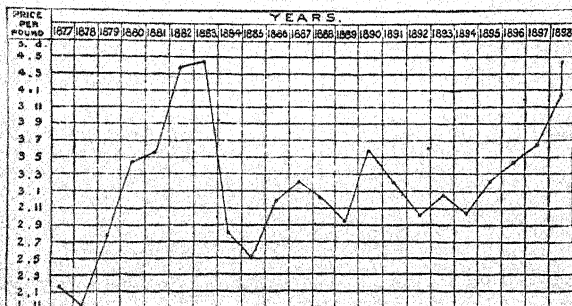


FIG. 98.¹

that in 1877 the price per lb. was about 2s. 2d., and in 1878 1s. 11d., the lowest price in the series of years. In what years was a maximum price reached? 1883 and 1898. Between what two years was there the greatest change in the price? 1883-84.

Again, the price of any one kind of manufactured article varies with its size. Suppose a manufacturer of lantern screens wishes to publish a price list of various sizes. Suppose he has made screens 6, 9, 12, and 15 feet square, and arranged their prices thus:—

a screen 6 feet square	£1 5 0 = £1.25
„ 9 „	„ „ „ „	1 17 6 = 1.875
„ 12 „	„ „ „ „	3 10 0 = 3.50
„ 15 „	„ „ „ „	6 0 0 = 6.00

¹ Quoted by Gregory and Simmons, from the *Kew Bulletin*, in *Elementary Physics and Chemistry*.

Let him plot these sizes and prices on squared paper as in Fig. 99, and draw a curve through the points.

Any point on the curve then shows the size and the probable price of a lantern screen. The manufacturer can now read off

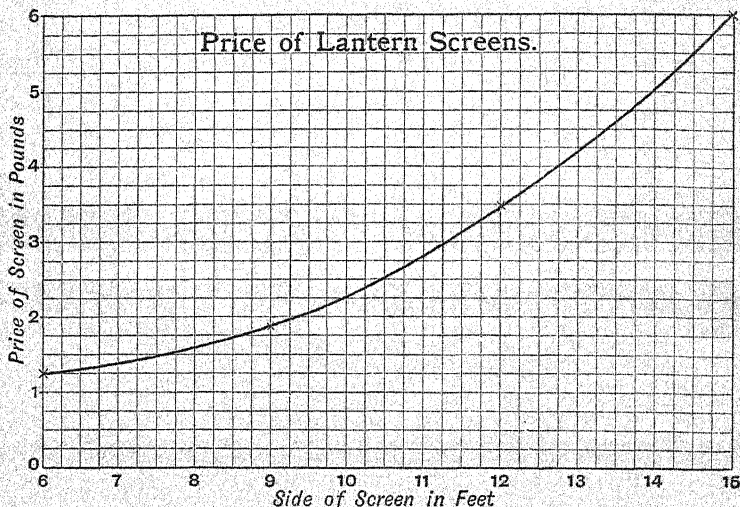


FIG. 99.

the probable prices of other sizes of screen ; for instance, he finds—

a screen	7½ feet square	should be priced	£1 9 0
„	10 „ „ „	„ „ „	2 6 0
„	14 „ „ „	„ „ „	5 0 0

This calculation of prices of screens in size lying *between* those of which the prices are known is another illustration of the use of interpolation.

Population.—The following numbers give the population of England and Wales at intervals of ten years during the 19th century :—

1801	8,893,000		1821	12,000,000
1811	10,164,000		1831	13,897,000

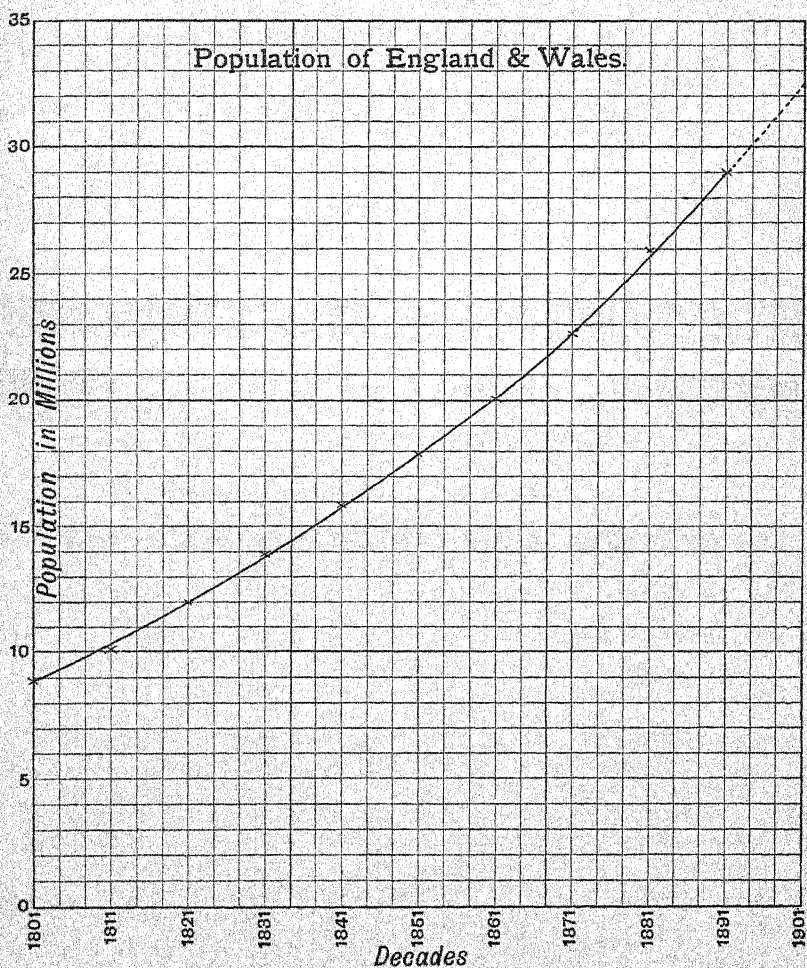


FIG. 100.

1841	15,914,000		1871	22,712,000
1851	17,928,000		1881	25,974,000
1861	20,066,000		1891	29,003,000

Plot these numbers on squared paper, and draw a curve lying evenly *among* the points marked, as in Fig. 100.

From the curve we can state what was the probable population of England and Wales in any particular year between 1801 and 1891. But also, by continuing the curve beyond the last plotted point corresponding to 1891, we can find out what it was likely to be in 1901, ten years after the last census in the Table. We see that in 1901 the population of England and Wales was likely to be about $32\frac{1}{2}$ millions.¹ Observe also that the *slope* of the curve becomes steeper and steeper as the years advance; this means, of course, that the *rate of increase* of the population becomes greater and greater.

Paper for curve plotting is manufactured in squares of $\frac{1}{10}$ inch side, with every tenth line thickened; also in square millimetres and centimetres; also in larger squares as exercise books by the educational stationers.

Paper ruled on the English scales, the square inch and its fractions, in great variety, of the best, is made by Messrs. Waterlow and Sons, Limited, 85 London Wall, E.C. Good ten-to-the-inch paper may be obtained from The Scientific Publishing Co., 53 New Bailey Street, Manchester.

Good millimetric rulings may be obtained through Messrs. Williams and Norgate, foreign booksellers, Henrietta Street, Covent Garden.

Some of the scientific instrument dealers stock useful qualities, *e.g.*, Harris of Birmingham, and Reynolds and Branson of Leeds. The paper sold by the educational stationery shops is less accurate. John Bellows, Gloucester, can do good work, to order.

Large sheets of "design paper," not quite so finely ruled, but in great variety of squares and oblongs, are used for lace designing in Nottingham, and may be obtained through stationers in that town, *e.g.*, Messrs. Mounteney.

¹ According to the census taken on April 1, 1901, the population of England and Wales is now (1901) 32,526,075.

EXERCISES : CURVE-PLOTTING

In the following exercises plot curves to represent the tabulated quantities. Those scales should be avoided which are obviously inconvenient. Scales should be so chosen that the plotted curve is not crowded in one corner of the paper. No exercise is completed until the scales and the names of the plotted quantities are clearly indicated on the paper.

1. Grams of Solid dissolved in 100 grams of water.

Temperature.	Barium Nitrate.	Potash Alum.	Sodium Sulphate.
0° C.	5.2	3.9	5.0
10°	7.0	9.5	9.0
20°	9.2	15.1	19.4
30°	11.6	22.0	40.0
40°	14.2	30.9	48.8
50°	17.1	44.1	46.7
60°	20.3	66.6	45.3
70°	23.6	90.7	44.4
80°	27.0	134.5	43.7
90°	30.6	209.3	43.1
100°	32.2	357.5	42.5

2. Pressure of Water Vapour above 100° C.

Temperature.	Pressure.	Temperature.	Pressure.
100° C. ...	1 atmosphere.	165° C. ...	7 atmospheres.
112° 2 ...	1½ "	198° 8 ...	15 "
120° 6 ...	2 "	201° 9 ...	16 "
133° 9 ...	3 "	204° 9 ...	17 "
144 ...	4 "	207° 7 ...	18 "
152° 3 ...	5 "	210° 4 ...	19 "
156° 2 ...	6 "	213° 0 ...	20 "

3. Density of Water between 0° and 20° C.

0° C. ...	0.99988	12° C. ...	0.99955
2° ...	0.99997	14° ...	0.99930
4° ...	1.00000	16° ...	0.99900
6° ...	0.99997	18° ...	0.99840
8° ...	0.99988	20° ...	0.99807
10° ...	0.99974		

4. *Price List of Bottles*, white glass, English, stoppered.

Capacity: oz.	$\frac{1}{2}$	1	2	3	4	6	8	10	12	16	20
Price per doz.	4/-	4/6	5/6	6/-	7/-	8/-	8/6	10/6	12/-	14/-	15/-

Does it appear that the price list has been correctly made out?

5. *Price List of Basins*, porcelain, Berlin, for evaporation.

Diameter: inches	$2\frac{3}{4}$	$3\frac{1}{4}$	$3\frac{1}{2}$	$3\frac{3}{4}$	4	$4\frac{1}{4}$	$4\frac{3}{4}$
Price, each...	-/4	-/6	-/8	-/10	1/-	1/2	1/6

6. *Price List of Rubber Corks*, solid.

Diam. small end, in inches	$\frac{3}{8}$	$\frac{1}{2}$	$\frac{5}{8}$	$\frac{3}{4}$	$\frac{7}{8}$	1	$1\frac{1}{8}$	$1\frac{1}{4}$	
Price per doz.	1/6	1/9	2/4	3/-	4/3	4/9	6/-	7/9
Diam. small end, in inches	$1\frac{3}{8}$	$1\frac{1}{2}$	$1\frac{5}{8}$	$1\frac{3}{4}$	2	$2\frac{1}{4}$	$2\frac{1}{2}$	$2\frac{3}{8}$	
Price per doz.	9/-	10/-	11/3	12/6	16/-	22/-	33/-	40/-

7. *Price List of Filter Papers*, Rhenish.

Diameter in cm.	4.25	5.5	7	9	11	12.5	15
Price per 1000	2/9	2/9	3/-	4/-	4/9	5/-	7/3
Diameter in cm.	18.5	24	27	32	38.5		
Price per 1000	8/9	15/9	22/-	26/10	31/6.		

8. *Population of Scotland*.

1811	1,806,000	1861	3,062,000
1821	2,091,000	1871	3,360,000
1831	2,364,000	1881	3,736,000
1841	2,620,000	1891	4,026,000
1851	2,889,000		

Estimate the population of Scotland in 1901.

9. *Premiums for Life Insurance*.—To insure the payment of £100 at death, the Government requires the following yearly premiums to be paid throughout life. (Change the shillings and pence into decimals of £1.)

Age of Insurer.	Premiums.	Age of Insurer.	Premiums.
15	1 8 6	40	2 18 0
20	1 13 0	45	3 9 6
25	1 17 6	50	4 4 0
30	2 3 0	55	5 4 0
35	2 9 6	60	6 10 6

What premium should be paid if the insurance is made at the age of 10 years?

10. *Conversion of Fahrenheit and Centigrade Temperatures.*

32° F.	=	0° C.
50°	=	10°
68°	=	20°
86°	=	30°
104°	=	40°
122°	=	50°

140° F.	=	60° C.
158°	=	70°
176°	=	80°
194°	=	90°
212°	=	100°

By interpolation find

- (a) What C. temperatures correspond to 41° F. and 77° F.
 (b) „ F. „ „ „ 15° C. „ 95° C.

11. *Logarithms of Numbers.*

Logarithm of 310	=	2.49136
„ „ 311	=	2.49276
„ „ 312	=	2.49415
„ „ 313	=	2.49554
„ „ 314	=	2.49693
„ „ 315	=	2.49831

By interpolation find the logarithms (to 4 places of decimals) of 310.4, 311.8, 312.55, 314.12, and 314.86.

12. *Logarithms of Numbers.*

Number.	Logarithm.	Number.	Logarithm.
1	0.000	6	0.778
2	0.301	7	0.845
3	0.477	8	0.903
4	0.602	9	0.954
5	0.698	10	1.000

By interpolation find the logarithms (to 2 places of decimals) of 1.5, 2.3, 5.8, 8.1, 9.26.

13. *Times for lighting Cycle Lamps.*

January 21	5.28 P.M.	July 21	9.2 P.M.
February 21	6.24	August 21	8.10
March 21	7.12	September 21	7.1
April 21	8.4	October 21	5.5
May 21	8.50	November 21	5.3
June 21	9.19	December 21	4.51

14. If a smooth curve can be drawn among a series of plotted points, it is probable that there is a relation between the two sets of quantities plotted. Find whether there is a relation between the *Distances of the Planets from the Sun, and their Times of Revolution round the Sun.*

	Distance in Millions of Miles.	Times of Revolution in Days.
Mercury	35	87
Venus	66	224
Earth	92	365
Mars	141	686
Jupiter	480	4,332
Saturn	881	10,759
Uranus	1,771	30,688
Neptune	2,775	60,181

15. Find whether there is any connection between the quantities tabulated below :—

	Specific Heat.	" Atomic Weight."
Lithium	0.94	7
Sodium	0.29	23
Aluminium	0.20	27
Potassium	0.166	39
Iron	0.112	56
Zinc	0.093	65
Silver	0.056	108
Tin	0.054	118
Iodine	0.054	127
Platinum	0.032	195
Mercury	0.032	200

16. *The Freezing-Point of Carbolic Acid.*—Fill one-quarter of a narrow test-tube with pure crystallised carbolic acid. Plunge the bulb of a thermometer into it and fix the tube nearly horizontally in a clamp. Warm the tube very gently until all the carbolic acid is melted, but do not let the temperature rise above 50° C. Then allow the tube to cool, and read the temperature accurately every half-minute, until it falls to nearly 30° C. Between the readings stir gently with the thermometer, but take care not to take the bulb out of the carbolic acid. Plot your observations, marking temperatures along the vertical axis. What temperature does the curve show to be the freezing-point of carbolic acid?¹

17. *The Freezing-Point of Sulphur.*—Proceed as in No. 16, but take readings every 15 seconds, as the temperature falls from 130° to 100° C.

18. *The Freezing-Point of Paraffin Wax.*—Proceed as in No. 16, but take readings every 15 seconds, as the temperature falls from 60° to 40° C.

Note.—Interesting additional exercises may be obtained from Whitaker's Almanac, Chambers's Mathematical Tables, the Nautical Almanac, and from the daily newspapers.

¹ The *freezing-point* of a substance is the temperature at which a substance remains as it is changing from the solid to the liquid state.

EXAMINATION PAPERS

The following papers have all been actually set in schools as terminal examinations upon the subject-matter of this volume. Not more than two hours were allowed in each case. The *average* age of the Forms examined was not less than $13\frac{1}{2}$ years nor more than 15 years.

(a)

1. What roughly represents (a) length of 1 decimetre?
(b) volume of 1 cubic centimetre?

How many grams are there in 1 ounce?

2. Add, expressing the answer in grams and decimal of a gram :—
1 Kgm. + 20 mg. + 3 C. + 5 D. + 200 mg.
3. Add, and express the sum in c.c. :—
 $\frac{1}{2}$ litre + 20 c.c. + 200 decilitres
4. A tin measures 150 mm. high, 12 cm. broad, and 24 cm. deep.
How many kilograms of water will it hold?
5. If 1 metre of wire weighs 120 centigrams, what will be the length of 60 milligrams of the same wire?
6. Write out a list of the weights in a box where the least is 0.01 gram and the greatest 100 grams.
7. Draw a portion of a burette stem including the figures 10, 11, and 12, and show the surface of water within it standing at 10.18 c.c.
8. Describe fully an experiment which shows that a liquid expands when heated.
9. Write a description of a mercurial thermometer. What numbers mark the "fixed points" on a Fahrenheit thermometer?
10. Draw the apparatus used to mark the boiling point on a thermometer.
11. Find (a) the C. reading corresponding to 122° F. and 14° F.
(b) ,, F. ,, ,, ,, 20° C. and -5° C.
12. Describe some way in which the volume of a glass stopper could be found.

(b)

13. If one end of a straw is dipped into some lemonade and you suck at the other end the lemonade ascends the straw. How do you account for this?

14. Describe in detail how to set up a mercurial barometer. What will be its usual height in inches?

Why does the height of a barometer change sometimes,

(a) in the course of a railway journey?

(b) at the same place in the course of a few hours?

15. What differences are there between the two flames which can be obtained with a Bunsen burner? State exactly what kind of flame and what part of the flame you would use

(a) to bend a glass tube,

(b) to smooth the sharp edge of a glass tube.

16. A beaker is full of powdered ice and a thermometer is plunged in it. State fully

(a) what changes you will see,

(b) what temperatures the thermometer will show if a flame is left under the beaker until all the ice is melted, and the water formed finally boils.

17. Explain what is meant by the statement that when certain substances are heated "physical changes" may take place. Give one example.

18. Draw a Liebig condenser in section.

(c)

19. Find (a) the F. temperature corresponding to 18° C.

(b) " C. " " " " 23° F.

20. How have you proved that there is a vacuum above the mercury in a barometer tube?

Why does the mercury stand so high? Why does it not reach up to the top of the tube?

21. Describe what occurs on heating

(a) tin in an iron spoon,

(b) mercury in a glass test-tube,

(c) olive oil in a flask.

22. Explain fully why the tumblers of water on the dinner table sometimes become bedimmed with dew.

23. Define or explain the terms—

(a) the "melting-point" of a solid,

(b) the "solvent action" of a liquid,

(c) the "latent heat of vaporisation" of a liquid.

24. Draw carefully a wash-bottle.

25. What do you understand by the terms "hard" and "soft" as applied to water? What would be the effect on the "hardness" of a "hard" water of—

- (a) shaking it with washing soda,
- (b) shaking it with gypsum,
- (c) boiling it,
- (d) distilling it?

26. 10 c.c. of a solution of salt were found to weigh 11.86 grams, and to leave 3.05 grams residue after evaporation. Calculate the amount of salt dissolved in 100 grams of water.

(d)

27. Define the "relative density" of a substance?

You are provided with an irregular lump of granite. Describe in detail two methods by which you could find its relative density. Which will give the more accurate result?

28. A glass stopper weighs 60 grams in air, 35 grams in water, and 40 grams in spirits. Find the relative density of the spirits.

29. A clear liquid is said to be pure water. Describe all the experiments you would make in order to prove whether it is so or not.

30. You are provided with some surface soil, which contains a small quantity of nitre. Describe in detail how you would obtain from it pure crystals of nitre.

31. What is "water of crystallisation"? Instance two kinds of crystals which contain it, and two kinds which do not.

32. Calculate the solubility of potassium chlorate at the ordinary temperature if 16.14 grams of solution on evaporation are found to contain 1.01 gram of the substance.

(e)

33. Plot the following barometric heights in pencil on squared paper:—

June 1	30.16 ins.	June 9	30.35 ins.
„ 2	29.94 „	„ 10	30.25 „
„ 3	30.24 „	„ 11	30.25 „
„ 4	30.14 „	„ 12	30.11 „
„ 5	30.18 „	„ 13	30.13 „
„ 6	30.23 „	„ 14	29.99 „
„ 7	30.30 „	„ 15	30.05 „
„ 8	30.35 „		

34. Explain clearly how it is that (a) *sunshine*, and (b) *wind* increase the rate at which a wet road dries.

35. When a long frost is followed by a few warm days the walls inside a house are often found dripping with moisture. Account for this.

36. You are provided with some muddy water. Describe in detail how you would prepare some pure water from it.

37. Define a *calorie*, and the *specific heat* of a metal.

Calculate the specific heat of a metal if :—

Weight of metal	25 gm.
Volume of water	50 c.c.
Temperature of hot metal	100° C.
„ cold water	11° C.
„ water after experiment	15° C.

38. Draw carefully the apparatus which was used to measure the latent heat of vaporisation of water.

39. If the L. H. of liquefaction of ice is 80, and the L. H. of vaporisation of water is 540, calculate the amount of heat required to change 1 gram of ice at 0° C. into steam at 100° C.

(f)

40. Four experiments were made to find the percentage of water in a crystal :—

(a)	2.64	grams lost	0.65	when heated
(b)	2.97	„ „	0.75	„ „
(c)	2.33	„ „	0.58	„ „
(d)	2.06	„ „	0.52	„ „

Calculate in each case the percentage loss in weight, and also the *average* result.

41. If you were given a muddy liquid containing a soluble and an insoluble constituent, describe in detail the steps you would take :—

- (a) to separate the *whole* of the soluble portion,
 (b) to obtain a specimen of the soluble portion
 in *crystals*, if possible.

42. Describe how you would distinguish between :—

saltpetre and sal-ammonia,
 caustic soda and washing soda,
 sulphuric acid and nitric acid.

43. Draw the apparatus required for the preparation of hydrochloric acid.

Name the chemicals used.

44. What do you call the substance which can be made from hydrochloric acid and caustic soda?

If the acid was in strong solution, and the soda in solid sticks, explain in detail how you would prepare a sample of the pure compound free from any admixture of the original substances.

45. What are some of the properties of sulphuric acid? What is the effect of warming it with some nitre crystals?

PRACTICAL EXAMINATIONS

The following practical tests have all been terminal examinations which have been actually set in schools. The average age of the Forms examined was not less than $13\frac{1}{2}$ years nor more than 15 years.

(a) 1 HOUR

1. Make a list of, and add up the weights on the pan of the balance.
2. Read the level of water in the burette.
3. Read the temperature of the running water.
4. Measure the distance between the 1st and 10th lines on your paper
(a) in inches and decimal,
(b) in centimetres and decimal.
5. Add water to the measuring jar so that it contains exactly 86 c.c.
6. Weigh the glass stopper.
7. Weigh the beaker: then measure into it 5 times by means of the pipette 100 grains of water. What does the water weigh? Calculate what 1 grain of water weighs.

(b) 1 HOUR

8. Read (a) the level of the liquid in the burette,
(b) the volume of water in the measuring jar.
9. Select a cork to fit the test-tube. Fit it with a glass tube bent twice at right angles.
10. You are provided with a substance A.¹
 - (1) Describe its appearance.
 - (2) Describe the effect of heating a small quantity in a dry test-tube.
 - (3) Find whether it is soluble in water. State how you proceed.
 - (4) Find the effect on the temperature of 100 c.c. of water of adding 30 grains of A to it, and stirring.

¹ AmCl.

(c) 1½ HOURS

11. Find the relative density of the glass stopper.
12. Describe the appearance of B.¹
Heat some of B in a dry test-tube. Describe in detail what you observe.
13. Find the percentage loss in weight when about 2 grams of B are heated in a crucible upon a sand-bath.
14. Find the action of water, both hot and cold, upon B.
Prepare from B as large crystals as the time permits—both clean and dry. Describe your method.

(d) 1½ HOURS

15. Heat C² in a dry test-tube. Describe what occurs.
16. Identify the two solids D and E, and the two liquids G and H.
17. You will be provided with a stick of solid caustic potash. Prepare a good specimen of the salt which can be obtained from it and sulphuric acid.

(e) 1½ HOURS

18. You are provided with a solution of a certain substance³ :—
- (a) Find the relative density of the solution.
- (b) Prepare by crystallisation some of the dissolved substance.
Leave a specimen for inspection.
- (c) Describe the appearance of the crystals obtained.
- (d) Heat some of the crystals in a dry test-tube, and describe in detail what you observe.
- (e) Find the percentage loss in weight when about 2 grams of the crystals are heated in a crucible upon a sand-bath.

(f) 1½ HOURS.

19. Find the percentage loss in weight when about 2 grams of K⁴ are heated in a crucible.
20. What is the action of water upon K? State what experiments you try.
21. Separate from K a sample of each of the two substances which it contains. Describe your method.

¹ Powdered ZnSO₄ · 7H₂O.² HgI₂.³ ZnSO₄ · 7H₂O.⁴ 50 per cent. SiO₂ + 50 per cent. CuSO₄ · 5H₂O.

ANSWERS

EXAMPLES I.

Metric Prefixes : Length. (Page 11.)

- | | |
|------------------|------------------|
| 1. 50 cm. | 10. 12,345·6 cm. |
| 2. 100 mm. | 11. 40. |
| 3. 10,000 m. | 12. 5000. |
| 4. 1,000,000 mm. | 13. 4 cm. |
| 5. 0·1 dm. | 14. 2 Km. |
| 6. 0·000001 Km. | 15. 120 cm. |
| 7. 4·56 dm. | 16. 4. |
| 8. 234,000 cm. | 17. 40,000. |
| 9. 0·9 cm. | 18. 4. |

EXAMPLES II.

Metric Square Measure : Area. (Page 14.)

- | | |
|-------------------|--------------------------------|
| 1. 100 sq. dm. | 6. 100 sq. cm. |
| 2. 10,000 sq. cm. | 7. 64. |
| 3. 60 sq. cm. | 8. 25,000. |
| 4. 24 sq. cm. | 9. 196 $\frac{2}{3}$ cwt. |
| 5. 625 sq. mm. | 10. 1 $\frac{19}{81}$ sq. yds. |

EXAMPLES III.

Metric Cubic and Capacity Measure. (Page 18.)

- | | |
|-------------------|---------------------------|
| 1. 1000 c.c. | 9. 56,780 c.c. |
| 2. 27,000 c.c. | 10. 1200 c.c. |
| 3. 15,625 cu. mm. | 11. 600 cu. mm. |
| 4. 1000 cu. mm. | 12. 4320 c.c. |
| 5. 1,000,000 c.c. | 13. 2 litres. |
| 6. 3000 c.c. | 14. 400. |
| 7. 200 c.c. | 15. 324,000 litres. |
| 8. 1·234 litre. | 16. 8 $\frac{1}{3}$ days. |

EXAMPLES IV.

Metric System: Weight. (Page 19.)

- | | |
|---------------|----------------|
| 1. 10 dgm. | 11. 25.202 gm. |
| 2. 10 mgm. | 12. 1023.4 gm. |
| 3. 7.891 Kgm. | 13. 25 c.c. |
| 4. 0.2 gm. | 14. 27 gm. |
| 5. 0.05 gm. | 15. 1150 gm. |
| 6. 6540 gm. | 16. 8000 gm. |
| 7. 0.1 gm. | 17. 1350 gm. |
| 8. 34 mgm. | 18. 55 c.c. |
| 9. 250 mgm. | 19. 5 gm. |
| 10. 3.25 gm. | 20. 1 gm. |

EXAMPLES V.

Metric System: Miscellaneous. (Page 21.)

- | | |
|------------------------|------------------------|
| (a) | (c) |
| 1. 20. | 13. 1920 c.c. |
| 2. 50. | 14. 8075. |
| 3. 20 mgm. | 15. 6 gm. |
| 4. 1370 gm. | 16. $\frac{1}{2}$ 125. |
| 5. 2777.7 cm. per sec. | 17. 5 m. |
| 6. 0.01 gm. | 18. 23.7 Km. |
| (b) | (d) |
| 7. 1450.632 m. | 19. 2400.71 gm. |
| 8. 100 sq. cm. | 20. 100 gm. |
| 9. 4.32 cu. dm. | 21. 20 times. |
| 10. 0.4 gm. | 22. 50 gm. |
| 11. 0.5 gm. | 23. 1370 gm. |
| 12. 30.48 cm. | 24. 37.037 ounces. |

EXAMPLES VI.

Burette Reading. (Page 35.)

- | | |
|-----------|-------------|
| 1. A. 27. | 2. I. 1.68. |
| B. 43. | K. 13.32. |
| C. 6.7. | L. 10.68. |
| D. 830. | M. 23.2. |
| E. 114. | O. 46.8. |
| F. 136. | P. 632. |
| G. 235. | Q. 868. |
| H. 265. | R. 680. |

EXAMPLES VII.

Interpolation: Conversion of Thermometric Scales. (Page 46.)

1. 63.	10. 50° C.
2. 1°31.	11. 68° F.
3. 5°58 $\frac{1}{2}$ p.m.	12. 23° F.
4. 1°361.	13. 40°·1 F.
5. 77° F.	14. 4°·4 C.
6. 41° F.	15. - 10° C.
7. 176° F.	16. - 20° C.
8. 10° C.	17. 12° R.
9. 35° C.	18. 68° F.

EXAMPLES VIII.

Relative Densities of Liquids. (Page 53.)

(a)

1. 147 gm.	6. ·074 c.c.
2. 800 gm.	7. 1111 $\frac{1}{4}$ c.c.
3. 0·862.	8. 1·25 c.c.
4. 1·023.	9. 50 c.c.
5. Water.	10. 25 c.c.

(b)

11. 13°5; ·07407.	18. 150 c.c.
12. 38°09 c.c.	19. 250 c.c.
13. 70°85 gm.	20. ·923.
14. 36°2 gm.	21. 2800 gm.
15. ·940.	22. ·98.
16. 1·026.	23. 20 sq. mm.
17. 60 c.c.	

EXAMPLES IX.

Relative Densities of Solids. (Page 57.)

1. (1) 8·95.	5. 8.
(2) 7·2.	6. 22·8 gm.
(3) 0·25.	7. 1791 $\frac{80}{4}$.
(4) 11·4.	8. 20 oz. 11 $\frac{1}{4}$ z.
2. 2·7.	9. 50 oz.
3. 0·6.	10. 0·001 c.c.
4. 2·5; 2·14; 10.	

EXAMPLES X.

Relative Densities of Solids. (Page 59.)

- | | |
|---------|------------|
| 1. 2.5. | 3. 11.4. |
| 2. 7.5. | 4. 145 gm. |

EXAMPLES XI.

Relative Densities of Solids. (Page 63.)

- | | |
|------------------|-------------|
| 1. 17.5. | 6. 37.6 gm. |
| 2. 2.72. | 7. 7.7 lbs. |
| 3. 2.00. | 8. 2.4. |
| 4. 2.62. | 9. 0.8. |
| 5. Copper: 8.91. | 10. 8.6. |

EXAMPLES XII.

(a) Specific Heat. (Page 117.)

- | | |
|---------------------|-----------------|
| 1. i. 33° 5 C. | 4. 0.034 + |
| ii. 32° 41 C. | 5. 0.5. |
| iii. 25° C. | 6. 3° 75 C. |
| iv. 56° 8 C. | 7. 0° 94 + C. |
| 2. i. 475 calories. | 8. 51° 2 + C. |
| ii. 403.75 " | 9. 2° 9 + C. |
| iii. 60 " | 10. 154.3 + gm. |
| iv. 8800 " | 11. 333.3 gm. |
| 3. 0.057 +. | |

(b) Latent Heat. (Page 118.)

- | | |
|----------------------|-------------------|
| 12. 80,000 calories. | 15. 186.2 gm. |
| 13. i. 78.4. | 16. i. 520. |
| ii. 78.5. | ii. 515. |
| iii. 80.6. | 17. 717 calories. |
| 14. 65° 4 C. | |

EXAMPLES XIII.

Solubility Curves. (Page 180.)

- | | |
|---------------------|-----------------------|
| 1. 16, 54, 74 gm. | 6. 1 gram per degree. |
| 2. 16°, 52°, 95° C. | 7. (i.) 14° 5 C. |
| 3. (i.) 19 gm. | (ii.) 61° C. |
| (ii.) 19 gm. | 8. (i.) 200 gm. |
| 4. 77° C. | (ii.) 50 gm. |
| 5. (i.) Nitre. | 9. 33 gm. |
| (ii.) Salt. | |

EXAMINATION PAPERS. (Page 194.)

2.	1000.75 gm.	28.	2.5.
3.	20,520 c.c.	32.	6.6.
4.	4.32 Kgm.	37.	0.094.
5.	5 cm.	39.	720 calories.
11.	(a.) 50° C. and - 10° C.	40.	(a.) 24.62 per cent.
	(b.) 68° F. and 23° F.		(b.) 25.25 "
19.	(a.) 64.8 F.		(c.) 24.89 "
	(b.) - 5° C.		(d.) 25.24 "
			Average = 25.00 per cent.

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