HIGH SCHOOL CHEMISTRY.

A. P. KNIGHT.

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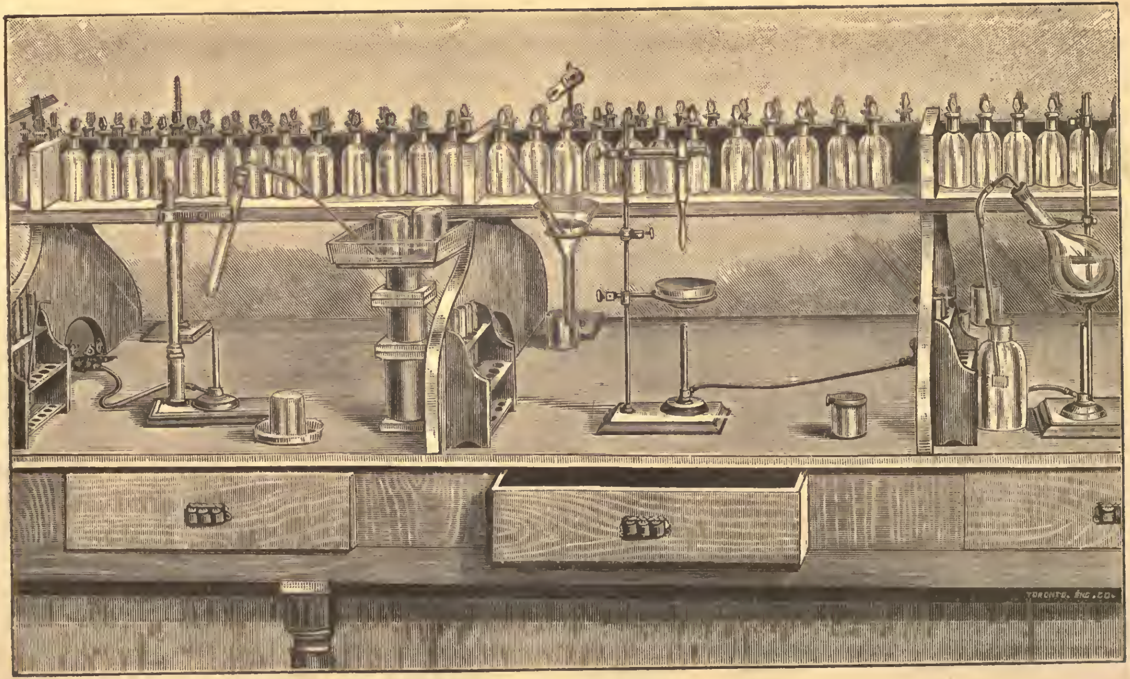
TORONTO ENG. CO.

HIGH SCHOOL CHEMISTRY,

CONSISTING OF DIRECTIONS FOR PERFORMING

A SERIES OF EXPERIMENTS,

WITH TEST QUESTIONS ON THE EXPERIMENTS,
AND SIMPLE

PROBLEMS FOR INVESTIGATION.

BY
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Authorized by the Department of Education for Ontario.

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1887.

TO

M. R. M

AND TO

N. F. DUPUIS, M. A., F. B. S., EDIN.
Professor of Mathematics, Queen's College, Kingston, Canada,

I DEDICATE THIS BOOK,

In grateful acknowledgment that what is of most worth
in my life and work has come to me
through theirs.

INTRODUCTION.

A knowledge of chemistry is acquired—

1. By experiment;

2. By accurate observation of the phenomena revealed by experiment, and by systematically recording these observations as they occur;

3. By reasoning based upon our observations;

4. By investigation; that is, by originating new processes or employing those with which the student is already familiar, and applying them to the solution of new problems;

5. By a knowledge of chemical theory, which is itself based upon a knowledge of the facts of chemistry, and becomes in its turn an aid in making further progress in the science.

All these means of acquiring a knowledge of chemistry have been steadily kept in view in the preparation of this manual. I have long held the opinion that a text-book for beginners in chemistry should consist mainly of directions for performing a series of experiments. These experiments should be so graded and arranged that all the prominent facts and principles of the science could be re-discovered, as it were, by any intelligent pupil who performed the experiments and reasoned upon the phenomena they presented.

How far the teachers of chemistry have departed from these principles in the preparation of text-books can be seen by opening the first book on the subject which you may chance to lay your hand on in any book store. Here is an extract from one of these books by a very distinguished chemist :—

> "EXPERIMENT.—Let us burn our taper in a clean glass bottle with a narrow neck ; after it has burnt for a few minutes we notice that the flame grows less and less, and in a short time the taper goes out. This is the first thing we have to observe. We next have to discover why the taper goes out. For this purpose let us see whether the air in the bottle is now the same as it was before the candle was burnt. How can we tell this ? Let us pour some clear lime-water first into a bottle filled with air in which no candle has burnt, and then into the one in which our taper burnt. You see the difference at once ! In the first bottle the lime-water remains clear, in the second it becomes milky. Hence we see that the air has been changed in some way by the burning of the taper. This milkiness is nothing else than chalk, and chalk is made up of lime and carbonic acid. Carbonic acid is, like common air, a colourless invisible gas which we cannot see, but which we find turns the lime-water milky, and puts out a burning taper. Part of the wax has been changed by burning into this carbonic acid gas ; that is, the carbon or charcoal of the burnt wax is to be found again in this invisible gas. Some of this carbon you may notice going away unburnt as smoke or soot ; and if you quickly press a sheet of white paper on to the flame so as not to burn the paper, you will see that it becomes stained with a black ring of soot or carbon."

Now, the observations, *(a)* that the flame grew less and went out, *(b)* that the lime-water became milky, *(c)* that the carbon of the wax passed off unburnt as soot, *(d)* that in doing so, it blackened a sheet of white paper, should all be made by the pupil, without assistance from either text-book or teacher. If this be not done, the pupil becomes "the mere recipient of another's observations." In the same way the conclusions, *(a)* that the air was changed in some way by the burning candle, and *(b)* that the something produced by the burning candle "is, like common air, a colorless, invisible gas," should also be reached by the pupil without assistance of any kind. Only when every pupil in a class fails to make an important observation, or to draw a legitimate conclusion, should the teacher step in and give his class assistance. This remark, of course, does not apply to much valuable information, the result of long and difficult experiments by great chemists. It applies only to the numerous experiments which may be easily performed in any fairly well equipped school laboratory.

Any assistance which the pupil requires can easily be conveyed by a series of well-considered questions on each experiment performed. This is the object in the large number of questions found throughout this book. Observations and conclusions have been designedly omitted. These the teacher will readily find in the thousand and one works on chemistry which have come from the printing press during the past few years. Here and there I have inserted a few simple problems, which are intended to stimulate in the pupil a desire for original research, as well as to test his power of applying the knowledge he has already acquired to the solution of new problems. I agree with the Editor of the *Chemical News* that even beginners should aim at doing original work.

"Once more the laboratories and lecture-rooms of our colleges are thrown open, and once more professors, demonstrators, and students are preparing to resume their duties. Once more, also, in accordance with our custom, we take it upon ourselves to address a few hints to those who are beginning or continuing the study of chemistry. *We have always advised the student to qualify himself for original research,* believing that such will prove the best course of training, whether for pure chemistry or for its industrial application."—*Chem. News, Sept. 18th, 1885.*

The book is, however, purely a pupil's text-book. Its use in any school implies that the teacher possesses a thorough knowledge of elementary chemistry. It implies, also, that the pupils have access to the best works of reference on the subject, so that in case of doubt they may appeal to the highest authority.

The book is intended to cover two years of school work. In the first half, the aim has been to furnish young students with a stock of facts upon which to found, subsequently, a fuller knowledge of the principles of the science. It is scarcely necessary to say that the calculation of chemical constants is very difficult work for the junior pupils of our secondary schools. Such pupils should not, in my opinion be allowed to read beyond Chapter XXVIII. during their first year in chemistry; and they might perhaps with profit omit Chapters IX. and XIV. Beginning at Chapter XXIX., the plan of the book was somewhat changed. In that and in a few succeeding chapters an attempt was made to acquaint senior students with the principal theories of chemistry—the knowledge previously acquired being used as a foundation. Then follows a brief study of some of the more important metals, theory and practice being kept prominent, but questions regarding mere observation of facts being almost entirely omitted. A brief outline of Mendelejeff's classification of the elements comes in naturally towards the end of the book, after the student has acquired some knowledge of the principal elements and their characteristic compounds

A few words may not be out of place as to my own practice in teaching chemistry. All the simpler experiments—probably two-thirds in number—are performed by the pupils, each working at a separate table. The more difficult ones are made by myself in presence of the class. In both cases the pupils are required to record neatly, in note books, a description of the operations and of the phenomena revealed; and, lastly, a conclusion based upon the facts of the experiment, or of it and others which preceded it. A little more than half of the time devoted to the subject is allotted to experimentation and to recording results; the rest of the time is spent in systematic drill in class, in the shape of oral questions on the observations made and the conclusions reached.

Prof. Huxley, the late Prof. Miller, and other great teachers, all advocate the method of teaching science that has been adopted in this book. Not only so, but the same method is advocated by writers who are considered high authorities on the principles and practice of teaching.

> "It is becoming more generally accepted every day by good teachers not only of chemistry, but of physics, that the best teaching is given in the laboratory rather than in the lecture room. It is not merely by seeing experiments tried, but by trying them, that the properties of objects, their structures and organization are best learned. But here it must be borne in mind that the discipline you want to give must be definite and exact; it is not seeing and handling only, but careful measurement if it be mechanics, careful observation if it be botany or physiology, and whatever it be, careful notes and recordation of the results of each experiment as it is made."—*Fitch's Lectures on Teaching.*

> "In manhood, when there are no longer teachers at hand, the observations and inferences required for daily guidance, must be made unhelped; and success in life depends upon the accuracy and completeness with which they are made . . . To *tell* a pupil this and to *show* him the other, is not to teach him how to observe, but to make him a mere recipient of another's observations: a proceeding which weakens rather than strengthens his powers of self-instruction, and deprives him of the pleasures resulting from successful activity."—*Education, by Herbert Spencer.*

My own teaching for the past ten years has been along the lines indicated, but it was not until October, 1885, that I found time to enlarge my lesson notes and place them in the hands of the Copp, Clark Company for publication. At that time I intended that the book should be through the press before the beginning of the New Year (1886), but for one reason or another its issue has been delayed until now.

The publishers are indebted to Messrs. J. & H. Berge, dealers in chemical apparatus, 95 John St., New York, for the use of a few of the illustrations used in this book.

I have to express my indebtedness to Professor Goodwin, of Queen's College, and to Professor Waddell, of the Royal Military College, for valuable suggestions in preparing this manual, and for kindness in reading most of the proof-sheets; also, to Mr. W. Lochhead, B.A., of Cornell University, for assistance in preparing my manuscripts for the press. This gentleman is the joint author of the articles on sodium, iron and lead; and the sole author of the chapter on the elements of the silver family.

The works of the following named authors have been freely consulted in the preparation of this hand-book: Roscoe and Schorlemmer, Remsen. Muir, Wurtz and Reynolds.

KINGSTON, 1887.

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PRACTICAL CHEMISTRY.

TO THE STUDENT.

Before beginning your studies, it becomes necessary to give a few general directions regarding chemical manipulations. The following will be found useful at every stage of your studies :—

1. Work carefully and note all phenomena that occur, and record them in your note-book.
2. When heating a liquid in a test-tube take care that the flame does not strike the tube at the upper end of the liquid.
3. Never hold the mouth of a test-tube, while heating it, towards yourself, or towards another.
4. Never heat beakers, evaporating dishes or flasks without placing a piece of wire gauze or a sand-bath between them and the flame.
5. Use sulphur matches ; parlor ones are dangerous.
6. Do not lay down the cork of a reagent bottle while pouring out a solution, but hold it between the first and second fingers.
7. Have a place for every piece of apparatus and every reagent you use, and when you are done using them, clean the apparatus and put it and the reagents back in their respective places.

PRACTICAL CHEMISTRY.

CHAPTER I.

§ 1.—Solution.

Let us begin our study of chemistry by some experiments.

Experiments.

1. Fill a small beaker half-full of water, and then suspend in the water by a thread a piece of white sugar, or a crystal of sulphate of copper.

2. Place about a tablespoonful of water in a small evaporating dish, stir into it a small quantity of salt or alum, and after it has entirely disappeared, heat the solution over a spirit lamp, as in Fig. 1, until all the water is driven off.

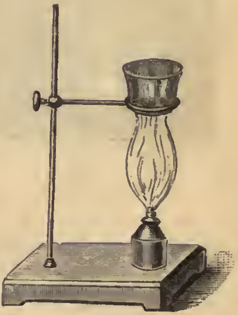


FIG. 1.

3. Repeat the preceding experiment, using weighed quantities of salt and of water, and then weigh the solution.

FIG. 2.

4. Take an ounce each of powdered alum, washing soda, and copper sulphate, and dissolve each in a fluid ounce of water in three separate beakers. Stir with a glass rod, and observe the extent

to which each one dissolves. Now heat the solutions until they boil, stir again, and notice whether there is any variation in solubility. Pour each hot solution into a separate plate, and allow them to cool slowly. After some time, examine carefully and compare the geometrical forms which the solids have taken with those in Fig. 2.

5. Ascertain whether calcium chloride is soluble in alcohol; and sulphur, in carbon disulphide.

In order to test the accuracy of the observations which you have made in performing the preceding experiments, try to answer the following

FIG. 2.

QUESTIONS.

1. Where did the solids go to?

2. Devise experiments to find out how a solid may be got most quickly into solution.

3. How may a solid be obtained from the liquid in which it is dissolved?

4. How does the weight of a solution compare with the weight of the solid and liquid which form it?

5. Mix thoroughly some powdered salt and sulphate of copper together, and then devise means of separating them, basing your device upon the preceding experiments.

6. Ascertain, by evaporation, whether there is any solid matter dissolved in a sample of river or spring water.

7. What influence has temperature on the quantity of a solid held in solution?

8. Find out a substance insoluble in water, but soluble in carbon disulphide.

9. How can salt be obtained from sea-water. Define solution.

Saturation.—When a liquid has dissolved all that it can of a given solid, at a given temperature, the liquid is said to be saturated.

---

§ 2.—Filtration.

Experiments.

1. Half-fill a beaker with water, and into it drop a piece of quick-lime the size of a bean. Stir the mixture, and then filter it. This is done by folding a circular piece of filtering paper—white blotting paper will do well enough—across twice, so as to form a quadrant, and fitting it to a funnel, as in Fig. 3. A glass stirring rod should be held vertically above the funnel, and the solution poured down the rod. Catch the liquid that passes through the filtering paper, and taste it. Evaporate some of it upon a piece of platinum foil, or upon a sheet of mica.

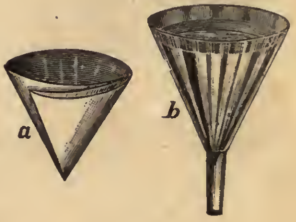


FIG 3.

2. Ascertain whether salt or sugar dissolved in water can be removed bv filtration through paper.

QUESTIONS.

1. What kind of matter is usually removed from a liquid, by filtration? What, by evaporation.

2. Did the quick-lime dissolve in the water, or remain in suspension, or did it do both? Give reasons for your answer.

---

§ 3.—Solution of Liquids.

Experiments.

1. Fill a small test-tube to the depth of 4 centimetres with water, and then add chloroform, or sulphuric ether, to the depth of 1 centimetre. Shake well, and allow

the mixture to stand for a few minutes, after which examine it carefully.

2. Repeat experiment 1, using water and coal oil, or water and mercury.

3. Mix, in a test-tube, 1 part of sulphuric acid with 20 parts of water. Taste the mixture, and say whether the acid has dissolved in the water.

4. Repeat experiment 2, using equal parts of common alcohol and water.

QUESTIONS.

1. Record your observations on the four preceding experiments in a tabulated form, stating clearly the cases in which the liquids appeared to dissolve each other.

2. Write out a conclusion, basing it upon the facts you have observed.

3. How would you define the solution of a liquid in a liquid?

---

§ 4.—Gases in Solution.

Experiments.

1. Fit a Florence flask, capable of holding about a litre, with cork and delivery tube, as in Fig. 4. Fill both flask and tube with spring water, and then gradually heat the flask until the water boils briskly. Collect the gas that comes off, in a tube filled with water, and placed over the end of the delivery tube.

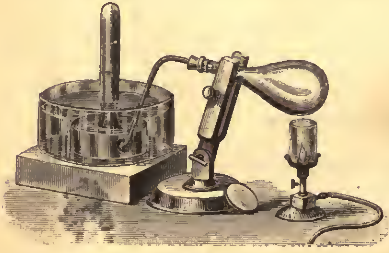


Fig. 4.

2. Pour a little soda-water or ginger-ale into a test-tube and then warm gently over a spirit lamp.

3. Repeat experiment 1, using the same liquid as in experiment 2.

QUESTIONS.

1. Where must the gases have come from which collected in the graduated tube?

2. What causes the effervescence in soda-water and other foaming liquors?

3. What effect has increase of temperature upon the quantity of gas that may be dissolved in water?

---

§ 5.—Theory of Solution.

Solution is explained on the theory that all substances are built up of very minute and indivisible particles, called molecules. When a solid, therefore, goes into solution, it is supposed that its molecules diffuse themselves among those of the liquid, and thus become invisible. This is mere theory however, and the student must always discriminate carefully between theories and facts.

PROBLEMS.

1. Find out whether sand and commercial iodine are separately soluble in common alcohol.

2. Ascertain whether a mixture of sand and ammonium carbonate is volatile when heated.

3. Devise plans for separating mixtures of *(a)* potassic chloride and manganese dioxide, *(b)* sand and ammonium carbonate, *(c)* sand and sugar.

4. Gunpowder consists of charcoal, sulphur, and nitre. Separate these substances in a given sample of gunpowder, using water and carbon disulphide as solvents.

CHAPTER II.

6.—Physical Change.

Fig. 5.

Experiments.

1. Heat a small piece of platinum wire in the flame of a spirit lamp or gas burner.

2. Fill a test-tube full of water, and invert it over a tumbler or beaker half-full of water, as in Fig. 5. Heat the upper end with a spirit lamp until steam forms, and then allow the test-tube to cool.

QUESTIONS.

1. What color did the platinum wire take while in the lamp flame?

2. On removing the platinum wire from the flame, had it undergone any change?

3. What changes did you observe during the gradual heating of the water? What caused the change? What caused the steam to go back to the form of water?

4. Was the water after cooling different in any way from what it was before being heated?

5. Have the weights of the substances altered?

6. The changes undergone by the water and platinum while being heated as above, are called physical. Give a definition of a physical change, and mention other examples of it which you have seen in preceding experiments.

---

§ 7.—Chemical Changes.

Experiments.

1. Take a piece of magnesium wire six inches long, weigh it accurately, place it in a crucible loosely covered,

and heat in a lamp flame, until a visible change is effected. Weigh the product.

2. Heat a piece of bright copper wire in a lamp flame.

3. Pulverize a crystal of blue vitriol and heat it upon a piece of mica. When cold put a drop or two of water on it.

4. Heat a small crystal of alum on a piece of platinum foil or mica, until ebullition ceases. Weigh before and after heating.

QUESTIONS.

1. What was the color of the light emitted by the magnesium when burning?

2. Did the product of its combustion weigh more or less than the original piece of magnesium? If more, where did the additional matter come from and what force caused its union with the magnesium?

3. Was the change in the magnesium a temporary or a permanent one?

4. What change in color and character did the vitriol undergo on being heated?

5. What color did the polished copper assume while hot? After cooling? Was the change temporary or permanent?

6. Did the alum after cooling weigh more or less than before it was heated? If less, where did the lost matter go? If more, where did the additional matter come from?

7. What opposite changes did heat produce in the alum, and in the magnesium?

8. What force caused the copper wire to change its appearance?

9. What force caused the changes in the alum?

10. The changes undergone by the substances in the above experiments are called chemical ones. Define a chemical change.

Chemical changes may be caused by any one of the following *agents* or *forces*:—*Heat*, *Light*, *Electricity*, *Mechanical force*, *Chemism*, and the *Vital force* of animal or vegetable organisms. Our next experiments will furnish instances of chemical changes produced by these forces.

§ 8.—Mechanical Force.

Experiments.

1. Take two or three *small* crystals of chlorate of potash, and half the quantity of sulphur. Grind the mixture, at first lightly, in a mortar, and afterwards with considerable force.

[This experiment is dangerous unless carefully performed, and upon small quantities.]

2. Mix the ingredients lightly as before, but instead of continuing to rub them forcibly in the mortar, heat the mixture on a tin plate, or a piece of mica.

QUESTIONS.

1. What phenomenon occurred on forcibly rubbing chlorate of potash and sulphur together? What caused it?

2. How did it differ from that which took place on heating the mixture? Name the force which produced the chemical change in the second experiment.

---

§ 9.—Light.

Experiments.

1. Dissolve a few crystals of nitrate of silver in water. Then dip a piece of common white paper into the solution, and afterward into a similar solution of common salt, and dry in the dark. Tear the paper into two pieces, keeping one in a dark room, and exposing the other to sunlight.

2. Dip a piece of filtering paper into a saturated solution of bichromate of potash, and dry in the dark. Then lay a leaf on it, press the leaf under a plate of glass, and expose the whole to direct sunlight for half an hour.

QUESTIONS.

1. What change took place in each piece of paper?
2. Did air cause the change in the piece exposed to light?
3. What is the force chiefly employed by the photographer?
4. What change, if any, took place in the filtering paper, experiment 2.

§ 10.—Electricity.

Experiments.

1. Pass a current of electricity through water in a decomposition-of-water apparatus.

2. Attach a bright piece of iron, three or four inches long, to the terminal wire connected with the zinc of a three or four celled electric battery. and then immerse both the iron and the other terminal wire in a solution of sulphate of copper contained in a glass beaker or clean wooden trough, as in Fig. 6.

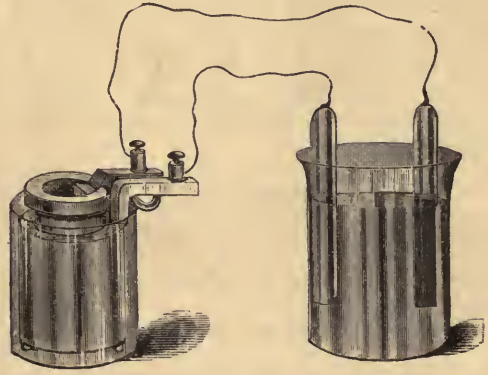


Fig. 6.

QUESTIONS.

1. What phenomena arose on passing a current of electricity through water?

2. What change took place in the piece of iron immersed in the copper sulphate?

3. What force caused the changes in the water, and on the surface of the piece of iron?

§ 11.—Chemism, Chemical Attraction or Chemical Affinity.

Experiments.

1. Put two dry Seidlitz powders together, then pour water on the mixture.

2. Dissolve them separately in water, and then mix the solutions.

3. Dissolve a few crystals of iodide of potassium and of lead acetate in separate test-tubes and then mix the solutions.

4. Wet the inside of a glass beaker with strong aqua ammoniæ, and the inside of another beaker with a solution of strong hydrochloric acid ; cover the first beaker with a glass plate and invert over the other. Then draw out the plate.

5. Cut a thin slice from the end of a stick of phosphorus. Dry it well and place on a plate, and then sprinkle over it a little powdered iodine. Cover with a wide mouthed bottle.

QUESTIONS.

1. What took place on mixing the two dry Seidlitz powders ? What on mixing their solutions ?

2. What two visible changes took place on mixing the solutions in experiment 3.

3. What two visible changes took place on mixing the gases, experiment 4.

4. Give a definition of chemical affinity and mention its characteristics as illustrated in the foregoing experiments.

CHAPTER III.

§ 12.—Mechanical Mixture.

Experiment.

Mix thoroughly some fine iron filings with two-thirds of their weight of flowers of sulphur, and carry out the following experiments :—

(a). Examine carefully with a magnifying glass.

(b). Pass a magnet through the mixture.

(c). Throw some of the mixture into a tumbler of water.

(d). Drop some of it into the liquid called carbon disulphide, and after shaking for a moment or two pour off the liquid without disturbing the sediment.

(e). Heat some of it in a test-tube until it glows.

QUESTIONS.

1. How can the sulphur be separated from the mixture? What is the color of the mixture before being heated? Afterwards?

2. What does the magnet extract from the mixture before heating? After heating?

3. How does the water affect the mixture?

4. Can any sulphur be seen by the aid of the magnifying glass before or after the heating?

5. What does the carbon disulphide separate from the mixture?

6. The force which unites the particles of a mechanical mixture is called *adhesion*. Define a mechanical mixture, and give other examples of it.

---

§ 13.—Chemical Compounds.

The substance formed in the last of the foregoing experiments when sulphur and iron filings were heated is called a chemical compound. The two next experiments will illustrate the formation of other chemical compounds.

Experiments.

1. Counterpoise upon your chemical balance a small crucible and then place in it a weighed quantity of tin-foil. Heat the crucible strongly for some time, then allow it to cool, and afterwards weigh it.

2. Repeat the foregoing experiment, using very fine iron filings, or better still, *ferrum redactum*.

QUESTIONS.

1. Did the ash produced by burning the tin-foil weigh more or less than the tin-foil did? If more, what must have been the source of gain?

2. Did the iron change in weight? What must it have combined with?

3. Name another metal which you burned and which also changed weight during combustion.

---

§ 14.—Decomposition of Substances.

The first question which a chemist asks himself regarding any natural substance is, *of what is it composed?* He proceeds to get an answer to this question by subjecting the substance to experiment, applying, if necessary, all the forces at his disposal in order to analyze or decompose it into simpler constituents. For example, it is quite possible for a chemist to take the ashes that are formed in burning the magnesium, or iron, or tin-foil, and to analyze these ashes and to obtain from them the metals—magnesium, or iron, or tin, as the case may be; but the processes he would adopt in doing this would be too difficult for a beginner to understand, and consequently we shall have to take some simpler illustrations of the breaking up or decomposition of what we call compound bodies.

Experiments.

1. Take a piece of the ore of lead called galena. Powder it in a mortar. Then take a piece of wood charcoal, scoop out a hole in it, and place in the hollow some of the powdered galena. Then heat with a blow-pipe until a metallic bead is formed.

2. Place a splinter of dry wood in a test-tube and heat by holding in the flame of a spirit-lamp.

3. Put about 8 grains of silver nitrate in a small test-tube, and heat very *strongly*.

[NOTE.—In the third experiment a simple substance remains in the test-tube, and a simple and a compound one pass off in the form of gas.]

QUESTIONS.

1. What was left on the charcoal? What, if anything, appeared to be given off during this experiment?

2. What visible phenomena took place in heating the wood? What was left at the bottom of the test tube after heating, in experiments 2 and 3?

3. A simple substance or element is one that has not been decomposed into two or more dissimilar ones.

4. Define a chemical compound, basing your definition upon the above experiments.

5. Heat a little red oxide of mercury in a test tube for ten minutes, and then say whether it is a simple or a compound substance.

§. 15.—Combination of Elements.

In the two next experiments simple substances or elements unite to form chemical compounds.

Experiments.

1. Put a few drops of mercury into a mortar, and into the mercury drop a few fragments of iodine. Rub them together.

2. Mix intimately some copper filings and flowers of sulphur in about equal proportions. Put in a test-tube, and heat until they glow.

QUESTIONS.

1. What state of matter results from the union of the mercury and the iodine? What is its color? If mercury and iodine be elements, what general name will be given to the substance formed from their chemical union?

2. What appearance had the compound which was formed in the second experiment?

3. Develop a definition of a chemical compound from the above experiments. Mention the names of any compounds which you may know. How can chemists distinguish elements from compounds?

4. The name synthesis is given to the process of forming chemical compounds from elements.

§ 16.—Elements and Compounds.

There are between sixty and seventy substances which are, at present, admitted to be simple substances or elements. Others may hereafter be discovered, and some now considered elements may hereafter be proved to be compounds. Compound substances are almost innumerable, many being found as constituents of the earth's crust, and many more as component parts of animal and vegetable organisms.

REVIEW QUESTIONS.

1. Is the change a physical or a chemical one that takes place in :—
    (1.) The melting of wax. Grinding of chalk to powder on a blackboard.
    (2.) The freezing of water.
    (3.) The vaporization of sulphur.
    (4.) The rusting of iron. The heating and hammering of iron on a blacksmith's anvil.
    (5.) The burning of a lamp.
    (6.) The solution of sugar in water.
    (7.) The heating of sugar in a red hot crucible.

2. Name the forces generally employed in producing chemical changes.

3. Specify those used in each of the foregoing experiments. Mention an experiment in which chemical affinity is observed to act only when the particles of substances are free to move.

4. Name one metal which on being heated undergoes a chemical change, and one, which does not.

5. What forces may be employed in exploding *(a)* gunpowder, *(b)* dynamite.

6. Contrast a mixture of carbon and sulphur with a chemical compound of the same two elements. How do the two differ in (*a*) physical state, *(b)* color *(c)* taste, *(d)* smell, *(e)* combustibility? Are there any other points of agreement or difference?

7. Tell how you would experiment to find out if a substance is, or is not, soluble in water or other solvent.

8. Tell how you would find out how much sand there is in 100 grains of the sugar of which your teacher will supply you with a sample.

CHAPTER IV.

§ 17.—Air.

We shall now proceed to study the chemical properties of air, first asking the question: *Of what is air composed?* Our experiments with magnesium, iron filings, and tin-foil, when these were burnt in air, shewed us that the ashes formed from these metals weighed more than the metals themselves. Whence came this increase of weight? Evidently these metals, when heated, united with the air, or with something in the air, and this air, or its gaseous constituent has taken the solid form in the ashes. We can easily devise an experiment which will reveal the source of the gain in weight, as well as answer, partially at least, the question: *Of what is air composed?*

Experiments.

1. Wet the inside of a pickle bottle with water, and drop into it some fine iron filings. Then shake the bottle so that the inside may become closely sprinkled with the filings. Place the bottle, mouth downward, over a soup-plate filled with water, and allow the whole to stand for a day or two.

Without awaiting the result of the preceding experiment, proceed with the following ones.

2. Cover a cork about two inches in diameter with a piece of tin and float on a soup plate full of water. Take a piece of phosphorus about the size of a pea, and place it on the cork. Now set fire to the phosphorus, and then cover it quickly with a beaker or small bell jar, placing it mouth downwards, as in Fig. 7. Allow it to stand thus for 15 or 20 minutes.

FIG. 7.

3. Prepare the gas as before, using a large bell jar, then transfer it by means of a pneumatic trough to four separate beakers, and proceed to perform the following experiments:—

(*a*). Smell the contents of the first beaker.

(*b*). Plunge a burning taper into the second.

(*c*). Pour some lime-water into the third, and then shake it.

(*d*). Try to pour the gas in the fourth *downward* into an "empty" beaker.

(*e*). Try to pour the contents of the fifth *upward* into an "empty" beaker.

(*f*). Test the result of (*d*) and (*e*) by plunging lighted tapers into the vessels into which you tried to pour the gas.

QUESTIONS.

1. Compare your results in experiment 2 with those of other students, or repeat the experiment, using more phosphorus than at first.

2. Is air a simple or a compound substance? Give reasons for your answer.

3. By how much was the original air reduced in volume? Where did it go to?

4. If air is composed of two things, how can one be distinguished from the other?

5. What color had the substance which was formed by burning the phosphorus? Where did it go to?

6. What effect had the gas on the burning taper? On the lime-water? Is the gas heavier or lighter than air?

7. What gas escaped on covering the burning phosphorus with the bell jar? Why?

The gas that remained in the bell jar is known as **nitrogen**; that which united with the phosphorus, magnesium, iron filings, and tin foil is called **oxygen**. The latter is often called a supporter of combustion, and the former, a non-supporter.

---

§ 18.—Burning.

Several interesting inquiries are suggested by these experiments. Are all cases of ordinary combustion due to the oxygen in the air? Will air burn in an atmosphere of coal gas, just as coal gas burns in the air? Is matter ever lost or destroyed in combustion? If the combustion of a substance, like magnesium, and the increase in weight which takes place, are due to union with the oxygen of the air, does the union of the oxygen and magnesium take place in definite, or in variable proportions by weight? These inquiries will be partly answered in the following experiments.

Experiments.

1. Light a short piece of candle and float it on a flat cork in a soup plate filled with water. Then cover the floating candle with a jar, as in Fig. 8, pressing the jar down into the water.

2. Take a small lamp-chimney and place it over the lighted candle used in the preceding experiment. Press the chimney firmly upon the cork so that no air can enter from below. The candle must, of course, be placed upon a table in this case.

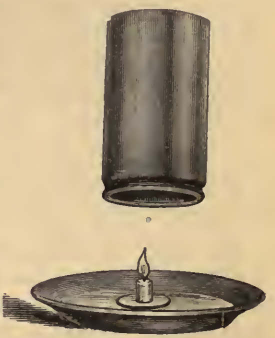


FIG. 8.

3. Cut a piece of cardboard so as to fit the upper two-thirds of the chimney, place it inside, and then repeat the preceding experiment.

QUESTIONS.

1. In which of these experiments did the candle go out? Why?

2. What substance was supplied to the burning candle in the third experiment which was not supplied to it in the first or second?

3. Criticize the following conclusion, supposing it to be based upon the foregoing experiments.—All things, when they burn, unite with the oxygen of the air, and increase in weight.

---

§ 19.—Burning of Air.

Experiment.

Take a small lamp-chimney and fit into its lower end a good cork carrying two tubes. The upper end of one of these must merely pass through the cork, and its lower end must be attached to the gas supply. The other tube must move easily through the cork, and must be long enough to project above the top of the chimney; its lower end must be attached to a tube supplying a steady stream of air. When the apparatus is all ready, turn on the gas, and after a minute or two, light the gas at the top of the chimney. Then turn down the gas until there is only a small flame; move up the long tube to the top of the chimney; turn on the air supply and when the air is alight, withdraw the air tube until the flame of burning air is about the middle of the chimney.

§ 20.—Matter Indestructible.

In most cases of ordinary combustion there appears to be an utter destruction of the substances burnt. Wood, coal, oil, wax, and many other things burn and leave little or no trace behind them. Is the matter of which they are composed really destroyed, or does it also unite with the oxygen of the air, take on new forms, and increase in weight just as magnesium did? We shall put this question to Nature in the shape of an experiment.

Experiments.

FIG. 9.

1. Pass the product of the combustion of a candle or of a coal oil lamp through the apparatus Fig. 9, attaching, if necessary, an aspirator to the U tube D, in order to draw a current of air through it. Weigh the candle before and after the experiment, and weigh the U tube D also, and the bent tube C. The tube D contains pieces of caustic potash, and the middle part of the tube C is immersed in water, in order to keep it cool.

2. Repeat this experiment first putting enough lime water in the tube C to merely cover the bend.

QUESTIONS.

1. What change took place in the weight of the candle? In the weight of the tubes?

2. What relation is there, if any, between the change in weight of the candle and the change in weight of the two tubes?

3. What property of matter is taught by the first experiment?

4. What inference may now be fairly drawn regarding the combustion of all bodies? Is the product of combustion always visible?

5. Since your experiment with burning air, would you now look upon coal gas as a "supporter of combustion?" Define this phrase, giving it a more extended meaning than was given to it in § 17.

6. What change took place in the lime-water? The product of the burning candle which produced this change is called CARBON DIOXIDE, or sometimes, CARBONIC ACID GAS.

---

§ 21.—Law of Constants.

The next inquiry which we must make is, *Does the union of the oxygen with magnesium, and with the elements of wood, coal, &c., in combustion, take place in certain fixed and invariable proportions by weight, or are the proportions indefinite and variable?*

Experiments.

1. Weigh a small crucible and its lid on a good chemical balance. Then take 500 centigrams of magnesium wire, and coil it up so that it will be inside the crucible and near its mouth. Place the wire inside the crucible with one end projecting a little, and ignite this end by holding it in a lamp flame. While the magnesium is burning hold the lid over the crucible loosely, so that as little as possible of the white smoke which forms may escape. When the wire is all burnt place the cover on the crucible, and allow it to stand until perfectly cool; then weigh. State your results as follows :—

Weight of crucible and wire = ... c. gs.
Weight of crucible alone = ... "

Weight of wire taken = ... c. gs.

Weight of crucible and ash = ... c. gs.
Weight of crucible alone = ... "

Weight of ash = ... c. gs.

2. Counterpoise on a chemical balance, a small, dry, clean test-tube of *hard* glass. Remove from balance, and in it place 500 centigrams of pure nitrate of silver. Heat in the flame of a Bunsen lamp, cautiously at first, then *strongly*, until no red fumes are evolved. After cooling, weigh carefully the test-tube and the metal remaining in it. Calculate what weight of silver you have obtained from the 500 centigrams of nitrate, recording your results as before.

3. Repeat the foregoing experiment in every particular, and when you have done, write down what the results of your experiments appear to indicate.

QUESTIONS.

1. What different result, if any, was obtained on carefully repeating the second experiment?

2. If 170 centigrams of silver nitrate were used in repeating the second experiment, what weight of metallic silver would be left?

3. What conclusion might we come to if these experiments, repeated a number of times, gave uniformly the same results?

4. What inference might fairly be drawn from such experiments, in regard to the weights of substances which compose other chemical compounds?

5. Give a definition of the law of chemistry, illustrated in the above experiments. What is meant by a *Law of Nature?*

§ 22.—Weight and Volume.

Chemical union may take place in certain proportions by *weight*, and when the substances exist in the gaseous condition in certain proportions by *volume;* but in both cases the combination is regulated by certain laws called laws of chemical combination. It is, of course, very difficult for a beginner to prove accurately the correctness of these laws. The reason of this is that the full proof requires the employment of some of the most delicate and difficult processes in Chemistry. However, a careful repetition of the two preceding experiments will greatly aid the student in gaining a correct conception of one of the fundamental laws of the science. This law has not been *proved* to hold good in the case of all known compounds, for there are too many of them; but it has been proved true for a large number of chemical compounds, and, as there is no known exception to it, we *believe* it to hold good in all cases.

CHAPTER V.

§ 23.—**Oxygen.**

From our previous experiments we have learnt that oxygen exists in air; we have learnt also that when metals like magnesium and tin-foil burn in air, the ash or rust that forms is a chemical compound, consisting of the burnt metal and the oxygen of the air. It ought, therefore, to be possible to obtain oxygen again from the white ash, or as we shall now call it, *oxide* of magnesium. Oxygen can be obtained from this oxide, but the process of getting it would be a circuitous one. From the oxide of other common metals, for example, from that of mercury, the gas can easily be obtained. This oxide of mercury—the "red oxide"—is formed whenever metallic mercury is heated for a long time in air.

Experiments.

1. Weigh out 1 gram of red oxide of mercury and place in a test-tube fitted with a cork and delivery tube as in Fig. 10. Heat uniformly at first so as not to break the glass. Oxygen is driven off, and collects in the upright glass vessel standing in the "pneumatic trough." As soon as the disengagement of gas slackens, you should raise the end of the delivery tube from the water, and *then* extinguish the lamp. The gas that collects at first should be thrown away. Raise the tube containing the oxygen out of the water and plunge a glowing splinter into it. (HgO) heated = $Hg + O$

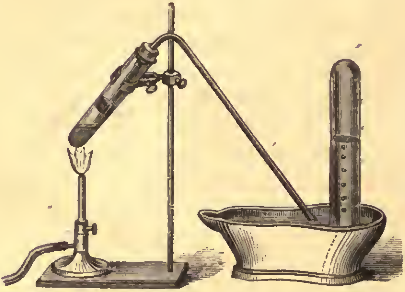


FIG. 10.

2. Heat some red oxide of lead in the same manner as in the preceding experiment. Test as before. $Pb_3O_4 =$

3. The best method of preparing oxygen in moderately large quantities in a small laboratory is the following :— Into a mortar put 10 grams of potassic chlorate and 2½ grams of manganese dioxide Powder them, and place in a dry Florence flask. Use chemically pure material only, as serious explosions have occurred from organic matter or carbon being mixed with the manganese dioxide. Heat strongly but carefully, and collect four or five small jars of the gas; or better still, collect in a gas-holder. Manganese dioxide is mixed with the potassic chlorate to cause the latter to part with its

$4KClO_3 +$ [illegible] $= 4KCl +$ [illegible] $+$

oxygen at a lower temperature than it otherwise would. The manganese dioxide remains unchanged at the end of the experiment.

4. Into a jar of oxygen plunge a piece of glowing charcoal. After combustion has ceased, pour some water into the jar, shake, and then test with blue litmus.*

5. Draw the temper from a piece of fine watch spring by passing it slowly through a lamp flame; then tip with sulphur, and ignite. Place in a second jar of oxygen, and it will burn with beautiful scintillations. Add water, then taste, and test, first with blue litmus, then with red litmus solution. acid

6. Put a little sulphur in a deflagrating spoon, ignite, and place in a third jar of oxygen. Pour some water into the jar after the combustion has ceased. Test with blue litmus paper; also taste the solution. [illegible]

7. Clean the spoon used in the last experiment, and place a piece of phosphorus on it about the size of a pea. Ignite, and place in a fourth jar of oxygen. After combustion has ceased, remove the spoon, and pour some water into the jar. Shake up the water with the product of the combustion, and then add some blue litmus solution. Also taste the solution. acid

8. Cut zinc foil into fine strips, tip them with sulphur ignite and place in a fifth jar of oxygen. After combustion has ceased, add water; taste as before, and test with *reddened* litmus paper. alkali

9. Place a piece of sodium in a deflagrating spoon, hold it in a lamp flame until it begins to burn, and then plunge into a sixth jar of oxygen. Add water, taste, and test with reddened litmus.

---

*To make red and blue solutions, "steep some solid litmus in water or weak alcohol. Divide the liquid which you pour off from the sediment, into two parts; to one add a few drops of weak sulphuric acid; to the other add a little solution of caustic soda. You will then have red and blue litmus solutions, and if you add them to the products of your experiments with oxygen, you will be able to test whether a new compound has been formed in case other evidence of a chemical change is wanting."

10. Repeat the preceding experiment, using, first of all, potassium, and then magnesium.

11. Pour a little potassium hydroxide into a jar of oxygen. Shake, and if no change in volume takes place, add a small quantity of pyrogallic acid. Shake again, and note any change.

QUESTIONS.

1. With what substance is the oxygen mixed at first as it comes off from mercuric oxide, manganese dioxide, or potassic chlorate. Do any of these substances change color on being heated?

2. How do we separate the manganese dioxide from the potassic chloride as they occur mixed in the residue from preparing oxygen?

3. Into what three classes may oxides be divided? How would you distinguish them? Base your answer upon the actions which the compounds formed in combustion had upon red and blue litmus solutions.

4. Make a list of the properties of oxygen.

5. Contrast the properties of oxygen with those of air?

6. What would you consider the most important property of the gas?

7. Give a reason for concluding that the gas is only slightly soluble in water?

8. What effect had oxygen on pyrogallic acid solution? The change that takes place is generally considered a *test* for free oxygen.

---

§ 24.— Other Properties.

Oxygen is found combined with organic bodies and minerals, forming nearly $\frac{1}{3}$ of the weight of solid earth. It is soluble to the extent of four per cent. by volume in water, a fact of great importance in relation to aquatic plants and animals. By cold and pressure combined, it has been made to take the liquid, and even the solid, form; it unites with all the other elements, except fluorine; and is that element in air which sustains animal life, hence called *vital air*. In respiration we simply take oxygen into the system, and this causes slow combustion of the tissues, and consequently gives rise to animal heat. Ordinary *combustion* or *burning* is simply chemical action, attended by great heat and light, chemical compounds being formed, by uniting with the oxygen of the air. Before this union can take place, however, a substance must be heated to its kindling point.

An **oxide** is a compound formed by the union of oxygen with some other element.

CHAPTER VI.

§ 25.—Nitrogen.

The simplest method of obtaining nitrogen is by burning phosphorus, hydrogen, or some other combustible in a bell jar over water, the oxygen being burnt out and nitrogen remaining. The student should repeat the second experiment of § 17. After doing so, he may try the following one.

Experiment.

The apparatus represented in Figure 11 may be used for preparing nitrogen by passing air over red-hot copper. A is a U tube filled with calcic chloride, B is a straight tube filled with fine, bright copper filings, and C is a large-mouthed bottle used as a gas holder, one of its tubes passing through the cork, the other passing to the bottom of the bottle; D is a piece of rubber tubing attached to one of the glass tubes in order to convert it into a syphon and draw off the water from the bottle. On starting the experiment the bottle must be full of water. When the copper has been made red-hot, the syphon must be made to act *very slowly* and as the water flows out of the bottle, air is drawn through the

U tube and passes over the red-hot copper. The nitrogen is collected in the bottle.

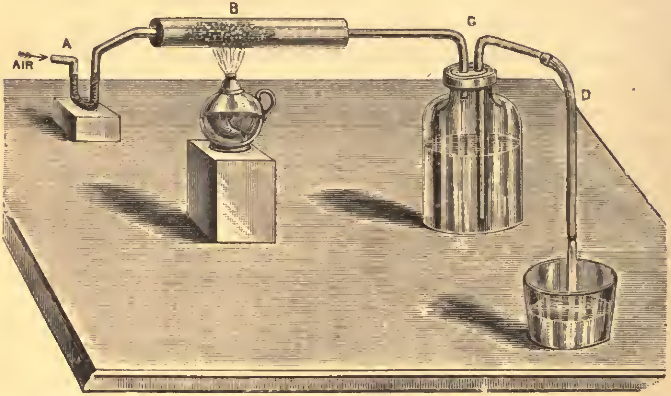


Fig 11.

QUESTIONS.

1. How does the method of preparing nitrogen explained in the preceding section differ from that in section 17 ?

2. What effect has red-hot copper on air ?

3. Give tests by which nitrogen may be distinguished from oxygen.

4. How can you form an estimate as to the extent to which nitrogen is soluble in water ?

5. Write out an account of the physical and chemical properties of nitrogen.

§ 26.—Other Properties.

Nitrogen is a little lighter than air ; will not support life ; has been liquefied by a cold of —146°C. and under a pressure of 33 atmospheres ; soluble in water to the extent of two per cent. of its volume. It unites readily and directly with few other elements, but indirectly it forms very important compounds with oxygen, hydrogen and carbon. Nitrogen is found in plants, being a constituent of some of the strongest poisons, such as prussic acid and strychnine. It is a component also of bread, milk, and the flesh of animals.

§ 27.—The Atmosphere.

The atmosphere is an ocean of gases pressing with an average weight of 14·7 lbs. upon every square inch of the earth's surface. This pressure is varying all the time. The instrument, known as the *barometer* (for a full description of which you must consult some good work on Physics), is used to measure this varying pressure. In this instrument a column of mercury 760 mm. (or 30 ins.) in height, at O°C. represents the standard pressure of the air at sea level. Whether the gases constituting the air are united in definite proportions by weight and volume, or whether they are merely mixed, your experiments have not yet revealed to you. Nor do you yet know whether there may not be a number of other substances present besides oxygen and nitrogen. After you have studied some other elements and compounds, you may resume your study of the air, and will then be able to understand fully its composition. Meanwhile, however, there is one other constituent of the atmosphere whose presence you can easily detect.

Experiment.

Dry the outside of a clean tumbler and fill it half-full of water. Put a piece of ice in the tumbler and then place it in a warm room.

QUESTIONS.

1. What appeared on tne outside of the tumbler? Where did it come from. Give reasons for your answer.
2. Mention other cases in which the moisture in air becomes visible. What is its source?
3. Is the vapor of water in air heavier or lighter than air? In what forms does it fall to the earth?

§ 28.—Analysis of Air.

The proportions by *weight* in which oxygen and nitrogen exist in air have frequently been determined by carefully conducted experiments, patiently and laboriously carried on for a long time. The processes employed in the determination are, however, too difficult for beginners to understand, so we shall content ourselves, for the present, with determining roughly the proportions by *volume* in which these two elements exist in air.

Experiment.

1. Take a graduated tube about 30 centimetres long and 3 centimetres wide, and invert it over a vessel 7 or 8 centimetres in depth, filled with water. Fix the tube ina support, taking care that the water stands at the same level on the inside as it does on the outside of the tube. Then pass up into the tube a piece of phosphorus attached to a copper wire, as in Fig. 12. To attach the wire to the phosphorus, fuse it under water in a test tube, introduce the end of the wire into it, and then let it cool. Leave the whole twenty-four hours, then withdraw the phosphorus and adjust the level of the water inside and outside the tube; read off the volume of the gas remaining in the graduated tube. The volume of gas at the beginning and at the end of the experiment must be reduced to that at standard temperature and pressure. Nitrogen four-fifths – $\frac{1}{5}$ ox

FIG. 12.

CHAPTER VII.

§ 29.—Pure Water.

Your experiments in evaporating spring, river, and well water have made you familiar with the fact that water from any of these sources is not pure. Now, it is of the utmost importance to the chemist that the water he uses should be perfectly free from any kind of solid, liquid, or gaseous impurity Water is, in short, an almost universal solvent, but different substances vary greatly in the extent to which they are soluble in it. How then can pure water be obtained ?

Experiment.

Fill a glass retort about half-full of rain or river water, attach it to a Liebig's condenser and apply heat until the water boils. Throw away the water which first collects in the receiver, and do not continue the action after four-fifths of the water have been vaporized, unless you wish to find the amount of dissolved matter in the water. The apparatus required for the distillation of water is represented in Fig. 13.

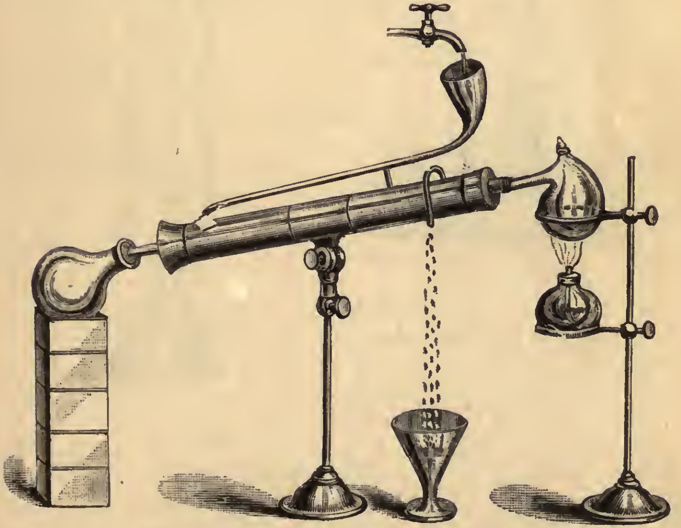


FIG. 13.

2. Repeat this experiment, first dissolving a little sulphate of copper in the water. What is the purest kind of water which we meet with in nature? Why?

§ 30.—Analysis of Water.

The ancients supposed that there were only four elements in the universe. viz., fire, air, earth, and water. We have already seen that air is not an element. Is water one? We shall see.

Experiments.

1. Take a test-tube about 2 centimetres in diameter, and 10 or 12 centimetres in length. Fill it with acidulated water (1 to 60), and invert it over a beaker containing water. Under the mouth of the test-tube, place the terminal wires of a galvanic battery, as in Fig. 14. The ends of these wires should consist of platinum, and should not touch each other when placed under the mouth of the test-tube.

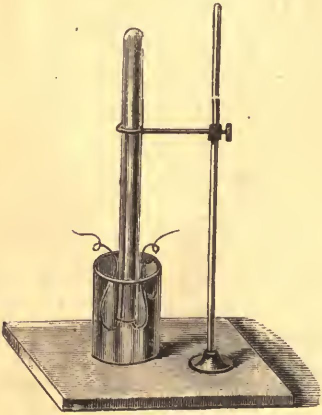


FIG. 14.

FIG. 15.

QUESTIONS.

1. What appeared to collect at the top of the tube? Where did it come from?

2. Did both wires appear to yield an equal quantity?

3. What became of the water as the gas collected?

4. What force produced the chemical change?

2. When all the water has been expelled by the accumulated gas in the preceding experiment, raise the tube keeping it mouth downward, and apply a lighted match to it.

3. Repeat experiment 1, using two test-tubes full of acidulated water, inverted over a soup plate. Place the end of a wire under each tube. Each wire must be coated with paraffin or sealing wax where it touches the water, except about 1 centimetre at the end. When gas has filled one of the test-tubes, stop the current, and examine the gases. Put a glowing splinter of wood into the one with least gas in it, and apply a lighted taper to the full one. By the use of the apparatus, Fig. 15, this experiment may be easily and neatly performed.

This process of decomposing water is called the **Electrolysis** of water.

QUESTIONS.

1. Why coat the wires with wax?

2. What difference is there in the amount of gas collected in each tube?

3. Name the gas of which there was the smaller volume. Give a reason for your answer.

4. What effect had the gas in the other tube on the burning splinter?

5. How did the two gases differ in behaviour? Mention three properties which they have in common.

The gas of which there was the larger quantity is called HYDROGEN.

---

§ 31.—Synthesis of Water.

We have now proved that two gases, oxygen and hydrogen, may be obtained from water, by passing a current of electricity through it; our next step is to reverse the process, and prepare water from the union of these same gases. This process of forming water is called the **synthesis** of water.

Experiments.

1. Take a graduated tube called a eudiometer, fill it with mercury, and invert it over mercury in a soup-plate or saucer. Then pass into it a known volume of oxygen and twice the volume of hydrogen, measuring both at the same temperature and pressure. Pass a spark from a Leyden jar, or from a Ruhmkorff's coil, through the mixed gases, Figure 16. Before igniting the gases, press the eudiometer firmly on a rubber pad placed at the bottom of the plate or saucer. After ignition examine the top of the eudiometer with a good lens.

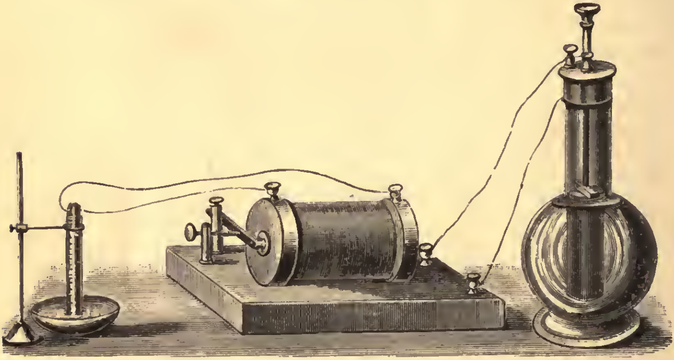


Fig. 16.

2. Repeat this experiment using equal volumes of the two gases. Test any gas that remains.

3. Try the experiment again, using twice as much oxygen as hydrogen. Test as before any gas that remains.

QUESTIONS.

1. What two phenomena occurred on igniting the mixed gases—one perceived by the hand, the other by the eye?

2. Did the mercury rise or fall in the tube after the gases were ignited? Explain.

3. What substance was perceived at the top of the eudiometer after ignition? In what physical state?

4. What gas remained in excess in the eudiometer after ignition in experiments 2 and 3? How did you know?

5. In what proportion by volume did the gases combine?

---

§ 32.—Volume of Steam.

We have now to find out how many volumes of water-gas, or steam, will be produced by the union of two volumes of hydrogen and one of oxygen.

Experiment.

Fill a eudiometer one-third full of a mixture of hydrogen and oxygen gases—using two volumes of the former to one of the latter. Cover the eudiometer with a larger tube, into the top and bottom of which pass tightly-fitting corks perforated with tubes, admitting steam at the top and giving exit to it at the bottom, Fig. 17. The wires from the battery to the eudiometer should pass into the jacket through its upper cork. After the steam has been admitted, mark the height of the mercury above that in the trough, and also the volume of the contained gases, then explode them, as in experiment § 31. After explosion depress the eudiometer, until the mercury in the tube stands the same height above that in the trough as before. Then measure the volume of the water-gas (steam) in the eudiometer, and compare this volume with that of the original mixture.

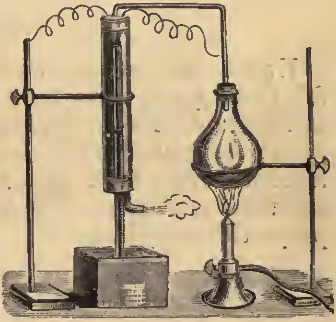
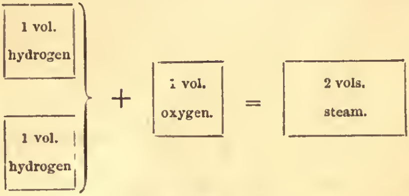


Fig. 17

If we represent equal volumes of oxygen and of hydrogen by equal squares, and then place in these squares the first letter of the name of these elements, we can represent to the eye, by another figure, the volume of water-gas or steam formed, and the condensation which occurs after union. Thus:



QUESTIONS.

1. Was there expansion or contraction of the original volume? To what extent?

2. What was the volume of the steam formed as compared with that of the hydrogen? Of the oxygen?

CHAPTER VIII.

§ 33.—Hydrogen.

Having decomposed air and water, we naturally ask: *May not oxygen, nitrogen, and hydrogen also be compound substances?* To this question we can only answer that as yet we have been unable to decompose them by any force which we have at our command, and we therefore call them elements. We shall now proceed to study the properties of hydrogen.

Experiments.

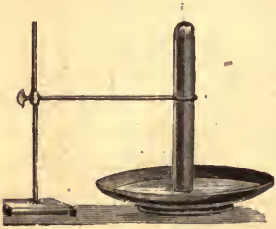


FIG. 18.

1. Make a mixture of water and strong sulphuric acid in the proportion of 6 to 1. Fill a test-tube with this mixture, and invert it over a saucer half full of the same mixture. Below the test-tube, which must always be kept with its mouth below the level of the acid and water, place some small pieces of pure zinc, as in Fig. 18.

$Zn + H_2SO_4 + H_2O = ZnSO_4 + H_2O + H$

2. Prepare two or three test-tubes full of this gas in this manner, and then perform the following experiments with them .—*(a).* Smell the gas in one of the test-tubes. *(b).* Raise a second out of the water, and, keeping it mouth downward, apply a lighted match to it.

QUESTIONS.

1. What became of the zinc? How did you know that the gas was not air?

2. Tell how you know that this gas is not soluble to any great extent in water.

3. Why keep the tube with its mouth downward?

4. Devise an experiment to prove that the hydrogen could not have come from the water.

5. What then was the object in using water in experiment 1? Prove the correctness of your answer by an experiment.

6. What forces came into play in experiment 1, and what was their respective results?

Experiment.

Take a wide-mouthed bottle and fit it with a good cork perforated by two glass tubes, one of which passes nearly to the bottom of the bottle, and has on its upper end a funnel-like expansion; the other tube merely passes through the cork and is bent at right angles, and has a rubber tube attached to it for conveying the gas to a "pneumatic trough" as in Fig. 19. Place some clippings of zinc in the bottle, fill it about one-third full of water, and then pour down the funnel tube about a teaspoonful of sulphuric acid. The gas begins to form quickly, and is collected in bottles previously filled with water and kept mouth downward in the water

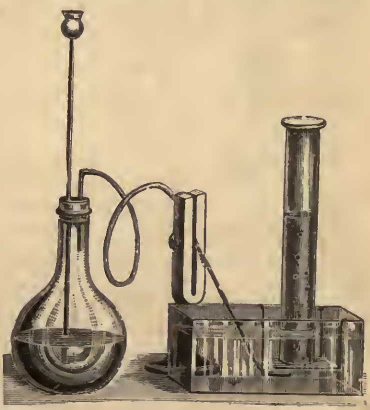


FIG. 19.

in the pneumatic trough. Before collecting the gas for future experiments, it is well to allow it to escape for a few minutes in order to carry with it the air that previously filled the bottle. Collect two or three bottles or large test tubes full of gas and preserve them for future experiments. You must preserve also the liquid which remains in the bottle after the gas has ceased to come off. Filter the liquid and then either evaporate it over a lamp flame, or allow it to stand in an open vessel for a day or two. Then examine carefully.

---

§ 34.—Properties of Hydrogen.

Experiments.

Fig. 20.

1. Keep the mouth of one of the bottles downward, and plunge a lighted taper upward into it, as indicated in Fig. 20. Then withdraw the taper slowly, allowing the burnt end to remain a moment or two at the mouth. If you cannot answer the following questions, your observations have not been accurately made, and you should repeat the experiment.

QUESTIONS.

1. What happened to the flame of the taper when first put into the bottle? What phenomenon accompanied this?

2. What was the color of the hydrogen flame?

3. With what must the hydrogen have united in burning?

4. What happened to the taper as it was withdrawn? Why?

5. In collecting this gas by upward displacement, did the bubbles rise through the water in a straight line? If not, explain how they rose, and why?

2. Take the second jar, and keeping its mouth downwards, quickly (Fig. 21) turn its mouth upwards under the mouth of a similar jar filled with air. Let it remain thus for a few seconds, and then apply a burning taper to the mouth of the upper vessel.

3. Pass the gas from the generating apparatus into soap-suds, and send up a few bubbles.

4. Pass the gas also into a collodion balloon.

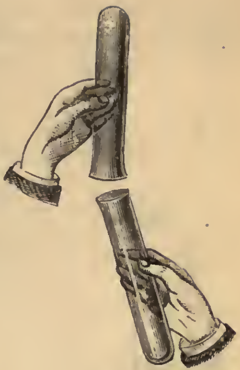


FIG. 21.

QUESTIONS.

1. Is this gas heavier or lighter than air?

2. What phenomena are observed when the lighted taper is brought to the mouth of the jar, experiment 2? How do they differ from those of the first experiment?

3. What use can be made of hydrogen gas?

§ 35.—Another Method.

There are **other ways** of preparing hydrogen, and you might now try one of them.

Experiments.

1. Prepare hydrogen gas, using the apparatus Fig. 22, and putting into it granulated zinc and hydrochloric acid. Pass the gas through a tube containing fragments of calcic chloride for the purpose of drying the gas, and through the cork which should tightly fit the end of this tube, pass a glass tube drawn to a fine point. After the gas has escaped for three or four minutes,* apply a lighted match to the gas-jet.

Fig 22.

[* NOTE,—If the lighted match be applied to the gas-jet before all the air is expelled from the flask, a dangerous explosion will result.]

$Zn + HCl = ZnCl + H$

2. Introduce a piece of iron or steel wire into the flame. Try the effect of the flame on platinum wire also.

Fig. 23.

3. *The Chemical Harmonicum.*—Bring down over the jet a tube about 2 centimetres wide and 30 or 40 centimetres long. Use tubes of different diameters and different lengths, and move them slowly up and down.

4. Invert a long, dry, wide-mouthed bottle (a pickle bottle will do well) over the jet, as in Fig. 23.

QUESTIONS.

1. What is the use of passing the hydrogen through calcic chloride?

2. What is the colour of the flame? Try whether this colour would change if a platinum or a copper tube were used instead of the glass one.

3. Is this flame hotter than that of a spirit lamp? Devise an experiment to show that your answer is correct.

4. What phenomena occurred in experiment 3? How was the pitch of the note made to vary? Why did the upper end of the tube become dimmed? Did the shape of the flame change? How?

5. Where did the moisture on the inside of the bottle (Exp. 4) come from? Give a reason for concluding that it could not have come from the generating flask.

§ 36.—Water from Hydrogen.

We have now to try to ascertain what weight of water is produced by the combustion of a known weight of hydrogen.

Experiment.

Weigh a U tube filled loosely with small pieces of calcic chloride, and then pass the product of the combusstion of the hydrogen through this tube by means of an *aspirator.* The fittings must be perfectly air-tight. Carefully weigh the U tube after the action has been continued for about 10 minutes. The generating flask—a small one capable of yielding a constant stream at pleasure—must also be weighed before and after burning the gas. Then the weight lost by the flask must be compared with the gain in weight of the U tube.

The apparatus suitable for conducting this experiment is represented in Fig. 24.

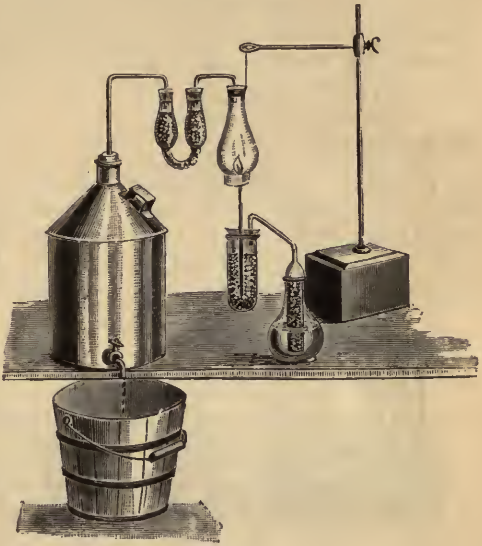


Fig. 24.

QUESTIONS.

1. If only 1 centigram of hydrogen passes out of the flask, and 9 centigrams of water are produced, where must the additional 8 centigrams have come from ?

2. What is the proportion by weight in which hydrogen unites with the oxygen which it takes from the air ?

3. What was the proportion in which magnesium united with this same constituent of air in experiment 1, § 21 ?

4. Point out an error in the reasoning, if I conclude from this experiment that the substance which united with the hydrogen is the cause of all cases of ordinary combustion.

5. In what proportion by volume were oxygen and hydrogen proved to combine in § 31 ?

6. How many times, therefore, is oxygen heavier than hydrogen if we take equal volumes of each ?

7. Calculate from this experiment and from experiment 1, § 32, the weight of any volume of steam as compared with the weight of an equal volume of hydrogen.

§ 37.—Hydrogen from Steam.

Your experiments with air must already have suggested to you that as red-hot copper took out oxygen from the air, so it or some other metal might also, under favorable conditions, take oxygen from water, and thus liberate hydrogen. We shall put this conjecture to the test of experiment.

Experiments.

1. Take an iron tube about $\frac{3}{4}$ of a meter in length, and 2 centimeters in diameter (an old gun barrel will do well); fill it nearly full of clean iron filings; fit each end with a tightly fitting cork and tube, the one leading to a pneumatic trough; the other connected with a flask containing boiling water as shewn in Fig. 25. Place the iron tube in a charcoal fire, built upon bricks. When it is red-hot, boil the water in the flask, and force steam through the tube. After the tube has cooled, turn out the iron filings and examine them carefully.

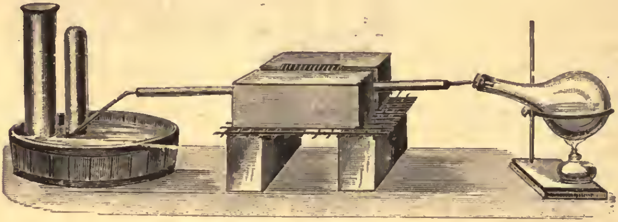


FIG. 25.

2. Collect a few jars of the gas, prepared in the foregoing manner and test for hydrogen as in former experiments. Then fill a bottle one-quarter full of the gas, and the rest of it with air; and explode the mixture.

3. Repeat the last experiment, filling two-thirds of the bottle with hydrogen and the other third with oxygen. But before exploding this mixture, wind a towel round

the bottle for fear it may be broken by the violence of the explosion.

4. Take two rubber bags, and fill one with hydrogen and the other with oxygen. Subject both bags to an equal amount of pressure between two boards, or otherwise, then connect them with the apparatus known as an oxy-hydrogen blow-pipe, and having turned on the hydrogen, ignite it, and then carefully and very gradually turn on oxygen gas from the other bag. A complete blowpipe apparatus is represented in Fig. 26.

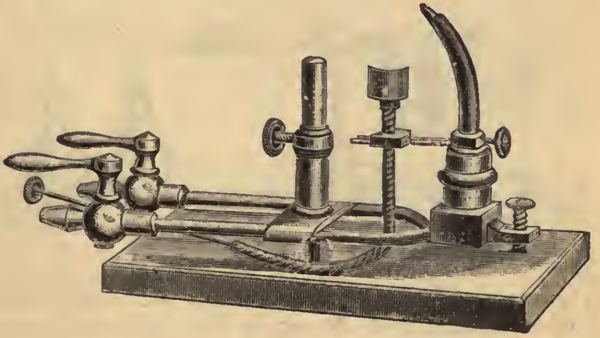


Fig. 26.

5. Introduce into the oxy-hydrogen flame, separately, a piece of platinum wire, of steel, of zinc and of quicklime.

§ 38.—Steam from Hydrogen.

Our observations on the experiments in the preceding section suggest to us to inquire what the result would be if we reversed the process; that is, if were to pass hydrogen over heated black oxide of copper, or red oxide of iron.

Experiment.

Take a glass or iron tube about a centimeter in diameter and 25 centimeters long; and about its middle, place a small quantity of black oxide of copper. Connect it with a tightly fitting cork and tube to a hydrogen generating apparatus. The gas must be dried by passing it through sulphuric acid (second flask from the right in Fig. 27) and through calcic chloride, before it reaches the tube with the oxide of copper in it. After allowing the hydrogen to escape for a few minutes, so as to drive out all the air from the apparatus, heat the oxide of copper, but not before

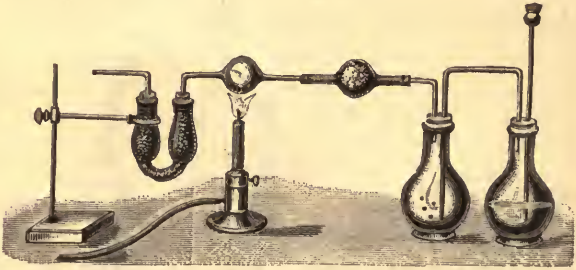


FIG 27.

The water, of course, passes off in the form of steam, but is condensed and caught in the drying calcic chloride tube, D. This experiment has been used by MM. Dumas and Boussingault in determining the composition of water, and deducing, in this way, the law of union in definite proportions by weight.

QUESTIONS.

1. What compound is formed in passing steam over red-hot iron in a tube? What gas passes out of the tube?

If I reverse the process, and send hydrogen into a tube containing red-hot copper oxide, what compound is formed?

2. Ascertain by experiment, or by consulting some standard textbook, whether red-hot copper has the same action on steam that red-hot iron has.

3. A reducing agent is one that extracts, from any compound, the oxygen which it contains. How has the reducing power of hydrogen been used in determining the composition of water by weight?

§ 39.—Hydrogen from Cold Water.

The next inquiry which the student will naturally make is: *Are there other metals besides iron which will extract oxygen from water and allow the hydrogen to pass off?*

Experiments.

1. Take pieces of the metals sodium and potassium of the same size and drop them into separate saucers of water, as in Fig. 28. Bring a lighted taper down close to the sodium on the water. After all chemical action has ceased, test the water with reddened litmus paper.

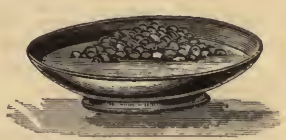


FIG. 28.

2. Cut off with a pen-knife a piece of the metal potassium about the size of a pea. Then fill a test-tube with water and invert it, mouth downwards in a soup-plate of water as in Fig. 29, and fasten in position as shown. Be careful to admit no air into the test-tube. Now hold the piece of potassium firmly in a pair of for-

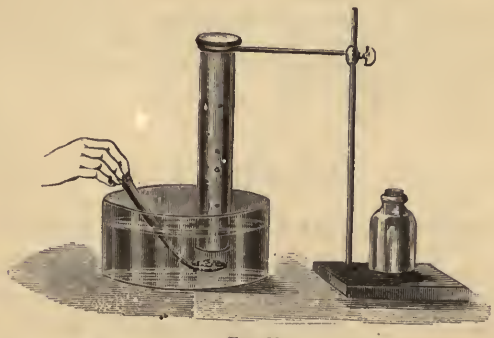


FIG. 29.

[NOTE.—Both potassium and sodium should be kept in naphtha or coal oil If exposed to air or water they soon lose their metallic form.]

ceps, and very *quickly* plunge it into the water under the test-tube. Wait until all action has ceased, then raise the tube out of the water and apply a lighted match to the gas that has collected in the test-tube. Test the water with a slip of reddened litmus paper. Taste the water. caustic

3. Take a piece of the metal sodium about the size of a bean, and perform an experiment similar to the last in all respects.

QUESTIONS.

1. What phenomena were observable on placing each metal in the water? In what respects did they agree? In what differ?

2. What was the color of the flame that surrounded the potassuim? What caused the flame?

3. Why was there none around the sodium at first? Was the color of the flame around each metal the same?

4. To answer the first part of question 3, place a piece of sodium on blotting paper and float the two on water.

5. If sodium forms a compound with water when thrown upon it, where is that compound? If a solid and dissolved in the water, how can you extract it? Devise an experiment to test the correctness of your answer.

6. These experiments illustrate a chemical change known as *decomposition by substitution*. Give another example of it.

7. Ascertain by experiment or by consulting some standard text-book whether other metals besides sodium and potassium will decompose water.

CHAPTER IX.

§ 40.—Physical Characters of Gases.

In taking up the study of the physical characters of oxygen, nitrogen and hydrogen, we find ourselves on the border land of the two great sciences of Physics and Chemistry. Our experiments on solution, evaporation and distillation were founded on principles with which our previous studies in Physics had made us familiar. But we are now about to draw still more largely on our knowledge of this science.

Besides possessing the chemical properties which you have already seen illustrated, oxygen, nitrogen and hydrogen possess these two very important physical ones.

1. "They are affected in the same way, and to the same extent, by equal alterations of pressure ;"

2. "They are affected to the same extent by equal alterations of temperature."

The statements of fact embodied in (1) and (2) are often spoken of as laws ; the first being generally known as Boyle and Mariotte's Law ; the second as that of Gay-Lussac & Charles. They may be stated as follows :

§ 41.—Boyle's Law

"The temperature remaining the same, the volume of a given quantity of a gas varies inversely as the pressure which it bears." [Consult Ganot's Physics, edited by Atkinson, 1881, §§ 174-5].

§ 42.—Gay-Lussac & Charles' Law.

"All gases expand the $\frac{1}{273}$ of their volume at 0° C., for every increase of temperature of 1° C., provided the pressure remains unchanged." [Consult Ganot's Physics, edited by Atkinson, 1881, §§ 331-333].

When explaining the facts of solution reference was made to the theory of the molecular constitution of matter. This theory was, in fact, first propounded by Avogadro, a celebrated Italian physicist, in order to explain the facts expressed in the two foregoing laws. Avogadro's explanation was unheeded at the time, but it was re-announced by Ampère many years afterwards, and is consequently, sometimes called Ampère's law. It may be stated as follows :—

"All substances in the gaseous state, and under like conditions of temperature and pressure, contain an equal number of molecules in equal volumes."

§ 43.—Theories and Facts.

"Generalizations from observed facts, so long as they are uncertain and incomplete, are called **hypotheses** and **theories**; when tolerably complete and reasonably certain, they are called **laws**. The attention of the student should be constantly directed to the keen discrimination between, *facts*, and the *speculations* founded upon these facts; between the actual evidence of our trained senses brought intelligently to bear upon chemical phenomena, and the reasonings and abstract conclusions based upon this evidence; between, in short, that which we may know and that which we may believe."

"If we admit the hypothesis that gases, like other bodies, are made up of small independent masses called molecules, and that heat causes these molecules to separate from one another, whilst cold or pressure causes them to approach, we are led to the assumption that in equal volumes of different gases there exist the same number of molecules." On the other hand, "if we admit the law of Avogadro, we see at once why gases are equally expanded by heat, why they are equally contracted by cold or pressure, and why they combine together, according to the discovery of Gay-Lussac, in simple proportions by volume."

§ 44.—Molecules and Atoms.

A **molecule** is the minutest particle of a compound or of an element capable of independent existence.

An **atom** is the smallest part of an element that can enter into the molecule of a chemical compound.

When chemical combination takes place it is supposed that one or more atoms of one element unite with one or more atoms of another element, and as these atoms have definite weights, it follows that the compound formed must always contain the elements in the proportions of their atomic weights. In other words, the molecular weight of a compound is the sum of the atomic weights of its constituents.

---

§ 45.—Specific Gravity of Gases.

From our experiments with water we were led to believe that one volume of oxygen is 16 times as heavy as one of hydrogen. We can test the accuracy of this conclusion by actually weighing equal volumes of these gases.

Experiment.

Take a globe, capable of holding about half a gallon, and fitted with a good stop-cock. Attach it to a good air pump and exhaust, as far as possible, all air from it. Weigh the empty globe, then fill it with hydrogen gas from the jar, Fig. 30, by turning the stop-cocks both of the jar and of the globe. Adjust the level of the water inside and outside of the jar and then close the stop-cock. Then weigh again carefully. Exhaust it a second time, then fill with oxygen gas and weigh.

Fig. 30.

[Consult Ganot's Physics, edited by Atkinson, 1881, §§ 335-6].

QUESTIONS.

1. What was the weight of the empty globe? What was the weight of the globe when filled with hydrogen?

2. What then must have been the weight of the hydrogen?

[Note.—The specific gravity of all gases may be determined in a similar manner.

3. Similarly, what was the weight of the oxygen?

16 times heavier

4. How many times is oxygen heavier than hydrogen? How many times is one molecule of oxygen heavier than one of hydrogen?

5. Try how the weight of each compares with that of air? Of nitrogen?

6. What connection exists between the weight and the volume in which hydrogen and oxygen combine with each other to form water?

§ 46.—Atomic Weight.

The Atomic Weight of an element is the number representing how many times its atom is heavier than an atom of hydrogen. How can the atomic weight of oxygen be found? Applying Avogadro's law to the results of the last experiment, it is quite clear that one molecule of oxygen weighs sixteen times heavier than one of hydrogen; and connecting this conclusion with the results of experiment, § 32, we can easily form an opinion of the relation of the weight of an atom of oxygen to one of hydrogen. It will be remembered that—

2 vols. hydrogen + 1 vol. oxygen = 2 vols. steam;

∴ 2 mols. hydrogen + 1 mol. oxygen = 2 mols. steam.

PROBLEMS.

1. How much oxygen must there be in one molecule of steam?

2. Deduce from the last experiment and from experiment, § 32, a rule for finding, theoretically, the atomic weight of an elementary gas.

It must be distinctly borne in mind that we know nothing whatever about the absolute size, or the absolute weight of the atoms of different elements.

§ 47.—How to Find Atomic Weights.

The atomic weights of elements like oxygen and hydrogen which exist in the gaseous state, and of others which are easily vaporized, may be found by simply calculating their specific gravities as in experiment § 45.

The atomic weights of hydrogen, nitrogen, and oxygen have been found in this way. But we have to resort to a different method of finding the atomic weights of elements, which, like magnesium, silver, zinc, and the metals generally, exist in the solid state at all ordinary temperatures. Here again we have to call in the aid of Physics. In short, we have to know the specific heats of such elements, before we can calculate their atomic weights.

The specific heats of the following bodies have been determined by carefully conducted and frequently repeated experiments.

	Sp. Heat.		Sp. Heat
Bromine	0·0843	Platinum	0·0324
Copper	0·0939	Potassium	0·1655
Iodine	0·0541	Silver	0·0570
Magnesium	0·2475	Sulphur	0·1780
Mercury	0·0332	Zinc	0·0956
Phosphorus	0·1740		

§ 48.—Dulong and Petit's Law.

"The same quantity of heat is needed to heat an atom of all simple bodies to the same extent." [See Ganot's Physics, edited by Atkinson, 1881, § 458.]

The atomic heat of an element (nearly 6.0), when divided by its specific heat, gives the atomic weight very nearly; or, the atomic weight of an element multiplied by its specific heat gives the constant quantity called atomic heat. The atomic heat is nearly the same for all elements, namely, 6.0.

§ 49.—Displacement.

The student must not imagine that these two are the usual methods of finding atomic weights. The fact is complete analysis of a compound, or at least ermination of the proportions by weight in which ts displace each other from chemical compounds, precedes the calculation of atomic weights, and that we generally utilize our knowledge of specific weight and of Dulong and Petit's law to correct or corroborate our results. Displacement of one element by another has been illustrated in the preparation of hydrogen, but we must now make a much more accurate experiment than any which we have yet tried in connection with this phenomenon.

Experiments. $Pb(C_2H_3O_2)_2+Zn$

1. Take about 160 centigrams of acetate of lead, and dissolve in about 40 cubic centimetres of pure distilled water. Clean and dry a porcelain crucible and counterpoise it on a chemical balance. Pour the solution into the crucible. Then immerse in the solution as in Fig. 31, 130 centigrams of pure bright zinc in fine strips, and allow them to remain there until they have all disappeared. Pour off the liquid solution, and carefully wash by decantation, dry, and weigh the metal that remains in the crucible. Try whether the metal will make a mark on white paper. Compare the mark, if any, with that made by the metal lead.

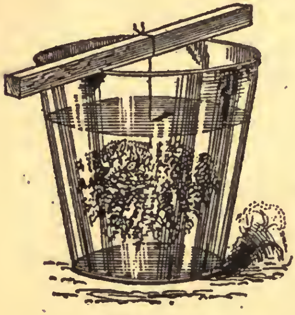


Fig. 31.

2. Repeat this experiment, using a solution of nitrate of silver and 12 centigrams of magnesium. Wash, dry, and weigh the residue carefully as before, and then rub it with a polished knife blade. Compare with silver.

[NOTE.—These experiments must be very carefully performed, especially as regards washing, drying and weighing. A thoroughly good chemical balance must be used.]

QUESTIONS.

1. What change did the zinc undergo? The magnesium.

2. What weight of lead was obtained? Of silver?

3. Perform one of these experiments a second time very carefully and see if you get a different result.

4. Try to displace copper from copper-sulphate by using iron. Then try to displace the iron from the solution by using zinc.

§ 50.—Chemical Equivalents.

The results of the above quantitative experiments, and of others which we have performed thus far in our course, show that:—

170 centigrams of silver nitrate yielded 108 of silver.

207 centigrams of lead were displaced from 325 of acetate of lead by 65 of zinc.

108 centigrams of silver were displaced from 170 of nitrate of silver by 12 of magnesium.

8 centigrams of oxygen unite with 1 of hydrogen to form water.

These numbers are called chemical equivalents because they represent the proportion by weight in which each of the substances named enters into or leaves chemical compounds. As hydrogen has never been known to unite with a less weight of any substance than its own, it has for this reason been taken as a standard by which to measure the equivalences of all other substances.

DEFINITION.—Chemical equivalents are numbers representing the proportions by weight in which different substances displace one another in chemical compounds, one part by weight of hydrogen being taken as the unit.

§ 51.—Equivalents *not* Atomic Weights.

By means of experiments similar to the preceding, the equivalents of a large number of the elements have been determined either directly or indirectly. Now the atomic weights of the elements are either the same as the equivalents or are simple multiples of the equivalents, and the specific weights of gases and the specific heat of elements are used in determining this point.

§ 52.—Electro-Chemical Series.

As a result of many experiments similar to the two last ones, the elements have nearly all been arranged in what is called an electro-motive series. In this series, soluble compounds of one metal are decomposed by another metal immersed in the solution, the first metal being displaced by the second, and forming a tree-like appearance or falling to the bottom of the containing vessel in the form of a powder. The displacing metal is said to be electro-positive to the displaced metal which itself is said to be electro-negative. The following is a short list of thirteen elements arranged in electro-chemical order :—

Electro-positive.

1. Zinc.
2. Cadmium.
3. Tin.
4. Lead.
5. Iron.
6. Nickel.
7. Bismuth.
8. Antimony.
9. Copper.
10. Silver.
11. Gold.
12. Platinum.
13. Graphite or Carbon.

Electro-negative.

The student is advised to take soluble compounds of these elements, and perform experiments with them similar to the two preceding ones, but without weighing the electro-negative element.

Each substance is electro-negative to those above it, and electro-positive to those below it on the list.

CHAPTER X.

§ 53.—Elements and Symbols.

There are 64 elements or simple substances, according to some chemists 69, and from one or more of these, every substance in nature is built up. A list of the elements is given at the end of this book, with their symbols and atomic weights.

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§ 54.—Metals—Non-Metals.

The elements are usually divided into two classes—metals and non-metals, there being 51 of the former, and 13 of the latter.

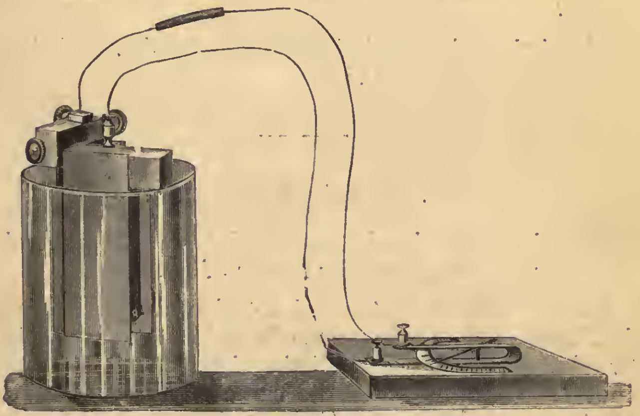


Fig. 32.

Experiments.

1. Take a piece of roll sulphur, and of iron or copper wire; hold them, one in each hand, and dip them into a vessel containing boiling water. Note the one along which the heat travels quickest to the hand.

2. Insert these same substances, in turn, into the circuit of a galvanic battery. Attach a galvanometer to the circuit, as in Fig. 32, and by its aid note which substance acts as a conductor of electricity. A toy compass will do for the experiment, as well as an expensive galvanometer.

3. Compare the surface appearance of copper, silver, and other metals with that of sulphur and phosphorus.

QUESTIONS.

1. How does the capacity of copper or iron, to conduct heat and electricity, compare with that of sulphur?

2. Supposing the above-named elements to be fair types of metals and non-metals respectively, mention two or three points of difference between the two kinds of elements.

3. How do the hydrates of metals and of non-metals differ from each other? [illegible]

4. Mention some valuable practical properties which certain metals possess?

No distinct line of difference can be drawn between metals and non-metals.

§ 55. Chemical Notation.

Chemical Notation is the art of designating chemical elements or compounds by means of *symbols.*

In our present system of notation, a **symbol** is the first letter of the Latin name of an element. Sometimes two letters are used to distinguish one element from another beginning with the same letter; for example, H stands for hydrogen, C for carbon, and Cl for chlorine.

Each symbol stands also for a definite weight of each element, called its *atomic weight; e. g.*, H always stands for 1 part by weight of hydrogen, and Cl for 35.5 parts by weight of chlorine.

The symbol of an element stands for three* distinct things :—

(1) The name of the element ;
(2) One atom of the element ;
(3) The atomic weight of an element.

Thus, the symbol O stands for : (1) the name oxygen ; (2) one atom of oxygen ; (3) 16 parts by weight of oxygen.

A small numeral written at the lower right hand corner of a symbol denotes that the atom is doubled, tripled, etc., *e. g.*, O_2.

If we take 1 centigram as our unit of weight, then H signifies 1 centigram of hydrogen, and O, 16 centigrams of oxygen, and so on.

§ 56.—Formulae.

A chemical formula consists of two or more symbols written side by side, and denotes that the elements for which the symbols stand have united to form a chemical compound. The symbol of the most electro-positive constituent of a compound stands first in its formula.

The formula of a compound substance stands for :

(1) The name of the compound ;
(2) One molecule of the compound ;
(3) The molecular weight of the compound ;

The molecular weight of a compound is found by taking the sum of the atomic weights of its constituent elements.

A numeral placed before a formula multiplies every atom and atomic weight in it, as far as the first comma, plus sign, or period. For example, in $4H_2O$, the 4 multiplies both the atoms and the atomic weights.

[*A symbol stands for these three things taken collectively—not separately. The same remark applies to formulae.]

The formula, H_2O, expresses the following facts:

(1) That water consists of hydrogen and oxygen;
(2) That its molecule consists of 3 atoms: 2 of hydrogen and 1 of oxygen;
(3) That its molecular weight is 18.

The formula, $C_{12}H_{22}O_{11}$, signifies:

(1) That cane-sugar consists of carbon, hydrogen, and oxygen;
(2) That its molecule consists of 45 atoms: 12 of carbon, 22 of hydrogen, and 11 of oxygen;
(3) That its molecular weight is 342

The molecule, O_2, consists of 2 atoms.
" " H_2, " " "
" " P_4, " 4 "
" " C_2H_6O, (common alcohol) consists of 9 atoms.

Remembering your studies on the physical characters of oxygen and hydrogen, § 42, you will easily understand how each of these four molecules, although differing widely from each other in the number of atoms composing them, occupies the same space or volume in the gaseous condition, according to the law of Avogadro.

§ 57.—Empirical Formulæ.

An empirical formula expresses the proportions by weight in which the constituents of a substance unite to form it, and the proper empirical formula for each compound substance is fixed by accurate chemical analysis. We have now to consider how empirical formulæ are calculated after the substances have been carefully analysed by methods which the young student cannot as yet understand.

A substance upon analysis yields the following percentage composition: Potassium, 28·73; oxygen, 47·02; sulphur, 23·52. Calculate its empirical formula.

In solving this question proceed as follows:

Potassium ....$28{\cdot}73 \div 39{\cdot}1 = {\cdot}73$
Hydrogen................. $\cdot73 \div 1 = {\cdot}73$
Oxygen ....$47{\cdot}02 \div 16 = 2{\cdot}93$
Sulphur ....$23{\cdot}52 \div 32 = {\cdot}73$

Now the smallest of these quotients is ·73, and dividing each of them by this, we obtain one for K, one for H, one for S, and four for O. The empirical formula is therefore, $KHSO_4$, and the substance is hydro-potassic sulphate.

Why do we divide the percentage of each substance by its atomic weight? Evidently, to get the number of atoms or parts of an atom of each constituent of the compound, if 100 parts of the compound be taken.

To solve all similar problems observe the following rule:

1. Divide the percentage amount of each constituent element by its own atomic weight.

2. Divide each of the quotients thus obtained by the lowest of them and the numbers obtained will express the number of atoms of each element in the compound.

The problems in the following exercise can all be solved in a similar manner.

EXERCISE.

1. Carbon. 42·86; oxygen, 57·14. Ans. CO, carbon monoxide.
2. Hydrogen, 2·73, chlorine, 97·27. Ans. HCl, hydric chloride.
3. Hydrogen, ·83; sodium, 19.17; sulphur, 26·66; oxygen, 53·33. Ans. $HNaSO_4$ hydro-sodic sulphate.
4. Sodium, 39·31; chlorine, 60·69. Ans. NaCl, sodic chloride.
5. Nitrogen, 82·35; hydrogen, 17·65. Ans. NH_3, ammonia.
6. Phosphorus, 91·17; hydrogen, 8·83. Ans. PH_3, phosphine.
7. Carbon, 26·67; hydrogen, 2·22; oxygen, 71·11. Ans. $C_2H_2O_4$, oxalic acid.
8. Carbon, 75; hydrogen, 25. Ans. CH_4, marsh gas.
9. Carbon, 12: calcium, 40; oxygen 48. Ans. $CaCO_3$, calcic carbonate.

§ 58.—Percentage Composition.

Sometimes we are given the formula of a substance and are asked to calculate the percentage composition. This is easily done. Proceed as follows :—

1. Find the molecular weight of the compound by taking the sum of the atomic weights of its constituents.

2. Then apply the "Rule of Three," or the "Unitary Method" as in the following example.

Calculate the percentage composition of sulphate of copper, $Cu\,SO_4$.

Copper		63·5
Sulphur		32
Oxygen	(16 × 4)	64
		159·5

If, by weight in 159½ parts of sulphate of copper, there are 63½ parts by weight of copper, how many parts by weight of copper will there be in 100 of sulphate.

Sulphate
159½ yield .... 63½ of copper.

∴ 1 will yield .... $\frac{63\frac{1}{2}}{159\frac{1}{2}}$

and ∴ 100 " $\frac{63{\cdot}5}{159.5} \times \frac{100}{1} = 39{\cdot}81$ per cent.

The percentage of sulphur and oxygen in this compound may be found in the same way.

EXERCISE.

What is the percentage composition of each of the following named substances :

1. Arsenious oxide, As_2O_3. ANS. 75·75 arsenic ; 24·25 oxygen.
2. Chloride of gold, $AuCl_3$. ANS. 35·09 gold ; 64·91 chlorine.
3. Arseniuretted hydrogen, AsH_3 ANS. 96·15 arsenic ; 3·85 hydrogen.
4. Potassium ferrocyanide, $K_4\,FeC_6N_6$. ANS. 42·39 potassium ; 15·22 iron ; 17·56 carbon ; 22·83 nitrogen.
5. Magnesium sulphate $MgSO_4$. ANS. 20 magnesium ; 26·67 sulphur ; 53·33 oxygen.

§ 59.—Rational Formulæ.

A rational formula besides expressing the proportions by weight in which the elements are united, expresses also the way in which the elements are supposed to be grouped within the molecule of a compound. For example,

$\left.\begin{matrix}HO\\HO\end{matrix}\right\}SO_2$ is the rational formula for sulphuric acid.

§ 60.—Atomicity or Valence.

We have already seen that the elements are divided into two classes—Metals and Non-Metals.

There is another classification of the elements that is even more important than the foregoing. The basis of this second classification is the number of atoms of hydrogen or chlorine that will unite with *one* atom of any of the other elements. According to this principle all the elements may be divided into six classes. To the first class will belong all those which unite atom for atom with hydrogen. Such elements bear the name of MONADS. To the second class will belong all elements one atom of which will unite with two of hydrogen. Such elements are called DYADS. If elements unite with three, four, five, or six atoms of hydrogen they are termed, TRIADS, TETRADS, PENTADS, and HEXADS, respectively. The above statement is true in a *general* way if the union of hydrogen or chlorine with other elements takes place by volume instead of by atoms.

The following table gives this classification in detail so far as the most important elements are concerned. It will be found useful to commit it to memory :—

	MONADS.	DYADS.	TRIADS.	TETRADS.	PENTADS.	HEXADS.
NON-METALS.	Bromine.	Oxygen.	[illegible]	Boron.	Nitrogen.	Sulphur.
	Chlorine.		[illegible]	Carbon.	Phosphorus.	
	Fluorine.			Silicon.		
	Hydrogen.					
	Iodine.					
METALS.	Potassium.	Calcium.	Bismuth.	Aluminum.	Arsenic.	Chromium.
	Sodium.	Copper.	Gold.	Cobalt.	Antimony.	
	Silver.	Magnesium.		Iron	Bismuth.	
	[illegible]	Mercury.		Lead.		
		Strontium.		Manganese.		
		Zinc.		Nickel.		
				Platinum.		
				Tin.		

Several of these elements exhibit varying atomicities but when they do so, they change by two degrees at a time. For example, triads may act as monads, or as pentads.

All the above pentads also act as triads.

Manganese and iron sometimes act like hexads, and are classified with them.

Sulphur usually acts as a dyad, and forms compounds resembling those of oxygen in chemical properties.

The **Atomicity** of an element is the number of atoms of hydrogen which an atom of the element will displace in a compound. Chlorine also may be taken as the unit by which to measure the valency of elements.

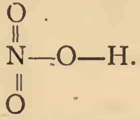
Artiads—atoms of elements of even atomicity are termed artiads.

Perissads—atoms of elements of uneven or odd atomicity are called perissads.

§ 61.—Graphic Formulæ.

Graphic formulæ express more fully than rational formulæ the manner in which we suppose atoms to be associated in forming compounds. For example, the graphic formula for water is H—O—H.

For nitric acid the empirical formula is HNO_3; the rational formula is NO_2 (OH), and the graphic formula,



$$\underset{\underset{O}{\|}}{\overset{\overset{O}{\|}}{N}}\text{—}O\text{—}H.$$

In graphic formulæ, the lines indicate the manner in which the atomicities of each element are disposed of; nitrogen being joined to the other elements by five links, oxygen by two, and hydrogen by one.

§ 62.—Chemical Nomenclature

is the system of naming chemical compounds.

In naming **binary compounds**, or compounds of two elements, we attach both prefixes and affixes to the the names of the elements. Generally the most *electro-positive* constituent stands first in the formula, that is, on the left-hand side.

1. **-ic** is generally attached to the name of the first element.

2. **-ide** is attached to the name of the second element.

For example, KCl is named potass-IC chlor-IDE.

-uret is an old ending, sometimes used instead of -IDE.

3. The number of atoms of the right-hand or *electro-negative* element is indicated by using the prefixes **mono-**, **di-**, **tri-**, **tetra-**, **penta-**, and **hex-**.

4. A special mode of designating the compounds of oxygen is by using the endings **-ous** and **-ic**, both being

attached to the first element; the former, when a smaller quantity of oxygen enters into the compound; the latter, when a larger quantity. Etymologically, the affix -OUS signifies full of, *e.g.* ferrous, full of iron. The affixes -OUS and -IC are used in a similar way to denote binary compounds of chlorine and sulphur with other elements. They are used also in naming acids, as we shall find later on.

5. Sesquioxide and suboxide are old terms.

EXERCISE.

Name the following binary compounds: $NaCl$, CuS, N_2O, N_2O_2; N_2O_3, N_2O_4, N_2O_5, HCl, H_2S, CO, CO_2, $CaCl_2$, CS_2, P_2O_5, CuO, HgO, K_2O, FeO, Fe_2O_3, As_2O_3, As_2O_5.

---

§ 63.—Equations.

A chemical equation consists of signs and formulæ, and expresses the fact that certain substances do, of themselves, or by means of some force applied to them, decompose and re-arrange their atoms so as to form other substances.

For example the chemical equation—

$$\begin{array}{c} H_2 + O = H_2O \\ \underbrace{2 + 16} \quad 18 \end{array}$$

expresses the fact that 2 centigrams of hydrogen unite with 16 centigrams of oxygen and form 18 centigrams of water. Also, that 2 volumes of hydrogen unite with 1 volume of oxygen and form 2 volumes of steam or water gas. In the same way, the equation

$$\begin{array}{c} CaCO_3 + 2\ HCl = CaCl_2 + H_2O + CO_2 \\ \underbrace{100 + 73} \qquad \underbrace{111 + 18 + 44} \end{array}$$

may be thus translated: mix 100 grams (or ounces) of marble with a solution of 73 grams of hydrochloric acid and they will yield 111 grams of calcic chloride, 18 grams of water, and 44 of carbonic anhydride.

The atoms on one side of an equation must all be accounted for on the other. The chemical equation thoroughly understood, enables us to calculate the amount of material required to produce a given weight of any substance; or, the quantity of the substance produced by the decomposition of a known weight of the material. Its importance, therefore, in working out chemical problems cannot be over-estimated.

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§ 64.—**Signs.**

The sign +, *plus*, placed between the formulæ of two substances, means that the two substances are *mixed* together.

The sign =, in chemistry, means "yields."

Chemical union, as has already been explained, may take place in certain porportions by *Weight*, or when the substances exist in the gaseous condition in certain proportions by *Volume;* and in both cases, the reactions taking place may be represented by an equation. The first equation may be reversed in order, thus :—

$$H_2O = O + H_2,$$

and now the meaning will be this :—18 centigrams of the compound substance, water, may be decomposed into two elementary ones, viz., 16 centigrams of oxygen, and 2 centigrams of hydrogen.

QUESTIONS.

1. Complete the following chemical equations and explain their meaning. They represent, in part, the reactions that took place in all the experiments described in the preceding part of this book. Three are completed as examples: the student should try to complete the rest, and in doing so, should apply the principle of atomicity.

(1.) $HgO = Hg + O.$

(2.) $Pb_3O_4 = 3PbO + O.$

(3.) $KClO_3 + MnO_2 = KCl + MnO_2 + 3O$.

(4.) $C + 2O = CO_2$

(5.) $3Fe + 4O = Fe_3O_4$

(6.) $S + 2O = SO_2$

(7.) $2P + 5O = P_2O_5$ -

(8.) $Zn + O = ZnO$

(9.) $SO_2 + H_2O = H_2SO_3$

(10.) $P_2O_5 + H_2O = 2HPO_3$

(11.) $N_2O = 2N + O$

(12.) $N_2O_3 + H_2O = 2HNO_2$ ([illegible]

(13.) $2K + O = K_2O$

(14.) $K_2O + H_2O = 2KHO$ ([illegible]

(15.) $Mg + O = MgO$ (magnesia)

(16.) $MgO + H_2O = Mg(HO)_2$ magnesium [illegible]

(17.) $H_2SO_4 + Zn = ZnSO_4 + H_2$ (zinc sulphate)

(18.) $2HCl + Zn = ZnCl_2 + H_2$

(19.) $3Fe + 4H_2O = Fe_3O_4 + 8H$

(20.) $CuO + 2H = Cu + H_2O$

(21.) $H_2O + K = KHO + O$

(22.) $H_2O + Na = NaHO + H$

§ 65.—Molecular Equations.

If it be true that atoms do not exist alone, "perfect consistency would require that no equation should ever be written in such a manner as to represent less than a single molecule of an element in a free state as either entering into or issuing from a chemical reaction. Thus instead of $2H + O = H_2O$, we should write $2H_2 + O_2 = 2H_2O$. We do not propose to conform to this theoretical rule for three reasons. First, because many equations, representing chemical reactions, must be multiplied by two, in order to bring them into conformity with the hypothesis concerning molecular structure; the equations are thus rendered unduly complex; secondly, because, in undertaking to make chemical equations express the molecular constitution of elements and compounds, as well as the equality of the atomic weights on each side of the sign of equality, there is imminent danger of taking the student away from the sure ground of fact and experimental demonstration into an uncertain region of hypothesis based only on definitions and analogies; thirdly, because we are ignorant of even the probable molecular symbol of most of the elements. Of all the elementary substances recognized, we have reason to believe that eleven, when in the gaseous state, are made up of molecules each containing two atoms; that two (arsenic and phosphorus) contain four atoms; and that three (mercury, cadmium, and probably zinc) contain only a single atom to the molecule. Of the molecular structure of the remaining elements, numbering three-fourths of the whole, we at present know nothing."

CHAPTER XI.

§ 66.—Carbon.—C, A't W't, 12.

We have learnt that neither air nor water is a simple body. We look around us, and seeing many common substances, such as wood, sugar, meat, and bread, we are naturally led to inquire whether these are simple or compound. It getting an answer, we shall be introduced to one of the most important of elements, and one which is extensively distributed throughout organic compounds.

Experiments.

1. Partly fill a narrow test-tube with white paper, dry saw-dust, or pieces of dry wood. Hold the tube in a nearly horizontal position with its lower end in a lamp flame. As the heating goes on, notice whether any odor is evolved. Place separate pieces of blue and red litmus paper within the mouth of the tube. When nearly all action has ceased, turn out and examine what remains in the test-tube.

2. Place a little of the residue on platinum foil, or on sheet of mica, and heat over a lamp flame for some time. Observe what is left.

3. Clean as well as you can the test-tube used in the preceding experiment, and then repeat it, using a piece of woolen cloth, silk cloth, or dry lean meat.

4. Heat on a sheet of mica the residue obtained in experiment 3.

5. Heat some sugar upon a sheet of mica

§ 67.—Properties of Carbon.

The following experiments illustrate the principal properties of carbon, so far as these can be illustrated by simple experiments.

Experiments.

1. Place a wet filter paper inside a funnel, and then cover the inside of the filter paper with a thick coating of animal charcoal, or bone black. Now filter through the paper a wine-glass full of ale or porter.

2. Place a piece of charcoal in a test-tube and then pour upon it a little strong sulphuric acid. Observe whether the charcoal changes in any way. Try whether an alkali will produce any change in the charcoal.

3. Wet the inside of a large test-tube with liquor ammoniae. Now drop into the tube some wood charcoal previously heated in a covered crucible. Cork the tube and after a few minutes remove the cork and ascertain by smelling it whether there is any ammonia left in the test-tube.

QUESTIONS.

1, How is wood charcoal prepared? Animal charcoal? Is wood an element? Give reasons for your answer.

2. Mention one difference between the liquid produced in the destructive distillation of wood and that obtained in the destructive distillation of lean meat.

3. Write out an account of the properties of carbon as illustrated in the foregoing experiments.

4. Specify as many of the uses of charcoal as you can think of.

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§ 68.—Allotropic Forms.

Charcoal is an impure form of carbon ; graphite (plumbago or black-lead) is another ; but diamond is the element in an almost pure state and crystalline form. If you can procure a specimen of diamond, you should compare it with charcoal and with graphite, noting differences in *(a)* the mark they make on paper, *(b)* their color, *(c)* hardness, *(d)* specific weights, *(e)* power of conducting heat, and *(f)* power of conducting electricity.

(a) **Pit Coal** is composed of carbon in large proportion, oxygen in smaller quantity, hydrogen in smaller, nitrogen in still less, and a variable proportion of saline and earthy matter. It has been formed by the submersion of huge forests under the sea long ages ago, the wood being slowly changed into coal by the combined action of the pressure of water upon it, and of moderate heat from the interior of the earth.. There are two principal varieties of pit coal, **anthracite** and **bituminous**. Anthracite coal contains about 90 per cent. of carbon. When bituminous coal is heated in closed iron cylinders free from air, a large quantity of *gas* and *tar* is formed, containing the oxygen, hydrogen, nitrogen, and some of the carbon of the coal ; the residue is called **coke**. This is how coal gas is manufactured. This process of destructive distillation may be applied to wood also, when an inflammable *gaseous* product will be given off ; **wood-tar, vinegar,** and **wood-naphtha,** are the liquid products ; the black porous mass left behind is called wood *charcoal.*

(b) **Lamp-black,** the basis of printer's ink, is another form of carbon.

(c) **Animal Charcoal** or *ivory black* is made by heating the bones and flesh of animals in iron retorts.

The word **Allotropism** is used to express the fact that some elements exist in very unlike states or with very different properties, but all the while preserve their fundamental chemical identity.

PROBLEMS.

1. Ascertain what effect animal charcoal has upon solutions of *(a)* litmus, *(b)* indigo, *(c)* potassium permanganate.

2. From a sample of coarse brown sugar prepare some that will be pure and white.

3. Devise an experiment to ascertain how long charcoal will retain its decolorizing and deodorizing power when in use as a deodorant.

4. Using apparatus similar to that in Fig. 11, devise an experiment to show that when air or oxygen is passed over cold charcoal, no change takes place, but when, over burning charcoal a change does take place. To show this change, pass the air or gas from the heated charcoal through clear lime-water.

CHAPTER XII.

§ 69.—Carbon Dioxide.

We have been hitherto chiefly engaged in eliminating oxygen, nitrogen, hydrogen, and carbon from compounds which contain them, and in studying their prominent properties. We must now prepare and study some of the compounds formed by the union of these elements.

Carbon dioxide (Carbonic Anhydride, carbonic acid gas, choke damp); Formula $C\ O_2$ *; molecular weight, 44; specific weight, 22. 11.2 litres weigh 22 grams.*

Experiments.

1. Take the apparatus used in preparing hydrogen (Fig. 19), and in it place some powdered limestone or white marble. Then cover the marble with water, and pour down the thistle-tube some hydrochloric acid. Collect, over the pneumatic trough, two or three beakers full of the gas that comes off, and perform the following experiments:

a. Remove one of the beakers in the usual way and place it mouth upward on the table. Slip to one side the glass cover and insert a burning taper in the gas. St.

b. Remove the second beaker from the trough. Pour a little *fresh* water into it; cover with the palm of the hand and shake vigorously. Note the effect on the hand; then test the water with blue litmus. Taste it.

QUESTIONS AND PROBLEMS.

1. Devise an experiment to find out whether carbon dioxide is heavier or lighter than air.

2. Write out an account of the properties of carbon dioxide, noting its *(a)* color, *(b)* solubility in water, *(c)* effect on a burning candle, *(d)* effect on lime-water.

3. Devise experiments to show that carbon dioxide is formed in *(a)* burning a candle, *(b)* burning a lamp, *(c)* burning coal gas.

4. Show, by means of the lime-water test, that this gas is given off during respiration.

§ 70.—Decomposition of Carbon Dioxide.

Before this you will probably have learned that the gas formed in problem 4, section 68, was carbon dioxide. In that case, the action was one of combination; but we can easily reverse the process, and by using a modification of the same apparatus, bring about the decomposition of the gas.

Experiment.

Pass a stream of the gas first through a tube containing calcium chloride and then through another tube containing pieces of the metal sodium gently heated by the flame of a spirit lamp or gas burner. How does this experiment prove the presence of carbon in this gas?

§ 71.—Other Properties.

Carbonic anhydride may be condensed to a liquid by applying cold and a pressure of 40 atmospheres, or it be generated as a liquid in strong iron tubes; it may even be frozen to a snow-white solid, which, when mixed with ether, produces a freezing mixture of – 75°. It accumulates in old pits, wells, and mines, and issues sometimes from fissures in the earth and from the craters of volcanoes. It exists very abundantly in combination with lime and magnesia.

CHAPTER XIII.

§ 72.—Carbon Monoxide.

Besides the compound already studied, carbon and oxygen form another—not so well known, nor so important, as carbon dioxide—but still important enough to merit some consideration.

Carbon Monoxide (Carbonic Oxide): Formula, CO; molecular weight, 28; specific weight, 14. 11.2 litres weigh 14 grams.

Experiment.

Into a Florence flask put 8 or 10 grains of oxalic acid, and about 50 c. cm. of sulphuric acid. Fit with a tight cork and tube and attach, as in Fig. 33, to a wash bottle containing a strong solution of caustic potash. From the wash bottle a delivery tube should pass to the pneumatic trough. Apply heat cautiously to the flask, regulating it so that the gas may come off in a steady stream. After the air has been expelled from the apparatus, collect three bottles of the gas, and allow these bottles to stand over water for some time before using them. Meanwhile substitute for the delivery tube one whose end has been drawn to a fine point. Apply a lighted match to the jet.

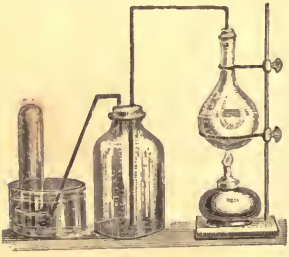

Fig. 33.

(a). Raise one of the bottles of gas from the water, and apply a lighted taper to its mouth. burns with [illegible]

(b). Try to pour the gas from one bottle to another, then test the result with a lighted taper.

(c). Purify thoroughly the gas in a third bottle, by shaking it up well with caustic potash or caustic soda solution, then test the gas with clear lime-water. [illegible]

QUESTIONS AND PROBLEMS.

1. Write out an account of the physical and chemical properties of this gas, so far as they have been illustrated in your experiments.

2. If carbon monoxide burns in air, try to find out whether vapor of water is formed during the process of combustion? Whether carbon dioxide is formed?

3. On what data can you conclude that carbon dioxide contains more oxygen than carbon monoxide?

4. Hydrogen and carbon monoxide are both combustible. Can you conclude from this resemblance that carbon monoxide and oxygen will form an explosive mixture? Test the accuracy of your conclusion.

5. If we can form carbon dioxide from carbon monoxide by supplying it with oxygen, it ought to be possible to prepare carbon monoxide from carbon dioxide by supplying it with carbon. Test the accuracy of this induction by using apparatus similar to that in Fig. 11, and passing dry carbon dioxide over red-hot charcoal.

§ 73.—Other Properties.

Carbon monoxide is an exceedingly poisonous gas; is slightly soluble in water. It is formed by ordinary combustion in our coal stoves, causing the characteristic "blue blazes" that are seen flickering over the top of coal fires (explain their formation); also by passing steam over red-hot charcoal.

§ 74.—Law of Multiple Proportion.

We have had numerous illustrations of the truth of the law of definite proportion. In studying the compounds of carbon and oxygen we meet with the first illustration of the law of multiple proportion. Accurate analyses of these compounds show that their composition is as follows :—

Carbon dioxide.	Carbon monoxide.
C = 12	C = 12
O_2 = 32	O = 16

By comparing the proportions by weight in which oxygen unites with carbon in each compound, the student can easily understand and formulate the law for himself.

CHAPTER XIV.

§ 75.—Units of Volume.

[illegible] H weigh / gram

Different writers on Chemistry adopt different units of volume for measuring gases. [illegible]

I. Dr. Hoffman's unit.—This is one *litre.* The weight of this volume of hydrogen at O°C. and 760 millimetres mercury he calls one *crith* (·08986 grams).

II. Dr. Williamson's unit.—This is 11.2 *litres*, which at O°C. and 760 millimetres mercury weighs one gram in the case of hydrogen, and in the case of all other elementary gases the number of grams represented by their atomic weight.

III. Dr. Reynold's unit is one *vol.* or 112 cubic centimetres. Its weight is one centigram in the case of hydrogen; and in the case of all other elementary gases it is the number of centigrams represented by their atomic weights.

As can be seen at a glance, this unit is exactly the hundredth part of Dr. Williamson's.

According to Dr. Williamson's unit :—

11.2 litres of oxygen weigh 16 grams
11.2 " nitrogen weigh 14 grams
11.2 " hydrogen weigh 1 gram

Before illustrating the application of the equation in making chemical calculations, it may be necessary to remind the student of the formulæ employed in reducing gases to standard temperature (0° C.), and pressure (760 m.), according to the laws of Boyle and of Charles.

The formula for applying Boyle's law to the measurement of gases is as follows :—

$$\text{Volume of any gas at } 0^\circ \text{ C.} = \frac{273 \times \text{volume of gas at } t^\circ \text{ C.},}{273 + t^\circ}$$

where t° denotes the temperature in degrees C., at which the gas is measured.

EXERCISE.

1. 1000 cubic centimeters of a gas at 12° : what volume at 75° ?

2. 50 cubic centimeters of a gas are measured at 10° C. : how much will it measure at 24° C. ?

3. I measure 100 cubic centimetres of a gas at 12° C. ; the gas is heated until it becomes 145 cubic centimetres. Find the temperature to which the gas has been changed.

4. 75 c. c. of hydrogen at 20° : what volume at − 10° ? = 67 [illegible]

5. 1500 c. c. of oxygen at 50° ; at what temperature will it measure 1000 c. c. ? = − 57 $\frac{2}{3}$

The formula for calculating changes in the volume of a gas due to change of pressure is

$$V = \frac{p\,v}{760}$$

where v denotes the observed volume, p the observed pressure, and V the volume at the standard pressure of 760 millimetres of mercury.

EXERCISE.

1. 100 c. c. of air when barometer = 750 mm. : find the volume at standard pressure. 98 $\frac{1}{14}$ c.c

2. 250 c. c. of hydrogen when bar. = 765 mm. : find the volume when the bar. = 745 mm. = 256 $\frac{[illegible]}{149}$

3. 150 c. c. of oxygen when bar. = 740 : find the volume at 770 mm. = 141

4. 200 c. c. of air at 10° C. and 750 mm. barometer pressure : what will be its volume at standard temperature and pressure ? = 190 [illegible]

5. 1000 c. c. of gas at 18° and 765 mm. : find the pressure when the temperature has fallen to 8° and the gas has expanded to 1250 c. c. ?

= 587 $\frac{133}{291}$

§ 76.—Chemical Calculations.

When the volume of a gas is mentioned in the following problems, it is supposed to be at the standard temperature and pressure unless the contrary is expressly stated.

What weight of hydrogen can be evolved from 392 grains of sulphuric acid?

To solve this and all similar chemical problems, pupils *must* know the chemical equations representing the reaction that takes place when an elementary or compound substance is evolved from others. The composition of sulphuric acid is:—

$$\begin{aligned} H_2 &= \ \ 2 \\ S &= 32 \\ O_4 &= 64 \\ \hline \text{Total} &= 98 \end{aligned}$$

Now the question is, if 2 grains (or oz., &c.) of hydrogen can be obtained from 98 of sulphuric acid, how many can be obtained from 392? Rule of Three: 98 (of H_2SO_4) : 392 (of H_2SO_4) : : 2 (of H) : x (of H).

Ans.—8 grains.

$$\begin{array}{ccccccc} 65 & & 98 & & 161 & & 2 \\ Zn & + & H_2SO_4 & = & ZnSO_4 & + & H_2 \end{array}$$

I.

1. What weight of zinc sulphate and hydrogen will be formed by acting on 100 lbs. of zinc with 98 lbs. of sulphuric acid? = 3

2. How many grams of hydrogen will occupy 224 litres at the standard temperature and pressure? = 20

3. Steam is passed through a tube containing red-hot iron filings, 18 litres of hydrogen pass out at the other end. What volume of steam entered the tube, and how much are the iron filings increased in weight?

4. What weight of water and potassium must be taken to produce 561 ounces (Troy) of caustic potash? What weight and volume of hydrogen will be produced? 180 of water & 391 Potassium

5. How much sulphuric acid and zinc must be taken to form 112 litres of hydrogen at 7° C? $477\frac{3}{4}$ H_2SO_4 & $316\frac{7}{8}$ of Zn

6. In 285 grains of caustic potash how many grains of potassium? of hydrogen?

7. What weight of sodium must be taken to obtain 20 grains of hydrogen from a litre of water?

8. A reservoir of hydrogen gas holds 89.6 litres. What weight of water will be formed in burning the gas in air? What volume of air will be required for the combustion, assuming that oxygen forms $\frac{1}{5}$ of the volume of air?

9. How many pounds of potassic chlorate must be taken to obtain 144 lbs. of oxygen?

10. I want 220 grams of oxygen. If I obtain it from potassic chlorate, how much of it must I use? If from water, how much? If from mercuric oxide, how much?

11. A gas bag is capable of containing 56 litres, how much potassic chlorate must be taken to procure enough oxygen to fill it at 35°C. and 750 mm.? $\frac{273}{308} \times \frac{56}{1} \times \frac{750}{760} \times \frac{16}{11.2}$

12. 25 litres of oxygen are exploded with 36 of hydrogen. What volume of gas (if any) remains? What volume of steam is produced? 36 And what is its weight? $\frac{36}{1} \times \frac{273}{273} \times \frac{11.2}{8}$

13. How much oxygen can be obtained from 435 grams of manganese dioxide by heating it to a red heat? What volume will it occupy at 30°C. and 780 mm.?

14. What volume will 80 grams of oxygen occupy at the standard temperature and pressure?

II.

1. How much potassium will be required to decompose 110 grams of carbon dioxide?

2. If 10 litres of carbon dioxide be passed ever red hot charcoal, what gas, and how many litres of it, will be formed at 30° C? What weight of it?

3. 20 litres of carbonic oxide are burned in oxygen gas. What gas is produced, what volume at 40°C. and weight of it?

4. How much carbon can be obtained from 264 grams of carbon dioxide?

5. What volume of oxygen at 10°C. is required to burn 66 grams of carbon?

6. In question 5, what volume of air would be needed for its combustion at 0°C.?

7. What volume do 110 grams of carbon dioxide occupy at 760 m. pressure and 0°C.?

8. What volume do 140 grams of carbonic oxide occupy at standard temperature and pressure?

9. What weight of carbon dioxide can be obtained from 250 grains of pure limestone by treating with hydric chloride?

10. What weight of carbonate of lime and hydric chloride must be decomposed to produce 352 grams of carbon dioxide?

11. What volume will 98 grams of carbonic oxide occupy at 720 m. pressure and 40° C.?

12. If 270 grams of oxalic acid be decomposed by sulphuric acid, find the volume of the gases produced at 750 m. pressure and 20° C.

CHAPTER XV.

§ 77.—Hydrocarbons.

A large number of compounds of carbon and hydrogen are known under the general name of hydrocarbons. So numerous are these compounds, and those which carbon forms with oxygen, nitrogen, sulphur, and phosphorus, that their mere names would fill a small volume. The study of the carbon compounds forms a distinct branch of chemistry under the name of organic chemistry. Formerly this name included the study of those compounds which, it was supposed, were formed only by the agency of life; but it was soon found that there was no essential difference between chemical substances whether of animal, vegetable, or mineral origin. The division of chemistry, therefore, into organic and inorganic is a pure matter of convenience. The last named division of the science treats of the composition and properties of air, earth and water.

Marsh gas is the first of a series of hydrocarbons known as the marsh gas series. Each member of it differs from the following one by CH_2. There is also a difference of 30° between the boiling points of successive members. All are inflammable. There is also a regular increase or decrease of other physical properties. Such series are called homologous series. The general algebraic formula for the series is $C_n H_{2n+2}$.

---

§ 78.—Methane.

Methane (Marsh Gas, Light carburetted hydrogen, "Fire-damp"), CH_4; *molecular weight*, 16; *specific weight*, 8. 11·2 *litres weigh* 8 *grams.*

Experiments.

1. Take a hard glass test-tube or Florence flask, and fit with a cork and fine delivery tube. Place in the test-

tube 2 grams of acetate of sodium $NaC_2H_3O_2$, 8 grams of sodium hydroxide and 2 grams of finely powdered quicklime CaO. Heat. After collecting a beaker or two of the gas, light the jet and observe the color of the flame.

2. Fill a small soda water bottle with a mixture of one part of the gas and two parts of oxygen. Ignite the mixture. Express the reaction by an equation. $CH_4 + 2$ =

3. Take a stoppered bottle and fill it with a mixture of equal volumes of marsh gas and chlorine. Expose to sunlight for a day, then test the contents with blue litmus. Note any change in color. $CH_4 + Cl = HCl$

QUESTIONS AND PROBLEMS.

1. Devise an experiment to ascertain whether marsh gas is acid or basic in reaction.

2. Devise simple experiments to prove that the gas contains carbon and hydrogen as constituent elements.

3. Prove that the gas is lighter than air. How would you distinguish the gas from air?

4. Find out whether the gas is soluble in water.

---

§ 79.—Sources.

Methane is generated in marshes by the decomposition of vegetable matter containing carbon and hydrogen. Formed in coal mines also, and on being mixed with air and ignited, causes fearful explosions. To prevent these, Sir H. Davy invented his celebrated **Safety Lamp.**

---

§ 80.—Olefiant Gas.

This is the old name of another gas that may be taken as a type of a second series of the hydrocarbons. The general formula of the homologues of this series is C_nH_{2n}.

Ethylene (*Ethene. Olefiant gas. Heavy carburetted hydrogen*), C_2H_4; *molecular weight*, 28; *specific weight*, 14. 11·2 *litres weigh* 14 *grams.*

$H_2SO_4 + C_2H_6O = C_2H_4 + H_2O +$ [illegible]

Experiments.

1. Into a Florence flask pour 50 or 60 c.c. of strong sulphuric acid and half that volume of alcohol. Insert a tightly fitting cork and delivery tube. Place the flask on a retort stand and heat gently. After the air has all been expelled, collect two or three jars of the gas.

[The gas when prepared in this way is mixed with other hydrocarbons.]

2. Ascertain whether the gas will burn or not. [illegible] Has it any taste?

3. Devise simple experiments to prove that the gas contains hydrogen and carbon.

4. Find out whether the gas is heavier or lighter than air.

5. Remove a jar of the gas, let it drain well, then turn it mouth upward and place over it another jar of the same size filled with chlorine. After standing for some time, note whether the color of the chlorine changes. Observe closely what forms at the bottom of the lower jar—"Dutch liquid." oil [illegible]

6. Ascertain whether the gas will explode when mixed with air or oxygen. it will

§ 81. Coal Gas.

Coal gas is formed by the distillation of coal in large iron retorts. The process of manufacturing it may be illustrated by strongly heating some powdered coal in a common clay pipe. The mouth of the pipe should be closed with kneaded clay. The average composition of coal gas (for coal gas is really a mixture of many gases) is about as follows :—

Hydrogen	45·
Marsh gas	35·
Carbonic oxide	7·
Olefiant gas	4·
Butylene	2·4
Hydric sulphide	0·3
Nitrogen	2·5
Carbon dioxide	3·8
Total	100· vols.

§ 82.—Coal Tar.

Coal tar is one of the many bye products obtained in the distillation of coal. From it are manufactured the beautiful aniline dyes so extensively used in recent years.

CHAPTER XVI.

§ 83.—Combustion.

Now that we have learned something of carbon, hydrogen, oxygen, and a few of the compounds which they form, we are in a position to study somewhat more fully than we have hitherto done, the subject of combustion. The first question which we shall try to answer is : *What gives a flame its luminosity?*

Experiments.

1. Sprinkle into the flame of an alcohol lamp or Bunsen burner, some fine iron filings.

2.. Rub together over the top of any non-luminous flame, two pieces of charcoal.

3. Hold a piece of platinum wire or a piece of lime in a flame of hydrogen gas.

QUESTIONS AND PROBLEMS.

1. What change took place in the character of the flame in experiments 1 and 2?

2. From what source did the light emanate in all three experiments?

3. Mention one way of changing a non-luminous flame into a luminous one.

4. Explain the source of the black mark formed on a white saucer by holding it horizontally across the flame of a candle, or of a coal oil lamp. What is the black substance?

5. How do you explain the facts observed in your experiments with oxygen, viz., that sulphur and phosphorus give more light when burned in oxygen than in air.

§ 84.- -Incandescent Gases.

Besides depending upon the circumstances just illustrated, the illuminating power, or brightness of a flame depends also upon the density of the incandescent gases, which form the flame. Increase the density and the luminosity of the flame will increase also

§ 85.—Structure of Flame.

Experiments.

1. Spread out the wick of a candle or alcohol lamp, light it, and then thrust into the middle of the flame the phosphorus end of a friction match.

2. Float a small cork upon a little common alcohol or methylated spirits placed in the bottom of a small saucer. Place a few grains of gunpowder upon the cork, and then ignite the alcohol.

3. Bring a piece of wire gauze down horizontally upon the flame of a candle, of a coal oil lamp, or alcohol lamp.

QUESTIONS AND PROBLEMS.

1. If the phosphorus did not take fire while the end of the match was in the middle of the flame, explain why it did not?

2. Why did not the gunpowder explode the moment the alcohol took fire?

3. If a splinter of wood be placed across the flame of an alcohol lamp where should the wood begin to burn first? Base your conjecture upon the results of the three preceding experiments, and then test its accuracy by actual experiment.

4. What shaped mark will a candle flame make upon a piece of white letter-paper when pressed for an instant horizontally upon the flame? Test your conjecture. Press the paper down almost to the wick, and remove quickly.

---

§ 86.—Candle Flame.

Experiments.

1. Light a candle and observe its flame carefully. Note how many parts there are in it. Take a narrow bent glass tube, about four or five inches long, and thrust one end of it into the dark cone in the middle of the flame, as in Fig. 34. Try to light the vapors which rise through the tube. It is customary to describe four parts in the flame of a candle. If you observe these four parts, describe each one in your own words, and then compare your description with that in some standard work on chemistry.

Fig. 34.

2. Compare the number of zones in an alcohol lamp flame with those in a candle flame. Compare also the relative areas of the zones in each.

3. Examine in a similar manner the flames of a Bunsen burner, first, when the air holes are closed, then, when open. Note consequent changes in the temperature and luminosity of the flame. Pass in nitrogen, or carbon dioxide through the air holes instead of air, and compare changes thus produced with these that took place in admitting air.

QUESTIONS.

1. What changes does a blowpipe make in the flame of a candle, (1) as regards its size, and (2) as regards its heat ?

2. What effect has a lamp chimney on the luminosity and temperature of a coal-oil lamp flame ?

3. "Flame is incandescent gas." "Only gases burn with a flame." Examine these statements in the light of the experiments you have just made, or of observations which you have made on flames in coal or wood stoves.

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§ 87.—Temperature of Ignition.

The temperature at which a substance begins to burn is called its temperature of ignition, or its kindling point. Do all substances ignite at the same temperature ?

Experiments.

1. Pour a little carbon bisulphide into a large test-tube. Close with the thumb, and shake well, so that the vapor will fill the tube. Then warm a glass rod and place it in the vapor. If there be no result, heat the rod more strongly and again place it in the vapor.

2. Try to light coal gas with a rod similarly warmed ; then with an iron rod nearly red hot ; and lastly with an iron rod at a white heat.

QUESTIONS AND PROBLEMS.

1. How does the ignition point of carbon bisulphide vapor differ from that of coal gas ?

2. Why is it harder to light a coal fire than a wood one ?

3. If you cool a burning substance will it cease to burn ? Investigate this point by making a small helix of copper wire and holding it in a candle flame. Heat the helix to redness and again place in the flame.

4. Investigate the principle of the Davy lamp, used to prevent explosions in coal mines. The "fire damp" burns on the inside of the fire gauze which surrounds the flame. Why does not the gas outside take fire ?

5. Why does blowing on a flame "*put it out ?*"

CHAPTER XVII.

§ 88.—Sulphur.

Symbol, *S* ; *atomic weight*, 32 ; *specific weight in the form of crystals*, 2·05.

Sulphur, known also as brimstone, is found native in many volcanic regions; it occurs also in the ores of some of the common metals. Iron pyrites FeS_2, galena PbS, cinnabar HgS, gypsum $CaSO_4 + 2H_2O$, and heavy spar $BaSO_4$, all contain sulphur. For a description of the methods of obtaining the *roll* and *flour* sulphur of commerce, you must consult some good work on technological chemistry. The preparation of a small quantity of the element from iron pyrites may be illustrated as follows:—

Experiment.

Powder some iron pyrites and place in a *hard* glass test-tube. Hold the test-tube nearly horizontally in the lamp flame and heat the lower end strongly for some time. Test the residue with a magnet. Observe what collects in the cool part of the test-tube.

§ 89.—Allotropic Modifications.

Sulphur is known in three different forms; of these, two are modifications of the shape in which sulphur crystallizes; the third one, known as *plastic* sulphur, is prepared as described in the next experiment, and is used for making moulds of coins, &c.

Experiment.

Place 15 or 20 grams of sulphur in a large test-tube and heat slowly over a lamp flame until the sulphur boils. As the heating goes on, note changes in the appearance of the substance. When it begins to boil, pour it into a vessel of cold water. When it has cooled, remove and examine Keep in a dry place for a few days, and then examine again.

QUESTIONS AND PROBLEMS.

1. What physical property of sulphur is illustrated in the first experiment? Where does the sulphur collect?

2. The residue from the iron pyrites has the composition Fe_2S_3. Write the equation expressing the reaction that took place.

3. How long does plastic sulphur retain its plastic property?

4. Prepare some sulphur crystals by melting 40 or 50 grams of flowers of sulphur in a porcelain or earthenware dish, and then allowing it to cool slowly. As cooling goes on, break the crust which forms, and pour out the liquid sulphur.

5. Try to prepare some more crystals by dissolving sulphur in carbon disulphide and allowing the solution to evaporate.

§ 90.—Sulphur Dioxide.

Sulphur and oxygen unite to form four compounds, but of these only two are important.

Sulphur dioxide (Sulphurous Anhydride): Formula, SO_2 *; molecular weight, 64; specific weight, 32.*

Experiments.

1. Heat a little sulphur on a piece of tin or zinc plate until it catches fire. Note the color of the flame and the smell emitted.

2. Put 20 grams of copper turnings into a Florence flask, fitted up like a hydrogen generating apparatus. Just cover the copper with strong sulphuric acid, and heat the flask very carefully. Collect the gas that escapes, by upward displacement of air. Compare its odor with that of burning sulphur. Collect three jars of the gas for further experiments. Keep them covered.

(a). Pour some water into one of the jars and then shake. Is the gas soluble in water? Test the water with blue litmus solution. Taste it.

$Cu + H_2SO_4 =$ [illegible]

(b). Hang a red rose or other high-colored flower in the second jar. If any change takes place in the flower, remove it and place in pure air.

(c). Pour the third jar of sulphur dioxide into an "empty" jar with a burning candle at its bottom.

QUESTIONS AND PROBLEMS.

1. Devise a means of collecting sulphur dioxide when formed by burning sulphur in the air or in oxygen

2. How does the weight of this gas compare with that of air?

3. Try to liquefy this gas by passing it into a cool test-tube.

4. Explain why the fire in a chimney may be extinguished by burning some sulphur in a stove connected with the chimney.

5. Suspend a moistened wheat straw in a jar of this gas, and observe what occurs.

§ 91.—Bleaching Explained.

The bleaching power of sulphur dioxide is supposed to be due to its affinity for oxygen. According to this theory the dioxide unites with the oxygen of the water which is used to moisten the article to be bleached; the hydrogen of the water then combines with the coloring matter and forms colorless compounds.

§ 92.—Other Properties.

Sulphur dioxide dissolves in water at O°C to the extent of nearly 80 times its own volume; may be condensed to a colorless liquid of sp. weight of 2·21. It is also a good disinfectant in case of contagious diseases like scarlet fever, &c.; it prevents the decay of animal and vegetable substances that is, it is an *antiseptic*.

§ 93.—Sulphurous Acid.

We have already seen that the sulphur dioxide dissolves readily in water, forming with it a new compound which has a sour taste and turns blue litmus red. This new compound is called sulphurous acid. The reaction that takes place may be symbolized as follows :—

$$SO_2 + H_2O = \underbrace{H_2S_2O_3.}_{\text{Sulphurous acid.}}$$

An **anhydride** is an oxide capable of forming an acid when united with water.

§ 94.—Sulphur Trioxide, $S.O_3$.

Unlike the preceding one, this oxide is of no practical importance. It may be prepared by passing sulphur dioxide and oxygen over platinised asbestos in a highly heated porcelain tube. When this is done, there issues from the end of the tube hot dense white fumes of sulphur trioxide. This substance is a solid, having an intense affinity for water and being in consequence very unstable. The reaction is :—

$$SO_3 + H_2O = \underbrace{H_2SO_4.}_{\text{Sulphuric acid.}}$$

§ 95.—Sulphuric Acid.

This is the strongest of all acids and may appropriately be spoken of as the "king" of acids. It is used very extensively, either directly or indirectly, in all arts and trades, and is the most important re-agent which the chemist possesses. Its compounds include many of the most important substances in commerce.

Experiments. $SO_2 + H_2O$

Sulphurous anhydride, steam, air, and nitric oxide, are passed into an immense chamber. "The nitric oxide in presence of oxygen, immediately becomes nitrogen peroxide, and this, when mixed with sulphurous anhydride and a large quantity of water, furnishes sulphuric acid and nitric oxide. The sulphuric acid remains dissolved in the water, while the nitric oxide, by absorbing oxygen from the air, again becomes nitrogen peroxide; this combines with fresh sulphurous anhydride, which, when acted on by water, becomes sulphuric acid, the nitric oxide being again liberated, to go through the same series of changes with fresh portions of oxygen and sulphurous anhydride as long as any remain in presence of each other uncombined.

$$NO_2 + SO_2 + xH_2O = NO + (x-1)\, H_2O + H_2SO_4."$$

The process of manufacturing sulphuric acid may be illustrated by the use of the apparatus shewn in Fig. 35.

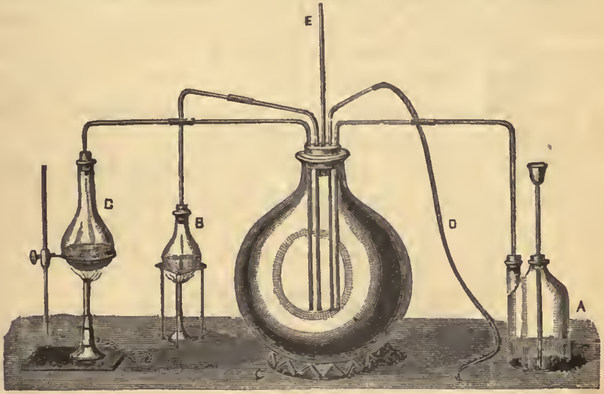

Fig. 35.

A is a flask containing copper filings and nitric acid for generating nitrogen dioxide and nitrogen trioxide.

B is a flask containing water for generating steam.

C is a flask generating sulphurous anhydride.

D is a tube through which air is forced into the condenser F.

E is an escape pipe for waste gases. It should be connected with a good ventilating shaft.

F is a glass globe used as a condensing chamber. Sulphuric acid collects at the bottom of the globe.

Sulphuric Acid: Formula, H_2SO_4 *; molecular weight 98; specific gravity of liquid, 1·846.*

Experiments.

1. Pour some commercial sulphuric acid into twenty or thirty times its own volume of water in a test-tube. Taste the solution.

2. Half-fill a small test-tube with blue litmus solution and add to it some of the diluted sulphuric acid. Add some of the acid to some red litmus solution also.

3. Place some "bread soda" in a test-tube and pour some of the dilute acid upon it.

4. Observe the color of some chemically pure sulphuric acid. Compare its weight with that of water, taking equal volumes of each. 4 to 3

5. Make a mixture of equal volumes of water and acid, and immerse in the mixture a glass mercurial thermometer, or a small test-tube filled with spirit, to note the change in temperature. from 54 to 190

6. Pour some strong acid on a little sugar. What element is shown to exist in the sugar?

7. Shake up some acid and oil in a test-tube. What action has the acid on the oil? no action

§ 96.—Solubility of Salts.

The sulphates are all soluble in water, except those of lead, barium, and strontium. The sulphate of calcium is slightly soluble.

§ 97.—Test.

The presence of any soluble sulphate may be detected by the addition of a few drops of a solution of barium nitrate or barium chloride. The color of the precipitate formed, accompanied by the fact that it is insoluble in any acid you may chose to add, is an invariable test for the acid or its salts.

§ 98.—Sulphuretted Hydrogen, SH_2.

This compound of sulphur and hydrogen is of great importance in chemical analysis. It is usually prepared by the action of an acid upon a sulphide.

Experiments. $FeS + H_2SO_4 = SH_2 + FeSO_4$

1. Take a hydrogen generating apparatus, and place in it some powdered sulphide of iron. Cover the sulphide with water, cork the apparatus tightly, and then add a few drops of sulphuric acid. Collect two or three bottles of the gas over warm water. Don't allow much of the gas to escape into the room. It is poisonous.

2. Bubble some of the gas through a solution of blue litmus. [illegible]

3. Attach a piece of glass tubing drawn to a fine point to the generating apparatus and try whether the gas will burn. [illegible]

4. Devise an experiment to ascertain whether the gas is soluble in water. [illegible]

5. Fill a bottle with chlorine gas and another with sulphuretted hydrogen. Invert the mouth of the chlorine bottle over that of the other and observe what takes place. Smell the gas that forms in the bottles. Infer the reaction, and write the equation. [illegible]

§ 99.—Sulphides.

A solution of sulphuretted hydrogen in water is much used in mineral analysis, inasmuch as the insoluble sulphides which it forms with soluble salts of many of the metals have characteristic colors that enable the analyst to recognize the presence of particular metals.

Experiments

1. Polish a five or ten cent piece of silver, and then place a drop of sulphuretted hydrogen water upon it. [illegible]

2. Prepare in separate test-tubes solutions of lead nitrate or acetate, copper sulphate, zinc sulphate, and ferrous sulphate. Add a few drops of hydrochloric acid to each, and then pass into each solution some gas from the generating apparatus. Take care to wash the delivery tube before passing it from one test-tube to the other. Tabulate your results, and memorize them. Try to write the equations which symbolize the reactions.

---

§ 100.—Test.

Experiments.

1. Moisten some paper in a solution of acetate of lead, and then bring it into contact with the gas ?

2. If the gas comes off in quantity, its smell is sufficiently characteristic to enable anyone to recognize it.

QUESTIONS.

1. Why is silver plate so easily blackened in the air of towns ?

2. Why do silver coins or watches when carried in the pocket along with matches change their color ?

3. How would you find out whether a sample of vinegar contained sulphuric acid or not ?

4. Sewer gas nearly always contains sulpuretted hydrogen. If sewer gas found its way into a house how would you disinfect it ?

CHAPTER XVIII.

§ 101.—Acids.

Chemists have found it convenient to classify a large number of chemical compounds as acids, bases, and salts.

In order to learn some of the general properties of acids, repeat the first three of the experiments in section ninety-five, using any three or four substances labelled acids, which you can find upon your working table. As you make each experiment, tabulate your results as follows :—

Name of Acid.	Taste.	Action on Red Litmus.	Action on Blue Litmus.	Action on Bread Soda.	REMARKS.
[illegible]	[illegible]	[illegible]	Changes to the red		

§ 102.—Bases.

Experiments.

1. Repeat the preceding experiments, using solutions of *(a)* quicklime, *(b)* slaked lime, *(c)* magnesium oxide, and *(d)* sodium hydroxide (caustic soda). Tabulate your results as before.

§ 103.—Hydroxides.

Experiments.

1. Take a piece of the metal potassium, about the size of a pea, place it in an iron spoon, and heat it over a spirit lamp until it has ceased to burn. Then add a little water, and test the solution with red and with blue litmus, as before.

2 Repeat this experiment, using the metals, magnesium and sodium.

3. Take a small jar and pour about ¼ inch of water into it. In case you have no spoon with a long handle, take half of a chalk crayon, and wind a piece of wire round one end of it, and use this instead. On the top of this crayon place a small piece of phosphorus, ignite it, and lower into the jar as in Fig. 36. Cover the mouth of the jar, and when the burning has ceased, remove the spoon or crayon, and shake up the water with the gas or solid produced. Test as before with the two litmus solutions.

FIG. 36.

4. Repeat this experiment, using separately, fragments of charcoal, and of sulphur. Tabulate your results as in previous experiments.

QUESTIONS.

1. Write out an account of what you have learnt about *acids* and bases, and tell how you would distinguish them from each other.

2. If you know whether an element is a metal or a non-metal, can you foretell whether its oxide, when dissolved in water, will produce an acid or a base? Explain.

A **hydroxide** is a compound formed by the union of the radical OH with atoms of the elements, or with other radicals. The group of atoms, OH, acting as one atom, and being present in a series of chemical compounds is called a *compound radical*, and its compounds are called hydroxides. The name given to this radical is hydroxyl. The compound of hydroxyl, OH, with potassium is potassium hydroxide KHO; with sodium, it is sodium hydroxide NaHO.

The stronger bases are known as **alkalies.**

Hydroxides are called **hydrates** by some chemists.

104.—Salts.

Let us now examine some of the compounds called salts.

Experiments.

1. Take a piece of "caustic soda" (sodic hydrate) NaHO, about the size of a pea, and dissolve it completely in a test-tube of water, then add to it hydrochloric acid, drop by drop, until a piece of blue litmus paper placed in the solution slowly begins to turn red, pour half of this solution into an evaporating dish, place on a sand bath and heat until all the water is driven off. Carefully examine the residue. Taste it.

Pour the rest of the solution into a flat dish of any kind, and allow it to remain for a day or two in a warm room.

The residue in both cases is called a *salt.*

2. Perform similar experiments using potassium hydroxide KHO (caustic potash), and nitric acid: also, sodium hydroxide and sulphuric acid.

Write out, in the following tabulated form, a synopsis of the leading characteristics of acids, bases, and salts, including their taste, action on litmus, action on each other, and a method of obtaining the members of each of these three groups of compounds.

	ACIDS.	ALKALIES.	SALTS.
1. Taste..................			
2. Action on litmus... .. .			
3. Action on each other....			
4. How obtained......... ..			

What you will have learned about acids bases and salts from your experiments with these compounds, will not be true as regards every member of each of these groups, but your knowledge will be sufficiently accurate for the present.

§ 105.—Naming Acids and Salts.

In naming **acids** the terminations **-ous** and **ic**, and the prefixes **hypo-** and **per-**, are used, *e.g.:*

$HClO$ is called hypochlorous acid.
$HClO_2$ is called chlorous acid.
$HClO_3$ is called chloric acid.
$HClO_4$ is called perchloric acid.

For the least amount of oxygen present in the above compounds, **hypo——ous** is used; **-ous,** for more oxygen; **-ic**, for still more of it; and **per——ic**, for the greatest amount. Similarly:—

HNO is called hyponitrous acid.
HNO_2 is called nitrous acid.
HNO_3 is called nitric acid.

Salts are named chiefly from the acids which form them.

If the name of the acid end in **-ic**, that of the salt ends in **-ate**.

If the name of the acid end in **-ous**, that of the salt ends in **-ite**.

The prefix of the name of the acid is retained in naming the salt, *e. g.:*

ACID.	NAME.	SALT.	NAME.
$HClO$	hypochlor-OUS acid.	$KClO$	potassic HYPO-chlor-ITE.
H_2SO_3	sulphur-OUS acid.	Na_2SO_3	sodic sulph-ITE.
HNO_3	nitr-IC acid.	$AgNO_3$	argentic nitr-ATE.
$HClO_4$	perchloric acid.	$KClO_4$	potassic PER-chlor-ATE.

§ 106.—Formulas of Hydroxides.

The principle of atomicity may be employed in formulating the theoretical oxides and hydroxides of the metals by using water as a type, and substituting in a single molecule of water one atom of a monad metal for one atom of hydrogen to form a hydroxide, and two atoms of a monad metal for the two atoms of hydrogen, to form an oxide. In two molecules of water, we must substitute one atom of a dyad metal for two atoms of hydrogen to form the hydroxide, and two atoms of the dyad metal for the four atoms of hydrogen to form the oxide. For example:

TYPE.	HYDROXIDE.	OXIDE.
H_2O	KHO	K_2O
$2\ H_2O$	$Ca(HO)_2$	CaO

1. Apply this principle and formulate hydroxides and oxides of the folowing metals: Sodium, silver, mercury, magnesium, iron, tin, platinum.
2. Name the compounds thus formed.

CHAPTER XIX.

§ 107.—Nitric Acid.

Before studying the compounds of nitrogen and oxygen, we shall dwell at some length upon an acid, which besides being of great practical use, is important as forming a starting point from which all the oxides of nitrogen can be derived. This acid is known as nitric acid. Its old name is *aqua fortis.*

Nitric Acid: Formula, HNO_3*; molecular weight,* 63*; specific weight of liquid,* 1·52*; boiling point,* 84·5° *Freezing,* – 40°.

Experiment.

Put into a tubulated glass retort 30 grams of powdered nitrate of potash, KNO_3, and an equal weight of strong sulphuric acid, H_2SO_4. Place the end of the retort in a flask which is made to float on a basin of water as in Fig. 37. Apply heat to the retort. Soon a yellowish coloured liquid distils over and is collected in the cool flask. The reaction may be represented as taking place in two successive stages, the first requiring a comparatively low, the second a high temperature.

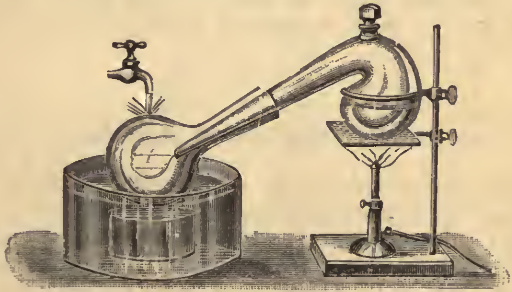

FIG. 37.

$$(a).\ 2KNO_3 + H_2SO_4 = HKSO_4 + HNO_3 + KNO_3.$$

On increasing the heat more acid comes off, the second reaction being represented as follows :—

$$(b).\ HKSO_4 + KNO_3 = K_2SO_4 + HNO_3.$$

Sodic nitrate, $NaNO_3$, may be used instead of potassic nitrate in the preparation of nitric acid ; in fact, sodic nitrate is generally used when this acid is to be manufactured on a large scale.

§ 108.—Properties.

Having prepared a sufficient quantity of the acid in the manner described, you may proceed to study its properties as follows :—

Experiments.

1. Immerse some wool, silk or other organic substances in a little of the acid. [illegible]

2. Add a few drops of it to a solution of indigo. [illegible]

3. Place some copper filings in the bottom of a test-tube, and then pour in some of the acid. When all action has ceased, evaporate to dryness the solution formed. [illegible]

§ 109.—Oxidizing Agent.

Nitric acid is said to be a powerful *oxidizing agent*, that is, it readily yields its oxygen to substances which have an affinity for that element This oxygen, at the moment it is liberated from the acid, is said to be in its *nascent* state. The three following experiments illustrate this oxidizing power of nitric acid. If you can procure some *very strong* acid—" fuming nitric acid " as it is called—you can successfully perform these experiments, but great care must be exercised or a dangerous accident may result.

Experiments.

1. Place a small piece of phosphorus in a saucer, then drop on it a little of the acid. Catches fire [illegible]

2. On powdered pieces of glowing charcoal, pour a little nitric acid. [illegible]

3. Place a little carbolic acid in a test-tube, and then pour in a few drops of the fuming acid. [illegible]

§ 110.—Other Properties.

Nitric acid when pure is colorless and has a specific weight of 1·5. The *commercial* acid contains about 30 per cent of water, and has a specific weight of 1·4; it is also yellow in color from containing some of the oxides of nitrogen in solution. Undiluted, it is a strong irritant poison when swallowed. It is **mono-basic,** that is, it contains only one atom of replaceable hydrogen in its molecule. The distinctive characteristic of a monobasic acid from a practical point of view is that it forms salts which have no acid reaction.

§ 111.—Uses.

It is used extensively in dyeing and as a medicine. In chemical analysis it is a solvent of great value. The nitrates are all soluble in water.

§ 112.—Sources.

Nitric acid is formed in air by lightning. It exists also in nature in the form of nitrates of potash, soda and lime, which are themselves the product of decomposition and oxidation of nitrogenous organic compounds with alkalies. "These nitrates are widely diffused in all surface soils, especially in hot countries such as India, where oxidation takes place rapidly. In the neighborhood of Indian villages, soil which contains considerable amounts of potash, thus becomes rich in nitre, or potassium nitrate. originating from the decomposition of the urea of the urine. It is from this source that the largest quantity of nitre imported into England is obtained."

§ 113.—Tests.

We are said to *test* for an element or for a compound, when we subject it to an experiment which reveals phenomena unlike those produced by any other substance under examination. For example, clear lime water becomes milky when carbon dioxide is passed into it; and this fact constitutes a characteristic test for carbon dioxide. Some one or more of the experiments performed with oxygen, nitrogen, hydrogen, carbon, and carbon monoxide will have enabled the student to distinguish these substances from each other, and from all other substances which he may hereafter meet with and which may resemble them, but to distinguish nitric acid from all other substances, we must perform a special experiment.

Experiment.

1. Dissolve a few crystals of ferrous sulphate, $FeSO_4$, in water in a test-tube. Add a few drops of sulphuric acid and allow the whole to cool. Then turn the test-tube sideways and gently pour nitric acid or a nitrate in solution down its side. The phenomenon which results will always enable us to recognize nitric acid or a nitrate.

QUESTIONS AND PROBLEMS.

1. Describe the physical and chemical properties of nitric acid?

2. Compare the action of nitric acid on wool and silk with its action on goose-quill clippings or on the finger nail?

3. Read up, in some standard work on organic chemistry, methods of manufacturing oxalic acid, gun-cotton, and nitro-glycerine.

4. Describe experiments to prove that this acid is easily decomposed.

5. Nitric acid acts on copper and forms the salt cupric nitrate $Cu(NO_3)_2$; find out whether it acts similarly on other common metals such as lead, zinc, iron and mercury.

6. The principle of atomicity may be employed in writing the formulas of salts from nitric acid, by replacing one atom of the hydrogen of the acid with one atom of a monad metal; two atoms of the hydrogen of the acid with one atom of a dyad metal, and so on. For example:

Acid.	*Salt.*	*Name of Salt.*
HNO_3	$AgNO_3$	Silver Nitrate.

(*a*) In the same way symbolize the salts which nitric acid may form with the following metals: Potassium, calcium, copper, lead.

(*b*) Name these salts.

7. Explain the action of nitric acid on a solution of sulphate of indigo.

CHAPTER XX.

§ 114.—Oxygen and Nitrogen.

Oxygen and nitrogen unite indirectly to form five well-known compounds. Their names and formulæ may be tabulated as follows:—

FORMULÆ.	OLD NAMES.	MODERN NAMES.
N_2O	Nitrous oxide..........	Nitrogen monoxide.
NO (N_2O_2).......	Nitric "	" dioxide.
N_2O_3.............	Nitrous anhydride....	" trioxide.
NO_2 (N_2O_4)......	Nitrogen peroxide.....	" tetroxide.
N_2O_5............	Nitric anhydride......	" pentoxide.

Only the first three will be considered in this book. The third and fourth are, of course, important from a theoretical point of view, but not sufficiently so to come within the scope of the present work.

Nitrous Anhydride, N_2O_3*; molecular weight, 76; specific weight, 38.*

Experiment.

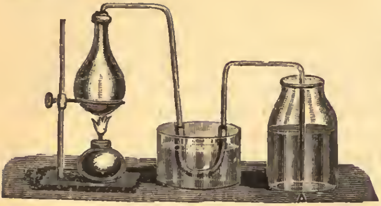

FIG. 38.

Fit a Florence flask with a cork and delivery tube and place on a retort stand, as in Fig. 38. To the delivery tube attach a U tube immersed in a freezing mixture of salt and snow. Connect the other end of the U tube with a glass tube leading to a vessel A, containing ice-water. Place 10 grams of starch in the flask and cover with nitric acid. Gently heat the generating flask and nitrogen trioxide will be plentifully produced, part of it being condensed in the U tube, and the remainder passing on into the ice-water.

Instead of starch, white arsenic, As_2O_3, may be used. The reaction in this case may be thus represented

$$2HNO_3 + As_2O_3 + 2H_2O = \overset{\text{Arsenic Acid.}}{2(H_3AsO_4)} + N_2O_3.$$

Notice the color of the gas. It is condensed to a liquid by a temperature of—18° C. Try to collect some of the gas over water. Has it any smell?

§ 115.—Nitrous Acid, HNO_2.

In preparing nitrogen trioxide, it is difficult to avoid preparing some nitrous acid at the same time. In fact, the ice-water in the preceding experiment already contains some of this acid dissolved in it. Its preparation may therefore be represented as follows :—

$$N_2O_3 + H_2O = 2HNO_2.$$

Experiments.

1. Pass some nitrogen trioxide into a cold solution of potassic hydrate until it is neutralized. Then evaporate to dryness. Try to symbolize the reaction that occurs. The salt prepared in this experiment is known as potassic nitrite.

2. Add a few drops of nitrous acid to a solution of potassium permanganate. Try to account for the phenomena which arise.

---

§ 116.—Other Properties and Tests.

Nitrous acid is an unstable compound decomposing, upon standing, into nitric acid, nitric oxide, and water Its salts are all soluble in water. When nitrites are acidulated with acetic acid they give a white precipitate with nitrate of silver. The next experiment furnishes us with a test for free nitrous acid.

Experiment.

Boil some starch in water so as to form a paste. Then add some iodide of potassium solution, and allow the whole to cool. The reaction, which occurs on adding free nitrous acid to the mixture forms, when taken in connection with the nitrite of silver test a sure indication of the presence of nitrous acid.

QUESTIONS AND PROBLEMS.

1. Write out an account of the physical and chemical properties, *(a)* of nitrous anhydride, *(b)* of nitrous acid.

2. Investigate some of the properties of potassic nitrite by *(a)* throwing some of it upon red-hot charcoal, *(b)* by placing a drop of any strong acid upon it.

3. Heat some nitrous acid solution, then test the residue for nitric acid.

4. Procure a nitrite and apply to it the test for nitrous acid.

---

§ 117.—Nitrogen Dioxide.

In preparing nitrogen dioxide in the following experiment, a gas is at first formed with which the student is already acquainted, but he must not be misled by its appearance.

Nitric Oxide NO (N_2O_2)*; molecular weight, 30; specific weight, 15. 11.2 litres weigh 15 grams.*

Experiment.

Fig. 39.

Place some copper filings in a hydrogen generating apparatus, similar to that in Fig. 39, add some warm water, and then pour down the funnel tube some strong nitric acid. The gas that first forms should be allowed to escape. The reaction may be thus represented :—

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

Collect over water four beakers full of the gas and perform the following experiments :—

(a). Allow the contents of the first beaker to escape into the air.

(b). Ignite a piece of phosphorus very slightly and plunge it into the second beaker.

(c). Allow another piece of phosphorus to burn strongly, and then place it in the third beaker.

---

§ 118.—Other Properties and Tests.

Nitric oxide condenses to a liquid at—11°C, and a pressure of 104 atmospheres. Unlike nitrous anhydride, it does not unite with water to form an acid. One test for this gas is its reaction with air or free oxygen; another is that with a solution of ferrous sulphate, as seen in the next experiment.

Experiment.

Pour a solution of ferrous sulphate, $FeSO_4$, into the fourth beaker full of the gas. Then hold the hand over the beaker's mouth and shake vigorously. Note the two phenomena that occur.

QUESTIONS AND PROBLEMS.

1. Name the reddish orange gas that forms at first in the flask in preparing nitric oxide? Name three gases that are mixed in the flask.

2. Try whether you can prepare this gas by using other metals than copper.

3. What gas is formed when it escapes into air? [Nitrogen peroxide is formed as well as the gas, whose name you are to give].

4. Explain the difference between the results of experiments *(b)* and *(c)*.

5. Is nitric oxide soluble to any great extent in water? Give a reason for your answer.

6. Investigate the proportions by volume in which oxygen and nitrogen are in this gas. To do so, fill with mercury a test-tube bent at an obtuse angle and invert it over another vessel containing mercury. Pass into the test-tube some of the gas. Then take a small piece of potassium or sodium and place it under the mouth of the test-tube. When the sodium rises to the top of the mercury in the tube, jerk the tube gently to one side so that the sodium is thrown on to its bent part; then heat the test-tube immediately below the sodium until it burns. After the tube has thoroughly cooled, note the change in volume. reduced

7. The specific weight of nitric oxide is found by actually weighing it to be 15, Should its formula be NO, or N_2O_2?

---

§ 119.—Nitrogen Monoxide.

"Laughing gas" is an old name for this compound. It derives this name from the fact that many persons after inhaling a mixture of the gas and air are compelled to laugh—*nolens volens.* On inhaling more of the gas, temporary unconsciousness is produced; it is therefore frequently used as an anaesthetic for minor operations in surgery.

Nitrous Oxide: formula, N_2O*; molecular weight, 44; specific weight, 22. 11.2 litres weigh 22 grams.*

Experiment.

Put 25 grams of commercial ammonic nitrate, NH_4NO_3, into an oxygen generating apparatus, connected with three bottles, as in Fig. 40. The first bottle should contain a solution of ferrous sulphate, the second, a solution of caustic potash, and the third, water. Heat the nitrate gently and nitrogen monoxide will be given off. Thus prepared, the gas will be found mixed with nitrogen dioxide, and chlorine gas. The first will be removed by passing through the ferrous sulphate solution, and the second by passing through the caustic potash solution.

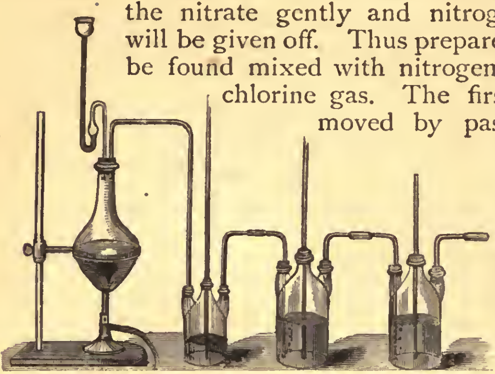

FIG. 40.

If the nitrate be chemically pure, the wash bottles may be omitted. In this case the reaction may be thus represented :—

$$NH_4NO_3 = 2H_2O + N_2O$$

Collect four jars of the gas over warm water and perform the following experiments :—

(a). Plunge a lighted taper into the first jar. burns

(b). Burn a piece of phosphorus, or carbon, or sulphur in the second jar. If burning strongly at first, they continue to burn. Write the equation expressing the reaction. burns more vigorously

(c). Explode a mixture of the gas with hydrogen. Write the equation. $N_2O + 2H = H_2O + N_2$

(d). Place the fourth jar, mouth downward, over *cold* water, and then shake. water rises, the gas dissolves in cold water

§ 120.—Other Properties.

Nitrous oxide is soluble in cold water to the extent of 130 per cent. of its own volume; it may be condensed to a liquid by a cold of 0°C, and a pressure of 30 atmospheres. Liquid nitrous oxide when mixed with carbon disulphide, CS_2, forms a freezing mixture capable of producing a cold of —140°C.

§ 121.—Nitrogen Acids.

It is interesting to note how hyponitrous, nitrous, and nitric, acids may be supposed to be formed from nitrous oxide, nitrous anhydride, and nitric anhydride respectively. Thus :—

$N_2O + H_2O = 2HNO$, Hyponitrous acid.
$N_2O_3 + H_2O = 2HNO_2$, Nitrous acid.
$N_2O_5 + H_2O = 2HNO_3$, Nitric acid.

The first of these has never been prepared in a free state, but its salts are known.

QUESTIONS AND PROBLEMS.

1. Has nitrous oxide any taste or smell? How can it be distinguished from oxygen?

2. Devise an experiment to ascertain whether it is heavier or lighter than air?

3. Pass some nitric oxide into air, and also into nitrous oxide. Explain the difference in the phenomena observed?

4. Ascertain the composition of nitrous oxide by volume. To do this, make an experiment similar to that by which you found out the composition by volume of nitric oxide.

CHAPTER XXI.

§ 122.—Nitrogen and Hydrogen.

There is but one known compound of nitrogen and hydrogen. This compound is called ammonia.

Ammonia: (Spirits of Hartshorn) Formula, NH_3; atomic weight, 17; specific weight, 8.5. 11.2 litres weigh 8·5 grams. 14 44

Experiments.

1. Mix in a mortar 30 centigrams of fine iron filings, with 2 centigrams of solid caustic potash. Then transfer the mixture to a test-tube fitted with cork and delivery tube, and heat until gas escapes. Collect some of this gas and test for hydrogen.

15 Fe + KHO + [illegible] = [illegible]

2. Repeat the experiment, substituting 2 centigrams of nitre for the 2 of potash. In this case test for nitrogen.

3. Now mix 30 centigrams of iron filings, 2 of caustic potash and 2 of nitre, place in a test-tube and heat as before. Smell the gas that is evolved; it is ammonia.

4. Make a mixture of nitrogen and hydrogen, taking these gases from gas-holders or bottles in which they have previously been stored. After the mixture has stood for some time, smell it. Does it smell like ammonia? No. no.

QUESTIONS AND PROBLEMS.

1. Describe briefly the physical properties of ammonia.
2. What is the relation of experiment 3, to experiments 1 and 2?
3. Try to account for the fact that ammonia has formed in the third experiment and not in the fourth.

§ 123.—Nascent State.

An element is said to be in its NASCENT state at the moment it is liberated from a compound. In this state its properties are more active than after it has been formed for some time. This greater energy is explained on the theory that elements, at the instant of their elimination from a compound, exist in the form of atoms and not of molecules. Soon after being eliminated, molecules are formed; and then the element loses its power of readily entering into union with other elements. Apply this theory of the nascent condition of elements to explain the relation to each other of the results of the four experiments in the preceding section.

§ 124.—Properties.

Experiments

1. Take about 20 grams of dry ammonic chloride and an equal quantity of dry quick-lime; powder them finely

$2NH_4Cl + CaO = 2NH_3 +$

in a mortar. Smell the mixture, and then transfer it to a flask with tightly fitting cork and long tube bent upwards. Heat gently. Hold a large test-tube over the delivery tube, and fill it with gas by downward displacement of air, as in Fig. 41.

2. Pass a lighted taper up into the test-tube full of gas. goes out

3. Pass some of the gas into reddened litmus. Upon the result of this, devise a means of knowing when a bottle is full of this gas. [illegible]

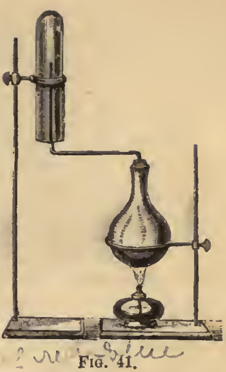

FIG. 41.

4. Moisten a glass rod with hydrochloric acid, and then bring it near the end of the delivery tube. Do the same with other acids.

5. Fill a narrow-necked bottle with the gas, by displacement of air as before, and then place the bottle mouth downward in water. Shake the bottle.

6. Pass the gas for some time into a long test-tube of ice cold water. Note any changes in temperature and volume of the water.

QUESTIONS AND PROBLEMS.

1. Describe the chemical properties of ammonia.

2. How can you tell when a jar is full of this gas, if you collect by displacement of air?

3. How do you know that ammonia is a base? What compound did it form with hydrochloric acid? In what physical state did this compound exist? $NH_3 + HCl =$

4. What became of the gas in experiment 5? How does the solubility of ammonia in water compare with that of oxygen? Of nitrogen? Of hydrogen?

5. Aqua ammonia is simply water with ammonia dissolved in it. Knowing this, try to devise a means of obtaining ammonia from aqua ammonia, and prepare some of the gas from this source.

6. Devise an experiment to ascertain whether ammonia is absorbed by freshly burned charcoal.

7. Powder some coal coarsely in a mortar. Then place in a hard glass tube and heat. Smell the gas that comes off. Is the liquid that forms acid or alkaline?

8. Ascertain, in a similar manner, whether animal matters, such as horn, dried flesh, glue, cheese, &c., will yield ammonia on being heated in a test-tube.

---

§ 125.—Ammonium Hydroxide.

Ammonium hydoxide is probably contained in *liquor ammoniæ* or *aqua ammoniæ*. Its formula is NH^4OH. For a description of the process of manufacturing liquor ammoniæ on a large scale, you must consult some treatise on chemistry. All that need be said here is that aqua ammonia is a by product from the manufacture of coal gas.

Experiments.


FIG. 42.

1. Pour a concentrated solution of ammonia into a flask. Through this solution pass a current of oxygen from a gas holder and tube, as in Fig. 42. Apply a lighted match to the mouth of the flask. gas burns bri

2. Heat to redness a spiral coil of platinum wire, and then place it quickly in the mouth of a bottle containing aqua ammonia. glows

yellowish gl

QUESTIONS.

1. What was the color of the flame produced in experiment 1? Try to write the equation expressing the reaction. $NH_4OH + O =$

2. Try whether a jet of pure *dry* ammonia will burn.

3. Why does the platinum wire continue to glow in experiment 2? Whence comes the heat?

4. Point out any resemblance you can see between the results of experiments 1 and 2.

5. Try whether you can expel all the gas from *liquor ammoniæ* by boiling it. Test the water afterwards with some of *Nessler's* solution (see test for ammonia).

---

§ 126.—Analysis of Ammonia.

Experiment.

Take a eudiometer, fill with mercury and invert over a small trough or saucer, also containing mercury. Heat some ammonium hydrate and pass 20 c. cs. of the gas into the eudiometer. Then pass a series of electric sparks from an induction coil through the gas, taking care to insert a Leyden jar in the circuit, so as to increase the heating effect. When the gas no longer expands, pass 29 c. cs. of pure oxygen into the eudiometer and explode. Depress the eudiometer so as to bring the mercury to the same level inside the eudiometer as the outside, and then note the volume of the gas remaining in it. 26 c. cs

QUESTIONS.

1. How much did the gas increase in volume by passing electricity through it?

2. How many c. cs. were there altogether in the eudiometer after passing in the oxygen? How many after the explosion?

3. Explain the cause of the reduction in volume?

4. What must be the composition of the remaining gases? What volume of each?

5. What then are the constituents of ammonia by volume?

6. How many atoms of each element must there be in the molecule of ammonia? What therefore must be its formula?

---

§ 127.—Ammonic Chloride.

You have already learned that ammonia is a base, and in your experiments with acids and bases you formed salts by neutralizing one of these compounds with the other. You might, therefore, fairly infer that a salt was formed when ammonia was brought into contact with hydrochloric or other acid. The salts thus formed are called ammonium salts, because they contain the radical NH_4—ammonium.

Experiments.

1. Pour about 10 c. cs. of aqua ammoniæ into an evaporator, and add hydrochloric acid until the solution is neutral. Then gently evaporate to dryness. The substance obtained is called ammonic chloride.

2. Place a small piece of the salt obtained in the preceding experiment upon a strip of platinum foil and heat for some time. remained the same

3. Prepare a solution of ammonic chloride with the remaining part of the salt; place in a test-tube, and add caustic soda or caustic potash, and heat. Smell.

Smells of [illegible]

QUESTIONS AND PROBLEMS.

1. Prepare ammonic chloride by using equal volumes of the two gases —ammonia and hydrochloric acid. $NH_3 + HCl$ = [illegible]

2. Prepare other salts of ammonia by using other acids than hydrochloric.

3. How was ammonia obtained from the ammonic chloride? Draw a conclusion as to the general manner in which ammonia may be obtained by the decomposition of ammonic chloride.

4. Ascertain whether ammonia can be obtained from other salts of ammonia beside the chloride. To do this, heat solutions of these salts, in a test-tube, with solutions of either caustic soda, caustic potash, or quicklime. At intervals, waft with the hand the vapour coming from the test-tube towards the face, so as to smell it. Warm a glass rod slightly, dip it into hydrochloric acid, and bring it over the mouth of the test-tube. Notice whether any phenomenon results.

5. Try to represent by equations the re-actions taking place in preparing the salts of ammonia. Thus :—

$$NH_4HO + HNO_3 = NH_4$$
$$2NH_4HO + H_2SO_4 =$$

6. Also, the reactions when these salts are decomposed by heating with caustic soda or caustic potash. Thus :—

$$NH_4NO_3 + KHO =$$

---

§ 128.—Formula of Ammonic Chloride.

We have already indicated that the formula of aqua ammonia is NH_4OH, or NH_3, H_2O. What is the formula of ammonic chloride? Is it NH_3, HCl, or is it NH_4Cl? The following experiment may help us to decide.

Experiment.

Place about 1 c. c. of mercury in a wide test-tube, and slightly heat it over a lamp. Drop into it, at intervals, two or three pieces of sodium, and heat until they have disappeared. Pour the amalgam thus formed into a strong solution of ammonic chloride, previously prepared. Allow the whole to stand for about a minute, then pour off the liquid and wash the residue with cold water. Smell the residue. Apply the flame test for sodium to the liquid which you poured off. After the residue has stood for half an hour or so, note whether any change has taken place in its volume. Test the liquids poured off for a chloride. Evaporate some of the liquids and ascertain what common substance is dissolved in it. Salt

QUESTIONS AND PROBLEMS.

1. What is an amalgam?

2. What change took place on throwing the amalgam into the solution of ammonic chloride?

3. With what element in the ammonic chloride did the sodium unite? How do you know?

4. When the sodium left the mercury, did the latter form an amalgam with any other substance? Give a reason for your answer.

5. What change gradually took place in the *volume* of the metallic looking mass which formed in the ammonic chloride solution?

6. If (Hg, Na) + NH_4Cl represents the substances and their relations before chemical action took place, how should the reaction which resulted be represented?

The compound represented by the formula NH_4 is often called ammonium. It is sometimes spoken of as a *metallic radical*, because it forms salts exactly analogous to those of sodium and potassium. Compare the hydroxides:—

KHO = potassium hydroxide
NH_4HO = ammonium hydroxide. Compare also the salts:—
KNO_3 = potassium nitrate
NH_4NO_3 = ammonium nitrate.

§ 129.—Hydroxides from Salts.

You have already learned one method of producing oxides. Hydroxides may be formed from some of these oxides by simply adding water to them. But hydroxides may be formed in another way, which we shall now illustrate.

Experiments.

1. Pour into a test-tube 1 c. c. of solution of ferrous sulphate, $FeSO_4$, and add to it gradually 1 c. c. of dilute ammonic hydroxide. form Hydroxide of Iron

2. To 1 c. c. of solution of zinc sulphate $ZnSO_4$, add 1 c. c. of ammonic hydrate. [illegible]

3. To 1 c. c. solution of silver nitrate, $AgNO_3$, add 1 c. c. of ammonic hydrate.

4. Filter the solutions obtained in each case, and dry and examine the precipitate. These precipitates are hydroxides.

5. Carefully evaporate the filtrate in each case, using for this purpose a sand bath. Examine any substance that may remain in the evaporator. [illegible] (5)

6. Prepare a number of solutions of salts of the metals and then add ammonic hydrate or caustic potash solution to each one in turn. Notice carefully the cases in which a precipitate forms.

QUESTIONS.

1. How are hydroxides of the metals obtained? How can you distinguish the hydroxides of one metal from the hydroxides of another?

X 2. What use can be made of these and similar experiments for analytical purposes?

3. Explain why a precipitate does not form in every case. [illegible]

4. Write equations representing the reactions taking place in each experiment.

5. Devise experiments to obtain ammonia from the salts left after evaporating the water in each case.

§ 130.—Sources of Ammonia.

It occurs in the urine and in some other products of animals; also in air as the result of the decay or decomposition of nitrogenous animal matter; hence it exists in rain water. Its compounds, ammonic chloride and ammonium carbonate, are found sparingly in nature.

§ 131.—Other Properties.

Ammonia is soluble to upwards of 700 times its bulk in water at 15°C; it becomes a liquid at—40°C, and may even be frozen at—75°C.

§ 132.—Uses.

Dilute liquor ammoniæ is used to neutralize acids spilled upon the clothes or upon the face. Its compounds are also used in medicine as stimulants in cases of fainting or of syncope from overdoses of chloroform, ether or laughing gas. They are also used in dyeing.

§ 133.—Test.

When present in minute quantities, as it frequently is in drinking water, ammonia is best detected by what is known as Nessler's test: "To a solution of potassic iodide add solution of mercuric chloride until the precipitate formed just ceases to be re-dissolved, then, add an equal volume solution of caustic potash, and allow the whole to stand until clear. A few drops of this solution will give a yellowish-brown precipitate, with even the slightest trace of ammonia."

PROBLEMS

1. Calculate what volume 51 grams of ammonia gas will occupy at 120°C?

2. 61 litres of ammonia gas are decomposed in a eudiometer, what volume will its constituent gases occupy?

3. What weight and volume of ammonia gas at 60° F. can be obtained from 214 grams of ammonic chloride?

4. If 85 grams of ammonia gas be decomposed in a eudiometer, and 22·4 litres of oxygen gas be added to the constituent gases, and the mixture exploded, what will be the volume of the resulting gases at 0°C?

5. What weight of quick-lime is required to decompose 107 grams of ammonic chloride, and what will be the weight of the calcic chloride and water produced? What volume of ammonia gas will be evolved at 150° F.

CHAPTER XXII.

§ 134.—Hydrochloric Acid.

Before entering upon the study of the next element, viz., chlorine, we shall consider at some length the properties of hydrochloric acid. As the name implies this substance is a compound of hydrogen and chlorine

Hydrochloric Acid ("Spirit of salt"): Formula, HCl; *molecular weight, 36.5; specific weight; 18.25. 11.2 litres weigh 18·25 grams.*

Experiments.

$NH_4Cl + H_2SO_4 = HCl +$

1. Place some ammonic chloride in a medium-sized test-tube, and add a few drops of sulphuric acid. Bring a lighted match to the mouth of the test-tube; also a piece of blue litmus paper; and lastly, a glass rod dipped in ammonium hydroxide. Smell very carefully.

2. Repeat this experiment, using a large test-tube or flask fitted with a cork or delivery tube, and substituting sodic chloride NaCl, (common salt) for the ammonic chloride. Use twice the weight of sulphuric acid that you do of salt, and apply heat very carefully. Collect some of the gas by passing the delivery tube to the bottom of an "empty" jar. Cover its mouth with a glass plate. Having filled the jar, remove the plate cover, and turn the jar mouth downward over some water coloured blue with litmus. Then shake slightly. Devise a means of finding out when the jar is full.

3. Fill a second jar with this gas as before. Place two or three globules of sodium, the size of a pea, in a deflagrating spoon, the handle of which passes through a cork that exactly fits the jar. Heat the sodium to ignition, and lower the spoon into the second jar of gas. After all action has ceased, withdraw the cork and quickly bring a lighted taper to the mouth of the jar.

§ 135.—Commercial Acid.

A solution of this gas in water is what is usually sold by druggists under the name of hydrochloric or muriatic acid. How can the gas be obtained from such a solution? The commercial acid is prepared as a by-product in the manufacture of common "soda" by Leblanc's process.

QUESTIONS AND PROBLEMS.

1. What other substance was prepared in previous experiments from ammonic chloride? Contrast the different properties of the two different gases thus differently prepared.

2. Write a description of the physical properties of hydrochloric acid gas.

3. Explain the effect of dry hydrochloric acid gas on air.

4. What common substance was formed by burning sodium in this gas? What element was liberated? How do you know?

§ 136.—Chlorides.

Procure some chemically pure hydrochloric acid and with it perform the following experiments. They illustrate the formation of chlorides.

Experiments.

1. Place a few pieces of zinc in a test-tube and then pour upon them about 2 c. cs. of hydrochloric acid. After all effervescence has ceased, remove the surplus zinc and evaporate the solution to dryness.

2. Place a few fragments of gold leaf in a test-tube, and pour upon them about 1 c. c. of hydrochloric acid. Warm slightly. After a minute or two, add a few drops of nitric acid.

3. Repeat the preceding experiment using small scraps of platinum instead of gold.

4. Place a little cupric oxide CuO, in a test-tube and pour some hydrochloric acid upon it. When the oxide ceases to dissolve, filter, and evaporate the solution to dryness.

5. Fill a tube with hydrochloric acid gas over mercury, and then pass a piece of quicklime CaO under the mouth of the tube.

6. Take 3 test-tubes: half fill the first with a solution of silver nitrate $Ag\,NO_3$; the second with a solution of mercuric nitrate $Hg\,(NO_3)_2$; and the third with a solu- of acetate of lead $Pb\,(C_2H_3O_2)_2$. Into each tube pour a few drops of hydrochloric acid.

Try to write the equations.

Aqua Regia.—A mixture of three volumes of hydrochloric acid and one volume of nitric acid is called aqua regia.

---

§ 137.—Acids on Metals.

The first effect of an acid on a metal is the replacement of the hydrogen of the acid by the metal; hydrogen is, accordingly, liberated. Hence in preparing hydrogen from zinc and sulphuric acid, we represent the reaction thus:—

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

In this case no heat is applied; but when we bring copper and sulphuric acid together at ordinary temperatures, no change takes place. We would expect, of course, the reaction to be:

$$Cu + H_2SO_4 = CuSO_4 + H_2.$$

and so it would be in all probability if any reaction did take place at ordinary temperatures. No reaction, however does take place unless we apply heat, and thus it happens that the hydrogen which is first liberated, acts upon the sulphuric acid, reducing it to sulphur dioxide and water. Thus:

$$H_2SO_4+H_2=2H_2O+SO_2.$$

Compare this with the action of nitric acid on copper, Here the reaction varies with the strength of the acid. but when this latter is very strong, the following equations will represent the reaction in the two stages in which it occurs.

$$(1).\ Cu+2HNO_3=Cu(NO_3)_2+H_2.$$

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

QUESTIONS AND PROBLEMS.

1. Try to form other chlorides by warming hydrochloric acid with metals, oxides or hydrates which you can find in the laboratory, and which you have not yet used.

2. Arrange under two headings—soluble and insoluble—the chlorides which you have formed.

3. Compare the action of hydrochloric acid on metals with that of sulphuric acid on metals.

4. How is the "solvent" power of hydrochloric acid increased?

5. Explain the effect of hydrochloric acid gas on quick lime. To observe the effect, fill a tube with hydrochloric acid gas over mercury, and then pass a piece of quick lime up under the mouth of the tube. Is this gas absorbed by charcoal? Find out by experiment.

---

§ 138.—Analysis of Hydrochloric Gas.

Experiment.

Take a bent tube like that in Fig. 43. Partly fill the tube, as indicated, with hydrochloric acid, and insert in the ends the terminal wires of a battery. These terminals should be carbon. Bring a lighted match to that end of the tube connected with the zinc of the battery. Moisten

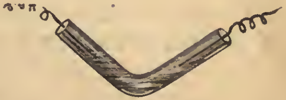

Fig. 43.

a piece of colored calico and place it over the other end of the tube. Color the acid with litmus solution.

§ 139.—Test for Chlorides.

Experiment.

1. Dissolve some common salt in half a test-tube full of pure water, and then add a few drops of nitrate of silver solution. Shake. Now pour half of the solution into a second test-tube; add a little nitric acid to the one test-tube and ammonium hydrate to the other. Boil the one to which you added nitric acid. becomes clea

2. Repeat this experiment, using any soluble chloride in place of common salt.

The gas may easily be distinguished from others which it resembles, by its behaviour in the experiments hitherto described.

QUESTIONS.

1. Suggest any reason why the terminal wires should be tipped with carbon.

2. You ought to be able to recognize one of the gases evolved in the bent tube. Name it and give a reason for your answer.

3. What was the colour of the precipitate thrown down by the nitrate of silver? How was it affected by nitric acid, and how by ammonium hydrate?

4. Suppose we have a solution containing salts of any number of metals, and add hydrochloric acid, what *three* metals will be removed by forming insoluble precipitates with the acid?

§ 140.—Other Properties.

Hydrochloric acid gas is soluble in water to the extent of 480 times its own volume; it may be liquefied.

EXERCISE.

The principle of atomicity may be employed in formulating the chlorides of the metals, by taking one molecule of hydric chloride as a type, and substituting for its atom of hydrogen one atom of a monad metal. One atom of a diad metal must be substituted for two atoms of hydrogen in two molecules of hydric chloride, to form the chloride of a diad metal and so on with triads, tetrads, &c. There are important exceptions to this application of the principle.

Type.	*Chloride.*	*Name.*
HCl	$NaCl$	Sodic Chloride.
$2HCl$	$CaCl_2$	Calcic Chloride.

1. Apply this principle and formulate theoretical chlorides of the following metals :—

Arsenic, gold, tin, manganese, iron, potassium, mercury, silver, zinc, calcium.

2. Name the compounds thus formed.

3. State as briefly as you can the reasons for believing that hydrochloric acid gas is composed of equal volumes each of hydrogen and chlorine.

4. What weight of common salt and sulphuric acid must be taken if it be required to eliminate 146 grams of hydric chloride ?

5. Calculate the amount of hydro-sodic sulphate that will be produced in generating 219 grams of hydric-chloride from salt and sulphuric acid at a moderate temperature.

6. Explain why we believe that hydrogen and chlorine are united in the proportions by weight of 1 to 35.5.

7. What volume will 73 grams of hydric chloride occupy at the standard temperature and pressure ?

8. What is the percentage composition of the gas ?

9. Calculate the weight and volume of hydric chloride at 30°C. that can be formed by heating to a moderate temperature 409.5 grams of common salt and 686 grams of sulphuric acid ?

CHAPTER XXIII.

§ 141.—Chlorine.

This element is never found uncombined in nature, but is generally obtained from the well-known substance, common salt.

Chlorine : Symbol, Cl ; atomic weight, 35.5 ; molecular weight, Cl_2*, 71 ; specific weight, 35.5. 11.2 litres weigh 35.5 grams.*

Experiments.

1. Into a test-tube put one part of manganese dioxide, two parts of salt, and three of sulphuric acid. Fit the test-tube with a cork and delivery tube. Heat gently and pass the gas that comes off into separate solutions of litmus and indigo. Cautiously smell the gas. Note its color. brown fumes

2. To prepare the gas on a larger scale, take a 4 oz. Florence flask and place in it about 20 grams of manganese dioxide and 100 c. cs. of strong hydrochloric acid. Use fittings similar to those in Fig. 44. Apply a *very* gentle heat. The delivery tube should pass almost to the bottom of the jar. Fill several jars, taking care that little or none of the gas escapes into the room. Afterwards pass the gas into a flask perfectly *full* of water; in about ten minutes place this flask aside for future use. Smell the water.

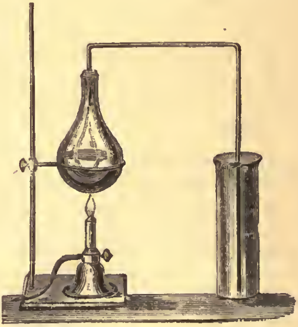

Fig. 44.

QUESTIONS AND PROBLEMS.

1. Write out an account of the physical properties of chlorine.

2. In the first experiment what gaseous compound is formed when the sulphuric acid acts on the salt? What becomes of this compound?

3. How does the first method of preparing the gas differ, if at all, from that employed in the second?

4. What class of oxides will cause hydrochloric acid to yield up its chlorine, and what class will not?

5. Treat the refuse from preparing oxygen from manganese dioxide and chlorate of potash with sulphuric acid. and observe what gas is evolved.

§ 142.—Properties of Chlorine.

Experiments.

1. Take the flask full of chlorine water, prepared in the last experiment, and fit it with a cork and tube. The

outer end of the tube must be drawn to a fine point. Insert the cork so that there is not a bubble of air left in the flask. Invert the flask as in Fig. 45, and expose to direct sunlight for a day. Then place the flask on the table, remove the cork, and quickly bring a glowing splinter to the mouth of the flask. Test the water in the flask with blue litmus solution. Taste it.

FIG. 45.

2. Lower very slowly a lighted taper into a jar of chlorine. At the same time suspend a piece of blue litmus paper at the mouth of the jar. Smell the gas that is formed during the combustion. Flame burns dull red black smoke rises

3. Wet a piece of blotting paper with oil of turpentine, $C_{10}H_{16}$, and then place it in another jar of the gas. The paper would take fire [illegible] dense [illegible]

4. Take a few pieces of the metal antimony and powder them; then place on a sheet of paper and shake the powder into a jar of chlorine. The antimony [illegible] burns [illegible]

5. Fill a small jar with hydrogen and then bring its mouth below the mouth of another jar of chlorine. Keep the jars mouth to mouth, and invert them several times, so as to mix the gases thoroughly. Then separate the jars, carefully corking one, and applying a lighted match to the other. Wrap a towel around the jar which you have corked, so as to exclude the light, and then carry it to where the sun is shining, either in a room or outside. Place it on the floor or on the ground and quickly unroll the towel so as to send the jar a short distance from you. Chlorine and hydrogen should never be mixed excepting in a *dim light.*

QUESTIONS AND PROBLEMS.

1. What element is liberated by the action of chlorine on water? What compound is formed?

2. Roughly speaking the wax of a taper consists of oxygen and hydrogen, united with much carbon. With which of these elements did the chlorine unite? What element was liberated?

3. Point out resemblances in the phenomena when a burning taper is put in chlorine, and when turpentine burns in the gas.

4. Try whether a piece of dry phosphorous will take fire when put into chlorine. takes fire quickly

5. Try the effect of chlorine on other metals besides antimony. Why do we not collect this gas over mercury?

6. Devise an experiment to ascertain roughly the extent to which chlorine is soluble in water.

7. Devise experiments to find out what effect chlorine gas has on solutions of logwood, cochineal, indigo, writing ink, or other colored liquids. Chlorine water may be used instead of the gas.

§ 143.—Bleaching Explained.

From the effect which chlorine is known to nave on water, and from the result of the following experiment the student will readily be able to discover the main cause of the bleaching power of this gas.

Experiment.

Prepare chlorine as before, and cause it to bubble slowly through some strong sulphuric acid in order to dry it. Collect a jar of the dry gas and place in it a piece of dry calico. Close the jar tightly. Allow the calico to remain in the jar for about fifteen minutes. Then open the jar; quickly remove the calico, wet it in water, and return to the jar for fifteen minutes more.

calico is not bleached when dry but when wet color slowly

§ 144.—Bleaching Powder.

Bleaching powder, or chloride of lime, is an important article of commerce which is extensively used in bleaching the coarser kinds of cotton and linen goods. Its manufacture is illustrated in the following experiment.

Experiment.

Cover the inside of a pint fruit jar with slaked lime, and then pass chlorine into it for some time. The chemical changes which take place may be thus represented:

$$2Cl_2 + \underbrace{2Ca(HO)_2}_{\text{Slaked lime.}} = 2H_2O + \underbrace{CaCl_2}_{\text{Calcic chloride.}} + \underbrace{Ca(ClO)_2}_{\text{Calcic hypochlorite.}}$$

It is this mixture of calcic chloride and calcic hypochlorite which forms the most important ingredients of what is popularly known as "bleaching powder."

§ 145.—Its Properties.

Experiments.

1. Remove from the jar the product obtained in the preceding experiment. Place it in a soup-plate and add about 100 c. cs. of water, stirring the mixture for five or ten minutes. Immerse in the solution thus prepared, a piece of printed calico. After a few minutes remove the calico and immerse it in a *very* dilute solution of sulphuric acid.

2. Make a solution of asafœtida by dissolving some of the substance in a little alcohol in a small beaker. Now add a few drops of sulphuric acid to a solution of bleaching powder, and then try what effect this solution will have on the solution of asafœtida. Does the smell change?

A **disinfectant** is a substance which arrests the spread of specific disease, by destroying the special agent that enters the body from without and causes the disease.

A **deodorant** is a substance that purifies the air or destroys bad smells.

Bleaching powder is used extensively as a disenfectant and deodorant.

QUESTIONS AND PROBLEMS.

1. If moistened calico be put into dry chlorine, what element of the moisture will the chlorine then unite with?

2. What element, in its nascent state, would then act on the coloring matter of the calico?

3. Why will not oxygen that has stood for some time act in a similar manner?

4. How would you remove the stains of writing ink from a pocket handkerchief?

5. What part does the sulphuric acid play in the bleaching process? Try whether any other acid will answer the same purpose.

6. Try the effect of chlorine water on any bad-smelling solution which you can secure, ammonium sulphide, for example.

7. Use the knowledge which you have obtained, about the bleaching power of chlorine, to explain its disinfecting and deodorizing power?

§ 146.—Chlorine Compounds.

The compounds of chlorine are numerous and important, but only a few of them can be referred to here. The preparation of these few can easily be understood from the preparation of chloride of lime.

Experiments.

1. Prepare chlorine as before, and pass it into a test-tube containing a dilute solution of caustic potash. The test-tube must be kept *cool.* After some time, discontinue the process, and then try whether the solution obtained will bleach, as did chloride of lime. The reaction in preparing this solution is exactly anologous to that in preparing bleaching powder. Thus:—

$$Cl_2 + 2KHO = \underbrace{KCl}_{\text{potassic chloride.}} + \underbrace{KClO}_{\text{potassic hypochlorite.}} + H_2O.$$

Try whether you will get similar results by passing chlorine into a cold solution of caustic soda.

3. Pass chlorine into a cold solution of sodium carbonate. Sodium chloride $NaCl$, and sodium hypochlorite $NaClO$, will be formed. Try to write the equation which will express the reaction that occurs. Try whether the solution obtained will bleach.

§ 147.—Potassic Chlorate.

This salt is used extensively in medicine and in making oxygen gas. Its preparation illustrates an important fact in experimentation, namely this, that if we vary the conditions under which an experiment is performed, we change the results.

Experiment.

Boil a strong solution of caustic potash in a test-tube, and pass into it a current of chlorine for half an hour. Evaporate the solution to a small quantity and then allow it to cool slowly. Both potassic chloride KCl, and potassic chlorate $KClO_3$, will be formed in the solution. The latter being the least soluble crystallises out *first*. The liquid that remains contains the potassic chloride in solution. Pour off this liquid. To purify the crystals, re-dissolve them in a little hot water, and allow them to reform.

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§ 148.—Its Properties.

All chlorates yield oxygen readily.

Experiments.

1. Taste these crystals.

2. Powder a few of these crystals with a little charcoal, and heat the mixture on a piece of mica, charcoal b[illegible] with the potassium

3. Powder some more of the salt with dry sugar. Place the mixture on a tin plate or piece of cardboard, and add a drop or two of sulphuric acid with a pipette. [illegible]

the mixture [illegible]

§ 149.—Tests for a Chlorate.

Experiments.

1. Heat a small crystal of any chlorate.

2. Place a *very small* crystal of a chlorate in a test-tube, and add four or five drops of sulphuric acid. Warm gently, directing the mouth of the tube away from any person. This experiment is dangerous.

[illegible]

3. Dissolve a crystal of a chlorate in water; add a little indigo solution, and then a few drops of sulphuric acid. Explain the cause of the change of color.

The Cl bleaches out the color

QUESTIONS AND PROBLEMS.

1. Try to prepare chlorate of soda by a method analogous to that by which you prepared chlorate of potash.

2. How can a chlorate be distinguished from a chloride?

3. How can the chlorate of potash be converted into the chloride?

4. What physical state do the compounds formed in an explosion usually assume?

5. Why cannot chloric acid be formed from chlorate of potash in a manner analogous to that of preparing nitric acid from nitrate of potash?

CHLORINE AND OXYGEN.

Oxygen forms with chlorine three well known oxides, and two hypothetical ones.

FORMULA.	NAME.	CORRESPONDING ACID.
Cl_2O.	Hypochlorous anhydride.	$HClO$ Hypochlorous acid.
Cl_2O_3.	Chlorous anhydride.	$HClO_2$ Chlorous acid.
$Cl_2O_4(ClO_2)$.	Chloric peroxide.	No corresponding acid.
Cl_2O_5.	Not eliminated.	$HClO_3$ Chloric acid.
Cl_2O_7.	Not eliminated.	$HClO_4$ Perchloric acid.

All the compounds of oxygen and chlorine are unstable, and most of them are explosive, breaking up into chlorine and oxygen.

Just as sodium nitrate $NaNO_3$, yields nitric acid, when treated with sulphuric acid; and sodic chloride $NaCl$, yields hydrochloric acid; so potassium chlorate $KClO_3$ yields chloric acid $HClO_3$; and potassium hypochlorite yields hypochlorous acid, $HClO$. Thus:

$$2NaNO_3 + H_2SO_4 = Na_2SO_4 + 2HNO_3,$$
$$2NaCl + H_2SO_4 = Na_2SO_4 + 2HCl,$$
$$2KClO_3 + H_2SO_4 = K_2SO_4 + 2HClO_3$$
$$2KClO + H_2SO_4 = K_2SO_4 + 2HClO.$$

As chloric and hypochlorous acids, however, break up as soon as formed, the student is advised not to attempt to prepare them in this way.

EXERCISE.

1. How much chlorine by weight and volume can be obtained from 1460 grams of hydric chloride?

2. How much chlorine can be liberated from 585 grams of common salt? What volume will it occupy at 60° F.?

3. What volume will 284 grams of it occupy at 80° F.?

4. What quantities of manganic sulphate, hydro-sodic sulphate, water, and chlorine, will be formed by the decomposition of 351 grams of common salt, with manganese dioxide and sulphuric acid?

5. If 142 grams of chlorine gas be passed into steam at a red heat, what substances will be formed, and what weight of each?

6. What weight of hydric chloride will 261 grams of maganese dioxide require for its decomposition?

CHAPTER XXIV.

§ 150.—Halogens. 1 Fluorine

This is the name given to three elements 2 chlorine, 3 bromine, and 4 iodine, all having relations to each other that mark them off as members of the same group or family of elements.

Iodine is never found native. It occurs as iodides of sodium and other metals in "kelp," the ashes of sea-weed. These iodides, being soluble in water, are removed from the kelp by a process known as "washing." The element and its compounds are used extensively in medicine.

Iodine:—Symbol, I; at. wt., 127; mol. wt., I_2, *254; sp. wt., (of solid), 4.95. Melts at 107° C.; boils at 180° C.*

Experiments.

1. Dissolve two or three crystals of iodide of potassium in a long test-tube half full of water. Then add a few drops of strong chlorine water. When a sediment has settled to the bottom of the tube, pour off the liquid carefully, and then heat the residue over a lamp. The substance obtained is iodine. KI + [illegible] = [illegible] Cl + I.

2. Place 10 centigrams of commercial iodine in each of 4 test-tubes To the first test-tube add about 4 c. cs. of water; to the second, 4 c. cs. of alcohol; to the third, 4 c.cs. of chloroform or ether; and to the fourth, 4 c. cs. of solution of iodide of potassium. Dilute the solution, thus formed—each to the same extent—with water, and observe the effect in each case.

QUESTIONS AND PROBLEMS.

1. Devise an experiment to change the physical state of iodine. Does iodine when heated combine directly with the oxygen of the air?

2. Prepare iodine from iodide of potassium, by a method analogous to that in which chlorine was prepared. Complete the equation:—

$$MnO_2+2KI+3H_2SO_4= \quad . \quad . \quad . \quad . \quad . \quad .$$

3. How can you separate iodine from water? How can you increase the solubility of this element? Solubility of [illegible]?

4. Write out an account of the physical and chemical properties of iodine.

5. Find out whether iodine water will bleach litmus solution. a

6. Try to form an iodide of mercury by direct union of the elements.

In performing the last experiment you will probably form two iodides. How may this be known?

§ 151.—Compounds of Iodine.

Hydriodic acid, HI, is the compound of hydrogen and iodine which is analogous to hydrochloric acid. There are two oxygen acids of iodine: Iodic acid, HIO_3, and periodic acid, HIO_4, whose resemblance in composition to corresponding acids of chlorine can be seen at a glance.

To prepare hydriodic acid:

Experiment.

1. A solution of this gas in water may be prepared, very simply, by making a solution of iodine in water, and then passing into this solution sufficient sulphuretted hydrogen to decolorize it. Hydriodic acid will be formed thus:

$$I_2 + H_2S = 2HI + S.$$

2. Expose some of this aqueous solution to sunlight and air for a day or two.

Hydriodic acid is a heavy, colorless gas, which unites with hydrates, oxides, or carbonates of many of the metals, and forms iodides—the compounds formed being exactly analogous to those formed by hydrochloric acid. Some of these iodides possess a bright color and are therefore useful in identifying simple solutions of some of the metals.

3. Take three test tubes and half fill each with solutions of mercuric chloride, silver nitrate, and (acetate or nitrate of lead) respectively Add to each a little hydriodic acid solution.

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§ 152.—Test for Iodine.

Experiments.

1. Heat a few granules of starch in half a beaker full of water until the starch thickens; then add enough water to make a thin "mucilage." turns [illegible] blue

2. Half fill a test-tube with some of this starch mucilage, and to it add a few drops of aqueous solution of iodine. Note the effect, and then heat the test-tube nearly to boiling. when [illegible] the color

3. To another test-tube, half filled with the starch mucilage, add a little aqueous solution of iodide of potassium. Does this change the mucilage as free iodine does?

4. Devise a test for an iodide by liberating iodine from it. Test an iodide with nitrate of silver and compare the result with that for a chloride

§ 153.—Bromine.

Bromine, the third member of the haloid family, is never found native. It is usually prepared from bromide of potassium, a salt which is extensively used in medicine, and which is found with bromide of sodium in "kelp." Both bromides exist in kelp in smaller quantity than the iodides.

Bromine:—Symbol, Br.; at. wt., 80; mol. wt., 160; sp. wt., 2.99; solid at − 22°C.; boils at 63°C.

Experiments.

1. Dissolve a few crystals of bromide of potassium in a test-tube partly filled with water. Add to this solution a few drops of chlorine water. Boil the solution. The vapor that passes off is bromine. Observe its color and smell, but do not breathe it freely as it is corrosive and poisonous. Try to complete the equation which expresses the changes that occur:—

$$2KBr + Cl_2 =$$ [illegible] + [illegible] . .

2. Try to prepare bromine from bromide of potassium, in a manner analogous to that in which you prepared chlorine from sodium chloride. Keep the flask cool into which the bromine vapor passes. Try to complete the equation:—

$$MnO_2 + 2KBr + 3H_2SO_4 = \quad . \quad . \quad . \quad . \quad . \quad .$$

The compounds of bromine are of great interest from a theoretical point of view. The well-known ones are:

Hydrobromic acid,	$H\,Br$.
Hypobromous "	$H\,Br\,O$.
Bromic "	$H\,Br\,O_3$.
Perbromic "	$H\,Br\,O_4$.

The first of these compounds is a colorless, irritating gas, forming with metals the salts called bromides, of which some are important in the arts, and in medicine. A mixture of bromide and bromate of potash may be prepared by passing bromine into a strong solution of potassium hydroxide, in a manner analagous to that in which chlorate of potash is prepared

The color of the vapor, and the color which the element imparts to ether, or to starch paste, are tests for bromine in a free state.

QUESTIONS AND PROBLEMS.

1. Devise an experiment to ascertain whether bromine water will bleach as did chlorine water.

2. We know from experiment that free chlorine will displace bromine and iodine from bromide and iodide of potassium respectively. Devise an experimeet to find out whether free bromine will displace iodine in a compound.

3. Try to separate bromine from water.

4. Try to form bromide of mercury by direct union of the elements. Use *small* quantities of each element and work cautiously.

5. As hydrochloric acid was obtained by the action of sulphuric acid on sodic chloride, try to prepare, in an analogous manner, hydriodic acid, and hydrobromic acids respectively. If you fail to obtain these acids, turn to the reaction of copper on nitric acid § 128 and then try to explain the results you obtained in this instance.

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§ 154.—Fluorine F. At. Wt. 19.

This element has never been satisfactorily examined, until lately. The reason for this is that as soon as eliminated it combines with almost every substance with which it comes in contact. It is a colorless gas. It occurs in fluorspar or calcic fluoride CaF_2.

Fluorine is generally classified as belonging to the chlorine family. In regard to the chemical conduct of the members of this family, it may be said that fluorine is the most active; then chlorine; then bromine; and lastly, iodine. The variation in the properties of these elements, particularly chlorine, bromine and iodine, seems to be *connected in some way with the variations in the weights of their atoms.* But of this we cannot as yet be certain. Do other families exhibit a similar peculiarity? We shall see later on.

Quite recently it has been prepared by the electrolysis of hydrogen fluoride. M. Moissan, who was the first to examine it, reports that "in contact with mercury it was completely absorbed, forming yellow mercurous fluoride; with water, it formed ozone; phosphorus spontaneously inflamed in it, forming phosphorus fluorides; sulphur heated and melted; carbon had no action; fused potassium chloride was attacked in the cold, evolving chlorine; crystallized silicium took fire in the gas, forming silicium fluorides. The metals burn in it. Organic matters are violently attacked by it. It combines spontaneously with hydrogen in the cold, and with detonation. Cork, alcohol, ether, turpentine and petroleum are at once inflamed in contact with it."

Try to point out relations between the halogens Note more particularly resemblances in *(a)* atomicity, and in *(b)* the compounds they form with oxygen and hydrogen. Do they exhibit any sequence in *(a)* specific gravity *(b)* atomic weights, *(c)* physical state, and *(d)* color?

§ 155.—Hydrofluoric Acid, HF.

This is a somewhat important compound of flourine, inasmuch as it is used in etching on glass.

Experiment.

Make a small cup out of a piece of sheet lead. If you have no sheet lead at hand, hammer a leaden bullet into a saucer-like shape. Place in the vessel thus formed about 10 grams of powdered calcic fluoride, and then add about 15 grams of strong sulphuric acid. Warm a piece of glass and cover one side of it with beeswax. When cold, scratch any design you please through the wax. Place the glass, wax side downward, over the leaden cup. Place the cup on a rising stand in a gas chamber, and apply gentle heat. After a few minutes remove the glass and clean it with a rag dipped in turpentine. Do not breathe the gas; it is exceedingly poisonous. Has it any color?

CHAPTER XXV.

§ 156.—Lime.

Limestone is another common substance. Of what is it composed? Marble is a purer kind of this same substance, and from marble we obtained carbon dioxide, by pouring hydrochloric acid upon it. As this acid contains neither carbon nor oxygen, it follows that carbon dioxide must be one of the constituents of marble or limestone.

Experiments.

1. Take a hydrogen generating apparatus and place in it some powdered chalk or oyster shells. Cover with water. Insert the cork tightly and then pour down the thistle tube some hydrochloric acid. Collect the gas that escapes and test it for carbon dioxide.

2. Take thin pieces of marble, chalk and oyster-shell and weigh them. Place them upon an old tin plate and heat in a fire to redness, for an hour or two. Remove them and try whether they will now yield carbon dioxide on the addition of hydrochloric acid. The residue is called **Quicklime**, or simply lime. Note differences in color and weight between lime and limestone. In what two ways was carbon dioxide expelled from limestone?

---

§ 157.—Burning Limestone.

The manufacture of quicklime is carried out on a large scale by heating limestone or marble to redness in special furnaces called lime-kilns. Some of these are so arranged that they can be kept in operation for years without intermission. Many common ones are hollowed out of a hill side and lined with some hard stone. The analysis of lime shows that it is an oxide of a metal called calcium. Its formula is CaO. Calcium forms a large part—probably a sixteenth—of the solid crust of the earth, and occurs in limestone, marble, oyster shells, chalk, calcspar, coral, and many fossiliferous rocks and minerals. What two different compounds have we obtained from limestone?

Experiment.

The following is a simpler and quicker experiment than the last one, when you want to prepare quicklime on a small scale; but you need a Bunsen burner to do it with. Take a piece of calc-spar and round it wind some platinum wire. Hold the calc-spar in the hottest part of the Bunsen flame, and in a short time you have quicklime.

2. Moisten a piece of red litmus paper and press it against a piece of limestone or marble. Observe whether any change takes place.

3. Now press a similar piece of litmus against a piece of recently burnt quicklime.

§158.—Calcic Hydrate.

Calcic hydrate is the compound formed by the union of water with quicklime. Thus:

$$CaO + H_2O = Ca(HO)_2.$$

Experiments.

1. Take a piece of recently burnt quicklime, about 30 grams in weight. Place it upon a common plate and add to it slowly, about 15 or 20 c. cs. of water. What change takes place in temperature? makes warm In volume? larger Calcic hydrate is sometimes called slaked lime.

2. Take a piece of slaked lime the size of a bean; drop it into a small beaker half-full of water. Stir with a glass rod, and then pass half of it through filtering paper. The liquid that passes through is called **lime-water**.

[illegible]

$CaCO_3 + H_2O$ [illegible]

CO_2 [illegible]

The soft paste of slaked lime, caught in the filter paper, is called **cream of lime**. The mixture of limewater, and fine particles of slaked lime remaining in the beaker is called **milk of lime**.

§ 159.—Artificial Limestone.

If limestone, as our experiments would seem to show, be made up of carbon dioxide and quicklime, it ought to be possible to combine these two substances so as to form artificial limestone.

Experiments.

1. Pour into a beaker some of the limewater prepared in the last experiment, and then pass through it some carbon dioxide until a precipitate is formed. Filter the product obtained. Remove the white powder which is caught on the filter paper; dry it and compare with *finely* powdered chalk or marble. Calcium Carbonate

2. Place a drop of hydrochloric acid upon it. What gas is evolved? What is this white powder?

§ 160.—Solution of Limestone.

The solution of limestone in water is an important and interesting fact, and calls for investigation.

Experiments.

1. Pass carbon dioxide into limewater, but instead of stopping the process when a precipitate first forms, continue to pass the gas until the liquid again becomes almost if not altogether clear.

2. Boil the liquid just obtained. What phenomena result? Apply the results of these two experiments to explain the origin of the incrustations which form on the inside of kettles and boilers; also the formation of stalagmites and stalactites.

§ 161.—Carbonates.

It is quite evident from our experiments with carbon dioxide that it unites with water and forms an acid. The reaction may be thus symbolized :—

$$CO_2 + H_2O = \underbrace{H_2CO_3}_{\text{Carbonic Acid.}}$$

No such acid as carbonic acid has ever been isolated ; it is known only in its weak aqueous solution ; but its salts are numerous and well known under the name of carbonates. Limestone or marble is a carbonate of calcium.

QUESTIONS AND PROBLEMS.

1. Can you conclude that oyster shells and chalk have the same composition as marble and limestone, because they all yield carbon dioxide ?

2. Devise an experiment to prove that the substance formed by passing carbon dioxide into limewater resulted from the union of the gas with something dissolved in the water, and not from the union of the gas with pure water.

3. Try to prepare other carbonates in the same way as you prepared carbonate of lime.

4. Complete the following equations :—

$CaCO_3 + 2HCl =$
$Ca(HO)_2 + CO_2 =$
$2NaHO + CO_2 =$
$2NH_4HO + CO_2 = (NH_4)CO_2 + H_2O$

5. Try to prepare calcic hydrate in another way than by adding water to quicklime.

---

§ 162.—Mortar.

Slaked lime is used for purifying illuminating gas, and for whitewashing ; but its chief use is in making mortar. This substance consists of one part of lime and three or four parts of sand, the two being thoroughly mixed with water. Mortar absorbs carbon dioxide from the air and slowly turns into limestone, hence its hardening. Express the reaction by an equation.

$Ca(HO)_2$ [illegible]

§ 163.—Gypsum.

Another important salt of lime is gypsum, $CaSO_4 + 2H_2O$. This substance is found native, and when ground to powder is used as "land plaster." When calcined it is called "plaster of Paris," and is used in making casts, and for giving firmness to writing paper. What is the object in calcining gypsum? Explain what is meant by its "setting." How can you prepare calcic sulphate $CaSO_4$?

$H_2SO_4 + \underline{CaO = CaSO_4} + H_2O.$

CHAPTER XXVI.

§ 164.—Phosphate of Lime.

Phosphate of lime $Ca_3(PO_4)_2$ is another well known salt of calcium. It is an importent article of export from Canada, considerable deposits of it being found in the County of Frontenac and along the Ottawa Valley. "The mineral is found in the igneous rocks, from whose disintegration our alluvial soils have been produced; hence, most soils must contain phosphates." Transformed by the wonderful energy of life, these phosphates pass from the soil to the plant, and from the plant to the animal. Phosphate of lime forms the chief constituent of the bones and teeth of animals. The greatest part of the phosphorus of commerce is manufactured from bones. For a full description of the process of its manufacture, you must look elsewhere than in this book. All that can be done here is merely to indicate how the element can be prepared.

Experiment.

Treat bone-earth or the mineral phosphate of lime with dilute sulphuric acid, and there will be formed an insoluble calcic sulphate and superphosphate of lime. Thus:

$$Ca_3(PO_4)_2 + 2\,H_2SO_4 = 2\,CaSO_4 + CaH_4\,(PO_4)_2.$$

Filter the product, and throw away the insoluble sulphate. Then evaporate the solution to dryness, mix the residue with charcoal, and distil in an earthen retort, the nozzle of which passes under the water. Carbonic oxide will be formed ; and phosphorus vapor will pass over, and will be condensed under the water. This reaction may be thus represented :

$$3\,CaH_4(PO_4{}_2 + C_{10} = Ca_3(PO_4)_2 + 6\,H_2O + 10\,CO + P_4.$$

Only two-thirds the phosphorus is liberated in this process, and what is obtained is more or less impure.

There are two well-known allotropic forms of phosphorus—the yellow and the red.

§ 165.—Phosphorus.

Symbol, P ; Atomic wt., 31 ; Molecular wt., P_4, *124.*

Great care must be taken in experimenting with phosphorus. It should be kept, and cut under water. Why ?

Experiment.

1. Cut off a very thin slice from the end of a stick of yellow phosphorus. Dry on a piece of blotting paper; then place on a piece of wood, and rub with a pestle.

2. Dissolve about 25 c. g. of yellow phosphorus in about 5 c. c. of carbon disulphide. Dip into the solution a piece of common paper ; remove the paper and hang it on an iron retort stand. Explain what takes place.

§ 166.—Red Phosphorus.

Red phosphorus is prepared from the yellow variety by heating it for several hours, to a temperature of 240° C., in a vessel from which oxygen is excluded. The young student should not try to prepare it. What is needed can easily be bought.

Experiments.

1. Try to ignite a little red phosphorus by using moderate friction. [illegible]

2. Try whether the red variety will shine in the dark. no Will the yellow? yes

3. Try whether it will dissolve in carbon disulphide.

4. Devise an experiment with small quantities of red and of yellow phosphorus, to determine which variety will ignite first, by applying equal degrees of heat to each.
yellow ignites first

---

§ 167.—Matches.

This element is extensively used in the manufacture of lucifer matches. A paste is made of phosphorus, potassic chlorate or potassic nitrate, glue and powdered sand, and the ends of the matches are tipped first with sulphur and then with this paste. The "safety" match is tipped with a paste of potassic chlorate and antimonious sulphide, and the "rubber," upon which it has to be rubbed before it ignites, is made of red phosphorus and antimonious sulphide.

The ordinary match, as well as the "safety" one, ignites by friction. The rubbing generates sufficient heat to cause the phosphorus to take fire, and when once ignited the combustion is supported by the oxygen supplied by the potassic chlorate or potassic nitrate. What are the products of the combustion of a friction match

---

§ 168.—Combustion of Phosphorus.

Phosphorus in burning in air or oxygen forms two oxides, phosphorus trioxide P_2O_3, and phosphorus pentoxide P_2O_5, the former being produced when the supply of oxygen is limited.

Experiment.

Place a dry piece of yellow phosphorus on a common plate. Ignite it, and cover with a bell-jar, taking care that the mouth of the jar fits closely to the plate. After the white substance has subsided remove the jar, collect some of the white solid and preserve it in a stoppered bottle. Observe any change which the remainder undergoes.

These oxides unite with water to form acids. Try to write the equations for the reactions.

NAME OF OXIDE.	FORMULA.	CORRESPONDING ACIDS.
Phosphorus trioxide.	P_2O_3.	H_3PO_3 Phosphorous acid.
Phosphorus pentoxide.	P_2O_5.	H_3PO_4 Orthophosphoric acid. HPO_3 Metaphosphoric acid. $H_4P_2O_7$ Pyrophosphoric acid.

§ 169.—Orthophosphoric Acid.

This substance has the formula H_3PO_4, and is generally known as phosphoric acid.

Experiments.

1. Place a small quantity of red phosphorus in a small evaporating dish, and then set it on a retort stand. Add a small quantity of re-agent nitric acid and heat gently, adding more acid if necessary until the phosphorus has entirely disappeared, and red fumes cease to come off. Continue to evaporate until the excess of nitric acid has been expelled. The residue is orthophosphoric acid.

2. Pour a small quantity of the acid thus prepared, into a test-tube and add a few drops of silver nitrate solution. A *light-yellow* precipitate soluble in ammonia, nitric acid, and acetic acid indicates phosphoric acid or one of its salts.

§ 170.—Metaphosphoric Acid.

This is the acid usually formed when phosphorus pentoxide is added to water. The reaction is probably:

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

It is formed when orthophosphoric acid is heated to about 400°. Reaction:

$$H_3PO_4 - H_2O = HPO_3.$$

In the market it is usually known as "glacial" phosphoric acid. A solution of this acid at ordinary temperatures slowly changes to orthophosphoric acid; but the change takes place rapidly if the solution be boiled. How does the one acid change into the other?

Experiment.

Make a fresh solution of "glacial" phosphoric acid, and add to it a few drops of silver nitrate solution. A *gelatinous-white* precipitate, soluble in nitric acid indicates metaphosphoric acid or one of its salts.

171.—Basicity.

Orthophoric acid H_3PO_4, is interesting as illustrating what is called a **tri-basic** acid; that is, an acid which is capable of forming three kinds of salts, according as one, two, or three of its hydrogen atoms are replaced by one, two, or three atoms of a monad metal. For example:

NaH_2PO_4 is called dihydro sodic phosphate.
Na_2HPO_4 is called hydro disodic phosphate.
Na_3PO_4 is called trisodic phosphate.

The **basicity** of an acid, therefore, depends upon he number of exchangable atoms of hydrogen which its molecule contains. Nitric acid, HNO_3 is **mono basic**; sulphuric acid is **di-basic**. Only one of the three atoms of hydrogen in phosphorous acid, H_3PO_3, can be replaced by a metal, and consequently this acid is mono-basic.

A **neutral** salt is formed when all the replaceable hydrogen of an acid has been displaced by a metal: an **acid** salt is formed when only part of the hydrogen has been so displaced.

---

§ 172.—Phosphorus and Hydrogen.

These elements form three compounds—all known as phosphuretted hydrogen. The substance whose formula is PH_3 is a gas; the one whose formula is PH_2 is a liquid; while the one whose formula is supposed to be P_2H is a solid.

Phosphuretted hydrogen (Phosphine): Formula, PH_3*; molecular weight, 34; specific weight, 17. 11.2 litres litres weigh 17 grams.*

Experiment.

Take a *small* generating flask and in it put a strong solution of potassium hydroxide. Add a few small pieces of yellow phosphorus. Insert cork and delivery tube. Now pass in carbon dioxide or hydrogen gas, so as to *expel all air* from the apparatus. The delivery tube should dip below the surface of some water placed in an evaporating dish.* The reaction is:

$$P_4 + 3\ H_2O + 3\ KHO = 3\ KH_2PO_2 + PH_3$$

The *vortex motion* of the ring-shaped clouds which form when this substance ignites spontaneously in air, is very beautiful. Try to symbolize the reaction

QUESTIONS AND PROBLEMS.

1. Describe the physical properties of phosphorus.

2. What property of the element is observed in the dark? This property is beautifully illustrated by placing a small piece of phosphorus in a narrow-necked flask, half filling the flask with water, and then boiling in a dark room. Use apparatus similar to that in Fig. 41. This is the principle of Mitscherlich's test for free phosphorus.

---

* If all the air be not expelled from the flask an explosion will result. When it is desired to stop the formation of the gas, enough warm water should be poured into the evaporating dish to fill the flask. The lamp should then be withdrawn. As the solution cools, the warm water is drawn u into the inside of the flask, and the evolution of gas ceases completely [illegible]

[illegible] $P_2O_5 + H_2O$ = [illegible]

3. Tabulate in opposite columns the resemblances and differences between red and yellow phosphorus.

4. How would you think it possible to turn red phosphorus back into the yellow variety?

CHAPTER XXVII.

§ 173.—**Sand.**

In treating of the chemistry of the earth, some reference must be made to sand, and to sandstone from which sand has been formed by disintegration. This substance occurs in varying proportions in nearly all good soils, but a soil consisting of pure sand is almost entirely unproductive. Analysis shows that sand is an oxide of the element silicon, the most abundant solid element in nature. So widely is this element distributed, that it constitutes from 22 to about 36 per cent. of the solid crust of the earth. It is never found uncombined, and like carbon is known in three modifications — amorphous, graphitoidal, and crystalline. The chemical name for pure sand is silicon dioxide or silica SiO_2. The modifications of this compound are numerous, as will be seen by a glance at the following list:—

"1. *Quartz crystals*, glassy hexagonal prisms terminating in hexagonal pyramids.
2. *Amethyst*, *smoky quartz*, *rosy quartz*, and *chrysoprase*, colored varieties of quartz.
3. *Quartzite*, a sedimentary rock.
4. *Sand* and *sandstone*, fine fragments of quartz more or less cemented together.
5. *Honestone* or *novaculite*, a fine-grained quartz rock.
6. *Chalcedony*, a mixture of crystalline and non-crystalline quartz.
7. *Agate*, consisting of layers of crystallized and amorphous quartz of various colors.
8. *Flint* and *chert*, a coarse variety of chalcedony.
9. *Opal*, a hydrated form of silica.
10. Various modifications of the above in which one form is passing into another."

§ 174.—Use of Sand.

Sand, or silica as we shall now call it, is that constituent of the soil which give firmness and stiffness to the stalks of grains and grasses, so that they are enabled to maintain the upright position.

§ 175.—Silicates.

Silica is insoluble in water and in all acids except one. Theoretically, it is supposed to unite with two molecules of water to form silicic acid. Thus :—

$$SiO_2 + 2H_2O = \underbrace{H_4SiO_4.}_{\text{Silicic acid.}}$$

No such direct union of water and silica has ever been obtained by experiment, but this acid can be prepared indirectly, and its compounds with metals, that is, its salts, are well known, and as we shall see presently, are important. These salts are called silicates. Silica may therefore be considered the anhydride of silicic acid. What other anhydrides of acids have you studied?

§ 176.—Clay.

Clay, like sand and limestone, is a very abundant substance. Of what is it composed? We found that limestone was made up of an oxide of calcium and carbon dioxide; similarly, it has been found that clay is composed of an oxide of the metal aluminum, Al_2O_3, united with silica SiO_2; it is, in short, a silicate of alumina. Aluminum, the metal of clay, is a white body with a brilliant lustre; it is 2.67 times heavier than water; may be rolled out into sheets, and drawn into fine wire; and it is not easily oxidised. The great cost of extracting it is the only reason why it is not extensively used in the arts. Every clay bank is a mine of aluminum, so that, next to oxygen and silicon, aluminum is the most abundant of elements. The metal gets its name from the fact that it is found in alum.

§ 177.—Alumina.

Alumina is an oxide of aluminum, Al_2O_3. A variety of this compound is known in nature as corundum, or emery, when coarse; but when crystallized it is known to jewellers as sapphire, ruby, oriental emerald, oriental topaz, and oriental amethyst.

Experiments.

1. Take a piece of ammonia alum and heat *very strongly* until all action has ceased. The residue is alumina.

2. Dissolve some alum in water; add ammonia; filter. Remove the filtrate from the paper, and heat very strongly on a piece of platinum foil. The residue is alumina. What compound is precipitated?

§ 178.—Soluble Glass.

Potassium silicate and sodium silicate are often called soluble glass. They are used in making artificial stone. Soluble glass may be prepared as follows:—

Experiment.

Powder 4 grams of fine white sand and place in an iron ladle. To this add 8 grams of caustic potash and about 70 grams of water. Boil the whole for three or four hours over a hot fire, supplying fresh water as it becomes necessary. Reaction:—

$$SiO_2 + 4KHO = \underbrace{K_4SiO_4}_{\text{Potassium silicate.}} + 2H_2O.$$

§ 179.—Glass.

"The various glasses of commerce are mixtures of a highly silicious silicate of sodium or potassium or of both, with silicates of other metals such as calcium, aluminum and lead." Thus, window glass is a double silicate of sodium and calcium; Bohemian glass (hard test-tube glass) consists of silicates of potassium and calcium; flint glass contains silicates of potassium and lead; and bottle glass is a mixture of silicates of alumina, calcium, iron and sodium.

§ 180.—Pottery.

The arts of moulding and pottery depend upon the plastic properties of various kind of clay, when mixed or puddled with water.

Experiment.

Take a little kaoline or any common clay and place upon a plate. Add a few drops of water and knead it into a thick paste. Mould it into the shape of a small cup, and try whether this cup will hold water. Dry thoroughly but slowly, and then place carefully in a *very* hot fire and bake for a few hours.

§ 181.—Soil.

"All soils have been produced by the disintegration of rocks, generally through the prolonged action of water, air, and frost. The character of a soil largely depends upon the character of the rocks from which it has been derived. Primitive and igneous rocks yield soils rich in potash; fossiliferous rocks produce soils rich in phosphoric acid. The principal ingredients of soils are sand, clay, carbonate of lime, and humus, or decayed vegetable matter. As each of these preponderate the soil is sandy, clayey, calcareous, or peaty."

[illegible] 32 8 + 15 [illegible] 40 + 10 − 25 [illegible] 10 + 10 + 45

CHAPTER XXVIII.

§ 182.—Analysis of Air.

We have already learned how to make a simple volumetric analysis of air. The result of our experiment in § 28 seemed to indicate that oxygen and nitrogen occurred in air in the proportions, roughly speaking, of one of the former to four of the latter.

A more accurate analysis of the proportions by volume in which oxygen and nitrogen exist in air may be made by means of a eudiometer.

Experiment. 230 div 115 div of H added

Put a known volume of pure air into the eudiometer, and then a volume of hydrogen sufficient to unite with all the oxygen in the air. Note the total volume. Now pass a spark of electricity into the eudiometer, taking all the precautions mentioned in § 31.

left 237 div.

230 + 115 = 345 − 237 = 108 div

QUESTIONS.

1. What volume of air was passed into the eudiometer? Of hydrogen? How much, therefore, of both together? After explosion, what volume of gas was left?

2. What was the diminution in volume? Why must ⅓ of this diminution represent the oxygen present in the air which was first passed into the eudiometer?

3. How then is the volume of the nitrogen in air found?

Twenty-eight volumetric analyses by Bunsen gave the following results:—

Oxygen	20.92	per cent.
Nitrogen	79.08	"

§ 183.—Composition of the Air by Weight.

Experiment 1, in § 25, illustrates a method of ascertaining the composition of air by weight. A known weight of air is passed over a known weight of red-hot copper. The gain in weight of the copper represents the weight of oxygen present in the air. Knowing the weight of the oxygen, that of the nitrogen is easily found. The gravimetric analysis of air, made on this principle, by Dumas and Boussingault, shews its composition to be :—

Oxygen	23 per cent.
Nitrogen	77 “

---

§ 184.—Specific Weight.

We have already seen that the composition of air by volume is, nitrogen 79·1 per cent., and oxygen 20·9 per cent. Knowing this, we can easily calculate its specific weight; thus :—

			a't. w't.		
Nitrogen	79·1	×	14	=	1107·4
Oxygen	20·9	×	16	=	334·4
	100 volumes.				1441·8
∴	1 volume.				14·42

This corresponds very nearly with its specific gravity as found by actual experiment, viz., 14·44. The number 14·44 will, of course, represent the weight of 11·2 litres of air in grams.

---

§ 185.—Air a Mixture.

Having analyzed air, the student is now in a position to answer the question : *Is air a mixture, or is it a true chemical compound?* To assist in deciding this point, the student should devise means of preparing air synthetically. In doing this he should observe carefully whether the oxygen and nitrogen on being brought together show any of the characteristies of a chemical change. He should watch particularly for (1) a change of temperature, (2) change of color, (3) change of volume, (4) change of state, or other evidences of chemical action. He should also inquire whether the analysis of air shows oxygen and nitrogen to be united in simple proportions by weight and volume, such as is the case in the union of oxygen and hydrogen to form water, and of carbon and oxygen to form carbon dioxide.

Experiment.

Expel the air from water and then analyze as in § 28. Compare the proportions by volume in which oxygen and nitrogen occur dissolved in water, with the proportions in which they exist in air. Would these proportions be the same in both cases if air were a chemical compound?

§ 186.—Impurities in Air.

Chemists have agreed to consider *pure dry air* as being made of nothing but oxygen and nitrogen; but we have already seen that there is present in ordinary air another substance, viz., vapor of water. Our experiments with burning carbon would certainly lead us to suspect the presence of yet another ingredient in air. Moreover, the fact that lime water turned milky when expired air was passed through it, might also suggest the presence of the same ingredient in the air. These conjectures we must now put to the test of experiment.

Before trying, however, -to ascertain whether our conjectures are right or wrong, we shall make a preliminary experiment which may aid us in our investigations.

Experiment.

1. Take a bottle such as is represented in Fig. 46. Partly fill with lime water. Then apply the mouth to the shorter tube limit and suck in air, drawing in fresh air through the longer tube. Do not continue this process for longer than a minute. Now reverse the process and blow air through the longer tube for about the same length of time.

FIG. 46.

2. Clean the bottle used in this experiment, and again partly fill with fresh lime water. Attach an *aspirator* to the end of the shorter tube, and draw air through the longer tube, so that you can count the bubbles as they pass. Continue the process for half an hour or thereabouts. If you have no aspirator, you should use a large bottle filled with water, and arranged as in § 25, when preparing nitrogen.

§ 187.—**Ozone.**

Symbol O_3*; molecular weight, 48.*

This substance is merely a modified form of oxygen, being one in which there are supposed to be three atoms in the molecule instead of two. It corrodes india-rubber; oxidizes silver, mercury, and many other metals, and, in doing so, changes to ordinary oxygen; possesses great disinfecting powers, and is a powerful bleaching agent. If breathed in small quantities it is said to be beneficial in the treatment of affections of the throat and lungs. City air contains little ozone, but in sea and country air it is present in considerable quantity.

Experiment.

Make a strong solution of potassium permanganate $K_2Mn_2O_8$, and place a little of it in a test-tube. Then add a few drops of strong sulphuric acid. Ozone is given off, and may be distinguished by its odor.

§ 188.—Organic Matter.

We have now mentioned all the substances usually enumerated in giving the composition of air, but it is evident that we have not exhausted the list. By more delicate means of analysis than you can be expected to employ as yet, chemists have discovered traces of ammonia and nitric acid in air, and, in the air of large towns more particularly, traces of sulphurous anhydride and sulphuretted hydrogen. Besides all these, there are few people who have not noticed the presence of substances that are peculiar to the air of certain localities. The air is supposed, and certainly not without reason, to be the medium through which the germs of certain infectious diseases are carried from one point to another. Minute particles of dust and smoke pervade the air of most towns.

§ 189.—Composition of Air.

The average composition of air is about as follows—

	Cub. centimetres.
Oxygen, including Ozone	206·1
Nitrogen	779·5
Aqueous vapor (invisible vapors clouds and fog)	14·0
Carbon dioxide	·4
Traces of ammonia, nitric acid and sulphuretted hydrogen	.0
Organic matter	·0
Dust and smoke	·0
Total	1000·0

§ 190.—Analysis of Expired Air.

Experiment.

Fill a small graduated tube with mercury and invert over another and larger tube also filled with mercury. Pass into the graduated tube, by the use of a curved pipette, as in Fig. 47, a measured volume of expired air. Immediately afterwards pass into the graduated tube a few drops of a strong solution of caustic potash or caustic soda, and note any change in the volume of confined air. Pass in also a little strong pyrogallic acid, and again note any change in volume after adjusting the level of the mercury inside and outside the graduated tube. 12.

Caustic potash [illegible] of CO_2

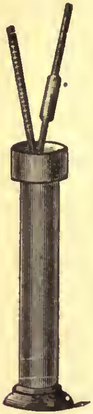

FIG. 47.

QUESTIONS

1. What effect has pure air on limewater . Expired air on lime water ? Explain the difference.

2. What effect has caustic soda or caustic potash on expired air ? Express in the form of an equation. What ingredient did the potash remove?

3. What volume of expired air did you put into the graduated tube ?

4. What ingredient did the pyrogallic acid remove ? What volume of it ? What gas must have remained in the tube ? What volume of it ?

5. Compare the proportions by volume in which oxygen and carbon dioxide respectively exist in common air and in expired air.

§ 191.—Plant Influence on Air.

We have learned how that, in respiration and in all ordinary combustion, air is being contaminated with enormous quantities of carbon dioxide. Do other animals besides man consume oxygen in respiration ? If so they are but increasing the amount of poisonous gas present in the air, and consequently lessening the time within which all the oxygen available for the support of life will finally disappear. The question therefore naturally arises, *why does not air soon become so impure as to be unfit to support human life ?* This question we shall endeavor to answer in our usual way.

Experiments.

1. Fill a large glass bottle with spring water, or better still, with drinking water, previously impregnated with carbon dioxide. Then invert it over a plate containing water as in Fig. 48. Put into the inside of the bottle a sprig of fresh mint or of water cresses, and expose the whole to strong sunlight for two or three hours. Examine carefully and test with a glowing splinter the gas which collects in the upper part of the bottle.


FIG. 48.

2. Repeat the experiment, keeping the plant in a dark room.

§ 192.—Diffusion of Gases.

Air, however, whether found on mountain tops or in valleys, whether taken from the green fields of England or from the vast prairies of America, is kept uniform in composition through the influence of a physical property which is characteristic of all gases.

Experiments.

1. Fill two separate bottomless glass bottles with oxygen and nitrogen respectively, and place one above the other, as shown in Fig. 49. Then insert a burning candle through the mouth of the upper bottle, and quickly lower it into and raise it out of the oxygen gas.

2. Repeat the experiment, but after bringing the two bottles together, wait for half an hour before testing with the candle.


FIG. 49.

FIG. 50.

3. Fit a new and clean porous cell of a galvanic battery with an India rubber cork and tube, and support in a beaker of water, as in Fig. 50. Then fill a large beaker with hydrogen and bring it mouth downward, over the porous cell.

QUESTIONS.

1. What effect have plants on carbon dioxide dissolved in water? What force produces the change? they absorb the (Sunlight)

2. Is it fair to assume that plants have the same effect on the gas, when in air?

3. What do plants do with the carbon of the carbon dioxide?

4. What other influences besides that of plants help to keep the air uniform in composition over the whole world.

+ 5. What visible phenomena occurred in experiment 3, § 192? Did the bubbles contain air or hydrogen? Give reasons for your answer?

6. Use other gases than air and hydrogen in this experiment, and note changes in the phenomena.

§ 193.—Graham's Law of Diffusion.

The rates of diffusion of two different gases in contact with each other is inversely proportional to the square root of their atomic weights.

§ 194.—Ventilation.

According to Dr. Parkes the air in our houses becomes unwholesome when carbon dioxide exceeds 6 parts in 10,000. If this be so, it becomes a matter of great importance to know when this amount is exceeded. Dr. Angus Smith, a celebrated authority on sanitary science, lays down the following rule :—" Let us keep our rooms so that the air gives no precipitate when a 10½ oz. bottle of the air gives no precipitate with half an ounce of clear lime water." Devise an experiment to test the purity of the air in your school room, employing Dr. Smith's test.

CHAPTER XXIX.

§ 195.—Impurities in Water.

The solvent action of water and the consequent impossibility of obtaining it pure in nature, have already been studied. We propose in this chapter to aid the student in testing for the impurities to be found in ordinary drinking water. Obviously our first step is to classify natural waters in some way, and to specify the impurities usually found in each class of water, and then to leave to the learner the task of discovering, by appropriate tests, what kind of impurity is present in the waters of the locality in which he lives.

§ 196.—Natural Waters.

Naturally occurring waters may be divided into four classes :—

1. **Rain water.**—This being the product of natural distillation, is the purest form in which we meet with water in nature. Even this, however, contains as impurities, the gases of the air, traces of common salt, ammoniacal salts, and organic matter of various kinds. If collected from the roof of a house it will, of course, contain other impurities besides those mentioned.

2. **Spring water.**—The impurities in this kind of water will vary greatly, and will be determined chiefly

by the character of the soil through which the water percolates before reaching the surface of the earth. It contains sodic chloride (salt), calcic sulphate, small quantities of magnesia carbonate and sulphate, silicates and a variety of other substances. Carbon dioxide is frequently found in spring water. Less frequently this water is impregnated with sulphuretted hydrogen. Well water is, of course, only a variety of spring water. In addition to the impurities which may be found in ordinary spring water, city well water is likely to contain ammonia, the nitrates and nitrites of calcium and sodium, and worst of all the drainage from animal refuse. These impurities are considered to be exceedingly deleterious to health, as they frequently give rise to various kinds of disease, such as typhoid fever. The well water of towns is more likely to be contaminated with sewage matter than that of country districts. Even in country districts, however, wells are often placed near stables and byres ; and the water used for drinking and culinary purposes is thus frequently rendered most unwholesome.

Springs, whose waters contain considerable quantities of salts of various kinds in solution, are called mineral springs. Such are Saratoga, in New York State, and Caledonia in Canada. Mineralized waters have in all ages been highly esteemed for their medicinal properties.

3. **River water.**—The suitability of river water for drinking purposes will depend very much upon the character of the organic matter suspended or dissolved in it, and this organic matter again will depend upon the region drained by the stream. Rivers usually contain a much less amount of salts in solution than spring water. If the district drained is low and marshy, and the course of the stream sluggish or nearly stagnant, the water is not likely to be wholesome. But if, on the other hand, the district drained be hilly and the current a rapid one, or if the bed of the stream be rocky or sandy, the water will likely be thoroughly good.

Running water is fitter for drinking than stagnant, because its motion exposes a fresh surface to the air, so that the oxygen is continually absorbed and oxidises the animal and vegetable matter in the water. In this way noxious compounds are often transformed into others which are perfectly harmless.

4. **Sea water.**—This is, of course, never used for drinking or culinary purposes unless first purified by distillation. It is the product of the washing of the land for long ages, and it contains, as might be expected, enormous quantities of soluble salts of various kinds. In what part of the ocean would you expect to find water of the greatest specific gravity? Messrs. Thorpe and Moreton analysed sea water from the Irish Channel, with the following result:—

Water	966.144
Sodium chloride	26.439
Potassium chloride	0.746
Magnesium "	3.150
" bromide	0.070
" sulphate	2.066
" nitrate	0.002
Calcium sulphate	1.331
" carbonate	0.047
Lithium chloride	traces
Ammonium chloride	"
Iron carbonate	0.005
Silica	traces
	1000 parts.

QUESTIONS AND PROBLEMS.

1. Catch some rain water in a large, clean earthenware or glass vessel, and find out how much gas is dissolved in it.

2. Test some of it for ammonia. Nessler -

3. Test a fresh sample of it for common salt. nitrate of silver

4. Mix together some alcohol and distilled water, and then try to prepare some pure water from the mixture by distillation.

5. Mix other liquids with pure water and then try to separate them by distillation.

fractional distillation.

§ 197.—Tests for Impurities.

The presence of ammonia, with chlorides and nitrites in drinking water, generally indicates contamination with sewage water. To decide, therefore, whether water is potable or not, the student should test for ammonia, then for chlorides, and lastly for nitrites. Water might however still be unfit for drinking, though containing none of these impurities. Organic matter of vegetable origin such as is found in the waters of some sluggish streams may render such water unfit for use. Evidently then, our next step should be to test for organic impurities. Two tests are usually given, one when little organic matter is present; the other, when *much* is present.

§ 198.—Organic Impurities.

Experiments.

1. Place the water to be tested in a Florence flask, and add to it, first, a few drops of sulphuric acid, and then enough of a solution of permanganate of potash to give to the whole a deep purple tint. Set to one side for an hour or two, in a warm place, and if the solution loses its color, organic impurities are present.

2. Fill a bottle with the water to be tested and cork it very tightly. Set aside in a warm place for a few days and then examine. An offensive odor indicates the presence of *much* organic matter.

§ 199.—Hardness.

Water that contains magnesium and calcium salts, and curdles soap, is said to be *hard*; water that does not contain these salts is *soft*. Hardness is usually considered as being of two kinds, viz., *temporary* and *permanent*. The former is due to the presence of calcic and magnesic carbonate, the latter to the presence of salts of calcium and magnesium other than the carbonates, such as the sulphates and nitrates.

§ 200.—Clark's Solution.

We usually test the degree of hardness by what is known as Clark's solution. This may be prepared as follows; "Dissolve 10 grams of castile soap in 1 litre of dilute alcohol, containing about 35 per cent. of pure spirit." The solution deteriorates by standiug. A *degree* of hardness represents .064 gm. of calcic carbonate or its equivalent in soap destroying power, in 4543.5 c. c. of water. In calculating the degree of hardness, we include, of course, both temporary and permanent hardness.

§ 201.—Temporary Hardness.

Experiments.

1. Half fill a large test tube with clear lime water, and then pass into it a stream of carbon dioxide. At first the usual result follows; but as the stream is continued the milkiness gradually disappears on account of the formation of a supposed bicarbonate of lime $H_2Ca(CO_3)_2$, which, unlike the carbonate, is soluble in water. Water with the bicarbonates of calcium of magnesium and of iron dissolved in it, is said to be *temporarily* hard.

2. Divide the water obtained in the preceding experiment into three parts and perform the following operations:

makes [illegible] soap,

(a) To the first part add some of Clark's soap solution.

(b) Boil the second part of the solution for half an hour. Then filter and test the filtrate with Clark's solution.
it curdles the soap

(c) Add carefully, to the third part, some clean lime water. Filter and test with Clark's solution. [illegible]

§ 202.—Permanent Hardness.

Experiments.

1. Fill a large test tube or small Florence flask about three-fourths full of pure water. Then add a gram or

two of calcic sulphate $CaSO_3$, and shake vigorously. Filter. The clear water contains calcic sulphate in solution and is *permanently* hard.

2. Divide the water thus obtained into four parts, and perform the same experiments with three parts of it that you did with the water which had bicarbonate of calcium dissolved in it.

3. To the fourth part add a solution of washing soda, Na_2CO_3. Note the result carefully.

Filter and test the filtrate as in the preceding cases with Clark's solution.

QUESTIONS AND PROBLEMS.

1. Devise an experiment to distinguish in a rough way, drinking water that contains organic matter of *vegetable* origin; from drinking water that contains organic matter of *animal* origin.

2. Explain how organic matter may be removed from drinking water?

3. Complete the two following equations. When complete they represent the reaction of washing soda on two kinds of hard water.

$$H_2Ca(CO_3)_2 + Na_2CO_3 = 2NaHCO_3 + CaCO_3$$
$$CaSO_4 + Na_2CO_3 = CaCO_3 + Na_2SO_4$$

4. Explain why some hard waters may be softened by boiling. Mention other methods of softening hard water.

5. Ascertain whether rain water that has been collected from the roof of a house contains any organic matter.

6. Ascertain whether spring water and river water contain any organic matter.

$$(H_2Ca(CO_3)_2)_{\text{heated}} = \underline{CaCO_3 + H_2O} + CO_2\uparrow$$

CHAPTER XXIX.

§ 203.—Definition of Chemistry.

We have now reached a point in our studies at which it will be profitable for us to pause, and take a general view of the subject, so far as we have studied it. We have had revealed to us something of the wonderful nature of many familiar substances, such as air, water, sulphur, charcoal, limestone, sand, and clay; and have observed the marvelous changes which these substances, and others derived from them, undergo when subjected to the influence of heat, light, electricity, or the still more subtle force, chemical affinity. At the beginning of our studies the fact was made prominent that one of the first questions which a chemist asks himself regarding any substance is: *Of what is it composed?* While this is quite true, you have now to learn that he also asks himself: *What can this substance do?* This second question is by all odds more difficult than the first; but, notwithstanding this, the chemist patiently and laboriously endeavors to answer it. In one case he seeks to become acquainted with the facts and principles concerned with *chemical composition*, in the other he endeavors to grasp the principles and mode of *chemical action.* Chemistry, therefore, may be defined as the science which deals with chemical action and its products, whether this action has taken place in past ages or in our own times.

§ 204.—Chemistry and Physics.

The student is now in a position to appreciate the distinction between physics and chemistry, and to understand their relation to each other. Physics is that branch of knowledge which deals with motion, light, heat, electricity, sound, magnetism, and all physical changes. Chemistry deals with the composition of substances, the changes in composition which substances undergo, and the conditions and laws which govern these changes.

Although physical and chemical changes appear at first sight to be quite unrelated, they are not really so. Motion when stopped turns into heat; heat generates motion. The locomotive moves through the influence of heat; but what causes the heat? Evidently the burning of the fuel. In this burning, however, a change in the composition of the fuel is actually taking place all the time, and this chemical change is generating heat, and the heat is causing the motion. We have frequently seen a chemical change generating heat, light and electricity; and we have seen that electricity, light, and heat may in turn produce a chemical change. Hence the relation of physics to chemistry. "Chemistry deals with those reactions between bodies wherein profound modifications in the properties of the bodies occur. Physics is concerned with the properties of this or that kind of matter considered for the most part apart from the action on it of other kinds of matter." In other words, the properties of the inter-acting bodies are not profoundly modified.

§ 205.—Chemistry and Biology.

Biology includes the two sciences usually known as botany and zoology. What relation do these studies bear to chemistry? The number and extent of the changes seen in the organic world are even more striking and interesting than those of the inorganic; and their causes are often so masked as to baffle all investigation. We usually attribute their origin and sequence to that secret, intangible and invisible power, which we, in our ignorance, call *life*. With the phenomena of life as such, the chemist does not primarily concern himself. He investigates the properties and composition of all substances formed by the agency of life, and seeks to build up in his laboratory the same compounds; but he does not concern himself with the relation of the life products to each other, except as chemical compounds; nor does he study the history of their birth, life and death. The compounds which he finds ready formed in the bodies of animals and plants—these are the objects of the chemist's studies, to the same extent and in the same manner exactly, as are the various objects of the mineral portion of the earth.

§ 206.—Chemical Action.

We have now to inquire whether the numerous chemical actions which we have observed cannot be classified in some comprehensive way so as to render their study more easy. Obviously, the simplest kinds of chemical action are those of (1) decomposition, and of (2) direct combination, of which the analysis and synthesis of water respectively furnish excellent examples. But there are other actions which cannot be classed as belonging to either of these. The preparation of nitric acid from nitre and sulphuric acid is a case in point. Here one element of the nitre is exchanged for one element of the sulphuric acid. Such a change as this is called one of *double decomposition* or *metathesis*. There is one variety of this kind of action to which a special name has been given, namely, *decomposition by substitution.* An example of this is furnished by the action of sulphuric acid on zinc. In this case the acid is decomposed, and zinc substituted for hydrogen. But as can be easily seen, this action is the same in principle as that which takes place between nitre and sulphuric acid, in which the potassium of the nitre is substituted for the hydrogen of the acid. Another variety of double decomposition is that sometimes called *reduction.* The action of hydrogen upon hot black-oxide of copper is an example. Here the copper of the oxide is exchanged for the hydrogen; or, in other words, the hydrogen is substituted for the copper. It is quite evident, therefore that the actions sometimes separately classified as *decomposition by substitution*, and as *reduction*, are really the same in kind as *double decomposition.*

Only one other kind of action needs be mentioned, *isomeric* action—in some respects a most remarkable one, and one which finds its fullest illustration in organic chemistry. Its chief characteristic is that, in a compound undergoing this change, the atoms composing a molecule appear to re-arrange themselves in some new way, and thus form another compound *differing in pro-*

perties from the original, but being *exactly the same in composition.* No examples of this kind of action have as yet come under our observation, but the following equation which represents a change of this kind may serve to make the meaning clear:

$$(NH_4)CNO \rightarrow CH_4N_2O.$$

The first of these compounds is known as ammonium cyanate, the second as urea; and the second may be derived from the first by simply heating it. Isomerism is analogous to allotropism, and both find their best explanation in the theory that the relative positions of the atoms in the molecule of one substance are different from what they are in the other, and this change in the relative position of the atoms is accompanied by changes in properties.

All cases of chemical action must belong to one or other of these four classes, viz., (1) decomposition, (2) direct combination, (3) double decomposition, and (4) isomeric; and the student is advised to refer every action coming under his observation to the particular class to which it belongs.

CHAPTER XXX.

§ 207.—Dalton's Theory.

The foundation laws of chemistry, usually known as the laws of *definite proportion* and of *multiple proportion*, have already been referred to. There is a third one, at any rate it is often called a third, which is really a corollary from the first law. This third is called the law of *reciprocal proportion*. As usually stated, it is this: "If two elements combine in certain proportions with a third, they combine in the same proportions with each other." For example, hydrogen and oxygen combine with each other in the proportion of 1 to 8, and oxygen and nitrogen in the proportion of 16 to 14; we would, therefore, expect that nitrogen and hydrogen would combine with each other in the proportion of 1 to 14. This we know to be the case, from the analysis of ammonia. To account for the laws of definite,* reciprocal, and multiple proportion, Dalton enunciated the theory that *all kinds of matter are made up of indivisible particles called atoms.* But he did not confine himself to merely propounding the theory; he endeavoured to find out the relative weights of the atoms. From the invariable composition of water by weight, Dalton imagined that it was composed of one atom of hydrogen united with one atom of oxygen, and as the proportion by weight in which these elements occur in water is as 1 to 8, Dalton concluded that an atom of oxygen was eight times heavier than an atom of hydrogen.

In 1809, a French chemist, Gay-Lussac, published a memoir in which he asserted that gases always combine in some simple proportion by volume. His words are: "They always combine in equal bulks, or one part of one by bulk, with two or with three parts of the other." Dalton rejected Gay-Lussac's conclusions, because he could not reconcile them with his own facts and theory. He regarded Gay-Lussac's experimental methods as untrustworthy.

Avogadro's theory, published in 1811, has already been explained. In some respects the conclusions reached by these three celebrated investigators contradict each other.

1. Dalton concluded from the known composition of water by weight that one atom of oxygen was eight times heavier than one atom of hydrogen.

2. Gay-Lussac proved that two volumes of hydrogen united with one volume of oxygen to form two volumes of steam.

3. Avogadro concluded that one atom of oxygen was sixteen times heavier than one of hydrogen.

"If Dalton's *definition* of atom and his rules regarding atomic synthesis are adopted, Avogadro's and Gay-

---

* Properly speaking, Dalton adopted and modified the old atomic theory of Lucretius

Lussac's statement that equal volumes contain equal numbers of atoms must be abandoned. The difficulty was removed by Avogadro, who introduced the idea of two kinds of atoms—'molécules integrantes,' or as we should now call them molecules; and 'molécules élémentaires,' or as we should now say atoms. The molecules of elements are decomposed in chemical processes, said Avogadro, and the atoms unite to form new compounds. The reaction between nitrogen and oxygen, inexplicable by Gay-Lussac's law, now becomes clear; each molecule of nitrogen and each molecule of oxygen divides into two parts, and these parts unite to form the new molecule of nitric oxide, hence there are twice as many molecules of nitric oxide produced as there were molecules of nitrogen or oxygen originally present."

The gain to the chemist in adopting the molecular theory has been immense, It affords the best known explanation of such phenomena as the diffusion of matter, diffusion of motion, diffusion of heat in gases, evaporation, condensation, electrolysis and spectroscopy.

Analysis by electricity

§ 208.—Division of the Molecule.

Let us now consider a few chemical reactions with which we have become familiar. Take first the one between oxygen and hydrogen in which steam is formed:

2 vols. of hydrogen unite with 1 vol. of oxygen to form
2 vols. steam;

therefore, since equal volumes contain equal numbers of molecules,

2 mols. of hydrogen + 1 mol. of oxygen = 2 mols. steam.
$\therefore$ 1 mol. hydrogen + $\frac{1}{2}$ mol. of oxygen = 1 mol. steam.

Hence, it becomes evident that the molecule of oxygen in this reaction must be divided into two parts. Again:

2 vols. of hydrogen unite with 2 vols. of chlorine to
form 2 vols. of hydrochloric acid,

and since equal volumes of gases contain the same number of molecules, it follows that

2 mols. of chlorine + 2 mols. of hydrogen =
4 mols. of hydrochloric gas,
∴ 1 mol. " + 1 mol. " =
2 mol. of " "

Now, since each molecule of the hydrochloric acid *must* contain both hydrogen and chlorine, it is evident that each molecule of hydrogen and each molecule of chlorine must have divided into it at least two parts, and from the union of these the new molecule of hydrochloric acid has resulted.

Again :—

2 vols. of nitrogen combine with 6 vols. of hydrogen to form 4 vols. of ammonia (gas),

$$\therefore 2 \text{ mols. of N} + 6 \text{ mols. of H} = 4 \text{ mols. of } NH_3,$$
$$\therefore \tfrac{1}{2} \text{ " of N} + \tfrac{3}{2} \text{ mol. of H} = 1 \text{ mol. of " }.$$

Here again it is evident that the nitrogen molecule must have been separated into two parts. And so, when various other reactions between gaseous substances are studied, the same conclusion is found to be unavoidable: the chemist is *forced* to recognize a smaller portion of matter than the molecule. This smaller portion he calls an atom. The student must here notice that the conclusion that there *are atoms* is reached by *reasoning*, based upon the results of our experiments.

Believing Avogadro's law to be true, and remembering that the hydrogen molecule divides into two parts in many chemical reactions, we arrive at the following practical definition of the molecular weight of a gas.

"The molecular weight of any gas is the weight of that volume thereof which is equal to the volume occupied by two parts by weight of hydrogen."

But it is evident that we have not yet learned all from Avogadro's law and our experiments that we may learn. In the first place, it is clear that the molecule of an element cannot contain less than two atoms, unless indeed the atom and the molecule should be identical; and it is also clear, that the molecule of a compound cannot contain *less* than one atom of each of its constituent elements. It follows, therefore, that we may find the atomic weight of any element by *determining the smallest part by weight of an element in the molecule of any of its gaseous compounds.* In other words, the atomic weight is determined by ascertaining two distinct classes of data:

(1) "The specific gravity of a series of gaseous compounds of the element in question;

(2) "Careful analysis of these compounds."

For example, the following data must be obtained in ascertaining the maximum atomic weight of oxygen:—

Compound.	Wt. of 2 Vols. of Gas. H=1.	Analysis of Compound.
Water,	17·99	15·96 of oxygen + 2 hydrogen.
Carbonic oxide,	27·96	15·96 " " + 11·97 carbon.
Carbon dioxide,	44·15	31·92 " " + 11·97 "
Nitrous oxide,	43·9	15·96 " " + 28·02 of nitrogen.
Nitric oxide,	30·0	15·96 " " + 14·01 of nitrogen.
Sulphur dioxide,	64·9	31·92 " " + 31·98 of sulphur.
" trioxide,	86·9	47·88 " " + 31·98 " "

From a study of the above data it becomes clear that 15·96, or in round numbers, 16 is the atomic weight of oxygen.

§ 209.—Unit of Volume.

In chapter XIV., we explained the various units of volume that have been adopted for the measurement of gases ; but inasmuch as 22.327 litres is the volume of all gases which weighs the units expressing their relative molecular weights, this volume is perhaps the one which is now in most general use. This volume of hydrogen at 0°C, and 760 mm. barometric pressure, weighs 2 grams in the metric system. In our English system of weights and measures the volume of hydrogen is 377 cub. feet, and, at 60° F., and 30 inches barometric pressure, weighs 2 lbs. As already intimated, chemists have agreed to take the least weight of any element found in such a molecular weight, as the weight of one atom of the element in question.

The practical advantage of this unit may be seen in the following table :—

22.327	litres	of	hydrogen	weigh	2	grams
"	"	"	oxygen	"	32	"
"	"	"	chlorine	"	71	"
"	"	"	nitrous oxide	"	44	"
"	"	"	carbon dioxide	"	44	"
"	"	"	ammonia	"	17	"
"	"	"	hydrochloric gas	"	36.5	"

and so on with all gases.

CHAPTER XXXI.

§ 210.—Classification of Elements.

We have already referred to the arbitrary division of the elements into metals and non-metals. This division is of very little practical importance, especially as it has been found impossible to discover any property that is common to all the substances called metals. It would seem better to divide the elements into two great classes, (1) *acid-forming* and (2) *base-forming*. Nearly one-fourth of the elements unite with oxygen and hydrogen to form acids, and the rest unite with these same elements to form bases. It has been found impossible to define a metal except as an element that replaces the hydrogen of an acid to form a salt, consequently, the division into acid-producing and base-producing practically coincides with that into non-metals and metals, but the former terms are much more suggestive. More important still is the classification of the elements into a number of natural groups or families. One of these families at least, the student must have already noticed while studying the acid-producing elements. Those which we have considered in whole or in part are:—

CHLORINE FAMILY.	SULPHUR FAMILY.	NITROGEN FAMILY.	CARBON FAMILY.
Chlorine.	Sulphur.	Nitrogen.	Carbon.
Bromine.	Oxygen. (?)	Phosphorus.	Silicon.
Iodine.	Selenium.	Arsenic.	
Fluorine	Tellurium.	Antimony.	
		Bismuth.	

Oxygen is generally classed with the sulphur family on account of the resemblance existing between the oxygen and the sulphur compounds, but oxygen in reality differs widely from the other members of the family. Selenium and tellurium are rare elements, and we can only glance at some of their analogous compounds and common properties.

§ 211.—Selenium Se, 79.

This element never occurs native, but is found as selenides of lead, silver and copper. Amorphous, crystalline, and vitreous modifications of it are known. In the form of a powder, it has a reddish colour, but *flowers* of selenium, analogous to flowers of sulphur, are scarlet. The formulæ of its compounds when tabulated shew their similarity to those of sulphur.

FORMULÆ.	NAMES.	REMARKS.
$H_2 Se$	Hydrogen selenide.	A poisonous gas.
$Se O_2$	Selenium dioxide.	
$H_2 Se O_3$	Selenious acid.	
$H_2 Se O_4$	Selenic "	

These acids form salts, called *selenites* and *selenates* respectively. Knowing how the analogous compounds of sulphur are formed, how would you expect to be able to form the above named compounds? Endeavour to answer this question, and then ask your teacher if you are right.

§ 212.—Tellurium Te, 125. (?)

Tellurium occurs native in very small quantities, and also as tellurides of some of the metals. It is a bluish-white, very brittle solid, with a metallic lustre, and specific gravity of 6.24. Its principal compounds are :—

FORMULÆ.	NAMES.	REMARKS.
H_2Te,	Hydrogen Telluride,	A poisonous gas.
TeO_2,	Tellurium dioxide,	A white solid.
H_2TeO_3,	Tellurious acid,	
TeO_3,	Tellurium trioxide,	Orange-yellow solid.
H_2TeO_4,	Telluric acid,	

The acids are formed by oxidizing the metal and uniting the oxide with water.

Note resemblances to the formulæ of compounds of sulphur and selenium. How would you expect to be able to form these compounds? What relation exists between the atomic weights of these three elements? Which atomic weight is a mean of the other two? Do the properties of these elements vary with the variation in their atomic weights?

The advantage of classifying the elements into groups naturally resembling each other, is two-fold. In the first place, it enables us to link together the almost innumerable facts of chemistry, and thus remember them; and in the second place, it enables us to grasp more readily the fundamental principles of the science.

CHAPTER XXXII.

§ 213.—The Nitrogen Family.

Two members of this family have already been fully considered; the other ones are arsenic, antimony and bismuth. In studying the sulphur family, we found that tellurium was a solid element possessing a metallic lustre, and for this reason sometimes classed as an *imperfect* metal. In the nitrogen family we have the same thing. Nitrogen and phosphorus are classed as metals; arsenic is said to be near the border line between metals and non-metals; antimony and bismuth are usually called metals. A very superficial examination of the formulæ of the commonly occurring compounds of these elements is sufficient to convince us of their natural resemblance to each other.

HYDRIDES.	OXIDES.	OXIDES.	CHLORIDES.	SULPHIDES.
NH_3	N_2O_3	N_2O_5	NCl_3 (?)	P_2S_3
PH_3	P_2O_3	P_2O_5	PCl_3	As_2S_3
AsH_3	As_2O_3	As_2O_5	$AsCl_3$	Sb_2S_3
SbH_3	Sb_2O_3	Sb_2O_5	$SbCl_3$	Bi_2S_3
	Bi_2O_3	Bi_2O_5	$BiCl_3$	

Arsenic occurs in nature in combination with metals; for example, with iron, copper, &c. The element possesses a metallic lustre: is not itself poisonous, but all its compounds are.

Symbol As, Atomic W't, 75; Molecular W't, As, 300; Specific W't, (as a solid), 5.6.

Experiments.

1. Take some arsenic trioxide As_2O_3 (known as "arsenic" in the drug shops) and make a pellet of it with some powdered charcoal and a drop or two of water. Place the pellet in the bottom of a hard glass test-tube, and heat to expel the water. Now make a loosely fitting stopper of chalk and insert so that it almost touches the pellet. Heat to redness. Arsenic forms above the chalk. Examine with a magnifying glass. What are its physical characters?

2. Heat any arsenic compound on a piece of charcoal before the reducing flame of a blow-pipe. What is the odor like?

The **trioxide** is formed by heating the metal arsenic in air or oxygen, the reaction, as might be expected, being similar to what takes place when phosphorus is heated in the air.

§ 214.—Arsenic Poisoning.

The frequency with which criminal poisoning has occurred by the use of "arsenic" has rendered it necessary for the chemist to have some thoroughly reliable test for its detection. Such a test he has in the behaviour of arsine, one of the compounds of the element. **Arsine,** or arseniuretted hydrogen AsH_3, is a compound analogous in composition to ammonia, and may be made by adding a compound of arsenic and oxygen to a mixture from which hydrogen is being evolved.

Experiment.

Fit up a hydrogen generating flask and attach to it a drying tube. Put some granulated zinc in the flask and pour some chemically pure sulphuric acid down the funnel tube. When the air has all been expelled from the apparatus, and the hydrogen lighted at the jet at the end of the drying tube, add slowly a little of a solution of arsenic oxide, As_2O_3, in dilute hydrochloric acid. What change takes place in the color of the flame? Are fumes given off? Hold over the flame a piece of porcelain or white stoneware. The formation of a black or brown shining spot or "mirror" on the porcelain is a test for the presence of arsenic. Attach a hard-glass tube to the end of the drying tube, and heat it near its middle point with a Bunsen burner. Just in front of the heated part there will form a layer of metallic arsenic. Explain its formation.

This method of testing for arsenic is known as Marsh's test. In applying it, great care must be taken that the chemicals used do not contain arsenic, otherwise, of course, the test would be of no value.

§ 215.—Antimony.

Antimony is another member of the nitrogen and phosphorus family. It usually occurs in combination with sulphur as stibnite, Sb_2S_3. The element is a bluish-white metallic looking substance; remains unchanged in dry air at ordinary temperatures, but when heated to redness, it burns with a white light, and forms the *trioxide* Sb_2O_3.

Antimony—Symbol, Sb. Atomic Weight, 120.

The metal is used principally in making alloys, but its compounds are used in many pharmaceutical preparations.

§ 216.—Antimoniuretted Hydrogen SbH_3.

This compound is analogous to arsine, and it is made in the same way. Another name for it is stibine.

Experiments.

1. Prepare some stibine, using the same apparatus as in preparing arsine, but adding tartar emetic solution instead of "arsenic" to the generating flask. The properties of stibine resemble those of arsine.

2. Introduce a piece of porcelain into the stibine flame and notice the antimony spot or "mirror." How does it differ from the arsenic mirror?

Powdered antimony combines directly and energetically with chlorine, bromine, and iodine. The terchloride of antimony can be prepared by dissolving finely powdered antimony in strong hydrochloric acid to which a few drops of nitric acid has been added. When completely dissolved, evaporate to a thick syrup. This is "butter of antimony." The *terchloride* decomposes on being thrown into water, forming hydrochloric acid and antimony teroxide. Try this experiment.

§ 217.—Bismuth.

This element occurs mostly native. It is obtained from its ores, *bismuth ochre*, Bi_2O_3, and bismuthite, Bi_2S_3, by heating them and allowing the molten metal to run off.

Bismuth—Symbol, Bi. Atomic Weight, 210.

Bismuth is hard, lustrous, and brittle, with a reddish tint. Its alloys *expand* on solidifying, and are, therefore, used in making delicate castings, and in electrotyping and stereotyping. "Fusible metals" are alloys of tin, lead and bismuth, and get their name from the fact that they melt at a very low temperature. The chief compounds of bismuth are the following :—

FORMULÆ.	NAMES.	REMARKS.
Bi_2O_2	Bismuth dioxide.	
Bi_2O_3	" trioxide,	[illegible]
Bi_2O_5	" pentoxide.	
$Bi(NO_3)_3+3H_2O$	" nitrate.	
$BiNO_3(OH)_2$	" sub-nitrate.	Used in medicine.
Bi_2S_3	" trisulphide.	

The trioxide is the chief ore of the metal. It is used as a pigment. Knowing how the oxides of phosphorus are formed, how would you expect to be able to form the **trioxide** of bismuth?

Experiments.

1. Dissolve a little bismuth in very dilute nitric acid. On evaporation, nitrate of bismuth forms in colorless deliquescent crystals.

2. Dissolve some of the crystals obtained in the last experiment in a *very* little water. To the solution thus obtained, add a large quantity of distilled water, and subnitrate of bismuth is precipitated as a white powder.

§ 218.—Test.

The result of the preceding experiment is, as far as it goes, a test for a soluble salt of bismuth. If, to the filtrate of the last experiment, sulphuretted hydrogen be added, a black precipitate is formed, soluble in hot nitric acid.

QUESTIONS.

1 What is the gradation in the specific gravity of the solid members of the nitrogen family?

2. Point out a similar gradation in their metallic character.

3. In what order does the chemical energy of these five elements stand as compared with the order of their atomic weights?

4. What physical property do antimony and bismuth communicate to their alloys?

5. Shew how the physical properties of these elements affect the physical properties of their compounds.

CHAPTER XXXIII.

§ 219.--Boron.

The next element which we shall consider is boron. It has been variously classified; oftenest with silicon and carbon, but sometimes with the nitrogen family; and still more recently, in accordance with the periodic law, it has been classified with aluminum. Boron is a non-metal.

As one of its commonest compounds (borax) is frequently used in testing for the metals by means of the blow-pipe, we shall begin our study of the element by investigating some of the properties of this compound.

§ 220.—Borax.

Borax occurs native in California, forming massive beds that are probably the bottoms of dried-up lakes of long ages ago. It is much used in the arts as a flux, because it prevents the formation of oxides when metals are soldered or welded together at a high temperature. It has the formula $Na_2 B_4 O_7 + 10 H_2O$.

Experiments.

1. Solder a short piece of platinum wire into a small glass tube about two inches long. Make a loop on the end of the wire; moisten in water, and then dip in powdered borax. Heat the loop and adhering borax in a Bunsen gas flame, or in the blow-pipe flame of an alcohol lamp, and continue the heat until a clear bead is obtained. Now remove from the flame, and when cool dip into a solution of a cobalt salt. Heat again. What color has been given to the bead?

2. Repeat the experiment with fresh borax, and when a clear, cool bead has again been obtained, dip it into a solution of a copper salt and heat. What colour does the copper salt give the bead?

3. Prepare new beads, dipping them each time into a different solution of some metallic salt. Use separately solutions of iron, nickel, manganese, and chromium. Tabulate the results obtained.

§ 221.—Boracic Acid.

The student can scarcely have failed to notice that the weaker acids are sometimes prepared from their salts by decomposition and substitution, the acid radicle of a strong acid displacing the weaker acid radicle of the salt, and giving rise to two new compounds—an acid and a salt. For example, nitric acid was prepared from nitre and sulphuric acid; hydrochloric acid was prepared from common salt and sulphuric acid; and carbonic acid H_2CO_3, from limestone and hydrochloric or sulphuric acid. This same principle is used in preparing boracic acid, or boric acid, as it is sometimes called.

Experiment.

Make a hot saturated solution of borax in an evaporating dish. Then add about half the volume of strong hydrochloric acid. Allow the mixture to cool. Collect and wash the crystals which form. They are boric acid, and when pure have the formula H_3BO_3.

Write the equation expressing the reaction. The acid, whose preparation has just been described, occurs dissolved in the waters of certain lagoons in Tuscany. In the vicinity of the lagoons are volcanic jets of steam, and the heat of these steam jets is used to evaporate the water which contains the acid. If boric acid be heated to redness for a long time, water is driven off and boric *trioxide* B_2O_3 is formed :—

$$2H_3BO_3 = 3\,H_2O + B_2O_3$$

Boron Symbol, B. ; Atomic Weight, 11 ; Specific Weight (Crystals), 2.5.

The element can be prepared from the *trioxide* B_2O_3 by fusing it at a high temperature with the metals potassium or sodium. Prepared in this way, the element is a dark-brown, odorless, tasteless powder, of no practical importance whatever.

§ 222.—Tests for Borates.

Experiments.

1. Make a solution of the free acid and immerse in it a strip of turmeric paper; then immerse the paper in dilute hydrochloric acid. Observe the change. Note how the change that takes place in this case differs from that which occurs when the turmeric paper is dipped in an alkali and then in hydrochloric acid.

2. Make a solution of borax in hydrochloric acid, and to the solution add alcohol in excess. Warm. Ignite the solution. The color imparted to the flame is a characteristic test.

§ 223.—Aluminum.

This element has already been dealt with in studying clay. Its classification, like that of boron, is difficult, some writers placing it by itself, and some assigning it to the iron, or manganese family. For the reason already assigned, we shall discuss it with boron. The element is without doubt a metal.

Its leading compounds not already referred to are the following :—

FORMULÆ.	NAMES.	REMARKS.
$Na_6Al_2O_6$(?)	Sodium aluminate.	
$Al_2(SO_4)_3$	Aluminium sulphate.	
$Al_2(OH)_6$	" hydroxide.	
$Al_2PO_4(OH)_3 + H_2O$	Turquois, a phosphate.	There are other phosphates.

There are besides these compounds many varieties of the *alums*. Of these the principal are :—

$K_2Al_2(SO_4)_4 + 24H_2O$; $Ag_2Al_2(SO_4)_4 + 24H_2O$; and $(NH_4)_2Al_2(SO_4)_4 + 24H_2O$.

The first of these, potash alum, and the last, ammonia alum, are the most important. The alums are prepared mostly from shale (a silicate of alumina) and iron pyrites, by burning them in heaps, when aluminic and ferrous sulphates are formed. From this mixture, alumic sulphate is dissolved out. To this, potassic or ammonic sulphate is added, according as it is desired, to make the one or the other kind of alum. The alum is purified by re-crystallization.

The alums are a group of compounds all similar in properties and in composition ; they are also isomorphous, and contain much water of crystallization.

Experiment.

Heat a small piece of alum on a sheet of mica, and observe the loss of its water of crystallization.

The **hydroxide of aluminium** is a weak acid as well as a base.

Experiment.

To a solution of alum add, little by little, some caustic soda. Aluminic hydroxide is first precipitated and then redissolved in the sodic hydroxide solution, forming an *aluminate of soda.*

§ 224.—Tests.

1. The precipitate formed on adding ammonia to a soluble salt of aluminum is characteristic. Ammonic sulphide gives a similar precipitate.

2. If the compound containing the aluminium be insoluble, it may be tested for by moistening the compound with cobaltous nitate $Co(NO_3)_2$, and heating with the blowpipe on charcoal. The color imparted is characteristic.

CHAPTER XXXIV.

225.—Extraction of Metals.

Before proceeding further with the consideration of the metals, it may be useful to refer briefly to some of the general processes of extracting them from their ores. In most cases they occur combined with other elements. Except the ores of the alkali metals, all ores are insoluble in water.

1 Copper, gold, silver, platinum, mercury &c. occur native.

2. Sodium, potassium, tin, zinc, iron, manganese, antimony, nickel and some others are obtained by heating their ores with coal or charcoal. This process is called *reduction.*

3. Lead, copper and bismuth are obtained by their ores being at first partially oxidized, and then subsequently fused.

4. Aluminium, magnesium and calcium may be obtained by heating with sodium or potassium.

5. Barium, calcium, strontium and lithium may be obtained by the electrolysis of their fused salts.

§ 226.—Compounds of Metals.

The derivatives of the metals are very numerous, and may be conveniently classified as follows :—

1. Compounds with chlorine, bromine, iodine and fluorine, known as *chlorides*, *bromides*, *iodides* and *fluorides* respectively.

2. Compounds with oxygen—*oxides*. These may be subdivided into three classes, according to their chemical properties—(a) base-forming oxides; (b) indifferent oxides; and (c) acid-forming. There are very few of the latter.

3. Compounds with oxygen and hydrogen — *hydroxides*.

4. Compounds with sulphur, and with hydrogen and sulphur, called *sulphides* and *hydrosulphides* respectively.

5. Compounds with nitric and nitrous acids—*nitrates* and *nitrites*.

6. Compounds with chloric and chlorous acids, &c., or the *chlorates*, *chlorites*, &c.

7. Compounds with sulphuric and sulphurous acids, or *sulphates* and *sulphites*.

8. Compounds with carbonic acids—*carbonates*.

9. Compounds with phosphoric and phosphorous acids, or the *phosphates* and *phosphites*.

10. Compounds with silicic acid, or the *silicates*.

11. Compounds with boric acid, or the *borates*.

You are more or less acquainted with the acids, and in acquiring your knowledge of them, frequent reference has been made to the salts which they form with the metals. In the remaining part of this book, prominence will be given to those compounds which illustrate general principles, or which are of special interest on account of their application to familiar processes.

CHAPTER XXXV.

§ 227.—Calcium Family.

This group is frequently spoken of as the metals of the alkaline earths. With them is frequently associated magnesium, especially for analytical purposes. Considered from the point of view of the periodic law, zinc also might be studied with the calcium family. We shall, however, first study calcium barium and strontium as a family, and afterwards treat of magnesium and zinc.

§ 228.—Calcium.

We have already studied to some extent a few of the compounds of this element. We have seen that it exists in carbonate and phosphate of lime. It is found also in gypsum as a sulphate $CaSO_4$, and in fluorspar as a fluoride CaF_2. The metal may be extracted with difficulty by fusing calcium iodide with metallic sodium in closed iron retorts.

Calcium—Symbol, Ca; Atomic Weight, 40.

The most important compounds of this metal are:

FORMULAE.	CHEMICAL NAMES.	COMMON NAMES.
$CaCl_2$	Calcic chloride	
CaO	" oxide	Quick lime
$Ca(OH)_2$	" hydroxide	Slaked lime
$CaCO_3$	" carbonate	Limestone
$CaSO_4$	" sulphate	Gypsum
$Ca(ClO)_2$	" hypochlorite	
$Ca_3(PO_4)_2$	" phosphate	In apatite.

Of these compounds we need study only the *chloride* and the *sulphate*, as we have already sufficiently considered all the other compounds when treating of phosphorus and lime.

Experiment.

Dissolve 15 grams of limestone or marble in hydrochloric acid. When all effervescence has ceased, evaporate to dryness. The residue is calcium chloride. Expose a few pieces of it to the air. What change takes place in it? To what use have you seen this substance put in the laboratory? Add some sulphuric acid to it and explain what takes place. Try to symbolize the reaction.

Calcium sulphate, $CaSO_4$, is found native in the form of selenite or gypsum, $CaSO_4 + 2H_2O$. This mineral, when powdered and heated, loses its water of crystallization and is known as *Plaster of Paris*—a substance capable of again uniting with water and forming a firm solid. The process of solidifying with water is known as "setting." The use of plaster of paris is as a finishing coat in plastering the interior of houses, and in making casts. Calcium sulphate, as we have already seen, is soluble in water, and imparts to it the property known as *permanent hardness.*

Experiments.

1. Prepare some calcium sulphate by adding sulphuric acid to calcic chloride or calcic carbonate. Evaporate the solution to dryness. Dissolve some of the sulphate thus formed in water. Is it difficultly or easily soluble in water? Add some carbonate of soda to the sulphate solution. What is precipitated? How can permanent hardness in water be removed?

2. Heat some powdered gypsum to a red heat in an open vessel. Examine what is left. Try whether it will "set" when mixed with a little water, so as to form a paste See whether gypsum will act in a similar manner.

Gypsum is frequently used as a fertilizer by farmers; but is not nearly so valuable for this purpose as calcium phosphate, $Ca_3(PO_4)_2$, a salt which was partially discussed in treating of phosphorus. The **normal** phosphate, that is, the one in which all the hydrogen of phosphoric acid, H_3PO_4, is replaced by calcium, is insoluble, and is, consequently, not easily taken up by plants; but when treated with a certain proportion of sulphuric acid, a mixture of calcium sulphate and mono-calcic phosphate, $CaH_4(PO_4)_2$, is formed—a valuable fertilizer known as superphosphate of lime.

$$Ca_3(PO_4)_2 + 2\,H_2SO_4 = 2\,CaSO_4 + CaH_4(PO_4)_2.$$

§ 229.—Test.

Experiment.

Dip a piece of platinum wire into a soluble salt of calcium, and then place in a non-luminous flame. The color imparted to the flame is distinctive.

§ 230.—Barium.

Heavy Spar, $BaSO_4$, is the most abundant ore of this metal. The ore is used for weighting paper and as a paint, but the metal itself is not used in the arts. Its symbol is Ba, and its atomic weight, 137. Its compounds closely resemble those of calcium, and may be tabulated as follows:—

FORMULA.	CHEMICAL NAME.	COMMON NAME.
BaO	Barium monoxide	
BaO_2	" dioxide	
$Ba(OH)_2$	" hydroxide	Caustic baryta
$BaCl_2$	" chloride	
$BaSO_4$	" sulphate	Heavy Spar
$Ba(NO_3)_2$	" nitrate	
$BaCO_3$	" carbonate	In Weatherite

Barium monoxide has, in recent years, been put to a peculiar use. When heated in air to a dull red heat, it takes up oxygen, forming the dioxide, and this when heated to a still higher temperature yields up its oxygen and returns to the mon-oxide. In this way oxygen has been *extracted from the air.*

Barium dioxide is interesting for another reason. It is used in making hydrogen dioxide. When treated with sulphuric acid the reaction is :—

$$BaO_2 + H_2SO_4 = BaSO_4 + H_2O_2.$$

This same reaction may be obtained with another acid—hydrochloric. Compare this action with that of manganese dioxide on hydrochloric acid.

§ 231.—Preparation of Compounds.

Knowing how calcium hydroxide was formed, how would you expect to be able to form barium *hydroxide ?* Given barium carbonate, $BaCO_3$, and other necessaries, how would you prepare barium chloride and barium nitrate ? The two latter compounds are used as reagents in chemical analysis.

§ 232.—Test.

The flame test for this element is very characteristic.

Experiments.

1. Make a solution of the chloride or nitrate of barium, dip into it your platinum wire loop, and then place in the non-luminous flame of a Bunsen burner or spirit lamp.

2. Take a little of the solution used in the last experiment and add to it a few drops of sulphuric acid. The color of the precipitate that is formed, coupled with the fact that it is insoluble in all acids, is a test for barium.

§ 233.—Strontium.

The minerals celestine, $SrSO_4$, and strontianite, $SrCO_3$, are the commonest sources of the element. Its symbol is Sr, and atomic weight, 87.2. Strontium may be isolated in two ways: by the electrolysis of its chloride, or by heating the chloride with an amalgam of sodium. The strontium amalgam which is thus formed is washed, dried, and then heated in a current of hydrogen. The metal is yellowish in color, malleable, oxidizable in air, and burns when heated, forming an oxide of strontium.

Its principal compounds are :

FORMULAE.	NAMES.	REMARKS.
$SrCO_3$	Strontium carbonate.	A mineral
$Sr(NO_3)_2$	" nitrate	

Strontium carbonate can easily be prepared in the laboratory by precipitation.

Experiments.

1. Dissolve some strontium nitrate in water and add to this a solution of sodium carbonate. Strontium carbonate will be precipitated. What other salt will be formed in solution. Write the equation.

2. Try to prepare the carbonate by using other soluble salts of strontium, with solutions of other alkaline carbonates.

3. The **nitrate** can easily be prepared from the carbonate by treating it with nitric acid. Try to do this as an experiment. Write the equation.

The **nitrate** is much used in making *red fire* on the stage.

Experiments.

1. Mix *very carefully* equal parts of finely powdered and thoroughly dried strontium nitrate and chlorate of potash, with an equal bulk of powdered shellac. No rubbing must be used, or *an explosion will occur.* The ingredients must be powdered separately and then mixed on paper. Ignite the mixture.

2. Repeat the foregoing experiment, using barium nitrate in place of strontium nitrate. What change in color is obtained?

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§ 234.—Test.

The flame test is usually sufficient to recognize a salt of strontium.

Experiment.

Moisten any chemically pure salt of strontium with hydrochloric acid and heat on a platinum loop in the non-luminous flame.

QUESTIONS.

1. What relation exists between the atomic weights of calcium barium and strontium?

2. Mix salts of barium and strontium and apply the flame test to the mixture. Which metal colors the flame first? Would you expect this result from a study of their atomic weights?

3. Would you expect analogous results if you mixed a calcium with a strontium salt? Try the experiment.

4. What weight of water will be needed to slake 100 grams of quicklime?

5. Explain how calcic carbonate acts as an antidote in poisoning by mineral acids.

6. How would you prepare baric nitrate from baric chloride.

7. Baric sulphate is often used by painters as a substitute for white lead. Why is it preferable?

8. What substances are formed when solutions of sodic sulphate and baric nitrate are mixed?

9. Compare by means of actual experiment the solubility in water of calcic sulphate, with that of barium sulphate, and with that of strontium sulphate.

CHAPTER XXXVI.

§ 235.—The Magnesium Family.

This family consists of but two elements, magnesium and zinc. In some respects magnesium resembles the calcium family, in other respects, it appears related to cadmium. It exists abundantly in nature. Its well known and naturally occuring compounds may be thus tabulated :—

FORMULAE.	COMMON NAME.	REMARKS.
$MgCO_3$.	Magnesite.	
$(MgCa)CO_3$.	Dolomite.	Mountain limestone.
$MgSO_4+H_2O$.	Kieserite.	
$(MgCa)SiO_3$.	Asbestos.	
$Mg_2H_2(SiO_3)_3$.	Meerschaum.	

The **sulphate** occurs as Epsom salts in some medicinal springs, and the chloride is a constituent of sea water. The metal is usually prepared by treating magnesium chloride with sodium at a high temperature. The sodium displaces the magnesium. Write the equation.

Magnesium, Mg; A't. W't., 24; Melting Pt., 750°.

Magnesium is a silver-white metal with a high metallic lustre; it changes slowly in air forming an oxide. It does not decompose water unless heated to 100°. How does this action of the metal compare with that of sodium and potassium? The metal is used in pyrotechny, photography, and signaling.

§ 236.—Important Compounds.

In addition to the naturally occurring compounds already mentioned, the following are important :—

FORMULAE.	NAMES.	REMARKS.
MgO.	Magnesia.	Used in medicine.
$MgCl_2$.	Magnesium chloride.	Used in dressing cotton goods.
$MgSO_4+7H_2O$.	Epsom salts.	Used in medicine.
$MgCO_3$.	Magnesium carbonate.	Used in medicine and as a face powder.

§ 237.—Preparation.

The **oxide**—a fine white variety of it called *magnesia usta*—is generally prepared by heating precipitated magnesium carbonate. A hydroxide of this is obtained by adding water :—$MgO+H_2O=Mg(OH)_2$. Compare these reactions with the burning of limestone and the slaking of quicklime.

Experiment.

Take a piece of magnesium ribbon ; set fire to it and catch the white powder that forms. What is the composition of the powder? Add water to it, and test with red litmus paper. How can you prepare Epsom salts from this white powder. Try it.

Epsom salts are generally prepared from the mineral kieserite by boiling with water. The insoluble kieserite ($MgSO_4+H_2O$) takes up more water and changes to the very soluble Epsom salts.

The **carbonate** may be prepared by precipitation.

Experiments.

1. Make a hot solution of Epsom salts and to it add a hot solution of sodic carbonate. Basic *carbonate* of magnesia will be precipitated. Filter, wash, and dry. Examine the salt.

2. Repeat this experiment, using cold solutions instead of hot ones. What different result is obtained?

3. Try to prepare the carbonate by passing carbon dioxide into a mixture of magnesium hydroxide and water.

4. Try to prepare magnesium *chloride* from the oxide and hydrochloric acid.

§ 238.—Tests.

Experiments.

1. To a soluble salt of magnesium add a solution of ammonic carbonate. If the precipitate which forms is soluble in acids and in ammonium chloride solution, magnesium is indicated.

2. A solution of di-sodic phosphate (with ammonia) added to a soluble salt of magnesium, gives a characteristic precipitate after stirring the mixture with a glass rod for a few minutes. The solutions must be dilute, and should be mixed in a test-tube.

A compound of magnesium insoluble in water must first be brought into solution, by dissolving some of it in hydrochloric acid. Then proceed as in (1) and (2).

§ 239.—Zinc.

The principal ores of zinc with their approximate formulae and chemical names are best seen in a tabulated form :—

FORMULAE.	COMMON NAMES.	CHEMICAL NAMES.
$ZnCO_3$.	Calamine or zinc spar.	Zinc carbonate.
Zn_2SiO_4.	Willemite.	" silicate.
ZnS.	Zinc blende.	" sulphide.
ZnO.	Red zinc ore.	
$ZnO.FE_2O_3$.	Franklinite.	

The metal is extracted by *reduction*. The ore is roasted and ground fine, mixed with coal dust, and heated in earthenware retorts. The metal, being volatile, distils over and is condensed in iron tubes. (Indicate the reaction by an equation.) Commercial zinc contains lead, arsenic and other metals as impurities. At different temperatures, it exhibits different properties. Brittle at ordinary, and at high temperatures, it is nevertheless quite malleable at 100° to 150°. It does not tarnish in dry air. As we have seen, it dissolves in the mineral acids, and evolves hydrogen. Its chief use is for *galvanizing* iron, that is, covering it with a thin coating of zinc. It is used extensively also in making batteries. Mixed with copper it forms *brass*, and this again when alloyed with nickel forms *German silver*.

§ 240.—Important Compounds.

FORMULAE.	NAMES.	REMARKS.
ZnO	Zinc white	Used as a paint.
$ZnCl_2$	“ chloride	A caustic in surgery.
$ZnSO_4 + 7H_2O$	“ sulphate (white vitriol)	Used in medicine and in dyeing.
ZnS	“ sulphide	
$ZnCO_3$	“ carbonate	

Knowing how magnesium oxide is formed, how would you expect to be able to prepare *zinc oxide?*

Experiments.

1. Place a small quantity of zinc carbonate in a small porcelain crucible and heat in a bunsen flame for 15 minutes. The residue is zinc oxide. Note the change in color when it has cooled.

2. Given pieces of zinc, hydrochloric acid, and all necessary apparatus, how could you prepare zinc *chloride?* Describe the salt.

3. Given metallic zinc, sulphuric acid, and necessary apparatus, how could you prepare zinc *sulphate?* Describe the salt.

4. Pass sulphuretted hydrogen into separate solutions of zinc chloride and zinc sulphate. What solid will be precipitated? Indicate the reaction by equations.

5. Given zinc sulphate and sodic carbonate, how could you prepare zinc carbonate? Try to prepare some.

All the soluble salts of zinc are poisonous. The chloride—used in soldering—is sometimes the cause of poisoning from being contained in canned goods. The sulphate is the analogue of magnesic sulphate. It is obtained on a large scale by *heating* the sulphide in contact with air. Write the equation.

§ 241.—Tests.

Compounds of zinc insoluble in water must be dissolved in hydrochloric or sulphuric acid, and then tested for as follows:—

Experiments.

1. To any soluble salt of zinc add ammonia. A white precipitate indicates zinc.

2. Ammonic sulphide gives a similar precipitate, soluble in dilute hydrochloric acid.

QUESTIONS.

1. How would you distinguish white vitriol from Epsom salts?

2. When sodium hydroxide solution is added to magnesic sulphate solution a white precipitate falls. What is this precipitate? Write the equation.

3. In what way does calcic carbonate act as an antidote to poisoning by zinc chloride?

4. How does the solubility of magnesic hydroxide in water differ from that of calcic hydroxide in water?

5. Point out resemblances between the compounds of magnesium and zinc (*a*) in composition; (*b*) in physical properties; (*c*) chemical behaviour.

CHAPTER XXXVII.

§ 242.—The Alkali Family.

The two best known members of this family are potassium and sodium, and as these metals are obtained from the alkalies, caustic potash and caustic soda, the family is often called the *metals of the alkalies.* The radicle ammonium NH_4, is generally classed with this group, as well as three comparatively rare metals: lithium, rubidium, and caesium. Of course ammonium is not to be considered a true metal. All that is meant by placing it in this family is that it behaves like the metals of this group and forms compounds with acids that are exactly analogous to those of sodium and potassium.

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§ 243.—Potassium.

Potassium compounds are derived indirectly from the soil; but as yet, no cheap and easy method has been discovered of separating them from other ingredients of the soil. Granitic rocks which contain silicate of potassium are the source of these compounds, and the disintegration of such rocks explains the presence of the derivatives of potassium in nearly all soils. From the soil they pass to plants, and from the ashes of these they are removed by a process of washing and evaporation.

Experiment.

Place some wood ashes in a tin vessel and add five or six times the bulk of hot water. Shake up well, then filter, test with litmus, evaporate to dryness. The residue is called potash; when purified it is known as *pearl-ash.* How would you prove that this residue is a *car-*

bonate? How would you prove that this salt is *hygroscopic*, and deliquescent? From this substance most of the potassium compounds are prepared.

Potassium—K. 39.1; sp. wt. 0.865. Melting pt. 62.5°.

Potassium is made from the carbonate by heating it to a very high temperature with charcoal, in iron bottles. The metal distils over as a green-colored vapor, which is not to be allowed to pass into air, but is passed into receivers containing vapor of naphtha. Try to write the equation. Metallic potassium is used as a reducing agent in preparing some of the rare metals.

Experiments.

1. Cut off a piece of the metal with a knife. Squeeze it between the fingers. Note the color of the newly cut surface, and any change which the color undergoes immediately after being cut. Explain such change.

2. Throw a small piece of it on some water placed in the bottom of a large bottle. What phenomena result. Explain them. Note the color of the flame and distinguish it from that of sodium.

3. Try the effect of pieces of potassium upon other liquids.

4. Devise an experiment to shew that potassium decomposes carbon dioxide and forms potassium carbonate, K_2CO_3, and carbon. Write the equation.

§ 244.—Potassium Compounds.

The following compounds of potassium are important:—

FORMULÆ.	NAMES.	REMARKS.
K_2CO_3	Potassium carbonate	Used as a medicine.
$HKCO_3$	Hydrogen potas'm carbonate	Sometimes called "bicarbonate."
KHO	Potassium hydroxide	Often called caustic potash.
KCl	" chloride	Occurs native.
KBr	" bromide	Used as a medicine.
KI	" iodide	" "
K_2SO_4	" sulphate	" "
$HKSO_4$	Hydrogen potas'm sulphate	" "
KNO_3	Potassium nitrate	Occurs native. Used as a medicine.
$KClO_3$	" chlorate	Used as a medicine.

There are some compounds of potassium with organic acids, such as the tartarate, oxalate, acetate, &c., but these we shall not notice at present.

§ 245.—Preparation.

Iodide of potassium is the analogue of chloride of potassium, and is made in the same way.

Experiments.

1. Place some iodine in the bottom of a porcelain dish, and cover well with warm water. Add gradually, warm caustic potash, until the color of the iodine has disappeared. Evaporate to dryness, and heat the residue strongly. Both iodide and *iodate* of potash are formed at first. The heating after evaporation decomposes the iodate into iodide and oxyg

2. Repeat this experiment, using bromine in place of iodine; and bromide of potash will be obtained.

3. Compare crystals of potassium *bromide*, and potassium iodide. Treat a crystal or two of each, separately, with a few drops of concentrated sulphuric acid. Explain what takes place.

Potassium hydroxide is prepared by heating the carbonate with slaked lime in a silver or iron vessel.

Experiments.

1. Dissolve 50 grams of potassic carbonate in about five or six times its own bulk of water. Heat to boiling, and add gradually about 25 grams of slaked lime. Stir with an *iron* spoon. After the solution has cooled, draw it off with a syphon: it is caustic potash.

2. Examine some sticks of the solid caustic potash. Dissolve in water and compare the solution with that obtained in the above experiment.

3. Throw some metallic potassium into water. Potassic hydroxide is formed in solution.

Potassium nitrate occurs in nature. The mode of its formation has been already explained. *Saltpetre plantations* illustrate almost exactly the manner in which nitre has been formed in nature. Heaps of refuse from stables or other sources, are mixed with lime and wood-ashes, and kept moist. On standing for some time, slow oxidation of nitrogenous organic matter occurs, giving rise to nitric acid, and then to nitrate of potash. The nitrate is extracted by water and purified by crystallization.

The **chloride** occurs as a mineral, sylvite, and is found also in some mineral springs It is used in preparing the nitrate by dissolving it with sodic nitrate in hot water. Explain the reaction.

It is not necessary to again refer to the preparation of the chlorate. How can the *sulphate* be formed?

§ 246.—Tests.

The flame test for potassium is very characteristic. Try it with a loop of *clean* platinum wire.

Experiments.

1. Make a solution of any potassium salt, say the iodide, in a test-tube. Add a little tartaric acid and stir with a glass rod. A granular-crystalline precipitate indicates potassium.

2. Apply the flame test to a mixture of potassium and sodium salts, and then look at the flame through blue glass or through a thin vessel filled with a solution of indigo.

§ 247.—Sodium.

Sodium is found abundantly in nature as chlorides, nitrates, carbonates, and borates. The chloride of sodium, or common salt, furnishes the great bulk of the sodium used in commerce. The sources of common salt are (1) the beds of *halite*, or rock-salt, (2) brine springs, and (3) sea-water.

At Goderich, Ont., both halite and brine springs occur. New York and Michigan furnish six-sevenths of the salt used in the United States.

Sodium is obtained from anhydrous sodium carbonate by heating it in iron retorts with finely divided charcoal to a full white heat. The sodium is vaporized and distilled in condensers, whence it is collected and preserved under naphtha.

Symbol, Na ; At. Wt., 23 ; Melting Pt., 95.6°.

Experiments.

Observe the behavior of sodium under the following conditions:

(*a*) Expose a freshly cut surface to the air.

(*b*) Put a small fragment, previously dried by pressure, between folds of blotting paper, on filter paper floated on water in a small dish.

(*c*) Put pieces of red and blue litmus paper into the water.

(*d*) Try to stick the fresh-cut surfaces together again.

Sodium is used extensively in the reduction of magnesium and aluminum from their ores.

§ 248.—Compounds.

The compounds of this metal are numerous and important. We shall notice the following :—

FORMULÆ.	NAMES.	REMARKS.
$NaHO$	Sodium hydroxide	Often called caustic soda.
$NaNO_3$	" nitrate	Found in nature.
Na_2SO_4	" sulphate	
Na_2CO_3	" carbonate	The most important.
Na_2HPO_4	" phosphate	
$Na_2S_2O_3+5\,H_2O$	" thiosulphate	
$Na_2B_4O_7$	" pyroborate	Occurs native.
$NaCl$	" chloride	Occurs native.

§ 249.—Preparation.

Caustic soda, $NaHO$, is prepared in the action of sodium carbonate on slaked lime.

Experiment.

Dissolve some sodium carbonate in seven parts of water, in an iron vessel, then apply heat ; while boiling add, with stirring, small quantities of slaked lime; after boiling a few minutes, allow the mixture to settle.

$$Na_2CO_3 + Ca(OH)_2 = 2\,NaOH + CaCO_3.$$

Sodium nitrate, $NaNO_3$, or Chili nitre is obtained in large quantities from Chili and Peru. It is used in the manufacture of potassium nitrate, sulphuric acid, nitric acid, and as a manure. It is quite deliquescent and is never prepared artificially for commercial purposes.

Sodium sulphate Na_2SO_4, or salt-cake, is prepared in the manufacture of soda.

Sodium carbonate, Na_2CO_3, or soda-ash, is used extensively in many industries, *e.g.*, glass, starch, sugar, glucose, oil-refining, candles, aniline dyes, earthenware and pottery, and dynamite. A very large percentage of the soda consumed in the United States is brought from England. The *alkali plains* of Wyoming abound in deposits of sulphate and carbonate of sodium. Commercial soda-ash is sometimes prepared by the Leblanc process from common salt by (1) treating it with sulphuric acid, (2) melting the *salt-cake*, the product of (1) in a furnace with chalk and coal, and (3) dissolving the soda-ash out with water. The reactions are:

$$2\,NaCl + H_2SO_4 = Na_2SO_4 + 2\,HCl.$$
$$Na_2SO_4 + 4\,C = Na_2S + 4\,CO.$$
$$Na_2S + CaCO_3 = Na_2CO_3 + CaS.$$

Another process which is coming more and more into practice is the *Solvay* or *Ammonia* process. The following experiment will illustrate the principle involved:

Experiment.

To a strong solution of ammonia add some common salt till the cold solution is saturated. Pass a slow stream of carbon dioxide through this pure saturated solution. Note the white precipitate:

$$NaCl + CO_2 + NH_4HO = Na\,HCO_3 + NH_4Cl.$$

The ammonia is recovered from the ammonium chloride by distillation with lime, and the carbon dioxide obtained on heating the *bi-carbonate of soda* is used in the decomposition of the salt. The above process is used in the Syracuse salt works.

Sodium phosphate, Na_2HPO_4, or phosphate of soda is used extensively in the laboratory as a re-agent, and in medicine as a mild purgative. It is prepared by the addition of sodium carbonate on phosphoric acid as long as any effervescence occurs.

Sodium thiosulphate, $Na_2S_2O_3+5H_2O$, or hyposulphite of soda, is extensively used in photography and paper manufacture. It may be prepared by decomposing calcium thiosulphate with sodium sulphate.

§ 250.—Test.

Experiment.

Heat any sodium compound on the loop of a platinum wire in the flame of a spirit lamp. Note the persistent yellow color. Test the color on viewing it through blue glass.

§ 251.—Resemblances.

They are all soft at ordinary temperatures; and volatile at high temperatures; unite readily with oxygen; unite with water to form basic (never acid) hydrates; the vast majority of their salts are soluble in water, and often isomorphous; the formulæ of their compounds are generally uniform, so much so that if a compound of one of these elements is proved to exist, the strong presumption is that analogous compounds of all the other members of the family also exist, and with similar, though not always identical, properties. Lithium's atomic weight is 7. What relation exists among the atomic weights of lithium, potassium, and sodium?

§ 252.—Ammonium, NH_4.

The ammonium salts are generally prepared in the laboratory by neutralizing acids with a solution of ammonia, but they are not always manufactured in this way. Liquor ammoniæ may be supposed to contain ammonium hydroxide, NH_4OH, in solution, and ammonium hydroxide acts towards acids exactly as sodic or potassic hydrate does.

Experiments.

1. Prepare separately, as indicated above, the *nitrate sulphate*, and *chloride* of ammonium. Evaporate the solutions and examine the salts.

2. Heat a few grains of each of the salts, separately, on a piece of platinum foil, or mica sheet. What remains?

3. How can the carbonate be prepared?

Ammonium sulphide $(NH_4)_2S$ is used extensively in making chemical analysis. It decomposes on standing and has to be prepared from time to time as needed.

Experiment.

Pass hydrogen sulphide through aqua ammonia until the solution will not precipitate magnesium sulphate. You have now a solution of ammonium sulphide. It is poisonous.

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§ 253.—Test.

Experiment.

Make a solution of any ammonium salt, add a little caustic soda solution and heat. Smell the gas that comes off.

QUESTIONS.

1. Name any insoluble sodium salts.
2. How would you prepare sal. sodæ from soda ash?
3. How would you distinguish Na_2CO_3 from $NaHCO_3$?

4. For the preparation of black ash, 224 lbs. each of salt cake and limestone are fused with 140 lbs. coal. What quantities are theoretically required for this weight of salt-cake? (Jones).

5. How is pure caustic soda obtained?

6. How can potassium nitrate be made from sodic nitrate?

7. Suppose you wanted potassium sulphate how could you prepare it if you had only sulphuric acid and wood-ashes?

8. Why cannot sodium nitrate be used in place of potassium nitrate in making gunpowder?

CHAPTER XXXVIII.

§ 254.—The Lead Family.

The lead family consists of two members, lead and tin. These elements, according to Mendelejeff, belong to the same natural family as carbon and silicon—the one pair being metallic elements, the other pair non-metallic.

§ 255.—Lead.

In nature lead is seldom found in a free state; its oxides are rare, but its sulphide, PbS galena, is found in large quantities both in Canada and in the United States. In Canada, the Galena occurs chiefly in veins cutting the Laurentian limestone, and in the shales of the Eastern Townships. Colorado, Idaho, Utah, Missouri and Nevada are the chief lead-producing districts in the United States.

Lead. Pb. At. Wt. 206.4, Melting P'., 334°. Sp. Wt. 11.3.

There are two distinct stages in the treatment of galena for lead. The first consists in roasting the ore in a furnace with a flux to remove the earthy impurities. In the second stage, the roasted residue is melted, when sulphur dioxide escapes and metallic lead remains.

The reaction which takes place is :

$$PbSO_4 + 2PbO + 2PbS = 5Pb + 3SO_2.$$

When the galena contains arsenic, a long flue is constructed to carry off the poisonous arsenical fumes. Antimony in crude lead is removed by its more rapid oxidation than lead. The same fact is taken advantage of in separating silver from lead.

The following simple experiment illustrates the treatment during the *first* stage.

Experiment

Place some powdered galena in an open glass tube ; heat the tube, inclined slightly. Place a strip of moistened blue litmus in the end of the tube. Notice any change.

§ 256.—Properties.

Experiments.

1. Place 5 grms. of lead in an iron spoon, then with a small pair of bellows blow air on the molten mass.

2. Heat a small fragment of lead upon charcoal in the oxidizing flame.

3. Secure a piece of metallic lead and scrape a portion of it with a knife. Examine the scraped portion after exposing it to the air for a day.

4. Place a bar of lead about 5 c.m. long on an iron plate, then with a hammer beat the bar into a thin sheet.

Describe the properties of lead, basing your answer on the preceding experiments.

Owing to the ease with which lead can be worked, it is largely used in the arts, as tubing for conveyance of gas and water, in the construction of water cisterns, and in roofing. Lead is used largely as an alloy in type-metal, pewter, shot and solder.

§ 257.—Lead Poisoning.

The following experiments will illustrate the action of lead on water, which contains various substances in solution. The presence of carbon dioxide, air, and ammoniacal salts in water passing through leaden pipes makes it very unwholesome. A basic carbonate of lead is formed which dissolves in excess of carbon dioxide. The action of *hard* water on lead pipes soon ceases because of the formation of a protective layer on the surface of the metal. All the salts of lead are poisonous.

Experiments.

1. Put 1 grm. of clean lead in fine shavings into 10 c.c. of water, and after 24 hours decant the clear liquid into a second test-tube, and pass a current of hydric sulphide through it for two or three minutes, then examine the contents of the tube by looking through it lengthwise. Why pass the sulphuretted hydrogen through it?

2. Take five tubes, into each put 1 gm. of clean lead, and 10 c.c. of water (distilled if possible); to the contents of No. 1 add a drop of solution of common salt; to No. 2, a drop of potassic nitrate solution; pass a current of carbon dioxide through No. 3; to No. 4 add a drop of calcium sulphate solution; and to No. 5 add a solution of sodic sulphate. After 24 hours, decant the water in each case into a clean tube, and pass sulphuretted hydrogen through each solution for two minutes.

Try to indicate all the reactions that occur by means of equations. Repeat these experiments with a very dilute solution of lead acetate.

§ 258.—Compounds of Lead.

The important compounds of lead may be thus exhibited in tabular form :—

FORMULÆ.	NAMES.	REMARKS.
Pb_2O	Lead suboxide	
PbO	" oxide	Litharge or Massicot.
PbO_2	" peroxide	
Pb_3O_4	red lead	Used as a paint.
$Pb(C_2H_3O_2)_2$	Lead acetate	Used in medicine and in dyeing.
$2PbCO_3 + Pb(OH)_2$	" carbonate (basic)	White lead, a paint.
PbS	" sulphide	Galena, an ore.
$PbCl_2$	" chloride	

Experiments.

1. Place some oxide of lead on moist red litmus paper.

2. Pour 2 or 3 grms. of litharge into a test-tube, then pour in diluted nitric acid, and heat over a flame. Pour the liquid into a porcelain crucible and evaporate until crystals begin to appear. Cool the liquid slowly and quietly. The crystals are nitrate of lead. Can you get this same salt by dissolving lead in nitric acid?

Red Lead or Minium. Pb_3O_4. This compound is known as a mineral. As its name implies, Minium has a red crystalline color. It is used largely as a paint and in the preparation of flint glass. Commercial red lead often contains litharge. This oxide is apparently a mixture of litharge and the peroxide :—$Pb_3O_4 = 2PbO + PbO_2$. When litharge is heated in air it changes into red lead; and this again when heated to a high temperature gives up oxygen and changes back again to litharge. This change is illustrated in the following experiment.

Experiments.

1. Powder very finely 10 grms. of pure oxide of lead, then place the powder in an iron tube about 2 c.m. in

diameter and open at both ends. Heat very carefully to a dull redness revolving the tube at intervals so as to expose the powder to the action of the air.

2. Heat the red oxide on charcoal before the blowpipe (1) alone, (2) with soda. Notice what substance is formed in both cases, also the colored coatings and what effect the reducing flame has on the coating.

Lead Acetate or **Sugar of Lead,** $Pb(C_2H_3O_2)_2$ is used in the laboratory extensively. It is obtained by the action of acetic acid on litharge. It has a sweet but very astringent taste.

Experiment.

Introduce some powdered litharge and some moderately strong acetic acid into a test-tube. Allow the former to gradually dissolve in the latter. Evaporate the solution so as to obtain a mass of small crystals.

Lead Chloride, $PbCl_2$ is easily prepared by the action of hydrochloric acid on litharge. It is a white crystalline substance and frequently used in the laboratory. It is soluble in boiling water.

Experiment.

Dissolve 5 grms. of lead nitrate in 10 c.c. of water, add hydrochloric acid. Heat the solution containing the white precipitate to boiling, then filter and allow to cool.

Lead Carbonate or **White Lead,** $2PbCO_3 + Pb(OH)_2$ is a basic carbonate of lead. It is used very largely as a white paint. There are several methods of manufacture, but all are based on the reaction of acetic acid with lead oxide and the introduction of carbon dioxide.

Experiment.

Heat *White Lead* in a porcelain crucible. Notice the evolution of moisture and carbon dioxide, and the color of the powder.

Lead Sulphide or **Galena,** PbS, is the chief source of the metal. It resembles the latter very closely in color. It generally crystallizes in cubes. Silver, arsenic and antimony are often present in the ore.

Lead Nitrate, $Pb(NO_3)_2$. This is prepared in commerce by dissolving litharge in hot dilute nitric acid. It is largely used in dyeing, and has an astringent taste. Try to prepare it.

Lead Chromate, $PbCrO_4$, used as a fine pigment, is obtained as a yellow precipitate when a solution of potassium bi-chromate is added to one of lead nitrate.

§ 259.—Tests.

We have already seen that nitric acid dissolves lead very readily. The malleability, streak and tarnish have also been noted, so that it is an easy task to detect lead in the solid form when not associated with other metals.

Experiment.

Make a solution of lead acetate or nitrate and add to parts of it, in separate test tubes, solutions of sulphuretted hydrogen, potassic bichromate and hydrochloric acid. Note the color of the precipitates formed; they are characteristic of lead.

§ 260.—Tin.

Tin occurs native in very small quantities; its chief ore being the oxide, SnO_2, known as *tin-stone.* It is obtained from this ore by reduction with charcoal.

The tin-stone is first ground to powder, and then roasted in revolving cylinders through which a blast of highly heated air is passing. [What effect will this air have upon the volatile and oxidizable impurities?] The roasted ore is then mixed with anthracite and heated in a blast furnace, when tin appears in a molten condition. The metal is purified by *liquation*, that is, by slowly melting it, the pure metal melting first and flowing away from its impurities. The principal tin mines of the world are in England and South America.

Tin, Sn, At. Wt. 118, Melting Pt., 230°.

Tin is a bright, white metal resembling silver in appearance. It does not tarnish in air, and is, therefore, used extensively in covering sheets of iron, forming the so-called "tin-ware."

Experiments.

1. Take a piece of bar tin and bend it.

2. Hammer a piece of it out on an anvil as thin as you can.

3. Try to dissolve some of it in ordinary concentrated nitric acid. Metastannic acid $H_{10}Sn_5O_{15}$ (?), a white powder is formed. Some authorities say that stannic acid is the product.

4. Place a bright piece of zinc in an alkaline solution of stannous chloride. Crystals of tin are deposited.

Write out an account of all the properties of tin so far as you have observed them, or seen them illustrated by experiment.

§ 261.—Compounds of Tin.

Tin forms two classes of compounds: (*a*) the stannous and (*b*) stannic. In stannous compounds, the tin appears to act as a dyad as may be seen in the formulæ, SnO, $SnCl_2$, and SnS; whereas in stannic compounds the element appears to act as a tetrad, as indicated by the formulæ, SnO_2, $SnCl_4$, and SnS_2, which respectively represent, in each case, the oxide, the chloride and the sulphide. *Solder* is an alloy of lead and tin. *Bronze* and *bell-metal* consist of tin and copper.

FORMULÆ.	NAMES.	REMARKS.
SnO	Stannous Oxide	A basic oxide.
SnO_2	Stannic "	An acid oxide, also weakly basic.
H_2SnO_3	Stannic acid	Analogous to carbonates and silicates.
$H_{10}Sn_5O_{15}$	Metastannic acid	
$SnCl_2$	Stannous chloride	Used in dyeing.
$SnCl_4$	Stannic "	" "
SnS	Stannous sulphide	A brown powder.
SnS_2	Stannic "	A pigment—*mosaic gold.*

Apart from the tin-stone, the chlorides are the most important compounds of tin.

Experiments.

1. Place some small pieces of tin in a dilute solution of hydrochloric acid until they dissolve. Warm slightly if necessary. *Stannous chloride* is formed.

2. Place some scraps of tin in nitro-hydrochloric acid and heat gently. Stannous chloride may be used in place of tin. *Stannic chloride* will be formed.

Both stannous and stannic chloride are decomposed when *much* water is added to them. Try the experiments and indicate the reactions.

Experiment.

Try to prepare the *nitrate* and the *sulphate* by using scraps of tin and the appropriate acids.

If you had soluble salts of tin and appropriate reagents and apparatus how could you prepare stannous and stannic sulphide? Try to prepare them.

§ 262.—Tests.

Insoluble compounds of tin should be reduced by the blow-pipe on charcoal. The product should then be dissolved in hydrochloric acid. A soluble compound having been formed, it is tested as follows

Experiments.

1. Make a solution of stannous chloride and pass into it hydric sulphide. Note the precipitate. Test its solubility in ammonic sulphide.

2. Make a solution of stannic chloride and pass hydric sulphide into it. Note the precipitate.

3. To a solution of either stannic or stannous chloride add mercuric chloride.

If a stannous salt be present, mercuric chloride gives a precipitate which turns from white to gray, and then black.

QUESTIONS.

1. Why is it so difficult to make lead wire?

2. Explain what takes place on allowing molten lead to cool (1) slowly (2) quickly.

3. Why is lead not serviceable for castings?

4. Why can leaden pipes and cisterns be used with greater safety for a *hard* water supply than for one that is *soft*?

5. What advantage would red lead have over litharge in the manufacture of flint glass when there is considerable carbon in the molten glass?

6. What are the chief uses of red lead?

7. Explain the peculiar action of lead with acid solvents.

8. What are the different methods of forming salts?

9. Test whether the lead salts formed are soluble (1) in water, (2) in the acid from which they are formed.

10. Shew the similarity in structure between $Pb(NO_3)_2$ and KNO_3.

11. How many kilogr. of litharge can be obtained from 37.1 kilogr. of lead, and what volume of oxygen would be absorbed in the process?

12. Explain the constitution of the blue line at the edges of the gums in cases of chronic lead poisoning.

13. Why are not plumbers, who handle metallic lead only, subject to lead poisoning?

14. What is the color of the precipitate when sulphuric acid is added to a lead salt?

15. What is the best chemical antidote for lead poisoning? Explain the reason why.

16. How can you distinguish a stannous from a stannic compound?

17. What danger is there in eating canned goods that have been closed with solder? Explain.

18. What is tin salt? What is it used for?

19. What are the members of the tin family? What relation, if any, exists among their atomic weights? Point out resemblances between the stannates, carbonates and silicates.

20. What is *block tin?* What is tin foil?

21. An inferior quality of tin ware is made from an alloy of tin and lead. What danger in using such tin ware for cooking utensils, or for cans for preserving fruits?

CHAPTER XXXIX.

§ 263—Platinum Family.

This family consists of four members, platinum, iridium, osmium, and gold. The first is by all odds the most important element to the chemist; "without plantinum," said Liebig, "the composition of most minerals would have remained unknown." It occurs mostly alloyed with small quantities of the rare metals iridium, palladium and osmium, and is found in the Ural mountains, in the Rocky mountains, in Australia and in South America. It is extracted from its ores by dissolving it in *aqua regia.* The platinum chloride thus obtained (contains some iridium) is precipitated with ammonic chloride forming a double chloride of platinum and ammonium $(NH_4)_2 Pt Cl_6$. This is fused in lime crucibles and yields metallic platinum:—

$$(NH_4)_2 Pt Cl_6 = 2NH_4Cl + Pt + 2Cl_2$$

Experiment.

Examine a piece of platinum wire, and describe its physical appearance. Test its solubility in *aqua regia.*

Platinum is used for making crucibles, and for some parts of delicate philosophical instruments. The wire and foil are in constant use in the chemical laboratory. Small weights are generally made of platinum.

Platinum Pt, At. W't. 195, Sp. W't. 21.15.

The metal resists the action of each of the mineral acids if taken separately; but it is attacked by the caustic alkalies when highly heated, and by the oxides and sulphides of lead, copper, bismuth and a few others. It possesses the wonderful property of condensing gases upon its surface, especially the *spongy platinum.* The difficulty of extracting and working platinum renders it very expensive. Its salts are numerous but comparatively unimportant.

§ 264—Test.

Experiment.

Dissolve the substance supposed to contain platinum in *aqua regia ;* expel any surplus acid, and add ammonic chloride. A yellow crystalline precipitate indicates platinum.

Spongy platinum may be obtained from this precipitate by heating it strongly. Direct a stream of hydrogen gas against some spongy platinum.

§ 265.—Gold.

Gold occurs very widely distributed and nearly always pure. It is found in sand or inclosed in quartz rock. From the latter it is removed by pulverization and by the addition of mercury which readily forms an amalgam with the gold. From sand it is removed by washing in sluices. Pockets containing mercury are placed at intervals along the sluices, and amalgamation takes place in this case as in that of pulverization. The amalgam is subsequently removed and heated, when the mercury distils away and is used over again, and the gold remains behind.

Gold Au, At. Wt. 196.7, Sp. W't. 19.3.

In a pure state gold is too soft to wear well, and has, therefore, to be alloyed with copper to harden it. It is the most malleable and ductile of all metals. It can be beaten into leaves $\frac{1}{10000}$ of a millimeter in thickness; 1 gram of the metal can be drawn out into wire 1¾ miles in length. Like platinum it is attacked by the caustic alkalies when hot. It combines directly with chlorine but not directly with oxygen. Gold forms two series of compounds, aurous and auric, resembling platinum in this respect which forms platinous and platinic compounds. Their formulæ, however, are different, being for the oxides, Au_2O and Au_2O_3 and PtO and PtO_2; while for the chlorides the formulæ are $AuCl$ and $AuCl_3$; and $PtCl_2$ and $PtCl_4$. As we have already seen gold dissolves in aqua regia. The compound formed is *trichloride* of gold.

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§ 266.—Test.

The test for gold is the formation of the compound known as the *Purple of Cassius.*

Experiments.

1. To a solution of trichloride of gold add a drop or two of stannous chloride.

Note—If the re-agent be in excess, purple of Cassius will not be formed.

2. Repeat this experiment, adding ferrous sulphate. A purple precipitate indicates gold.

CHAPTER XL.

§ 267.—The Iron Family.

This family includes iron cobalt and nickel, and according to Mendelejeff forms a part of a larger group of elements which includes the platinum and palladium families. The relation of the various families and groups of elements to each other will be made clearer when we come to study the periodic law. On account of its great importance iron will be discussed at considerable length; cobalt and nickel will not be touched upon.

§ 268.—Ores and Extraction.

Iron is obtained native, but chiefly from the following ores:—*Magnetite*, *Hæmatite*, *Siderite Clay*, *Iron-stone*, and *Limonite*.

Iron, Fe At. Wt. 56; Sp. Wt. 7.8.

Magnetite Fe_3O_4 occurs in beds and veins of great extent and thickness in the Laurentian rocks of Canada, in the counties of Renfrew, Lanark, Hastings, and Ottawa, and is the chief ore of commercial importance in Canada. In the United States, magnetite occurs in North Eastern New York, New Jersey, and California.

Hæmatite Fe_2O_3 occurs in many regions along with magnetite. In Canada, it is found chiefly in Renfrew and the Eastern Townships; in the United States it occurs in great abundance in Missouri, New York and Michigan. The small island of Elba in Europe has been long famous for its hæmatite mines. This ore is the one chiefly reduced in the United States. It assumes different forms under different conditions, as *Grape ore*, *Micaceous ore*, *Amorphous ore* and *Granular ore*.

Clay Iron-stone occurs in the coal measures. It is the chief ore of iron in England and the coal regions of Pennsylvania. This is a *spathic ore* containing clay or sand, and is obtained in bands or nodules.

Siderite Fe CO_3 is invariably mixed with calcspar or dolomite.

Limonite $2Fe_2O_3+3H_2O$. In its purest form limonite is fibrous in structure and sometimes forms concutionary masses, when it is known as *brown hæmatite;* while the less pure and more earthy varieties are known by the names of *bog iron ore* and *iron ochre.* The *bog ores* are found along the north shore of the St. Lawrence from Montreal to Quebec, in the Eastern Townships, and on the north shore of Lake Erie. It is very abundant in Connecticut, New York, Pennsylvania and North Carolina. In Europe, Limonite is worked abundantly at Erzberg in Styria, where the celebrated Styrian steel is manufactured.

The process of getting the cast iron from the ore is simple, chemically, but difficult to perform in a small laboratory. The *stages* only will be indicated here.

The **first stage** consists in the reduction of the ore to the state of *oxide* by roasting the broken ore along with refuse coal. Moisture and carbon dioxide are expelled in this way.

The **second stage** consists in the reduction of the oxide to cast iron. The product of the first stage is mixed with limestone and coal in a tall, conical-shaped furnace. The oxide is reduced to the metallic state by the gases arising from this mixture. The lime acts as a flux and forms a fusible slag with the silicates present in the oxide of the first stage. This slag can be easily run off at the bottom of the furnace, and the reduced iron, which unites with carbon at a high temperature, collects on the hearth of the furnace, whence it can be drawn off and run into moulds, forming the well known *pig iron,* or *cast iron.*

Wrought iron is obtained from cast iron by the puddling process, which is simply one of oxidation, by which the silicon, carbon, sulphur and phosphorus are removed.

Steel is obtained from pig or cast iron when the carbon is reduced in amount to 1.5 or 0.15 %.

Bessemer Steel is produced by adding pure cast iron containing about 5 % of carbon to molten cast iron from which the carbon and silicon have been burnt out

The following experiments will illustrate the chief blow-pipe and other characters of the iron ores.

Experiments.

1. Procure a small piece of magnetite, and test its hardness with a knife, its streak on a small piece of unglazed porcelain, and its influence upon the magnetic needle.

2. Heat strongly a small fragment of magnetite upon charcoal before the blow-pipe.

3. Bring a hot bead of borax in contact with a few grains of finely powdered magnetite and heat in the oxidizing flame of the blow-pipe. What color is imparted to the bead?

4. Heat a small fragment of hæmatite on charcoal before the blowpipe.

5. In the shallow cavity of a piece of charcoal place a mixture of hæmatite powder and soda ; heat the mixture carefully in the blow-pipe flame.

6. Heat a small fragment of siderite on charcoal before the blow-pipe. Test the residue with the magnet.

7. Introduce ½ grm. of siderite powder into a test-tube, add hydrochloric acid and apply heat.

8. Heat some limonite powder with hydrochloric acid in a test-tube.

§ 269.—Properties.

Pure iron is almost unknown. That used for piano wires is the purest. It is greyish-white in color.

Iron when placed in a moist place becomes covered with rust.

Experiments.

1. Sprinkle a few drops of water on a bar of polished steel, and then expose it to the air for a day or two. That which forms on the polished surface is *oxide* and *hydroxide* of iron.

2. Heat one end of an iron wire in the blowpipe flame.

3. Try to clinch a common cut-iron nail.

4. Heat a common cut-iron nail till it is red hot, cool it slowly, then try to clinch it.

Iron in its three conditions of *cast iron*, *wrought iron* and *steel* is used for so many purposes that it is impossible to give an account of them here.

When heated to a white heat, iron wire softens and partially fuses. *Welding* is a union of two surfaces of iron in this pasty condition.

§ 270.—Compounds.

Iron forms two classes of compounds, *ferrous* and *ferric*. The former invariably tend to change into the latter, simple contact with the air being sufficient to bring this about. The change, however, takes place very readily if an oxidizing substance, such as nitric acid or chlorate of potash, be supplied to ferrous compounds.

FORMULÆ.	NAMES.	REMARKS.
FeO	Ferrous oxide	Unimportant.
Fe_2O_3	Ferric "	Magnetite, loadstone.
Fe_3O_4	Ferroso-ferric oxide	
$FeCl_2$	Ferrous chloride	
Fe_2Cl_6	Ferric "	
$FeSO_4 + 7H_2O$	Ferrous sulphate	
$Fe_2(SO_4)_3$	Ferric "	
FeS	Ferrous sulphide	
FeS_2	Iron pyrites	Occurs native as *Fool's Gold.*
$Fe(OH)_2$	Ferrous hydroxide	
$Fe_2(OH)_6$	Ferric "	

Ferrous Chloride, $FeCl_2$. This substance is obtained when iron is dissolved in hydrochloric acid.

Experiment.

Introduce 1 grm. of iron tacks into a test tube; add 10 c.c. hydrochloric acid diluted with an equal bulk of water. Filter and evaporate to about half the bulk, then allow the solution to cool. Write the equation. Note the color of the solution. What is the black substance caught on the filter paper?

Ferric Chloride, Fe_2Cl_6. If we saturate a solution of ferrous chloride with chlorine gas we obtain ferric chloride—a most valuable reagent in the laboratory. It may be more conveniently prepared by adding, successively, a little nitric acid to ferrous chloride. The reaction is one of oxidation. $2FeCl_2 + 2HCl + O = Fe_2Cl_6 +$

H_2O. Try the experiment. Note how the color of ferric chloride solution differs from that of the ferrous.

Ferrous Sulphate, $FeSO_4 + 7H_2O$, or *green vitriol,* is used in dyeing and ink-making. It can be prepared by dissolving iron in sulphuric acid, filtering and evaporating slowly to dryness. Try to prepare it. Symbolize the reaction.

Potassium Ferrocyanide, K_4FeCy_6, usually called yellow prussiate of potash, and Potassium Ferricyanide, K_3FeCy_6, or the red prussiate of potash, are both useful reagents in the laboratory in the detection of iron.

Ferrous sulphide, FeS, is formed by the direct union of the elements.

Experiment.

Place a little flowers of sulphur and double the quantity of iron filings in a test-tube, and heat until sulphur fumes cease to come off. Iron sulphide is formed. How could you prepare sulphuretted hydrogen from this compound?

Iron pyrites, FeS_2, occurs native in golden yellow cubes. It is used extensively in preparing *green vitriol.*

Experiments.

1. Powder some iron pyrites and heat it in a long test-tube. What element is given off?

2. Heat strongly some powdered pyrites in an iron spoon. What becomes of the sulphur? What is left behind?

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§ 271.—Tests.

The blowpipe tests for the common ores have already been given; but for any iron salt in solution the following test will be valuable:—

Experiments.

1. Make a borax bead on the loop of a platinum wire, and moisten it in a solution of an iron salt; heat the bead again, (1) in the oxidizing flame; (2) in the reducing flame. Note the effect on the color of the bead.

Potassium Sulphocyanide (KCyS) is a valuable test reagent.

2. Add potassium sulpho-cyanide to a solution of (1) a ferrous salt, and (2) a ferric salt. Note the effects on each solution.

3. Add to separate solutions of a ferrous salt, (1) ferrocyanide of potassium; (2) ferricyanide of potassium and note results.

4. Add to separate solutions of a ferric salt these same reagents and note results.

The tests in (3) and (4) are the simplest for young students.

QUESTIONS AND PROBLEMS.

1. Compare the streak of magnetite with that of hæmatite or limonite.

2. Devise an experiment to prove that limonite contains water?

3. What happens when powdered limonite is heated on charcoal?

4. What are the chief physical characters of cast iron?

5. What are the chief varieties of cast iron? Wherein consists their difference?

6. How may steel be made brittle?

7. What effect has strong nitric acid on iron? Try the experiment.

8. Will iron rust in water that is free from air? Answer this question by making an appropriate experiment.

9 How would you convert iron which is in the *ferric* condition into the *ferrous* condition?

10. How would you prepare ordinary black writing ink?

11. In the preparation of *copperas* or green vitriol on a large scale, moistened iron pyrites are subjected to "atmospheric oxidation." Try to explain by an equation.

12. Given a solution containing silver, lead and iron, how would you separate them

13. Explain the manufacture of wrought iron and Bessemer steel?

14. What are chief impurities in iron and how are they removed?

15. What effect have titanium phosphorus and sulphur on the quality of iron?

16. Why must "reduced iron" be kept in well-stoppered bottles?

17. Devise an experiment to ascertain whether iron will conduct electricity.

18. Mention facts which you have yourself observed that go to prove that iron is malleable and ductile.

19. Heat a ferrous salt, say, $FeSO_4$ in a Bunsen flame. What oxide do you obtain?

20. Heat a ferric salt, say, $Fe_2(NO_3)_6$ on charcoal in the reducing flame. What oxide is obtained?

21. Try to prepare ferrous hydroxide, and ferric hydroxide from ferrous chloride and ferric chloride, respectively. How can these hydroxides be distinguished from each other?

22, How can it be decided whether $FeCl_3$ or Fe_2Cl_6 is the correct formulæ for ferric chloride?

23. What is rouge? How is it made?

CHAPTER XLI.

§ 272.—Manganese.

Manganese presents some points of resemblance to iron and aluminum; and it also resembles members of the chlorine family. It is with this latter family that it is classed in accordance with the periodic law.

This metal occurs in nature chiefly in the form of the oxides, of which the dioxide MnO_2 is the most abundant. Manganese is not used for practical purposes. It is extracted from any of its oxides by reduction with charcoal.

Manganese Mn. At. Wt. 55, Melts at a white heat.

Manganese is a reddish white metal, very brittle, oxidises readily in air, decomposes warm water, and imparts hardness to iron when mixed with it in small quantities.

§ 273.—Compounds.

Like iron, it forms two series of salts, the manganous and manganic, which differ from each other in very much the same way as ferrous and ferric compounds do from one another.

FORMULÆ.	NAMES.	REMARKS.
MnO	Manganous oxide	A grayish green powder.
Mn_2O_3	Manganic "	
Mn_3O_4	Manganoso-manganic oxide	May be considered as a union of first two oxides.
MnO_2	Manganese dioxide	The black oxide.
Mn_2O_7	Manganese heptoxide	
H_2MnO_4	Manganic acid	Not eliminated.
$HMnO_4$	Permanganic acid	
$KMnO_4$	Potassium permanganate	

Of the oxides the first two are basic; the next two are indifferent; and the fifth, Mn_2O_7 may be considered as the anhydride of permanganic acid.

$$Mn_2O_7 + H_2O = 2HMnO_4.$$

The **black oxide** known in the mineral form as *pyrolusite*, is used extensively in the manufacture of glass and in preparing chlorine for the manufacture of bleaching powder. To what other uses have you seen it applied? If adulterated with coal dust or charcoal powder it becomes explosive when used for preparing oxygen. Explain how.

In the manganous salts the element is bivalent. These salts are not easily changed to the manganic; in fact the opposite change is the one that takes place most readily.

Experiments.

1. Try to prepare manganous sulphate $Mn SO_4$ from manganese dioxide and concentrated sulphuric acid.

2. Try to prepare other manganous salts in a similar way.

3. Try to prepare an insoluble salt of manganese by precipitation. Write the equation and name the salt.

§ 274—Preparation and Properties

Both the manganates and permanganates readily lose oxygen; they are therefore powerful oxidising agents and as such become valuable disinfectants and deodorizers.

Experiments.

1. Fuse in a crucible, manganese dioxide with one and a half times its weight of caustic potash. Stir the mixture with a rod so as to expose all of it to the air. *Potassic manganate* is obtained. Mix with water and evaporate.

2. The *permanganate of potash* may be obtained from a clear *alkaline* solution of the manganate, by adding to it dilute sulphuric acid until the color changes to purple.

Condy's disinfecting fluid is chiefly a solution of the permanganate in water. Both sodic and potassic permanganates oxidize many chemical and organic substances. The following experiments illustrate this :—

Experiments.

1. Pass sulphur dioxide into a solution of potassic permanganate. Explain the change that takes place by the aid of the equation :—

$$2KMnO_4 + 5SO_2 + 2H_2O = K_2SO_4 + 2Mn\ SO_4 + 2H_2\ SO_4.$$

2. Similary, add ferrous sulphate, sulphuretted hydrogen, and ammonic sulphide to separate solutions of potassic permanganate.

3. Bubble air from the lungs through more of the permanganate solution. Note the change of color.

275.—Test.

The manganese acids may be distinguished by the color of their salts, and by yielding the reactions obtained as follows :—

Experiments.

1. Make a borax bead on a loop of platinum, add any salt of manganesse and heat in the oxidizing flame. Note the difference in color when the bead is hot and when cold.

2. Repeat, adding *soda* and a drop of nitric acid to the bead.

3. Add dilute aqua ammoniæ to a solution of a *manganese* salt. Note the precipitate. It is soluble in excess.

QUESTIONS.

1. Point out one prominent point of resemblance between manganese and the halogens.

2. Try to point out resemblances between the composition of the sulphates and manganates.

3. How would you find out whether it was safe to make oxygen from pure chlorate of potash and adulterated manganese dioxide? How would you remove the impurity supposing it to be carbon?

4. Will potassic permanganate disinfect the air? If not, how can air be disinfected?

5. How could you prepare manganous hydroxide? Try to do it experimentally.

6. What is *chameleon* mineral. Prepare some, add it to pure water, and explain the phenomena that occur.

CHAPTER XLII.

§ 276.—The Periodic Law.

We have quite frequently, when discussing the various families of elements, made reference to the periodic law. Mendelejeff, a Russian chemist, was the first to make a systematic attempt to trace the relation between the atomic weights and the chemical and physical properties of the elements. Hence it is often spoken of as Mendelejeff's law. Muir, in his *Principles of Chemistry*, thus states it: "We may confidently say that a large probability has been established in favor of the hypothesis that the properties of the elements and of the compounds of each element, are periodic functions of the atomic weights of the elements. Lothar Meyer puts the general statement of the periodic law in this form: "if the elements are arranged in order of increasing atomic weights, the properties of these elements vary from member to member of the series, but return more or less nearly to the same value at certain fixed points in the series."

Let the elements be arranged in the order of their atomic weights; let this list of elements be (broadly) divided into series of sevens; let the members of the second series be placed under those of the first, those of the third under those of the second, and so on; and let the elements contained in a vertical column be called a *group*, and those in a horizontal column a *series*." In such an arrangement of the elements it will readily be seen that each group comprehends for the most part one or sometimes more of the natural families already described. Many familiar relationships can now be traced out. These can be seen all the more easily if certain gaps are supposed to exist in the list of elements. These gaps represent the relative position and probable atomic weights of elements not yet discovered. If the symbol R. stands for any element, its oxides and hydrides will be denoted by RO and RH respectively.

GROUPS.

	I.	II.	III.	IV.	V.	VI.	VII.	VIII.
Series	R_2O	R_2O_2	R_2O_3	R_2O_4 RH_4	R_2O_5 RH_3	R_2O_6 RH_2	R_2O_7 RH	R_2O_8
1	H=1	——	——	——	——	——	——	
2	Li=7	Be=9	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27	Si=28	P=31	S=32	Cl=35·5	{ Fe=56, Ni=58·6 Co=59, Cu=63
4	K=39	Ca=40	Se=44	Ti=48	V=51	Cr=52	Mn=55	
5	(Cu=63)	Zn=65	Ga=69	——72	As=75	Se=79	Br=80	{ Rh=104, Ru=104·5 Pd=106, Ag=108
6	Rb=85	Sn=87	Y=89	Zr=90	Nb=94	Mo=96	——100	
7	(Ag=108)	Cd=112	In=114	Sn=118	Sb=120	Te=126	I=127	152—156. 4 Elements.
8	Cs=133	Ba=137	La=139	Ce=141	Di=144	——149	——150	
9	4 Elements 156 to 162				Er=166	——167	——169	{ Ir=192·5, Os=193 Pt=194, Au=196
10	——170	——172	Yb=173	——178	Ta=182	W=184	——190	
11	(Au=196)	Hg=200	Tl=204	Pb=207	Bi=209	2 Elements 212 to 220		
12	——	——	——	Th=232	——237	U=240	——245	
	3 Elements 220 to 230							

This table and indeed the whole chapter may be considered an extract from chapter III. of Muir's *Principles of Chemistry.*

Mendelejeff endeavored to show that the atomic volume, malleability, ductility, power of forming oxides and chlorides of definite proportions, fusibility, position in electrical series, etc., of the elements, vary periodically with variations in the atomic weights of these elements. The student may, in the case of atomic volume, verify this relation for himself. Let him turn to a table of specific gravities of the elements in the *solid form*, and let him divide the numbers which he there finds into the numbers representing the respective atomic weights of the elements. The quotients thus obtained will be found to vary from point to point in a more or less periodic manner. Carnelley has shown that the fusibility of the elements varies in a similar manner with their atomic weights. But, what use you will ask can be made of this law. In answer it may be said that at least two very important practical uses have been made of it. In the first place, the law not only recognizes the fact that there are elements yet undiscovered, but *it predicts the properties of such unknown elements.* At the time at which Mendelejeff first published the law there was no element known which could be placed opposite the atomic weights 69 and 44 in group III. The former of those elements Mendelejeff called Eka-aluminum, the latter Eka-boron in reference to their relations to the elements aluminum and boron. The properties of both these elements were predicted by Mendelejeff, with what exactness, in the case of eka-boron, can be seen by the following table, because in 1879, Wilson discovered a metal which he named *scandium*, and which is without doubt the same as eka-boron.

Eka-Boron.	Scandium.
At. Weight about 44.	At. Weight 44.
Oxide Eb_2O_3 soluble in acids ; sp. gravity about 3.5; analogous but more basic than Al_2O_3 ; less basic than MgO ; insoluble in alkalies.	Oxide Sc_2O_3 ; sp. gravity 3.8 ; soluble in strong acids ; analogous with but more decidedly basic than Al_2O_3 ; insoluble in alkalies.
Salts colorless, yield gelatinous precipitate with KHO ; K_2CO_3, etc.	Solutions of salts colorless and yield gelatinous precipitates with KHO and K_2CO_3.
Sulphate $Eb_2(SO_4)_3$ will form a double salt with K_2SO_4, probably not isomorphous with the alums.	Sulphate $Sc_2(SO_4)_3$ forms a double salt, not an alum, $Sc_2(SO_4)_3.2K_2SO_4$
Chloride $EbCl_3$, or Eb_2Cl_6, sp. gravity about 2, less volatile than Al_2Cl_6.	

In the second place, the periodic law has been successfully used as a guide in the comparative study of the properties of elements already known. We cannot, however, within the limits of an elementary text-book illustrate this use of the law, but must pass on to a brief study of the law of isomorphism.

§ 277.—Law of Isomorphism.

Two substances are said to be isomorphous when they have exactly the same crystalline form, that is when they both take the same shape whether cubes, pyramids, octohedra, or other geometric forms. Kopp limits the phrase "truly isomorphous crystals" to crystals of those salts, anyone of which is capable of growing, in unmodified form, when immersed in a solution of any other.

Experiment.

Make a saturated solution of chrome alum, and in it place a small crystal of common alum. Allow the whole to stand for a day or two and then examine the crystal of alum. After some time remove the crystal from the solution, dry, and break it.

The more important families of isomorphous elements are the following :—

I.—Fluorine, Bromine, Iodine, and Chlorine in all compounds. Partially Manganese.

II.—Sulphur, Selenium in all compounds.

III.—Arsenic, Antimony, Bismuth, Tellurium as elements, and the first three in all corresponding compounds.

IV.—Lithium, Sodium, Potassium in most compounds.

V.—Calcium, Barium, Strontium, Lead ; Magnesium, Zinc, Manganese, Iron ; e.g. in carbonates.

VI.—Aluminum, chromium, Manganese, Iron, in the sesquioxides and salts derived therefrom.

VII.—Copper, Silver in compounds of the type R''_2O.

VIII.—Palladium, Platinum, Iridium, Osmium, in most compounds.

IX.—Carbon, Silicon, Tin, partially in compounds of the type RO_2, and salts derived from the type H_2RO_3.

The determination of crystalline form is frequently an aid in fixing or correcting atomic weights, but this method can be used only tentatively and in a limited number of cases. If it be assumed that, as a general rule, those amounts of two substances which are crystallographically equivalent have analogous atomic constitutions ; and if we suppose that of two compounds exhibiting identical crystalline form, the atomic weights of the elements in one are known, it is evident in what way determinations of crystalline form may aid in fixing atomic weights. For example, suppose we have analysed green oxide of chromium and found its formula to be $n\mathrm{Cr_2O_3}$, and the atomic weight of chromium to be 52.4 ; and suppose further that this compound of chromium exhibits the same crystalline form as ferric oxide, we assume that this latter oxide should be represented by the formulæ $n\mathrm{Fe_2O_3}$, and as this formula is in keeping with analysis it is assigned to ferric oxide. On comparing these crystallographically equivalent quantities of the two oxides it is found that 52.4 × 2 parts of chromium are replaced by 55.9 × 2 parts of iron, and as 54.2 has been determined to be the atomic weight of chromium it is argued that the atomic weight of iron must be 55.9.

CHAPTER XLIII.

§ 278.—Silver Family.

This family consists of silver, copper and mercury.

Silver is often found native in Peru, Mexico, and Saxony, but the large proportion of the metal is obtained from the compounds, *Argentite* or *Silver glance* Ag_2S, *Horn-silver* $AgCl$, *Ruby silver* Ag_3SbS_3, and *Argentiferous* galena. Silver Islet was for a long time famous for its rich silver veins, but now Nevada and New Mexico are the chief silver-producing regions in the United States.

Symbol Ag. At. Wt. 108, Melting Pt. 1000°.

Three different processes are used in preparing metallic silver from the ore :—

1. *Amalgamation*—the most important process. The ore in this case is roasted and treated with a solution of salt ; then the product is mixed with mercury and copper sulphate, when the *amalgam* is formed. This is filtered and distilled and silver remains. The silver is purified by oxidation, in a current of air.

The following experiments will illustrate the principle of amalgamation :—

Experiments.

1. Introduce some pure silver into some mercury in a porcelain crucible. Note the effect—the silver dissolves and an amalgam is formed.

2. Apply heat to the crucible containing the silver amalgam. The mercury is driven off and pure silver is left.

2. *Cupellation*, used when silver is present in large quantities in galena. The ore is melted, and allowed to cool slowly, when some lead separates out. The silver with the remaining lead is heated in *cupels*, in a current of air. The lead is oxidized and silver remains.

Experiment.

Mix some silver chloride with about three times its weight of sodium carbonate; place the mixture in a clay crucible, and heat it very strongly in a strong current of air. When the crucible is cool, break it and examine the bead.

The reaction which takes place may be thus expressed :—

$$2AgCl+Na_2CO_3=2Ag+2NaCl+CO_2+O.$$

3. *Precipitation*, used very conveniently in reducing silver from its solution.

Experiment.

Into a solution of silver nitrate put a clean strip of copper. Note the effect. Silver is deposited on the strip, and can be easily removed.

Reaction : $2AgNO_3+Cu=2Ag+Cu(NO_3)_2$.

Silver is a fine white metal in its pure state, with a sp. gravity of 10.56. The bright surface is easily tarnished with sulphides. It is very malleable and ductile so that *silver leaves* and *silver wire* are quite common. Silver is used extensively in the *silvering* process by means of *silver amalgam.*

Experiment.

Fuse some pure silver on a clay crucible before the blowpipe. Allow it to cool. Note any phenomenon.

Silver dissolves easily in nitric acid, which is its best solvent.

Experiment.

Put a small piece of pure silver into some nitric acid. Note the result. Evaporate the liquid and preserve the crystals obtained.

§ 279.—Compounds.

Silver Nitrate, sometimes called *Lunar Caustic.* Usually prepared by dissolving silver in nitric acid. It is used largely in chemical analysis, in photography, in medicine, and as an indelible ink.

Experiments.

1. Dissolve some crystals of silver nitrate in water, and others in alcohol. Note how much water is required to dissolve the crystals.

2. Allow a drop of the solution of silver nitrate to fall on the skin. Note the dark stain. Explain it.

3 Add some gum arabic to a solution of silver nitrate. Use this as an ink, and write on cloth. Note the change in color as soon as exposed to light.

Silver Chloride $AgCl$, occurs in nature as the mineral *cesargyrite* or horn-silver. It is usuallly obtained by adding HCl to a solution of a silver salt. It is insoluble in HNO_3 but soluble in NH_4OH.

Experiments.

1. Heat a small fragment of the chloride on charcoal before the blowpipe.

2. To a solution of a silver salt add a few drops of HCl. Note the white curdy ppt. and the influence of the light on the color of the ppt.

Silver Bromide $AgBr$, found native in Chili as the mineral *Bromyrite.* Like the chloride, it is easily dissolved, or materially changed chemically under the influence of light, hence it is used in photography. It may be prepared by adding potassium bromide to a solution of silver nitrate.

§ 280.—Tests.

Hydrochloric acid precipitates salts of silver and gives a white ppt. as has been shown in a previous experiment. Sulphuretted-hydrogen is also used as a test reagent.

Experiment.

Pass sulphuretted-hydrogen through a solution of silver nitrate. Note the black ppt. of Ag_2S, which is soluble in KCY and HNO_3.

QUESTIONS.

1. Where are the silver regions of Canada?
2. Explain the chemical changes which take place in the three processes of the extraction of the metal from the ore.
3. Account for the *spitting* when fused silver is about to solidify? Explain the action of charcoal in preventing spitting.
4. How may silver be hardened without change of color?
5. How would you separate the silver from the copper in a silver coin?
6. Account for the constitution of the following compounds of silver:—
   Ag_2O_2, Ag_2O, Ag_4O,
   and $AgCl$, Ag_2Cl, and $AgNO_3$.
7. Is silver a monad or a diad element?

§ 281.—Copper.

Copper occurs *native* both in large masses, and as disseminated grains and sheets in beds of *trap* and *conglomerate.* It occurs in nearly every variety of ore. "Three quarters of the world's supply is derived from sulphuretted compounds of copper and iron which form a series of ores of definite composition."

Copper pyrites contains	34.6	p.ct. of copper.
Variegated copper ore contains	60.8	"
Copper glance	77.2	"

Copper pyrites is of most frequent occurrence. The chief regions of supply in North America are the Lake Superior region, Arizona and Montana.

Symbol Cu, At. Wt. 63, Melting Pt. 1090°.

The first product of roasting is termed *coarse metal.* The crushed and dressed ore is roasted to drive off much of the sulphur as sulphur dioxide gas; then the roasted ore is fused with limestone, coal and sand, when most of the iron, etc., is removed as silicates; the copper as sulphide is found at the bottom of the furnace.

The second product of roasting is termed *fine metal,* when the iron is removed by oxidation. By continued roasting and oxidation, the sulphide is converted into oxide, then into the metal. This process is long and difficult inasmuch as the pure copper is very liable to unite with oxygen and form dioxide.

Experiments.

1. Hold a piece of thin copper foil in the outer zone of the lamp flame until a black coating is formed; shake it up in a test-tube with 5 c.c. of water and test with litmus.

2. Hold a piece of copper wire in the inner zone of the lamp flame. Notice whether there is any difference in effect between this experiment and the preceding one.

Copper oxide is easily reduced to the metallic state.

Experiments.

1. Mix copper oxide thoroughly with about $\frac{1}{16}$ of its weight of powdered charcoal, and place the mixture in a small hard glass tube provided with a narrow delivery tube which bends down and dips into some lime water in a tes -tube Heat the mixture. Examine carefully the residue and the color of the lime water.

2. Heat 1 grm. of the black oxide of copper in an ignition tube through which a slow stream of hydrogen is passing, and hold the dry outer edge of a test-tube containing cold water close to the mouth of the ignition tube ; allow the tube to cool, then turn off the hydrogen and examine the contents of the tube.

Copper may be separated from a copper solution by the action of clean iron. Devise an experiment to prove this.

Copper may also be separated by a galvanic current.

Experiment.

Allow two platinum slips attached to the terminal wires of a battery to dip into a copper solution. Note the deposition of copper on the platinum slip connected with the negative pole of the battery.

Copper has a metallic lustre and is red in color ; quite malleable and ductile ; and is a good conductor of heat and electricity. Copper, on account of its toughness, is used in the manufacture of boilers. As long as the copper vessels are kept clean there is little danger of poisoning from the copper salts, but sometimes a greenish acetate is formed —*verdigris*—when the metal is exposed to vegetable juices, which is exceedingly dangerous.

Experiment.

Expose to the air six or seven copper sheets piled up in layers between woollen cloths soaked in vinegar. Examine after a few days and note the verdigris formed.

Copper is now used extensively in electro-plating. The alloys—brass, bronze, speculum metal and argentine—are of great importance.

Brass contains about 66 parts copper and 34 parts zinc.
Speculum metal “ 66 “ “ 34 “ tin.
Bronze contains copper, zinc and tin.
Argentine “ “ “ nickel.

§ 282.—Compounds.

Copper Sulphate or Blue Vitriol $CuSO_4+5H_2O$, is manufactured on a large scale by roasting copper pyrites $CuFeS_2$ at a low heat, and digesting the product with water. The copper sulphate along with the ferrous sulphate remains in solution. When all the water of crystallization is removed, the copper sulphate becomes a yellowish white powder, which becomes blue when moistened with water.

Experiment.

Heat gently in a test-tube 1 grm. of copper turnings and 3 c.c. of strong sulphuric acid. Allow the solution to evaporate in a porcelain crucible.

Copper Nitrate $Cu(NO_3)_2$ is prepared by dissolving copper turnings in dilute nitric acid. It is used as "an oxidizing agent in the processes of dying and calico-printing." Try to prepare it.

Cupric Sulphide CuS is prepared by passing sulphuretted-hydrogen through a solution of cupric sulphate, when a brownish black precipitate is formed.

Cuprous Sulphide (Cu_2S) occurs in nature as the mineral copper-glance, or chalcocite, and may be obtained from cupric sulphide by heating with an excess of sulphur in a hydrogen flame.

§ 283.—Tests.

The blowpipe tests for copper are generally characteristic. On charcoal, in the reducing flame, small reddish particles of copper can be obtained. To a borax bead copper gives in the oxidizing flame a green color while hot, but bluish when cold.

Many copper compounds give a green coloration to the flame.

H_2S gives a black precipitate of CuS soluble in HNO_3.

NH_4OH gives to copper solutions a blue color.

QUESTIONS.

1. Name as many ores of copper as you can, giving the formula for each.
2. Explain the process of the extraction of the metal from the ore.
3. What is the action of copper on each of the following bodies? (a) Sulphuric acid, (b) Ammonia, (c) Nitric acid, (d) Chlorine.
4. How would you prepare copper oxide? How would you reduce copper oxide?
5. Explain the formation of verdigris, Scheele's Green.
6. To what is the color of copper sulphate due? Prove your statement.
7. How is copper sulphate obtained commercially? What occurs (a) when it is heated, (b) when ammonia is added to a solution of it?

§ 284.—Mercury.

Mercury occurs native in small quantities. Its chief source is Cinnabar HgS. The most important cinnabar mines are at Almaden in Spain, at New Almaden in California, at Idria in Austria, in Peru and in Japan.

Symbol Hg; At. Wt. 200; Melting Pt. —40°.

For commercial purposes Cinnabar is heated in a furnace where strong currents of air are continually passing through the ore. The mercury is vaporized, and condensed in pipes or chambers which lead the liquid into the reservoir. The same principle is involved in the following experiments.

Experiments.

1. Heat a small fragment of cinnabar on charcoal before the blow-pipe.

2. Heat in an open tube some powdered cinnabar. Test the gas given off by its action on moistened blue litmus paper.

3. Scoop a small hole at the end of a piece of charcoal, and mix some powdered cinnabar with two or three parts of soda; heat carefully in the reducing flame of a candle; place the fused substance on a clean silver surface, and moisten with a drop of water.

Mercury is a liquid metal of a silver-white color. When subjected to a cold of −40° it becomes solid and contracts in volume. When solid, mercury is ductile and has a tin-white color. It is acted on slightly by the air.

Experiment.

Expose some mercury in a porcelain crucible for two or three days to the atmosphere and note the result.

Mercury, by uniting with other metals, forms *amalgams*, which are of considerable importance.

Experiment.

Pour a few drops of mercury on tin foil in a porcelain crucible. Apply heat gently. Note the result and test the fluidity of the new liquid.

Tin amalgams are used to form the reflecting backs of mirrors.

Experiments.

1. Procure a sheet of tin foil, stretch it on a flat smooth surface, and cover it with a layer of mercury. Lay a glass plate upon she mercury and squeeze out any excess of the latter. Examine the plate after an interval of five or ten minutes.

2. Boil a globule of two of mercury in a test-tube. Note the *bumping* and condensation.

§ 285.—Compounds.

Mercurous Chloride Hg_2Cl_2 or *Calomel*, is insoluble in dilute acids but soluble in aqua-regia. It is prepared by adding hydrochloric acid to mercurous nitrate.

Experiment.

Add some dilute nitric acid to some mercury in a crucible. Note that the mercury dissolves and *mercurous nitrate* is formed. From this prepare the chloride.

Mercuric Chloride $HgCl_2$ or *corrosive sublimate*, is soluble in dilute acids. It is prepared by heating a mixture of mercuric sulphate and sodium chloride.

Experiments.

1. Heat gently in the hood 20 grms. of mercury with 12 c.c. of concentrated sulphuric acid, stir frequently until a white substance ($HgSO_4$) remains.

2. Make a mixture of about equal parts by weight of mercuric sulphate and sodium chloride and put it into a flask provided with a long wide tube, and apply heat gently. The mercuric chloride sublimes on the tube.

3. Heat mercuric chloride in a matrass with dry sodium carbonate. Note the sublimate of pure mercury.

Mercuric Oxide HgO, or *Red Oxide of Mercury*, is extensively used in surgery. It may be prepared as follows :

Experiment.

Heat a mixture of two parts mercuric nitrate and one part mercury in a porcelain crucible with stirring. An orange-red powder remains when all the vapors are driven off. What gas have you seen prepared from this oxide ? How ?

§ 286—Tests.

Metallic mercury can be readily recognized by the properties already mentioned.

The blow-pipe characters have also been illustrated.

H Cl gives a white precipitate with a solution of any *mercurous* salt, which turns black with ammonia.

H_2S. gives a black precipitate with solution of any mercuric salt, soluble in agua-regia.

$SnCl_2$ gives a black ppt. in an aqua-regia solution of any mercury compound.

QUESTIONS.

1. What is the chief source of mercury ?
2. Explain the reaction when cinnabar is treated with quick-lime.
3. How would you prove that a molecule of mercury contains but *one* atom ?
4. Starting with mercury, how may the oxides and chlorides of mercury be prepared ?
5. Explain the colour changes. of the precipitate formed by passing H_2S through a mercuric solution.
6. What is the antidote for mercurial poisoning ?
7. Give the equations which express the reactions in the preparation of corrosive sublimate.
8. Why is mercury preferable to water in the construction of barometers and thermometers ?
9. How is mercurous nitrate prepared ?
10. In the preparation of mercury sulphate, supposing you add too much H_2SO_4, what becomes of the excess ?
11. How would you prove the presence of sulphur in cinnabar ?
12. What right has mercury to be classed with the metals ?
13. What difference does it make whether concentrated or dilute HNO_3 is added to mercury ? Under what conditions would the same reaction occur with heat and without heat ?
14. What change takes place when tin and lead are brought into contact with mercury ?
15. How is pure mercury obtained from the ore ?
16. Why should manganese dioxide be used with sodium chloride in the preparation of corrosive sublimate ?

CHAPTER XLIV.

§ 287.—Chemical Analysis

Chemical analysis is generally recognized to be of two kinds, (1) qualitative, and (2) quantitative. In making a qualitative analysis of an unknown substance, we are simply required to determine the various constituents which compose it; but in making a quantitative analysis, we must not only find out *what* the substance is composed of, but also the proportions by weight in which the constituents unite to form the unknown compound. Qualitative analysis is comparatively simple; quantitative is difficult and can be carried on only in a well equipped laboratory. Again, the substance submitted for analysis may be, (1) *pure*, or an individual chemical compound, like common salt, or it may be (2) *mixed*, that is, it may consist of common salt mixed with a number of other substances. The qualitative analysis of a simple substance, chemically pure, is easy; the qualitative analysis of complex mixtures is beyond the scope of the present work. It must be understood, therefore, that qualitative analysis, as discussed in this chapter, is confined to the case of a chemically pure *acid*, or *base*, or *salt*. Generally speaking, bases are tested for first, and the base having been determined, we proceed to test for the acid if the unknown substance be a salt. For the purpose of systematic testing the metals have been classified into six groups or families.

I.—*Lead*, *Silver*, and *Mercury* (mercur*ous* salts) are precipitated from their salts as *chlorides*, by *hydrochloric acid* as a group reagent.

II.—Lead, mercury (mercur*ic* salts) copper, cadmium, bismuth, antimony (arsenic), tin, gold, and platinum are precipitated from their salts as sulphides, by *sulphuretted hydrogen*, from neutral or acid solutions.

III.—Iron, chromium, aluminum, zinc, manganese, cobalt and nickel are precipitated from their salts as sulphides, in presence of an alkali, by means of *ammonic sulphide* as a group reagent.

IV.—Calcium, strontium, and barium, are precipitated as carbonates by *ammonic carbonate* in presence of ammonic chloride.

V.—Magnesium is precipitated as a phosphate by sodic phosphate as a group reagent.

VI.—Lithium, sodium, potassium, and ammonium (NH_4), have salts mostly soluble.

For the same purpose the acids may be arranged in four groups.

I.—Organic acids. These *char* when heated, e.g. tartaric, citric, etc.

II.—Inorganic acids, which form insoluble compounds with barium. These acids are, sulphuric, carbonic, phosphoric, oxalic, boric, sulphurous, etc. The group reagent for these is *barium nitrate*.

III.—Inorganic acids for which the group reagent is *argentic nitrate*. These are hypochlorous, hydriodic, hydrobromic, hydrochloric, and nitrous acids.

IV.—Acids which give no precipitate with the group reagents, viz., nitric, chloric, acetic, and in very dilute solutions oxalic, boric, nitrous and sulphurous acids.

On receiving a substance for analyzis the first step is to get it into solution. We should try to do this by testing the solubility of small portions of the substance in (1) water; (2) hydrochloric acid; (3) nitric acid; (4) nitro-hydrochloric acid. If insoluble in every one of these, the substance is probably one of the following:—*Baric sulphate, strontium sulphate, silica, calcic fluoride, alumina, stannic oxide, argentic chloride, plumbic sulphate, carbon,* or *ferric oxide,* and would have to be tested for separately, excepting in the case of the last two.

The following tables are taken (by permission) from Dr. Goodwin's *Text-book of Chemistry,* and are intended for the examination of solutions containing not more than one metal and one acid.

We start out with a soluble salt.

Analytical Tables.

TESTING FOR THE METAL.

TABLE A.

N.B.—*O. S.* means *Original Solution.*

<table>
<tr><td colspan="12">To O. S. add dilute HCl in excess.</td></tr>
<tr><td colspan="3">Wht. ppt. insoluble in excess.—Add NH_4OH.</td><td colspan="9">No ppt. (or wht. ppt. soluble in excess of the acid).*—Add H_2S.</td></tr>
<tr><td rowspan="5">No change
Pb</td><td rowspan="5">Dissolved
Ag</td><td rowspan="5">Blackened
Hg (-ous)</td><td colspan="4">Yellow or orange ppt.—Add $(NH_4)_2S$ to a part.</td><td colspan="5">Black or brown ppt.—Add $(NH_4)_2S$.</td></tr>
<tr><td colspan="3">Dissolved.—Add $(NH_4)_2CO_3$ to a second part.</td><td rowspan="4">Undissolved.
Cd</td><td rowspan="4">Dissolved
Sn (-ous)</td><td colspan="4">Undissolved.—Add a few drops H_2SO_4 to O. S.</td></tr>
<tr><td rowspan="3">Dissolved
As</td><td colspan="2">Undissolved.</td><td rowspan="3">Wht. ppt.
Pb</td><td colspan="3">No ppt.—Add NH_4OH to O. S.</td></tr>
<tr><td rowspan="2">Ppt. is yellow.
Sn (-ic)</td><td rowspan="2">Ppt is orange.
Sb</td><td rowspan="2">Blue ppt. soluble in excess.
Cu</td><td colspan="2">Wht. ppt.—Add NaOH to O. S.</td></tr>
<tr><td>Yellow ppt.
Hg (-ic)</td><td>White ppt.
Bi</td></tr>
</table>

If no ppt. with H_2S, pass to Table B.

* Indicating compounds of Sb, Bi, and Sn. A ppt. of sulphur appearing gradually indicates a *thiosulphate* or *persulphide*.

TABLE B.

To O. S. add NH_4Cl,* NH_4OH,† and $(NH_4)_2S$.						
Black ppt.—Add dilute HCl.			White or colored ppt.—Add NaOH to O. S.			
Dissolved. Fe To distinguish between *-ous* and *-ic* Salts, test O. S.	Undissolved.—Add to O. S. NH_4OH, drop by drop.		White ppt., soluble in excess. Divide the solution into two parts. To one part add H_2S, and to the other NH_4Cl.		Greenish ppt. soluble in large excess; re-precipitated on boiling. Cr	Buff or skin colored ppt., turning rust-colored when shaken up with air. Mn
	Blue ppt. soluble in excess, giving a *brown* solution. Co	*Green* ppt. soluble in excess, giving a *blue* solution. Ni	Wht. ppt. with H_2S. Zn	Wht. ppt. with NH_4Cl. Al		

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* To prevent the precipitation of magnesium oompounds.

† To neutralise any acidity.

TABLE C.

<table>
<tr><td colspan="8">To O. S. add NH_4Cl,* NH_4OH,† and $(NH_4)_2CO_3$.</td></tr>
<tr><td colspan="3">Wht. ppt.—Add $CaSO_4$ to O. S.</td><td colspan="5">No ppt.—Add Na_2HPO_4, and stir.</td></tr>
<tr><td rowspan="4">No ppt.
Ca</td><td rowspan="4">Wht. ppt. after some time.
Sr</td><td rowspan="4">Wht. ppt. at once.
Ba</td><td rowspan="4">Wht. ppt.
Mg</td><td colspan="4">No ppt.—Add NaOH to O. S., and heat.</td></tr>
<tr><td rowspan="3">Smell of ammonia.
NH_4</td><td colspan="3">No smell of ammonia.—Add $H_2\overline{T}$,</td></tr>
<tr><td rowspan="2">Granular wht. ppt.
K</td><td colspan="2">No ppt.—Test O. S. by flame test.</td></tr>
<tr><td>Yellow Flame.
Na</td><td>Colourless.
No common metal present.</td></tr>
</table>

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* To prevent the precipitation of magnesium compounds.

† To neutralise any acidity.

TESTING FOR THE ACID.

Evaporate a small quantity of the solution to dryness, in a porcelain capsule, and heat strongly. If it chars, proceed according to Table D; if not, pass on to Table E.

TABLE D.

To the solution (made neutral, if necessary) add $CaCl_2$, and shake well for some time. If no ppt. appears, heat to boiling.				
Wht. ppt. in cold, soluble in HCl.	Wht. ppt. on boiling, soluble in HCl.	No ppt.—Add Fe_2Cl_6 to O. S.		
Tartaric Acid.	*Citric Acid.*	Light brown ppt.—Evaporate a little O. S. to dryness, heat, and smell.		Red colour.—No ppt.
		Choking fumes.	Tarry or aromatic fumes.	*Acetic Acid.*
		Succinic Acid.	*Benzoic Acid.*	—
		Confirm. No ppt. with HCl.		*N.B.*—Acetates char very little, and sometimes not at all.

TABLE E.

Add $Ba(NO_3)_2$ to O. S. If no ppt., pass on to TABLE F.

Wht. ppt.—Add dilute HNO_3 to the ppt.						Yellow ppt. | H_2CrO_4
Undissolved. | H_2SO_4	Dissolved with effervescence. | H_2CO_3	Dissolved without effervescence.				
		Add to O. S. NH_4Cl, NH_4OH, and $MgSO_4$. Stir well with a glass rod. Wht. ppt. | H_3PO_4*	Add $CaCl_2$ to O. S. Wht. ppt. insoluble in acetic acid. | $C_2H_2O_4$	Evaporate some of O. S. to dryness, moisten some of the solid with H_2SO_4, and bring into the Bunsen flame on a platinum wire. Green flame. | H_3BO_3	Add $AgNO_3$ to O. S., and boil. Wht. ppt., turning grey. | H_2SO_3	

* *Arsenic acid* gives same result.

TABLE F.

Add $AgNO_3$ to the solution (made neutral, if necessary).

<table>
<tr><td rowspan="7">Black ppt.
H_2S</td><td colspan="7">Wht. or yellowish ppt.—Add HNO_3.</td><td colspan="3">No ppt.—Test for HNO_3 and $HClO_3$.</td></tr>
<tr><td colspan="7">Not dissolved.—Add Fe_2Cl_6 to O. S.</td><td colspan="3">Dissolved.—Add dilute H_2SO_4, and warm.</td></tr>
<tr><td rowspan="5">Blue.
$K_4Fe(CN)_6$</td><td colspan="6">Not blue.—Add Fe_2SO_4 to same portion.</td><td rowspan="5">Ppt. of sulphur, and smell of SO_2.
$H_2S_2O_3$</td><td rowspan="5">Nitrous fumes.
HNO_2</td><td rowspan="5">No result. Test for *Boric* and *Oxalic* acids as solutions may have been too dilute to give ppt. with $Ba(NO_3)_2$.</td></tr>
<tr><td rowspan="4">Blue.
$K_3Fe(CN)_6$</td><td colspan="5">Not blue.—Add HNaO to some portion, warm, and acidify with HCl.</td></tr>
<tr><td rowspan="3">Blue on acidifying.
HCN</td><td colspan="4">Not blue.—Add $Pb\bar{A}_2$, and boil.</td></tr>
<tr><td rowspan="2">Wht. ppt. turning black on boiling.
$HClO$</td><td rowspan="2">Yellow ppt. dissolved by heating with much water, and crystallising on cooling.
HI</td><td colspan="2">Permanent wht. ppt. dissolved on heating, &c. Add to O. S. MnO_2 and H_2SO_4, and heat.</td></tr>
<tr><td>Brown vapour.
HBr</td><td>Greenish yellow gas.
HCl</td></tr>
</table>

If no result so far, test for *Acetic* acid, and for *Nitrous*, *Boric* and *Oxalic* acids, as the solutions may have been too dilute for the discovery of these acids by the group reagents.

APPENDIX.

§ 288.—Metric System.

This is a decimal system, hence its advantages over the English one. It was devised by the French.

MEASURES OF LENGTH.

10 decimetres (dcm.) = 1 metre = 39·37 inches.
100 centimetres (cm.) = " = "
1000 millimetres (mm.) = " = "
1000 metres (m.) = 1 kilometre = 39370·79 inches.

MEASURES OF CAPACITY.

The unit is one cubic decimetre, called 1 litre = 1·76 Imperial pints, or 61·024 cubic inches.

1000 litres = 1 kilolitre.

MEASURES OF WEIGHT.

The unit is the weight of 1 cb. Cm. of distilled water at 4°C., called 1 gram = 15·432 grains.

The commercial unit is 1000 grams = 1 kilogram = 2·2046 lbs. Avoirdupois.

The subdivision of both the *litre* and the *gram* into tenths, hundredths and thousandths are named by prefixing deci-, centi- and milli- respectively to these names.

§ 289.—List of Elements.

The following is a list of those recognized by all chemists, with their symbols and atomic weights :—

NAME OF ELEMENT.	SYMBOL.	ATOMIC WEIGHT.
1. **Aluminium**	Al	27.3
2. Antimony	Sb (Stibium)	122
3. Arsenic	As	75
4. Barium	Ba	137
5. Bismuth	Bi	210
6. Boron	B	11
7. **Bromine**	Br	80
8. Cadmium	Cd	112
9. Cæsium	Cs	133
10. **Calcium**	Ca	40
11. **Carbon**	C	12
12. Cerium	Ce	138
13. **Chlorine**	Cl	35·5
14. Chromium	Cr	52
15. Cobalt	Co	58·7

Name of Element.	Symbol.	Atomic Weight.
16. **Copper**	Cu (Cuprum)	63·5
17. Didymium	D	147
18. Erbium	E	170·6
19. **Fluorine**	F	19
20. Gallium	Ga	70
21. Glucinum	Gl	9·3
22. Gold	Au (Aurum)	197
23. **Hydrogen**	H	1
24. Indium	In	113·4
25. **Iodine**	I	127
26. Iridium	Ir	193
27. **Iron**	Fe (Ferrum)	56
28. Lanthanum	La	139
29. **Lead**	Pb (Plumbum)	207
30. Lithium	Li	7
31. **Magnesium**	Mg	24
32. Manganese	Mn	55
33. **Mercury**	Hg (Hydrargyrum)	200
34. Molybdenum	Mo	96
35. Nickel	Ni	58·7
36. Niobium	Nb	94
37. **Nitrogen**	N	14
38. Osmium	Os	199
39. **Oxygen**	O	16
40. Palladium	Pd	106.5
41. **Phosphorus**	P	31
42. Platinum	Pt	197·5
43. **Potassium**	K (Kalium)	39·1
44. Rhodium	Ro	104·5
45. Rubidium	Rb	85.4
46. Ruthenium	Ru	104·5
47. Selenion	Se	79·5
48. **Silver**	Ag (Argentum)	108
49. **Silicon**	Si	28
50. **Sodium**	Na (Natrium)	23
51. Strontium	Sr	87·5
52. **Sulphur**	S	32
53. Tantalum	Ta	182
54. Tellurium	Te	128
55. Thallium	Tl	203·5
56. Thorium	Th	234
57. Tin	Sn (Stannum)	118
58. Titanium	Ti	50
59. Tungsten	W (Wolfram)	184
60. Uranium	U	240
61. Vanadium	V	51
62. Yttrium	Y	89·6
63. **Zinc**	Zn	65
64. Zirconium	Zr	89.6

Some of the more important ones are printed in full-faced type.

The Science Room.

The science room in a school should be one of the largest rooms in the building, well lighted and well ventilated. If used as a chemical, and also as a physical laboratory. the working tables of the students should be placed along one or two sides of the room. In the centre should be placed desks, or benches with arms, at which pupils may sit when making observations and writing notes on experiments performed by the science master in presence of the class.

Tank.—A large barrel will make a good tank for washing water. It should be placed in one corner of the room, 5 or 6 feet above the floor, and connected with one or more sinks by rubber, or better still, by iron-tubing.

Gas-Chamber.—In another corner of the room should be placed the gas-chamber. It may be made of pine boards, and should have sliding glass doors. Convenient dimensions are 3 ft. wide, 2 ft. deep, and 4 ft. high, the top of it being connected with a good ventilating or other chimney.

Working Tables.—The following are convenient dimensions: 12 ft. long, 3 ft. high, and 3½ ft. broad. Such tables are divided lengthwise along the middle, and transversely also into compartments 2½ ft. long; consequently each table affords working room for 10 students, 5 on each side. A cut of one of the tables in use in the Kingston Collegiate Institute is shown in the frontispiece. The tables, of course, when constructed in this manner are placed along the middle of the room.

Apparatus.—Each student's compartment should be supplied with the following apparatus :—

6 test-tubes, about 4 in. long, by ½ in. diam.
2 " " 8 " by about 1 in. diam.
1 rubber cork to fit large test-tubes, and pierced with one hole for bent delivery tube.
1 test-tube rack.
1 swab for washing test-tubes. } These two articles can easily be
1 glass stirring rod. } made by the student.
1 piece of platinum wire 3 in. long, with its end fused into a glass tube 2 in. long.
1 piece of platinum foil, or a sheet of mica.
1 blowpipe.
1 pair tongs for test-tubes.
1 funnel 3 in. in diameter.
1 package of filter paper to fit funnel.
1 florence flask, 4 oz. capacity.
2 glass beakers, 1 of 2 oz., and 1 of 4 oz. capacity. (Common tumblers will do).
1 evaporating dish, 3 or 4 oz. capacity. (A small saucer will do).
1 Bunsen burner or spirit lamp.
1 ring stand with piece of wire gauze 4 in. × 4 in.
1 sand bath (a tin saucer with sand in it).
1 piece of charcoal.
1 soup-plate, which does well for a small pneumatic trough.

Each student should provide himself with an apron, a towel and a box of matches.

In schools in which it has been found inconvenient to provide working tables for students, and where the teacher alone is expected to perform all experiments, the laboratory should, in addition to the foregoing list, contain the following pieces of *general* apparatus.

General Apparatus.

A chemical balance with weights running from 50 grs. to 1 gr. in brass, and from 500 mgs. to 1 mg. in platinum.

A good chemical thermometer.
A pipette of about 5 cc. capacity.
A Eudiometer.
A decomposition of water apparatus.
A spectroscope.
Bell jars : Pint and quart fruit jars answer well.
Large Beakers, assorted sizes.
Funnels, various sizes.
Funnel or thistle tubes.
Retorts, various sizes.
Liebig's condenser.
Iron mortar and pestle. Also a porcelain mortar and pestle.
Glass tubing, assorted sizes.
Mercury trough.
Gas-holder.
Rubber corks of various sizes.
   " tubing " "
   " Gas-bag.
Oxyhydrogen blow-pipe.
Crucibles ; porcelain, sand, or plumbago.
Test-tubes ; a large stock of assorted sizes.

Students' Reagents.

It will be found convenient to supply each student with the following named liquid reagents. None but pure chemicals should be used. The bottles for these should have a capacity of about 4 fluid ounces, and should have ground glass stoppers—paraffined to prevent them from sticking.

Sulphuric acid, H_2SO_4 ; oil of vitriol.
Nitric acid, HNO_3 ; aqua fortis.
Hydrochloric acid, HCl; spirits of salt.
Acetic acid, $H(C_2H_3O_2)$.
Ammonia, NH_3 (3 volumes water to 1 of strong am.)
Ammonic chloride, NH_4Cl, sal ammoniac.
Potassinm hydroxide, KHO ; caustic potash, (1 to 20 of water).
Sodium hydroxide, $NaHO$; caustic soda, (1 of salt to 10 of water).
Potassium iodide, KI, (1 of salt to 20 of water).

Calcium hydroxide, $Ca(OH)_2$; lime water. Add 30 or 40 parts of water and stir for 15 minutes. Throw away this solution, and add 300 parts of water and shake, and when still milky transfer it to a glass stoppered bottle, and allow it to stand until clear.

Magnesium sulphate, $MgSO_4$; Epsom salts, (1 of salt to 10 of water).

Mercuric chloride, $HgCl_2$; corrosive sublimate, (1 of salt to 16 of water).

Silver nitrate, $AgNO_3$; (1 of salt to 50 of water).

Lead acetate, $Pb(C_2H_3O_2)_2$; (1 of salt to 10 of water).

Alcohol, C_2H_6O.

Methylated spirits for lamp.

Turpentine, $C_{10}H_{16}$.

The student's table should also be supplied with the following named reagents, in the solid form. They are most conveniently kept in 2 oz. bottles with large mouths. Deliquescent solids should be placed in glass stoppered bottles:

Ferrous sulphate $FeSO_4+7H_2O$.
Copper sulphate $CuSO_4$.
Sodium carbonate Na_2CO_3.
Potassium chlorate $KClO_3$.
Metallic zinc.
Copper wire.
Copper filings.
Manganese dioxide.
Sulphur.

Besides the reagents already mentioned a number of others are needed by the student. These, along with working material, may be left on a side-table, and dealt out as required.

Material and Reagents.

The following is a full list of all the reagents and working material necessary for repeating the experiments detailed in this book, and no more need be ordered for schools in which the teacher performs all experiments. Where pupils have working-tables and are expected to try the experiments for themselves, the material should be ordered in much larger quantities than herein specified. Pupils, however, should be taught to use VERY SMALL QUANTITIES of the material in performing any experiment.

Acid acetic	1 oz.
“ hydrochloric	1 lb.
“ nitric	1 lb.
“ phosphoric	1 oz.
“ pyrogallic	1 oz.
“ sulphuric	1 lb.
“ tartaric	1 oz.
Alcohol	1 pt.
Alum, potassic	4 oz.
Ammonia	2 lb.
Ammonic chloride	4 oz.
“ nitrate	1 oz.
Arsenic metal	½ oz.
Arsenic white	1 oz.
Antimony sulphide	1 oz.
Barium peroxide	½ oz.
Beeswax	1 oz.
Boric acid	1 oz.
Borax	2 oz.
Calcium chloride	1 oz.
“ carbonate	2 oz.
“ fluoride	1 oz.
Carbon bisulphide	1 oz.
Chloroform	1 oz.
Charcoal, animal	4 oz.
“ wood	4 oz.

Coal, various kinds	
Copper, metal	1 oz. foil and wire.
" sulphate	4 oz.
" oxide	1 oz.
Ether, sulphuric	4 oz.
Gold leaf	
Galena	1 oz.
Gypsum	1 oz.
Indigo	1 oz
Iodine	⅛ oz.
Iron filings	2 oz.
" wire	1 oz.
" sesquioxide	1 oz.
" sulphate	1 lb.
" sulphide	1 oz.
" pyrites	1 oz.
Litharge	1 oz.
Lead acetate	1 oz.
" nitrate	½ oz.
" red	1 oz.
" peroxide	1 oz.
Litmus, pure	1 oz.
Manganese dioxide	1 oz.
Magnesium ribbon	4 ft.
" sulphate	2 oz.
Metals—specimens of each	
" —ores of each	
Mercury, metal	3 oz.
" oxide	1 oz.
" bichloride	1 oz.
Powder, bleaching	2 oz.
Phosphorus, white	1 oz.
Phosphorus, red	½ oz.
Platinum	
Potassium, metal	⅛ oz.
" bicarbonate	2 oz.
" bichromate	4 oz.
" bromide	1 oz.
" chlorate	4 oz.
" iodido	1 oz.
" nitrate	8 oz.
" hydrate	2 oz.
" permanganate	1 oz.
Quicklime	1 lb.
Sugar	2 oz.
Silicon dioxide—all varieties	
Silver nitrate	1 oz.
Starch	1 oz.
Sodium, metal	¼ oz.
" acetate	1 oz.
" phosphate	1 oz.
" carbonate	1 oz.
" hydrate	2 oz.
" chloride	2 lb.
" nitrate	1 lb.
Strontium nitrate	1 oz.
Sulphur (roll and flower	1 lb.
Tartar emetic	1 oz.
Turpentine	1 oz.
Watch-spring	
Wax candles and tapers	
Zinc, metal	1 lb.
" clippings	

A few reagents have been intentionally omitted, but only these which can easily be prepared in the laboratory.

Books of Reference.

Books of reference should be kept where pupils can consult them at pleasure. The following should be in every school laboratory .—

Treatise on Inorganic Chemistry, 3 vols., by Roscoe and Schorlemmer.

Experimental Chemistry, Parts I. II., by Dr. Reynolds.

The Chemistry of the Farm, by Warington.

Introduction to Chemical Philosophy, by Tilden.

University of Toronto.

I.

1. Chlorine was formerly regarded as a compound of hydrochloric acid gas with oxygen. Describe experiments, proving that this was an incorrect view.

2. 35·5 parts by weight of Chlorine, combine with 23 parts by weight of Sodium, 20 parts by weight of Calcium, and 3 parts by weight of Carbon, respectively. Explain why chemists consider the atomic weights of the three elements as follows :

$$Na = 23. \quad Ca = 40. \quad C = 12.$$

3. How may the following compounds of Sulphur be prepared *from the element*: Hydrogen Sulphide, Sulphur Dioxide, Sulphur trioxide, Sulphuric acid, Carbon disulphide?

4. Calculate the pressure produced by 1 grm. of hydrogen confined in a space of 100^{cc} at a temperature of 273°C, and calculate what weight of zinc would yield 1 grm. of hydrogen.

$$Z = 65.$$

5. Express by equations the following reactions:

(*a*) Chlorine on solution of potassium hydrate.
(*b*) Ammonium chloride on calcium hydrate.
(*c*) Heat on ammonium nitrate.

II.

1. What do you consider the objects of Chemistry as a science?

2. What facts lead us to the conclusion:

(1) That chlorine is an element?
(2) That the atomic weight of chlorine is 35.5?
(3) That the molecular weight of chlorine is 71?

3. What relations have been been found to exist between the volumes and weights of gases?

Calculate the weight of one cubic metre of each of the following gases at 0°.C under a pressure of one metre of mercury:

Oxygen, Hydrogen Iodide, Hydrogen Sulphide.

$$S = 32;\ I = 127;\ O = 16.$$

4. What explanation is given of the existence of the allotropic modification of oxygen known as ozone.

University College, Toronto.

I.

1. Give equations representing the following reactions:

(*a*) Hydrogen on Ferric Oxide. $2H + Fe_2O_3 = 2FeO +$
(*b*) Steam on Iron. $Fe + H_2O = FeO + H_2$
(*c*) Chlorine on a cold and (*d*) a hot solution of Potash. $Cl +$
(*e*) Nitric Acid on Tin. $HNO_3 + Sn =$

2. Describe the preparation of hydrogen sulphide. What volume of oxygen is required for the combustion of 1 litre of this gas, what will be the volume of the products formed?

3. How could carbon monoxide be shewn to contain half its volume of oxygen? CO

4. Chlorine is said to be monovalent, oxygen divalent, carbon tetravalent. Explain what is meant by these statements and give a method for the determination of the valency of one of these elements.

5. Describe the allotropic forms of sulphur. What proofs have we that these consist of nothing but sulphur?

6. Give a method for the preparation of hydrogen chloride, and explain fully the data on which the formula HCl is assigned to it.

$2NH_4Cl + H_2SO_4 = (NH_4)_2SO_4 + 2HCl$ | $2NaCl + H_2SO_4 = Na_2SO_4 +$

II.

1. Fownes, 12th ed., p. 252, states that "equal volumes of all gases contain an equal number of molecules;" and at p. 126, "the combinations (of the elements) are represented symbolically by the juxtaposition of the elementary atoms....thus the molecule of Hydrogen Chloride composed of one atom of hydrogen and one atom of chlorine is represented by HCl, the molecule of water by H_2O . . 3HCl denotes 3 molecules of hydrogen chloride, $3H_2SO_4$ 3 molecules of sulphuric acid."

In his account of the Chemistry of Chlorine, he represents chlorine in various equations by the following symbols: Cl, Cl_6, Cl_4, Cl_2. State fully what facts are conveyed by these symbols; which of them do you consider correct, and for what reasons?

For Honors.

Show that the statement that a certain gas contains one-fourth its volume of chlorine is impossible if the present atomic weight of chlorine is correct.

2. Show how the formula in use represent the per-centage composition of bodies, and from the following per-centage compostion calculate a formula:

S	56·637.
O	42·478.
H	·885.
	100·

3. Matter is said to be composed of the elements. Give several illustrations of the facts which lead to this theory, and explain why you consider a mixture of equal volumes of hydrogen and chlorine gases to be a "mixture" before explosion, and to be replaced by a "compound" after explosion.

4. Calculate the volume of $1^{\text{lb.}} = 453{\cdot}6^{\text{grms.}}$ of sulphur vapor at 500°C, and at 1000°C, and $760^{\circ\,\text{mm}}$ Bar.

III.

1. Illustrate the laws of combination with compounds of sulphur. Give a short account of the facts which convince us that:

(1) Sulphur is an element.
(2) The atomic weight of sulphur is 32.
(3) The molecule of sulphur contains two or six atoms.
(4) The atomicity of sulphur is two.

2. Describe a method of preparing ammonia. How could you show that this gas is a compound of the elements hydrogen and nitrogen. Quote experimental evidence that it contains half its volume of nitrogen and 1½ times its volume of hydrogen. $2NH_4Cl + Ca(OH)_2 = 2NH_3 + CaCl_2 +$

3. Hydrogen gas may be prepared by the reaction expressed by the following equation:

$$4H_2O + 3Fe = 4H_2 + Fe_3O_4.$$

What volume of hydrogen measured at 15°C and 745^{mm} Bar. will be formed by the decomposition of 400 grms. of water ?

What would be the volume if the molecule of hydrogen contained 3 atoms ?

4. How may oxygen be prepared ? By what properties is it characterized ?

The oxides of the non-metallic elements have been described as reacting with water to form acids. Illustrate this with the oxides of chlorine, sulphur, and nitrogen.

5. State fully the facts indicated by the formula S_2Cl_2. Why should it not be S Cl ?

IV.

1. " ..... the chemist finds himself obliged to divide substances into two great classes, (1) Compound Substances—those which he is able to split up into two or more essentially different materials, and (2) Elements or Simple Substances, those which he is unable thus to split up." —*Roscoe's Elementary Chemistry.*

Explain fully what is meant by the expression "split up."

2. A furnace uses 1 kilogram of coal an hour. What volume (in litres) of air at normal pressure and temper ture must pass through the furnace, assuming that only half of the oxygen is burnt, and that coal is pure carbon ?

3. Describe experiments showing that carbon monoxide contains half its own volume of oxygen.

4. In what forms does silicon chiefly occur in nature ? How is the element obtained in a free state ? Describe the characteristic properties of its oxide and fluoride.

5. How do you prove that air is a mixture ?

6. How is the element Sodium prepared ? Give a short account of the most important compounds, for what reason do we say that it is monovalent ?

7. Give a brief account of the Chemistry of Copper.

Calculate the percentage composition of cuprous chloride

$$Cu = 63{\cdot}5 \qquad Cl = 35{\cdot}5.$$

V.

1. What are the facts which convince chemists that matter is composed of a limited number of elements ?

2. The law of reciprocal proportion may be stated thus : if each of two elements combine with a third, then those quantities of the first two elements which combine with a fixed weight of the third are either those in which the first two elements combine or bear some simple ratio to them. Show that this is true of the elements mercury, chlorine and sulphur.

3. By what properties are the elements classified ? With what elements do you class iron ?

4. Describe the chemical reaction which occurs in the slaking of lime. Calculate the weight of water required to slake 1000 kilos. of lime. Do all oxides undergo this reaction?

5. If the equations

$$2PbS + 3O_2 = 2PbO + 2SO_2$$
$$PbS + 2PbO = 3Pb + SO_2$$

represent the chemical reaction of reducing lead, what volume of sulphur dioxide will be set free at 60° F. and 750^{mm} for each 100 kilos. of lead reduced?

6. If 200 million tons of coal were burnt last year, what volume of air must have been used, and what volume of carbon dioxide formed, taking an average temperature of 17° C. and normal pressure, and assuming coal to be pure carbon (·1 ton = 100 kilos. nearly)?

Second Class Teachers, 1885.

Examiner—JOHN SEATH, B. A.

1. Describe experiments to illustrate the general properties of acids, bases, and salts. Classify, if possible, the following under these heads, assigning your reason in each case:—

$$H_2S, KHO, CO, CaCO_3, H_2CO_3, CaO.$$

acid alkali acid acid acid [illegible]

2. Describe and explain fully one process by which you would disinfect a badly smelling drain.

3. State in each case the simplest mode of determining when a receiver is full, in the preparation of Ammonia, Chlorine, Carbon Dioxide, and Sulphur Dioxide. How would you transfer each of these gases from one receiver to another?

4. Describe experiments to show the nature and properties of Sulphur. How much air is needed to burn completely 4 oz. of Sulphur?

5. Fully describe and explain the following experiment:

(*a*) Some strong Sulphuric Acid is poured on a piece of zinc, and after the chemical action has ceased, water is carefully added.

(*b*) Carbon Dioxide is passed for some time through lime water. A portion of the clear solution thus obtained is boiled; another portion of it is exposed for an hour or so to the air; and, to another portion, lime water is added.

(*c*) Some distilled water is shaken up in each of the full receivers mentioned in 3 above.

(*d*) Some Chlorine gas is exposed to the air in an open receiver.

(*e*) One volume of Hydrogen is mixed with one volume and a half of Chlorine, and the mixture exposed to the action of diffused sunlight.

6. You are given a powder known to be Carbonate of Ammonia, Phosphate of Soda, Nitrate of Lead, or Chlorate of Potash. Describe the simplest mode of determining which it is.

First Class Teachers.—Grade C, 1885.

Examiner—John Seath, B.A.

1. State the principles that govern the relation of gases to pressure and to temperature.

One volume of Hydrogen is confined in a flask at 10°C under the ordinary pressure of the atmosphere, added to that of a column of mercury 60*mm* high. The flask is to be heated to 300°C without any increase taking place in the volume of the gas. How high must the column of Mercury then stand, supposing the atmospheric pressure to increase to 900*mm* ?

2. 0.3355 of an organic compound, containing only Carbon, Hydrogen and Oxygen, gave on combustion 0.6715 gramme of CO_2 and 0.2745 gramme of H_2O, and its vapor density was found to be forty-four times that of Hydrogen. Find its empirical and its molecular formula ; and express the latter in the graphic notation.

3. Make a list of (*a*) the impurities of city well-water, (*b*) the sources of such impurities, (*c*) the tests by which you would detect them, and (*d*) the means you would use to purify a given sample of impure water.

4. A powder is given you known to be Carbonate of Soda, Iodide of Potash, Bromide of Potash, Fluor Spar or Sulphate of Lime. Describe a *simple* mode of determining which it is.

5. 20 grammes of an aqueous solution of HCl were mixed with an excess of Argentic Nitrate. The precipitate, when collected, washed and dried, weighed 4.53 grammes. Calculate the percentage of HCl in the original solution ($Ag = 108$).

6. Fully describe and explain the following experiments :—

(*a*) Some white Arsenic is boiled with diluted Nitric Acid, and the gas given off is passed into water. To a portion of this solution is added a solution of Permanganate of Potash, and to another a solution of Iodide of Potash and Starch.

(*b*) A test-tube containing an aqueous solution of Chlorine is exposed to the strong rays of the sun.

(*c*) Oxygen which has been allowed to bubble through strong Sulphurous Acid is passed through a tube in which is heated some platinum sponge.

(*d*) Some Manganic Dioxide is boiled with an excess of strong HCl. The gas evolved is led into a strong aqueous solution of Potassic Iodide.

(*e*) Some Nitric oxide is mixed with an excess of Hydrogen and passed over moderately heated platinum sponge.

(*f*) A test-tube containing a piece of Phosphorus in an aqueous solution of fresh slaked lime is boiled for some time.

7. Explain fully what is meant by the statement that Silicon is an exception to Dulong and Petit's law.

Second Class Teachers, 1886.

Examiner—John Seath, B. A.

1. How would you demonstrate with $KClO_3$ the difference between physical change and chemical change?

2. With some water containing CO_2 in solution, is shaken up a mixture of pure sand and NaCl.

(1) How would you separate these four substances?
(2) How would you prove that you had separated them?

3. An organic body which is known to contain only C, O, and H, gives an analysis 27·58% of O and 10·35% of H. Its vapor density is 58, that of H being unity. What is its molecular formula?

4. You are given HCl and NH_3 (each in the form of a gas), litmus paper, and turmeric paper, and pure distilled water. How would you demonstrate the nature and properties of an acid, an alkali, and a salt?

5. A liquid is known to contain H_2SO_4, HI, HCl, KHO, or NH_4OH. Give a simple mode of determining which it is.

6. How would you demonstrate

(1) The resemblances and differences between H and Cl?
(2) The effects of heat upon a mixture of 4 vols. of H, one vol. of O, one of Cl, and one of N?

7. The water of a well is supposed to be contaminated by sewage. Describe the means you would take to determine the question?

8. (1) A glass rod moistened with strong H_2SO_4 is held very near a mixture of powdered $KClO_3$ and dry loaf sugar, *but so as not to touch it.*

(2) A glass rod moistened with strong H_2SO_4 is *brought into contact with* the same mixture.

Describe and explain what happens in each case, and state the general conclusion you would base on these and similar experiments.

Second Class.

NORMAL SCHOOLS, JUNE, 1887.

Examiner—John Seath, B.A.

1. What facts would lead you to the conclusion that matter is composed of elements?

Explain why, when equal volumes of H and Cl have been passed into a vessel, you consider the contents of the vessel to be a "mixture" before explosion and a "compound" after explosion.

2. Describe the allotropic forms of Carbon and Sulphur, and the experiments by which you would prove that they are allotropic forms.

3. Into separate test-tubes containing dilute HCl are put the following:—zinc, zinc oxide, marble, common salt, charcoal, gold. Explain, by means of equations, the chemical changes that take place.

4. If 112 litres of H weigh ten grammes, what is the weight in grammes of the same volume of Cl, and of the same volume of HCl, the, temperature and the pressure being the same?

Explain, in each case, the process by which you reach your conclusion.

First Class Teachers.—Grade C, 1887.

Examiners: { John Seath, B. A.
J. A. McLellan, LL.D.

1. How would you (1) prove and (2) explain Graham's Law of diffusion?

What volume of CO_2 will diffuse through a stucco plate in the same time as 2·464 grms of N_2O, both being at the standard pressure and temperature?

2. One gram of a certain metal when dissolved in dilute sulphuric acid liberates 200 c.c. of hydrogen gas. Find the combining weight of the metal.

3. By what experiments would you distinguish marsh gas and olefiant gas?

5 c.c. of a mixture of marsh gas and olefiant gas are exploded with 14 c.c. of oxygen: 9 c.c. of gas remain, of which 7 c.c. are absorbed by caustic potash. Find the volume of each of the gases in the original mixture.

4. (1) How would you prove that the gas obtained by pouring sulphuric acid upon ferrous sulphide contains both S and H?

(2) A solid substance contains both a carbonate and an easily dissolved sulphide. How would you prove the presence of these two bodies?

(3) A piece of sodium was completely converted into chloride by uniting with 200 c.c. of Cl at the standard temperature and pressure. What was the weight of the sodium?

5. Name the chief oxidizing agents with which you have experimented, and explain the theory of the action of each.

6 A small cage containing a live rabbit is placed in one pan of a delicate balance, and the instrument is then exactly counterpoised. If the whole be allowed to remain at rest for some time, a distinct diminution is seen of the weight of the cage and its contents. Explain this fact.

7. Name the elements in your course of study which form volatile compounds with hydrogen, arranging these elements according to their atomicities and giving the graphic formulæ of the hydrides.

8. You are given an unknown salt a common acid, and are required to identify the acid radicle. On analysis, the salt is found to be a phosphate. Give a full account of your analysis.

Second Class Teachers, 1887.

Examiners: { JOHN SEATH, B.A.
{ J. A. McLELLAN, LL.B.

1. The gas contained in a transparent receiver is known to consist of N and O. How would you determine whether it is a chemical compound or a mechanical mixture?

2. A glass is given you, containing muddy water impregnated with ammonia.

(*a*) How would you render the water pure?

(*b*) How would you prove that you had done so?

3. Some chlorate of potash and black oxide of manganese are heated in a test-tube till the evolution of gas ceases.

(*a*) How would you find the weight of the chloride?

(*b*) If the chloride weighs 10 grammes, what was the weight of the chlorate?

(*c*) How would you prove that the chloride is a different substance from the chlorate?

4. Find the quantities of lime and sal-ammoniac necessary to prepare 11.2 litres of NH_3.

5. Two receivers of the same size and shape, containing respectively Cl and H, are placed, for some time, mouth to mouth in diffused sunlight, the gases being separated by a thin plate of plaster of Paris.

(*a*) Describe minutely and explain what takes place in each receiver.

(*b*) What conclusions would you base on this experiment?

6. Into a receiver containing perfectly dry chlorine, is introduced some litmus paper.

Describe and explain what takes place (1) when the litmus paper is perfectly dry, and (2) when it has been moistened with water.

7. You are given an opaque receiver covered with a ground glass plate and known to contain N_2O, N_2O_2, HI, or Cl.

How would you determine most simply which it contains?

8. 3 litres of H, 2 of N, and 4 of O are measured at 0°C and 760^{mm} mercurial pressure, and an electric spark is passed through the mixture.

What is the volume of the gases after combustion, the measurement being made at 0°C and 760^{mm} mercurial pressure?

9. 100 grammes of nitre are distilled with sulphuric acid. What weight of ammonia will be needed to neutralize the distillate?

10. What volumes of the constituent gases would be obtained by decomposing three volumes of each of the following gases :—nitric oxide, ammonia, water vapor and hydrochloric acid?

From Williamson's Chemistry :—

I.

1. A litre of oxygen is required of the density of 100 at 0°C. What weight of potassic chlorate must be used for its preparation and what total pressure must be applied to it?

2. 100 cubic metres of hydrogen are supplied at 12°C. Wanted the weight of a balloon which when filled with the hydrogen would press upwards with a weight of 20 grammes.

3. What volume of air is required for the oxidation of that quantity of metallic copper which is reduced from its oxide by 10 grammes of hydrogen?

4. What weight of potassic chlorate is required for the evolution of that quantity of oxygen which is needed for the combustion of 10 litres of hydrogen?

5. An experimentalist fills with dry and pure atmospheric air at the temperature of 10°C, a bottle of 1 litre capacity at a height in the atmosphere such that a barometer stands at a height of 350 millimetres. What volume of nitrogen will he have at the normal temperature and pressure?

II.

1. What weight of nitrous oxide is obtained from a milligrammeo f ammonic nitrate? And what is the volume of the gas at the normal temperature and pressure?

2. What will be the dimensions of a flask capable of containing 20 grammes of ammonia gas at 12°C and 730 mm. pressure?

3. 3 grammes of nitric oxide are mixed with an excess of hydrogen and passed over moderately heated platinum sponge. What weight of ammonia is formed?

4. What weight of air is needed for the complete combustion of 1 milligramme of carbon?

5. What weight of carbon is contained in a litre of carbonic acid gas?

III.

1. What volume of carbonic acid must be passed over white-hot charcoal for the preparation of 10 litres of carbonic oxide?

2. What weight of ammonia would be neutralized by a kilogramme of hydric oxalate ($H_2C_2O_4$)?

3. What volume of air is needed for the complete combustion of a litre of marsh gas? What will be the volume (calculated at 0°C and 760 mm) of the steam formed by its combustion?

4. What volume of hydrogen would be obtained by passing a litre of hydrochloric acid gas over hot metallic iron?

5. What weight of potassic hydrate is needed for the neutralization of 100 grammes of hydrochloric acid gas?

IV.

1. 20 grammes of an aqueous solution of hydrochloric acid were mixed with an excess of argentic nitrate. The precipitate (argentic chloride) was collected, washed, dried, and weighed. It amounted to 453 grammes. Calculate the percentage of hydrochloric acid in the original solution.

2. What volume of chlorine can be made to combine directly with a litre of olefiant gas? What is the vapour-volume of the product?

3. What is the weight of a cubic centimetre of bromine vapor at 150°C and 760 min.?

4. What volume of hydriodic acid gas could be prepared from 10 grams of iodine?

5. What volume of oxygen is required for the combustion of a cubic centimetre of sulphuretted hydrogen? What is the volume of each product?

V.

1. What volume of hydrochloric acid would be obtained by replacing, by an equivalent weight of hydrogen, the boron in one litre of boric chloride (calculated at 0°C)?

2. What weight of boron is contained in a litre of boric fluoride?

3. What volume of oxygen is required for the combustion of 1 c.c. of phosphuretted hydrogen?

4. What volume of oxygen is needed for the complete combustion of 10 grams of carbon disulphide? What volume of each product is formed?

5. What volume of air is needed for the complete oxidation of 1 kilogramme of sulphur?

VI.

1. What weight of sulphurous acid is req uired for the neutralization of 100 grams of potassic hydrate?

2. What volume of nitric oxide is needed for the complete combustion of 10 grams of carbonic sulphide, CS_2?

3. What volume of hydriodic acid would combine with a litre of phosphuretted hydrogen?

4. What volume of phosphurous chloride could be obtained by combining one litre of chlorine with phosphorous?

EXAMINATION QUESTIONS.

The following questions have been prepared by Mr. Lochhead:—

1. Give the distinction between metallic and non-matellic elements. What reason are there for disregardfng the terms *non-metal* and *metal* as applied to the elements?

2. A substance on analysis gave the following percentage composition:

$$Mg = 9{\cdot}76, \quad S - 13{\cdot}01.$$
$$O - 26{\cdot}01 \quad H_2O\ 51{\cdot}22.$$

Calculate its formula.

3. Explain the chemical action which takes place in the formation of chalybeate spas.

4. How may the atomic weights of elements be determined?

5. Explain the formation of *stalactites, silicified wood,* and *coal.*

6. Required the weight of limestone needed to convert 50 tons of soda crystals into bicarbonate.

7. 109·6308 grams of mercury were obtained from 118·3938 grams of the red oxide. What is the atomic weight of mercury?

8. Explain Mendelejeff's classification of the elements. What modifications has Lothar Meyer made in his representation of the Periodic Law?

9. What are:—*Paris Green, Scheele's Green, Turnbull's Blue, White Lead, Blue Vitriol, Verdigris* and *fer ré duit?* Explain their composition.

10. What improvements did Simens and Dafur make on Bessemer's process?

11. Describe the chemical action which takes place when *mortar sets* and *lime slakes,* also when *Paris plaster sets.*

12. Explain the preparation of the *superphosphate of lime* fertilize from apatile.

13. Describe the chemical process of producing a photographic print

14. What is the composition of *kaolin, lapis lazuli, emery, clay,* and *ruby?*

15. Explain the process of *soldering.* What is the use of borax in the operation?

16. Write a short article on the occurrence of potassium in nature its uses and the part it takes in the vegetable and mineral kingdoms.

17. When gunpowder explodes, represent the chemical change by an equation.

18. What relation exist between the atomic weights of the aikal metals?

19. Explain a lime-kiln. What change takes place when lime is lef exposed to the air?

20. Why are the zinc plates in galvanic batteries amalgamated? What objection would there be to having a little corrosive sublimate mixed with the calomel which is used in medicine?

21. Name the chief zinc ores. How is the metal extracted from the ores?

22. How can potassium and sodium be detected when both are present?

23. Given that calomel vapor has the specific gravity 117·6. Is it proper to write its formula Hg_2Cl_2? What reasons have we in supposing that the sp. gr. is much higher?

24. How would you distinguish between mercurous and mercuric, ferrous and ferric, stannous and stannic compounds?

25. How would you prepare *mortar* as used by masons? What use is made of lime in agriculture?

26. Explain the composition of the following cements: *hydraulic, Roman, Portland.*

27. What substance is employed as an adulterant of *white lead?*

28. What is *dialysis?* How is the method employed in the separation of arsenic from organic mixtures?

29. How would you shew that "the heat capacity of an element is inversely as its equivalent?" Distinguish between the *equivalent* and the *atomic* weight of an element.

30. Enunciate Dulong and Petit's Law, and state what use it is in the accurate determination of atomic weights?

31. What is meant by *electro-positive* and *electro-negative* elements? Does this division correspond to the *non-metallic* and *metallic* division?

32. Distinguish between *empirical, rational* and *graphic* formulæ.

33. What experimental evidence is there for the existence of metal ammonium?

34. Explain the composition of the different *glasses.* What use does silica serve in the building up of plants?

35. What is the composition of *talc, meerschaum, serpentine* and *asbestos?* What uses are made of these minerals in the arts?

36. Distinguish between *Epsom* and *Rochelle Salts.*

37. What is the composition of *ordinary bricks* and *fire bricks?* Explain the cause of *slagging.*

38. How much iron sulphide will be required to yield ten litres of hydrogen sulphide at 0°C and 760 mm.?

39. Calculate the formula of soda felspar from the following analysis:

SiO_2 – 68·45; Al_2O_3 – 18·71; Fe_2O_3 – 0·27;
CaO – 0·50; MgO – 0·18; K_2O – 0·65;
Na_2O – 11·24 = 100·00.

40. What is a *mordant?* Name some compounds which are mordants.

The following sets of questions have been selected from those set at the examinations of the Science and Arts Department, South Kensington, England:—

I.

1. I add two volumes of oxygen to one volume of each of the following gases: What takes place and what effects will be produced if an electric spark be afterwards passed through each of the mixtures? Chlorine, hydrogen, sulphuretted hydrogen, nitrous oxide, nitric oxide, carbonic oxide, carbonic anhydride.

2. What takes place when carbonic anhydride is passed into—1st, distilled water; 2nd, baryta water; and 3rd, water containing some freshly precipitated calcic carbonate (carbonate of lime)?

3. State exactly how you would separate from each other, and individually detect, the following constituents of a solid substance given to

you for analysis :—Deroxide of mercury, soda, protoxide of iron, oxide of copper, magnesia, sulphuric acid, and hydrochloric acid.

4. How much hydrogen by weight and by volume (in litres) is required to reduce 25 grammes of ferric oxide (Fe_2O_3) to metallic iron ?

5. You have given to you distilled water, oil of vitriol, nitric acid, copper turnings, iron filings, and metallic lead. State what salts you could prepare from these materials, and describe briefly how you would make them. Give their graphic formulæ and explain the chemical changes by equations.

II.

1. You have given to you some sulphur, water, and nitric acid. Describe how you would make sulphuric acid from these materials.

2. You are required to make oxygen from chlorine and water. Describe exactly how you will do it, and give a sketch of the apparatus which you purpose to employ.

3. I pass sulphuretted hydrogen through an aqueous and acid solution of the following salts :—Merric Chloride, Leadinic Sulphate, Stannic Chloride, Chrome Alum, Potassic Chromate, Magnesic Sulphate.

4. You are required to perform a qualitative analysis of a solid substance containing the following ingredients :—Baric carbonate, sodic sulphate, cupric chloride, zincic chloride, and magnesic chloride. State exactly how you will separate and individually recognize such constituent.

III.

1. If I agitate with pure water a mixture of hydrogen, nitrogen, carbonic anhydride, and hydrochloric acid gases, what effect will be produced ?

2. I pass a stream of carbonic anhydride gas through an aqueous and ammoniacal solution containing the following salts:—Basic chloride, Sodic chloride, Calcic chloride, Potassic chloride, Strontic chloride, Magnesic chloride. Explain the chemical changes which occur and draw graphic formulas of the products.

3 Mention some of the chief compound medicals used in the notation of inorganic bodies, give their graphic formulas, and quote examples of their employment in the formulation of inorganic compounds.

4. If I allow a current of steam to blow through some iron nails heated to redness in a crucible what happens ? If I substitute copper nails for the iron ones in the above requirement, how will this affect the result ?

5. How would you separate :—(1) sand and sugar ; (2) chalk from salt ; (3) charcoal from nitre ; (4) glass from boric acid ; (5) sulphur from water ?

IV.

1. Classify the following elements into metals and non-metals, and into positive and negative elements :—Al, Ca, Cb, Cu, F, H, I, F, Pb, Mn, Hg, N, O, P, K, Si, Ag, Ua, S, and Zn.

2. Describe several methods of separating analytically manganese from iron.

3. How would you distinguish hydriodic acid in the presence of hydrochloric acid?

4. You have given you some cupric nitrate, stannic oxide, zincic nitrate, and charcoal. Describe how you would prepare from these materials an alloy containing the metals copper, tin and zinc.

V.

1. Describe the changes produced in brimstone by the continued application of heat.

2. What is the action of water upon the following substances:—hydrogen, carbonic anhydride, ammonia, sodic carbonate, chalk, and sodium?

3. What is a *crith* and how is it employed by chemists?

4. Metallic copper, sodic chloride, sulphuric acid, saltpetre, and solution of ammonia are given to you, and you are required to produce therefrom (*a*) nitrous oxide, and (*b*) nitric oxide.

5. Water is poured upon a powder consisting of equal weights of boric nitrate, common salt, sulphur, argentic nitrate and anhydrous magnesic sulphate. What acids and bases will be present in the filtered solution, and how would you identify them? What will be the nature of the insoluble substance?

VI.

1. A solid substance contains both a carbonate and an easily decomposable sulphide. State how you would prove the presence of these bodies?

2. How would you prepare PCl_3 and $AsCl_3$? Give a sketch of the necessary apparatus, and express by equations the action of water upon each of these compounds.

3. Water, sodium, hydrochloric acid gas, and iron being given to you, describe four methods of preparing hydrogen. Sketch the apparatus.

4. What chemical effect is produced when a strong solution of hydrochloric acid in water is added to each of the following substances:—Zinc, iron, chalk, silver, sodic carbonate and gold?

5. What takes place when a solution of mercuric chloride is added to one of potassic iodide?

VII.

1. To a solution of argentic nitrate, caustic potash is added in excess. The precipitate produced is washed with water and then washed with hydrogen dioxide. Explain the chemical changes occurring in these operations.

2. Explain what takes place when sulphuretted hydrogen is passed through bromine water.

3. What chemical changes occur when a solution of potassic iodide is added to each of the following salts:—Mercuric chloride, plumbic nitrate, sodic sulphate, argentic nitrate, and sodic sulphite.

4. Describe what takes place (*a*) when fully-burning sulphur is plunged into nitrous oxide, (*b*) when phosphorus in brilliant combustion is immersed in nitric oxide; (*c*) when carbonic anyhdride is passed over ignited sodium, and (*d*) when steam is passed over red-hot manganese.

VIII.

1. What changes takes place when zinc dissolves in dilute nitric acid?

2. I heat as strongly as possible, a fragment of each of the following substances in a glass tube through which a current of air is passing: Sulphur, phosphorus, boron, charcoal. What chemical changes take place?

3. What changes take place when the following salts are heated to redness in the open air:—Hydride disodic phosphate, microcosmic salt, magnesic carbonate and baric carbonate.

4. What happens when sulphuric acid is poured upon (1) rock salt, 2) chalk, (3) saltpetre. Give equations.

5. How would you prove that ammonia contains nitrogen and hydrogen?

IX.

1. Explain by equations and otherwise what takes place when chlorine gas is passed over red-hot (1) silicon, (2) diamond, (3) iron, (4) arsenic, (5) alumina.

2. A colourless solution gives a heavy, white precipitate with hydrochloric acid, soluble in hot water, and turned black by ammonia. What metal does this indicate, and how would you prove the indications to be correct?

3. By what tests would you ascertain whether a sample of carbonate of potash contained (1) chloride, (2) sulphate, (3) sulphide of potassium?

4. A sample of water contains in solution, in 100,000 parts, 16 parts of calcic carbonate, 5 parts of calcic sulphate, and 7 parts of common salt What is its hardness, and how much will this hardness be reduced by boiling the water for half an hour.

5. Explain by equations and otherwise, the reactions which occur when strong sulphuric acid is heated with (1) copper, (2) platinum, (3) fluor-spar, (4) chalk, (5) flint.

X.

1. Ammonia gas is lead into a solution of hydrochloric acid. What is produced? Give an equation. The resulting liquid is evaporated to dryness in a porcelain crucible, which is afterwards heated to redness. What will remain in the crucible?

2. Carbonate of ammonia and nitric acid are given you. How would you prepare nitrous oxide from these? Give equation. Draw apparatus used.

3. Describe minutely the test you would adopt for the detection of the presence of arsenic in a sample of green wall-paper.

4. A heavy white powder, insoluble in water, dissolves in hydrochloric acid with effervescence, and the solution tinges the flame green. What substance is indicated? How would you confirm the indication?

5. A glass of dirty water is given to you. Describe with sketch the plan you would adopt (1) to render the water clear, (2) to find out whether the clear water contains any solid matter dissolved in it.

XI.

1. Three pint bottles of chlorine are given to you, describe the experiments you would make to show clearly its characteristic properties·

2. Describe what you consider the best experimental proof that sulphur is a constituent of sulphuretted hydrogen.

3. A white powder is shaken up with distilled water; how whould you ascertain whether any of it dissolves?

4. A process similar to that used for preparing HCl does not answer if applied to the preparation of HI. Explain why, and tell how you would prepare the latter.

5. Metallic bismuth and metallic antimony are acted upon separately by nitric acid. Describe what takes place in each case.

XII.

1. If 3 c.c. of hydrogen diffuse across a porous plate in 90 minutes, what volume of oxygen would pass through it in 2 hours?

2. Compare the density of two gases, equal volumes of which pass through a porous plug in 1 hour and 4 hours respectively.

3. Compare the rates of diffusion of marsh gas and sulphur dioxide.

4. Assuming a sq. centr. of leaf in sunlight to decompose 0·11 c.c. of carbon dioxide in an hour, find the volume of oxygen liberated in 5 hrs. by leaves whose total surface is 2500 sq. cen.

5. How many grams of nitric acid containing 67·2 % of pure HNO_3, will neutralize 54·4 grams of ammonia containing 36 % of NH_3?

XIII.

1. How much potassium oxide is formed by burning in nitrogen monoxide potassium enough to liberate 3125 c.c. of nitrogen?

2. Water at 0°C. dissolves 1149 times its volume of ammonia gas, what mass of ammonia at 950 mm. would 625 grams of water at 0°C. dissolve?

3. Water at 15°C. dissolves its own volume of carbon dioxide, whatever the pressure, what mass of the gas can be dissolved under a pressure of 3 atmospheres in 25 litres of water at 15°C.?

4. A mine contains 104,000 cubic feet of marsh gas, and 1,000,000 cubic feet of atmospheric air at 0°C. and 760 m. m. pressure. What will be the volume of the gases at the same temperature and pressure after the explosion of the mixture.

5. Find how much iodine is liberated by adding 60 grams. of chlorine water containing 0·71 °/₀ of chlorine by mass to excess of solution of potassium iodide.

XIV.

1. What mass of chlorine would replace 32 grams of bromine in a compound of the latter?

2. How much solution of hydrogen chloride, containing 30 °/₀ of the gas, will neutralize 102 grams of solution of ammonia containing 35 °/₀ of ammonia gas?

3. One volume of water at 15° C. dissolves about 3·2 volumes of hydrogen sulphide; how much iron sulphide will be needed to prepare gas enough to saturate 375 litres of water at this temperature, and under 760 m. m. pressure, assuming no loss to occur.

4. If one pound of bone-ash contain 85 °/₀ of tri-calcic phosphate, Ca_3PO_4, how much phosphorus could be obtained from 100 pounds of bone-ash?

5. A mixture of marsh gas and olefiant gas is exploded with 9 c. c. of oxygen; on treating the 6 c. c. of gas which remain with caustic potash, their volume is reduced to 2 c. c. Find the composition of the mixture.

The author is indebted to Professor Goodwin, D.Sc. (Edin.), for permission to use the following questions. They are selected from Dr Goodwin's recent work, *Chemistry for Students of Medicine.*

1. Point out the errors in the following equation:—$2FeSo_4 = Fe_2O_3 + SO_3$.

2. In what respect does a chemical compound of iron and sulphur differ from a mixture of these two substances?

3. 100 grams of oxygen combine with 393·75 grams of copper. What is the equivalent of copper?

4. Iron forms two basic oxides. The percentage composition of one is,—*iron*, 77.78%; *oxygen*, 22.22%; and of the other,—*iron*, 70%; *oxygen*, 30%. Apply the law of multiple proportions.

5. A litre of nitrogen unites with 3 litres of hydrogen. The molecule of the compound formed contains 1 atom of nitrogen and 3 of hydrogen. What is the change of volume when combination takes place?

6. What is the composition of the substances represented by the following formulas:—NH_3, H_3PO_4, $SO_2(OH)_2$, Fe_2Cl_6, $CO(NH_2)_2$, and $K_4Fe(CN)_6$? How many atoms in the molecules represented as follows:—$C_2H_4(OH)_2$, $(NH_4)_2SO_4$, $FeSO_4$, $7H_2O$, and $Fe_2(OH)_6$?

7. What weights of the following gases measure, at standard temperature and pressure, 22.33 litres, viz.:—*carbon monoxide* (CO); *ethylene* (C_2H_4), *hydric sulphide* (H_2S); *phosphine* (PH_3), and *methylamine* ($CH_3.NH_2$)?

II.

1. Would you expect to find hydrogen uncombined in the air?

2. What volume of mixed gases, measured at 1000 mm. pressure and 10° C., is produced by the electrolysis of 100 grams of water?

3. "One atom of zinc sets free a molecule of hydrogen." What evidence can you bring in support of this statement?

4. How does it happen that the atomic weight of hydrogen is 1?

5. Why does hydrogen explode less violently with air than with oxygen.

6. From the following formulas deduce the atomicity of bismuth (Bi), phosphorus (P), tin (Sn), and sulphur (S): HCl, $BiCl_3$, PCl_5, H_2O, SnO_2, SO_3.

7. The specific weights of water vapour and oxygen are 9 and 16 (Hydrogen=1). If 100 c.c. of water vapour diffuse into a vacuum in a certain time, what volume of oxygen will diffuse under the same conditions in the same time?

III.

1. A flask filled with water and closed with a cork through which passes a narrow tube open at both ends is held upside down. The water does not run out. Why?

2. Air is found to become less and less dense as we ascend. Account for this.

3. Why do men need to breathe faster at great elevations than lower down?

4. Why does blood burst through any place where the skin is thin, when a very great elevation is reached, as in balloons?

IV.

1. Hold the moist stoppers of the ammonia and hydrochloric acid bottles near each other. Explain what you observe.

2. Warm some ammonic chloride solution with sodic hydroxide solution. Observe the smell. What substances have been formed? Write the equation.

3. In determining the composition of ammonia, how can the hydrogen be got rid of in order to measure the volume of nitrogen?

4. Lime water is an antidote to poisoning by nitric acid. Explain its action.

5. Calculate the weights of nitric acid required to neutralize 10 grams each of the following bases and oxides:—Sodic hydroxide ($NaOH$), potassic hydroxide (KOH), calcic hydroxide ($Ca(OH)_2$), magnesia (MgO), and litharge (PbO).

6. Write the formulas for nitrates of the following metals, referring to the table of elements for the valences:—Calcium, silver, iron, mercury copper, cobalt, barium, magnesium, and aluminium.

7. What experiments already made illustrate the direct replacement of hydrogen by metals?

8. What volume of nitrogen monoxide at 16° C. and 750 mm. will completely burn 1 gram carbon?

$$C+2N_2O=2N_2+CO_2.$$

9. What volume of air will convert 10 cubic inches of nitrogen dioxide into the tetroxide?

V.

1. Compare the halogens (1) as to their chemism for metals, and (2) as to their chemism for oxygen.

2. What weight of pure hydrochloric acid is there in 1 litre of the solution of sp. wt. 1.11?

3. Compare the bleaching power of chlorine (and water), and hypochlorous acid.

$$\begin{aligned} Cl_2+H_2O &= 2HCl+O. \\ HClO &= HCl+O. \end{aligned}$$

What weight of chlorine is equivalent to 100 grams of hypochlorous acid?

4. What weight of sodium chloride must be used to prepare chlorine enough to set free the bromine from 100 oz. of potassic bromide (KBr)?

5. Write the formulas of *magnesic*, *ammonic*, *ferric*, *cobaltic*, and *mercurous* chlorides; of *mercuric*, *argentic* and *baric* iodides; and of *ferric* bromide.

VI.

1. Show how the oxides of sulphur illustrate the law of multiple proportion.

2. 196 g. sulphuric acid in solution is mixed with 150 g. sodic hydroxide. Is the solution neutral, acid or alkaline?

3. What is the valence of sulphur in sulphur dioxide? In sulphur trioxide? In hydric sulphide? In sulphur hexiodide (SI_6)? In sulphuric acid ($SO_2{}^{-OH}_{-OH}$)?

4. What volume of chlorine is necessary to decompose 1 cubic foot of hydric sulphide gas?

5. Moisten a strip of filter paper with plumbic acetate and hold it in the mouth of the bottle of hydric sulphide solution. It is blackened. Explain

VII.

1. Compare nitrogen and phosphorus, as to their compounds.

2. In what dehydrating process is phosphorus pentoxide used?

3. Mix a solution of sodic phosphate (Na_2HPO_4) with one of plumbic acetate. What result? Try with phosphite and hypophosphite.

4. "The atomicity of phosphorus is 3 in some compounds, and 5 in others." Apply this statement.

5. What is the percentage of phosphorus in pure *apatite* ($Ca_3(PO_4)_2.CaCl_2$)?

VIII.

1. What causes the effervescence of soda water?

2. How would you prove by an experiment that carbon dioxide is formed by the burning of a candle?

3. Show that most of our materials for fires and lights are from a vegetable source.

4. Write a graphic formula for acetylene (C_2H_2). With how many atoms of chlorine would you expect its molecule to combine?

IX.

1. Compare silicon and carbon, (*a*) with regard to chemical, and (*b*) with regard to physical, properties.

2. What is *soluble glass?*

3. It is found that alkaline solutions eat away glass. Explain.

4. Mention an acid and a base that will dissolve sand.

5. How can it be proved that silica is an acid-forming oxide?

The following questions have been selected from Dr. Ira Remsen's *Elementary Course of Chemistry.* Messrs. Henry Holt & Co., of New York, are the publishers of Remsen's books in America, and from these gentlemen I have obtained permission to use the following selected questions and problems.

THE AUTHOR.

I.

1. What is meant by saying that physical and chemical changes are related?

2. Give some familiar examples which make these relations clear.

3. How does the steam-engine illustrate these relations?

4. Suppose a stone should fall upon some gunpowder and cause it to explode, which would be physical and which chemical change?

5. How can we distinguish chemical action from all other kinds of action?

6. How many elements enter into the composition of the things with which we generally have to deal?

7. In general, what is meant by chemical action?

8. What different kinds of chemical action are there? Give examples of each.

9. Which of the elements is most abundant? Which comes next in order?

10. Which is the principal element that enters into the composition of living things?

II.

1. Describe the changes which are produced in lead, zinc, and tin by heating them, and describe the experiments which taught you what these changes are.

2. How did you learn that the air had anything to do with these changes?

3. Did heat have anything to do with the changes? Suppose you knew that the bits of lead, zinc and tin increased in weight when heated in the air, and that they did not increase in weight when heated so that the air could not get at them, what would that show?

4. What familiar facts show that the air has something to do with burning?

III.

1. How can we get oxygen from the air? What happens to oxygen when it is much cooled down and compressed?

2. How does oxygen behave towards other substances at ordinary temperature? How do you know this? Does oxygen act upon anything at the ordinary temperature? Give examples.

3. What difference is there between the action of oxygen at the ordinary temperature and at higher temperatures? How did you learn this difference?

4. Is burning in the air the same chemical act as burning in oxygen? How can this be proved? Why do substances not burn as actively in the air as they do in oxygen?

5. What is meant by the kindling temperature? Explain why it is that a stick of wood burns gradually and not all at once.

6. Explain the connection between the heat and light produced, and the combustion of a substance.

7. What is meant by the expressions *chemical work* and *chemical energy?*

IV.

1. How are natural laws discovered?
2. What is a natural law?
3. What are the combining weights of the elements?

V.

1. How can nitrogen be obtained from the air by the use of copper? How does it differ from the oxygen in its conduct towards burning things? Could animals live in it? Why? Suppose there were no nitrogen in the air, how would our fires differ from the fires in the air?

VI.

1. How can it be shown that water is contained in wood? in meat?

2. Is water present in large or small proportion in animal and vegetable substances?

3 What is meant by water of crystallization?

4. What are efflorescent substances?

5. What are deliquescent substances?

VII.

1. Which are the common acids? What do they all contain? What takes place when they are treated with metallic elements?

2. What relation is there between the combining weights of hydrogen and oxygen and the weights of equal bulks of the two gases?

3. When we say that the combining weight of hydrogen is 1 and that of oxygen 16, what is meant? What is meant when we say the combining weight of iron is 56?

4. If we should call the combining weight of oxygen 100, what would be the combining weight of hydrogen?

VIII.

1. Explain exactly how the experiment with copper oxide teaches us what the composition of water is.

2. What is the character of the water found in mountain streams which flow over sand-stone? Why?

3. What is the character of water which flows over lime-stone?

4. How does water become salt?

5. What are effervescent waters? What is sulphur water?

6. What change takes place in river-water which has been contaminated with drainage?

7. Does combination take place more readily between those elements which are alike or between those which are unlike?

8. What change takes place in oxygen when electric sparks are passed through it?

IX.

1. From what kind of substances is ammonia given off in destructive distillation?

2. What is one of the products formed when substances containing carbon, hydrogen, and oxygen are heated? In what experiment which you have already performed is this shown?

3. Explain why ammonia is formed in gas-works?

4. What does the process of decay consist in?

5. What becomes of the nitrogen contained in animal substances when they decay?

6. What connection is there between saltpetre and nitric acid? between potassium nitrite and nitrous acid?

7. How is nitric acid formed in nature?

8. What is the appearance of pure nitric acid? What takes place when it is boiled? when the sun shines directly upon it? How does strong nitric acid act?

9, Why do substances burn in strong nitric acid? Describe some experiments which illustrate the power of nitric acid. What does the nitric acid give up to the substances upon which it acts?

10. What does nitric acid give up when a metallic element acts upon it? Describe what further takes place, and why?

X.

1. What is meant by saying that oxygen belongs to a family of elements?

2. What two steps are necessary in order to get chlorine out of sodium chloride or common salt? What resemblance is there between the process for making hydrochloric acid from common salt and that of making nitric acid from sodium nitrate? What is formed besides hydrochloric acid and nitric acid in each case?

3. What is the action of chlorine upon water?

4. What is disinfection? What is "bleaching powder"? What other name has it? Why is it valuable as a disinfecting agent?

5. Compare the action of hydrogen on oxygen and on chlorine. What are the products? What are chlorides?

6. What difference is there between the action of a mixture of chlorine and hydrogen and a mixture of hydrogen and oxygen?

7. What happens to hydrochloric acid when it is treated with a metallic element like zinc? when treated with an oxide like zinc oxide? When treated with substances which give up oxygen readily?

XI.

1. How can you determine what is formed when an acid acts upon a base? What is formed when hydrochloric acid acts upon caustic soda? nitric acid upon caustic soda? sulphuric acid upon caustic soda? hydrochloric acid upon caustic potash? nitric acid upon caustic potash? sulphuric acid upon caustic potash?

XII.

1. What takes place when animal and vegetable substances are heated to a high temperature? Why is this? What takes place when they are heated in the air?

2. What is coke, and how is it obtained? lampblack? bone-black or animal charcoal?

3. What are charcoal filters used for? What are bone-black filters used for?

4. What is the object of charring piles which are exposed to the action of air and water?

5. What different kinds of coal do we distinguish between? What is lignite? peat?

6. How can we form an idea in regard to the reason why one and the same thing can appear in different forms?

7. Explain what takes place when a mixture of charcoal and copper oxide is heated; when a mixture of arsenic and charcoal is heated. Is

there any resemblance between the action of hydrogen and of charcoal on heated copper oxide?

8. What important use is made of charcoal as a reducing agent?

XIII.

1. What are hydrocarbons? Are these easily formed in the laboratory? Under what circumstances are they easily and abundantly formed?

2. What are the final products of the oxidation of vegetable matter? What is the chief product of the reduction of vegetable matter?

3. Of what importance is the occurrence of marsh-gas in coal-mines?

4. What processes which are constantly taking place give rise to the formation of carbon dioxide?

5. Why does not carbon dioxide burn?

6. What is meant by the statement, "Carbon can do chemical work?" What resemblance is there between a piece of carbon and an elevated body of water?

7. What is choke-damp, and what does the name come from?

8. What resemblance is there between the food of animals and the fuel burned in stoves?

9. How does the carbon of animals and plants get back into the air again?

10. In what way is all life directly dependent upon the sun?

11. How are carbonates formed? What is the composition of sodium carbonate? of potassium carbonate? What, then, is the composition of carbonic acid? What takes place when carbonic acid is set free from carbonates?

12. What takes place when carbon dioxide acts upon potassium hydroxide? upon calcium hydroxide? Give the equations representing the action, and name the products.

13. Why is the use of water-gas sometimes objected to?

14. Why is carbon monoxide a good reducing agent?

15. Of what importance is it in the reduction of iron from its ores?

16. What is a flame? What is the difference between a candle and a lamp?

17. How can it be shown that when a burning gas is cooled down it is extinguished?

XIV.

1. Do we know why substances combine according to the law of definite and multiple proportions?

2. What is an hypothesis? a theory?

3. What theory has been suggested to explain the laws of definite and multiple proportions?

4. Show how this theory accounts for the facts?

5. How does Avogadro's law help us to determine the relative weights of the molecules of gaseous substances?

6. How are the atomic weights determined from the molecular weights?

7. What difference is there between chlorine, oxygen, nitrogen and carbon as shown by the symbols of their compounds with hydrogen? What is meant by the valence of an element?

8. What is meant by a univalent element? a bivalent element? a trivalent element? a quadrivalent element? Barium forms the compound $BaCl_2$; what is the valence of barium? Sodium forms the compound NaCl; what is the valence of sodium?

9. In the formation of potassium nitrate from nitric acid, how is the valence of potassium shown? When a bivalent element like calcium forms a salt with nitric acid, how does the displacement of hydrogen take place? Calcium is bivalent. What is the symbol of its salt with sulphuric acid? Explain this. What is the symbol of the sodium salt of sulphuric acid? What does this show with regard to the valence of sodium?

11. If magnesium, Mg, is bivalent, what is the symbol of its sulphate? of its nitrate? of its chloride? What is the basis for the distinction between acid-forming and base-forming elements? Give examples of the two classes. What are these classes sometimes called?

12. What is meant when we speak of a family of elements?

13. What are the families of acid-forming elements?

XV.

1. How is bromine obtained from sodium bromide? Give the equations representing the steps which must be taken, and explain what is meant by them.

2. What is hydrobromic acid, and how is it formed? What difference is there in the conduct of common salt and of sodium bromide towards sulphuric acid? How is this explained?

3. How is iodine obtained from sodium iodide? Give the equations representing the action, and explain what they mean.

4. What are the properties of iodine? Compare chlorine, bromine, and iodine, stating the points of resemblance and difference.

5. How can you easily tell whether a substance is an iodide or not?

6. What takes place when potassium iodide is treated with sulphuric acid.

7. What analogy is there between chlorine, bromine, and iodine, as far as the compounds which they form are concerned?

8. What relation is there between the atomic weights of chlorine, bromine and iodine?

XVI.

1. Why has sulphur been known for a long time? Where is it found in nature? What is the chief source of the sulphur of commerce?

Name some of the principal compounds in which sulphur occurs in nature?

2. Describe the process by which sulphur is extracted from its ores.

3. Can sulphur and hydrogen be made to combine directly? What is formed? Where is this compound found in nature? Under what conditions is it formed?

4. How is hydrogen sulphide made in the laboratory? Explain what takes place when sulphuric acid is used; when hydrochloric acid is used. How is the substance collected? How does it behave towards water? towards metals? What takes place when it is passed over heated iron? Is there any resemblance between this action and that which takes place when steam is passed over heated iron? Express both acts by chemical equations.

5. How can hydrogen sulphide be used for the purpose of learning what things are made of. What is formed when sulphur dioxide takes up more oxygen? What is the product of the action of sulphur trioxide on water? What relation is there between sulphurous acid and sulphuric acid?

6. Where is sulphur dioxide found in nature? How is it made in the laboratory? Explain the reactions, giving the equations.

7. How are sulphites made? What is the composition of sodium sulphite? What takes place when sodium sulphite is treated with sulphuric acid? with hydrochloric acid? Compare these reactions with those which take place when sodium carbonate, Na_2CO_3, is treated with sulphuric acid and with hydrochloric acid.

8. How does sulphuric acid act towards water? What change does it produce in wood? Explain the change. How does it act upon organic substances which contain hydrogen and oxygen?

9. What is a monobasic acid? a dibasic acid?

10. Define acid, normal and neutral salts.

XVII.

1. Explain what takes place when phosphorus and iodine are brought in contact with each other.

2. Mention some of the varieties of silicic acid, and point out the relation which exists between them.

3. What important manufactured product contains silicon?

XVIII.

1. What is meant by the name base-forming elements? What name is given to the elements which are not base-forming elements? How does the number of base-forming elements compare with the number of acid-forming elements?

2. What is meant by metallic properties?

3. What are the chief classes of metal derivatives?

4. What are minerals?

5. What is meant by the term metallurgy?

XIX.

1. Give an account of the manufacture of gunpowder, and explain its use as an explosive.

2. What is sodium amalgam, and for what is it used?

3. What is water-glass, and how is it made? What is it used for?

XX.

1. Mention some of the chief compounds of calcium found in nature.

2. Explain a lime-kiln. Explain the use of lime in making the lime-light. What change takes place when lime is left exposed to the air? What change takes place in it when it is treated with water?

3. What is bleaching-powder, and how is it made? Of what value is bleaching-powder? Under what conditions does it give up its chlorine?

4. Mention some of the principal varieties of calcium carbonate which are found in nature. What are stalagmites and stalactites?

5. Explain the difference between permanent and temporary hardness of water.

6. Of what importance is calcium phosphate to plants? What objection is there to the use of normal calcium phosphate as a fertilizer? What is superphosphate of lime, how is it made, and what is it used for?

7. How is mortar made? Why do freshly-plastered rooms remain moist so long? How can the process of drying be hastened?

8. What is common glass? What is the difference between sodium glass and potassium glass? What is flint-glass, and for what is it used?

Extract from the Curriculum of the University of Toronto Junior Matriculation:

CHEMISTRY.

ELEMENTARY INORGANIC CHEMISTRY. — This examination will be limited to the chemistry of the elements—hydrogen, chlorine, bromine, iodine, fluorine, oxygen, sulphur, nitrogen, phosphorus, arsenic, carbon, boron, silicon, and their characteristic compounds, including the laws of combination of the elements, and the meaning and use of the theory of the "Molecular and Atomic Structure of Matter."

ADDENDA.

The following was received too late for insertion in its proper place:

UNIVERSITY OF TORONTO.

At the Junior Matriculation Examination, Pass and Honor papers will be set in Chemistry and Biology; and the following prescription of work replaces that on page 11 of the Arts Curriculum, dated 1885 (this enactment comes into force at the Examinations of 1888):

I.—ELEMENTARY CHEMISTRY.

Definitions of objects of the science, relations of the physical sciences to Biology, and of chemistry to physics.

Chemical change: elementary composition of matter.

Laws of combination of the elements: atomic theory, molecules, Avogadro's law.

The determination of atomic weight, specific heat, atomic heat—nomenclature, classification.

The preparation, characteristic properties, and principal compounds of the following elements: hydrogen, chlorine, bromine, iodine, oxygen, sulphur, nitrogen, phosphorus, carbon, silicon.

The Honor Examination will include the chemistry of all the elements sufficient to illustrate the classification known as Mendelejeff's law.

The following is a fuller syllabus:

§ 1. Definitions of the objects of the science: its relations to physics; and the relation of the physical sciences (chemistry and physics) to biology.

§ 4. Definitions of matter in its three forms—gaseous, liquid, and solid; a chemist confines his attention to homogenous forms of matter; importance of mass (weight) as a measure of matter.

All matter, without an exception, is subject to chemical change; such changes may be classified into three divisions. Those in which each resultant form is lighter than (1), heavier than (2), or of equal weight (3) with the initial form. By continuing the chemical changes which result in a lighter form of matter, chemists are led to a limited number of forms which can not be made to give any lighter matter. These forms of matter have distinct spectra as gases. From these, in almost all cases, the original matter may be constructed. They are therefore called the elements.

§ 3. The names of the elements. The laws of combination of the elements in *Definite Proportion, Multiple Proportion, Reciprocal Proportion.*

Dalton's theory, that the elements are composed of atoms, explains these laws. The use and meaning of the term molecule. The use of symbols to denote atoms and molecules, and the use of equations to denote chemical change.

§ 4. Dalton's theory does not admit of practical application unless we have the means of measuring the number of atoms in a molecule. Dalton assumes that he knew this number, e.g. HO for water. Chemists solve this problem by Avogadro's law that "equal volumes of gases, measured at the same temperature and pressure, contain the same number of molecules, and therefore weigh in the ratio of the weights of these molecules," deduced from the physical laws of gases, and from their relative densities as compared with their combining weights, and also from the laws of combination by volume.

§ 5. The study of the combination of the elements hydrogen and chlorine gives proof that the molecule of hydrogen contains two parts. The study of the compound hydrogen chloride convinces chemists that these parts are indivisible and therefore atoms. Hydrogen is therefore represented by the symbol H_2.

§ 6. Hence that volume of any gas will weigh its molecular weight in any system of units, which weighs two such units when filled with hydrogen gas at the same temperature and pressure.

Thus 22.327 litres 0° C. and 760 m.m. *Bar* of hydrogen weigh 2 grams, and of any other gas its molecular weight in grams.

In like manner 377 cubic feet at 60° F. and 30 inches Bar of hydrogen weigh *2 lbs.*, and therefore this volume of any gas at the same temperature and pressure weigh its molecular weight in lbs. avoirdupois.

§ 7. Chemists have agreed to take the least weight of any element found in such a molecular weight as the weight of the atom.

§ 8. The law of the specific heat of the elements may be used to dedermine atomic weight.

§ 9. Classification of the elements by their atomic weight and by the chemical character of their compounds. Outlines of Mendelejeff's classification. Allotropic modifications of the elements. Valency.

§ 10. Acid. Salts. Bases. Nomenclature.

§ 11. The law of isomorphism, its application to the determination of atomic weight.

§ 12. The conditions of chemical combination. Heat as cause and result of chemical action. Electricity. Light.

§ 13. Many of the physical properties of bodies may be traced to the properties of the individual atoms. Molecular volume of solids and liquids.

This syllabus is intended to indicate to the student the most important theories of chemistry. The following selection of the elements, with their most characteristic compounds, may be studied in illustration of the outlines of Mendelejeff's classification of the elements:

Hydrogen	Sodium Potassium	Magnesium Zinc	Calcium Strontium Barium	Boron Aluminium

Carbon Silicon Tin Lead	Nitrogen Phosphorus Arsenic Antimony Bismuth	Oxygen Sulphur	Fluorine Chlorine Bromine Iodine	Manganese Iron Gold Platinium

The student is reminded that those facts are most necessary to be remembered which are of importance to the whole subject.

INDEX.

(The numerals refer to pages.)

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1000 c.c of air at 24° will give what volum
at −12° = $\frac{1000}{1} \times \frac{261}{297}$
1000 c.c. at −12° ~~atm~~ becomes 1500 c.c at wh
temperature = $\frac{1500}{1000} \times \frac{261}{1} = 391\frac{1}{2}$

$$\begin{array}{r} 391\frac{1}{2} \\ 273 \\ \hline 118\frac{1}{2}° \end{array}$$

Oxygen unites chemically with certain of the elements with the evolution of heat & light.

Law of Chemistry is that combinatio take place by definite & fixed proport

Law of Nature is a statement of fac of nature that invariably follows a fix order of nature

Potassium Chlorate when heated gently of off $\frac{1}{3}$ of its oxygen

$2(KClO_3)$ Heated $= 2KClO_2 + O_2$

heat further & the remainder of the Oxygen given off with explosion [illegible] (deflagration)

$2KClO_2 = KCl + O_2$

Deflagration when a crystal [illegible] flies about violently, it is said to [illegible]

electrical [illegible] oxygen [illegible]

<u>Metathesis</u> substitution [illegible]

H.	1	1
O.	2	16
N	3 or 5	14
C.	4	12
P.	3 or 5	31
S.	6 or 2	32
Cl.	1	35.5
I	1	127.
Br	1	80
F	1	19
Si	~~2~~ 4	28
B	4	11
As	3 or 5	75
Al		

27, Al [illegible]

PREFACE TO SUPPLEMENT.

I am indebted to F. W. Merchant, M.A., Principal of Owen Sound Collegiate Institute, for kindness in writing the following article on Mendelejeff's law. Mr. Merchant desires me to say that the article is based on *Meyer's Theories of Modern Chemistry*, and that he lays no claim to originality in its preparation. However this may be, I think my fellow teachers will agree with me when I say, that, in these supplementary pages, a difficult subject has been presented with great clearness and brevity.

A. P. KNIGHT.

THE PERIODIC LAW.

§ 1.—**A Natural Classification of the Elements.**

The fact that the properties of the elements are intimately connected with their atomic weights was first pointed out by Newlands; but Mendelejeff, a Russian chemist, was the first to make a systematic attempt to trace the relation between the atomic weights and the chemical and physical properties of the elements, and thus indicate a basis for a natural classification of the elements.

If the elements are arranged in a series, merely in the order of their atomic weights, ranging from the lowest upwards, their properties change from member to member, but recur to a greater or less extent for certain differences in the value of the atomic weights. No analogue of hydrogen, the first member of the series, is known; but the characteristic properties of the second member, lithium (Li = 7.01), recur, after an increase of 16 units in the atomic weight, in the ninth member, sodium, (Na = 22.99), and again after a similar increase in the 15th member, potassium (K = 39.03). Roughly speaking, there is a difference of 46 units between the atomic weight of potassium and the fourth alkali metal rubidium (Rb = 85.2), and again a difference of 47 units between the atomic weight of the latter and that of caesium (Cs = 132.7).

In this series of atomic weights a metal of the alkaline earths is associated with each of the alkali metals, *e.g.*, Beryllium (Be = 9.1.) with lithium, Magnesium (Mg = 23.94), with Sodium, Calcium (Ca = 39.9), with Potassium, Strontium (Sr = 87.2) with rubidium, and Barium (Ba

= 136.9) with caesium. The atomic weights of the two groups exhibit almost the same differences.

In the same way the metals of the alkaline earths are followed by other elements which also resemble each other and exhibit similar differences in their atomic weights.

If the series is interrupted at corresponding points, that is, at elements of a similar nature, several shorter series of analogous construction are formed, which may be so arranged that the elements follow each other in the order of their atomic weights in the horizontal rows, called "series," whilst the vertical rows called "groups," are composed of members of the natural families. (See table on next page). The first member of each horizontal row is directly connected with the last member of the previous one.

The elements occurring in the groups form a natural family, but the members of the group are not all equally closely related. In most groups of eight or nine members, four or five of the elements are more intimately connected together than the remaining three or four, which however again closely resemble each other. The five alkali metals, Li, Na, K, Rb and Cs, in the first group, bear a close resemblance to one another, whilst the three heavy metals, Cu, Ag, and Au, which possess similar properties in many respects, resemble the former metals only in certain points, namely, in the isomorphism of many of their compounds, and in their power of replacing a single atom of hydrogen in forming salts.

In the same way each of the remaining groups contains two distinct sub-groups, which are, however, connected to a certain extent.

In order to bring about a regular system, some spaces must be left vacant in the table.

GROUPS.

SERIES.	I.	II.	III.	IV.	V.	VI.	VII	VIII.		
1	H I									
2	Li 7·01	Be 9·08	B 10·9	C 11.97	N 14·01	O 15·96	F 19·06			
3	Na 22·99	Mg 23·94	Al 27·04	Si 28	P 30·96	S 31·98	Cl 35·37			
4	K 39·03	Ca 39·91	Sc 43·97	Ti 48	V 51·1	Cr 52 45	Mn 54·8	Fe 55·88	Co 58·6	Ni 58·6
5	Cu 63·18	Zn 64·88	Ga 69 9	? 72	As 74·9	Se 78·87	Br 79·76			
6	Rb 85·2	Sr 87·3	? Y 89 6	Zr 90·4	Nb 93·7	Mo 95·9	? 99	Ru 103.5	Rh 104·1	Pd 106·2
7	Ag. 107·66	Cd 111·7	In 113·4	Sn 117·35	Sb 119·6	Te 126·3	I 126·54			
8	Cs. 132·7	Ba 136·86	La 138·5	Ce 141·2	Di 145	? 151	? 152	Os 195 ?	Ir 192·5	Pt 194·3
9	? 165	? 170	Yb 172·6	? 176	Ta 182	W 183·6	? 185			
10	Au 196·2	Hg 199·8	Tl 203·7	Pb 206·39	Bi 207·5	? 210	? 211			
11	? 222	? 226	? 230	? Th 231·96	? 234	? U 239·8				

A careful study of these possible arrangements of the elements has led chemists to believe that the properties of the elements are closely connected with their atomic weights, or, in the language of mathematics, are *a periodic function* of the atomic weight. We shall now illustrate this in the case of some of the more important physical and chemical properties of leading elements :

§ 2.—Relation between Density and Atomic Weight.

One of the few properties which have been accurately determined in the case of most of the elements is the density in the solid state. This property clearly shows itself to be a periodic function of the atomic weight, since it increases and decreases regularly with an increase in the atomic weight. This connection is most conspicuous if we compare the ratio between the atomic weights and the densities, in other words, if we consider the volume occupied by the atomic weight instead of the weight contained in a unit of volume. This space is known as the atomic volume.

§ 3.—Relation between Atomic Volume and Atomic Weight.

At present the atomic volume cannot be measured absolutely, but a relative measurement may be made by taking such quantities of the different elements as are proportional to their atomic weight, and comparing the space occupied by these quantities. If, as usual, the density of water is taken as the unit and the space occupied by the unit weight of water for the unit of volume, then the values of the atomic volumes are represented by the quotient of the atomic weight by the density of the given element. The atomic weight of Li = 7.01, its density is 0.59 ; the atomic volume is therefore

$$\frac{7.01}{0.59} = 11.9\,;$$

or, expressed in the metric system of weights and measures, 7.01 grammes of lithium occupy a space of 11.9 cubic centimetres.

The table on the next page gives the atomic volumes and densities of the elements arranged in groups and series. The less known metals of the earths are omitted

RELATION BETWEEN ATOMIC WEIGHT AND ATOMIC VOLUME.

	I.	II.	III.	IV.	V.	VI.	VII.	VIII.		
D V	Li 0·59 11·9	Be 1·64 5·6	B 2·68 4·0	C 3·3 3·6	N ? ?	O. ? ?	F ? ?			
D V	Na 0·97 23·7	Mg 1·74 13·8	Al 2·56 10·6	Si 2·49 11·2	P 2·3 13·5	S 2.04 15·7	Cl 1·38 25·6			
D V	K 0·86 45·4	Ca 1·57 25·4	Sc ? ?	Ti ? ?	V 5·5 9·3	Cr 6·8 7·7	Mn 8·0 6·9	Fe 7·8 7·2	Co 8·5 6·9	Ni 8·8 6·7
D V	Cu 8·8 7·2	Zn 7·15 9·1	Ga 5·96 11·7		As 5·67 13.2	Se 4·6 17·1	Br 2·97 26·9			
D V	Rb 1·52 56·1	Sr 2·5 34·9	? Y ? ?	Zn 4·15 21·7	Nb 6·27 15·0	Mo 8·6 11·1		Ru 12·26 8·4	Rh 12·1 8·6	Pd 11·5 9·2
D V	Ag 10·5 10·2	Cd 8·65 12·9	In 7·42 15·3	Sn 7·29 16·1	Sb 6·7 17·9	Te 6·25 20·2	I 4·94 25·6			
D V	Cs 1·88 70·6	Ba 3·75 36·5	La 6·2 22.3	Ce 6·7 21·0	Di 6·5 22·3					
D V			? Yb ? ?		Ta 10·8 16·9	W 19·13 9·6		Os 22·48 8·7	Ir 22·42 8·6	Pt 21·5 9·1
D V	Au 19·3 10·2	Hg 13·59 14·7	Tl 11·86 18·1	Pb 11·38 18·1	Bi 9·82 21·1					
D V				Th 11·1 20·9		U 18·69 12·8				

A remarkable periodicity will be observed in the changes which the atomic volume undergoes with increasing atomic weight.

It is easily seen that both the density and atomic volume periodically increase and decrease. The first two series contains a whole period in which the density passes through the maximum from one minimum to another. After the third series each period embraces two series, the density increases and the atomic volume diminishes in the first whilst in the following columns the reverse obtain.

It will be noted too that similar elements frequently have equal, or nearly equal atomic volumes. Thus for example, the atomic volume of Cl, Br and I, are nearly equal, V = 26, approximately; V = 7 for Mn, Fe, Co and Ni; for Ru, Rh, Pd, Os, Ir, Pt, V = 9; for Ag and Au, V = 10. In other groups V increases with the atomic weight, as for example in the phosphorus and the sulphur families. In the groups of the alkali metals the atomic volumes are in the ratio, 1 : 2 : 3 : 4 : 5 : 6.

§ 4.—Relation between Fusibility and Atomic Weight.

The fusibility of the elements is closely connected with their atomic weights and atomic volumes. The following table contained in a research by T. Carnelly, shows, as far as possible the melting points of the elements in absolute temperature.

MELTING POINTS EXPRESSED IN ABSOLUTE TEMPERATURE.

I.	II.	III.	IV.	V.	VI.	VII.	VIII.		
H 70 ?									
Li 453	Be b 1230	B v h	C n f	N v l	O v l	F v l			
Na 369	Mg 1023	Al 1123	Si v h	P r 528 c 317	S 388	Cl 198			
K 335	Ca + Sr	Se ?	Ti n f	V n f	Cr b 2270	Mn 2170	Fe 2080	Co 2070	Ni 1870
Cu 1330	Zn 690	Ga 303		As a 773	Se 490	Br 266			
Rb 311	Sr + Ba	Y ?	Zr + Si	Nb n f	Mo v h		Ru 2070	Rh 2270	Pd 1775
Ag 1230	Cd 590	In 449	Sn 503	Sb 710	Te 728	I 387			
Cs 300	Ba 748	La a 710	Ce and b	Di 1273					
		Yb ?		Ta n f	W v h		Os 2770	Ir 2220	Pt 2050
Au 1310	Hg 234	Tl 563	Pb 599	Bi 540					
			Th		U F e				

NOTE.—In the above table n f = not fused; v h = very high; v l = very low; a = above; b = below; + higher than; — lower than; and in the case of phosphorus, r, = red, and c = colorless.

From this it will be seen that the fusibility of the elements considered as a function of their atomic weights exhibits a periodicity corresponding to that shown by the atomic volume.

By comparing this table with those on pages 287 and 289 the following relations are observed:

(1) Only those elements are easily fusible the atomic volume of which is larger than that of the element with the next smaller atomic weight. Those elements are difficultly fusible for which the reverse is the case.

(2) Either a very easily fusible or an infusible element occurs between every two groups of easily and difficultly fusible elements. This connecting link is wanting only in the case of carbon and nitrogen and between silicon and phosphorus.

(3) In most of the groups, the fusibility decreases with increasing atomic weight, but in the following groups it increases:—(a) the alkali metals; (b) apparently in the case of the alkaline earths. In the nitrogen bismuth group the fusibility decreases from nitrogen to arsenic, and increases from arsenic to bismuth. In the noble metals it increases from Cu to Ag, and again decreases from the latter to Au.

§ 5.—Relation between Volatility and Atomic Weight.

Volatility is intimately associated with fusibility, and hence with atomic volume and atomic weight.

The following table shows the absolute temperature of the boiling points of the elements under a pressure of one atmosphere:

TABLE OF BOILING POINTS.

Element .	H	Zn	Cd	Hg	Sn	Pb	N	P	As
Boiling Pt	b 70?	1200	1045	630	1870—2070	1870—2070	b 70?	563	700?
Element .	Sb	Bi	O	S	Se	Te	F	Cl	Br
Boiling Pt	1360—1870	1360—1870	b 70?	720	938	b 1600	b 100?	240	331
Element .	I	Na	K						
Boiling Pt	487	1130—1230	990—1000						

These imperfect data show the variations between the boiling points to resemble those exhibited by the melting points. In most groups the boiling point appears to increase with the atomic weight, but with the alkali metals and the zinc group the reverse is the case.

The preceding relations among atomic weight, atomic volume, fusibility and volatility, may be expressed by the following proposition :

Every element is easily fusible and volatile which possesses a greater atomic volume than the immediately preceding element with a lower atomic weight : its molecules can be easily separated. Inversely, each element is with difficulty fusible or volatile the atomic volume of which is smaller, or not larger than that of the preceding element, having a smaller atomic weight. All these elements melt and volatilize easily the atomic volumes of which would be diminished, if it were possible, by reducing their atomic weights to convert them into the element with the next lower atomic weight. On the other hand those elements fuse and volatilize with difficulty the atomic volumes of which would be increased, if it were possible, by reducing their atomic weights to convert them into the element immediately preceding.

§ 6.—Relation between Ductility, Internal Structure, etc., and Atomic Weight.

The malleability, ductility and hardness of the elements are intimately related to their atomic volumes and atomic weights. According to Boltone the hardness of elements is inversely proportional to their atomic volume; but his statement must be received with caution, since he found that his rule held good for indium, although he made use of the wrong atomic weight, namely 74, in his calculations. He also wrongly assumes that calcium is softer than sodium, etc.

Ductility, fusibility and volatility of the elements are closely connected with their internal structure, especially with their crystaline form and expansion by heat. These are therefore also periodic functions of the atomic weight.

The refraction of light by the elements is also essentially related to the atomic weight, but, as in the other cases, little is known of the causal connection.

The conductivity for heat and electricity of the elements is, as is well known, dependent on their ductility and malleability, and hence, like these, is a periodic function of the atomic weight, the periodicity of which coincides with that of the atomic volume.

The magnetic and diamagnetic properties of the elements appear also to be closely connected with their atomic weights.

The specific heat represents a function of the atomic weight; but it is not a periodic function, since according to the law of Dulong and Petit, the specific heats are inversely proportional to the atomic weights.

§ 7.—Relation between Chemical Properties and Atomic Weight.

Not only are the physical properties of the elements apparently dependent on their atomic weights, but the atomic weight has a distinct influence on their chemical properties. For example, valency is a periodic function of the atomic weight. It changes regularly from each to the next following family. Since the corresponding members of these families follow one another in the series of atomic weights, it follows that in this series also the valency changes regularly from member to member. Starting from the maximum of atomic volume we have the following example of this regularity:

Monovalent.	Divalent.	Trivalent	Tetravalent.	(Towards Positive Elements.)		
				Trivalent.	Divalent.	Monovalent.
Li $LiCl$	Be $BeCl_2$	B BCl_3	C CH_4	N NH_3	O OH_2	F FH

This same regularity is repeated in the compounds of these elements which follow fluorine:

Monovalent.	Divalent.	Trivalent.	Tetravalent.	(Towards Positive Elements.)		
				Trivalent.	Divalent.	Monovalent.
Na $NaCl$	Mg $MgCl_2$	Al $AlCl_3$	Si SiH_4	P PH_3	S SH_2	Cl ClH

Regularities of a similar character are found when, instead of the compounds with the positive element hydrogen, those with the negative element chlorine are considered. Thus, for the elements comprising the seventh series in the table page 285 we have the following:

$AgCl$	$CdCl_2$	$InCl_3$	$SnCl_4$	$SbCl_3$ $SbCl_5$	$TeCl_2$ $TeCl_4$	ICl ICl_3

The periodicity of the valency is shown most clearly when it is deduced from the compounds with oxygen. The greatest regularity is to be found in the composition of the oxides, which, with few exceptions change in a regular manner with increase in atomic weight. As a general rule it is found that the amount of oxygen combined with an atom of another element increases by half an atom from member to member of a series, but this increase never goes beyond four atoms, when it sinks again to half an atom. This relationship is best exhibited when, in order to avoid fractions of atoms, the formulæ are so written as to show the correspondence in composition without regard to the molecular weight. This is illustrated in the following table. The oxides inclosed in brackets are not known with certainty, or not in a state of purity, but from the analogies to nearly related compounds, there can be but little doubt of their existence:

OXIDES.

(Li_2O)	Na_2O	K_2O	Cu_2O	(Rb_2O)	Ag_2O	Cs_2O	....	Au_2O
Be_2O_2	Mg_2O_2	Ca_2O_2	Zn_2O_2	Sr_2O_2	Cd_2O_2	Ba_2O_2	Ng_2O_2	Hg_2O_2
B_2O_3	Al_2O_3	Se_2O_3	Ga_2O_3	Y_2O_3	In_2O_3	Ce_2O_3	Yb_2O_3	Tl_2O_3
C_2O_4	Si_2O_4	Ti_2O_4	....	Zr_2O_4	Sn_2O_4	....	....	Pb_2O_4
N_2O_5	P_2O_5	V_2O_5	As_2O_6	Nb_2O_5	Sb_2O_5	... .	Ta_2O_5	Bi_2O_6
O_2O_4	S_2O_6	Cr_2O_6	(Se_2O_6)	Mo_2O_6	Te_2O_6	....	W_2O_6	....
F ?	(Cl_2O_7)	Mn_2O_7	(Br_2O_7)	....	(I_2O_7)	....	....	....
....	....	$(Fe_2O_6$	....	Ru_2O_8	....	....	Os_2O_8	....
....	....	Co_2O_3	....	Rh_2O_4	....	....	Ir_2O_4	....
....	....	Ni_2O_3	....	Pd_2O_4	....	....	Pt_2O_4	....

It cannot well be doubted that the composition of the oxides not introduced in the table is dependent in a more or less simple manner on the atomic weight.

The composition of the oxides stands in a very close relationship to that of other compounds, for example, with that of the hydroxides and salts. The composition of the hydroxides, whether they are of a basic or acid nature, is, on the one hand, dependent upon that of the oxides, and, on the other, upon that of the hydrides, or hydrogen compounds. Instead of hydrides, the compounds with a monovalent hydro-carbon radicals *e.g.* those with ethyl C_2H_5 or Et, may be used as indicators, since the number of these radicals is throughout equal to that of hydrogen.

The following table exhibits the relationship among these compounds :

Oxide.	Hydroxide.	Hydride.	Ethide.
Na_2O	$NaOH$	....	$NaEt$
Mg_2O_2	$Mg(OH)_2$	....	$MgEt_2$
Al_2O_3	$Al_2(OH)_3$	....	$AlEt_3$
Si_2O_4	$Si(OH)_4$	SeH_4	$SiEt_4$
P_2O_5	$PO(OH)_3$	PH_3	PEt_3
S_2O_6	$SO_2(OH)_2$	SH_2	SEt_2
Cl_2O_7	$ClO_3(OH)$	ClH	$ClEt$
K_2O	$K(OH)$	....	KEt

Based on these regularities in the composition of chemical compounds, which might be illustrated by many other similar tables, the statement may be made that the valency of an element, as deduced from the composition of its compounds, is a periodic function of the atomic weight. The general chemical characters of the elements are, in like manner, seen to be periodic functions of the atomic weight, and the periods to correspond very nearly with those of the valency.

Hence in general:

§ .8—Periodic Law.

The chemical and physical properties of the elements are periodic functions of the atomic weight.

§ 9.—Uses of the Periodic Law.

The Periodic Law may be employed for the correction of doubtful atomic weights. For example, the atomic weight of molybdenum was, according to some investigators, found to be 96, and by others to be 92 (in round numbers). There was but little probability in favor of the correctness of this latter determination, for then molybdenum must be placed in an entirely false position in the system of atomic weights, viz.: before niobium. The more recent experiments of Liechti and Kempe have, in fact, shown the error of this last atomic weight determination.

The periodic law has been used to predict the probable existence and properties of undiscovered elements. The possibility of doing so is based on the conception that the properties of the elements in the series, and also in the groups, vary in a more or less continuous manner, although to different extents. The properties of an element are, as a rule, the means of those of its neighbors in the groups on the one hand, and of the series on the other. Mendelejeff, styles these four elements so related, "atomic analogues." The atomic analogues of Se are on the one hand, S and Te; and As and Br on the other. The relations among their densities, atomic weights and atomic volumes are exhibited in the following table, and they exhibit similar relations in their other properties.

Density.

	S 2·04	
As 5·67	Se 4 6	Br 2·97
	Te 6·25	

Atomic Weight.

	S 31·98	
As 74·9	Se 78·37	Br 79·76
	Te 126·3	

Atomic Volume.

	S 15·7	
As 13·2	Se 17·2	Br 26·9
	Te 20·2	

By similar considerations Mendelejeff was lead to predict the properties of the element situated between baron and yttrium, to which he gave the name 'Eka-boron,' and also those of the element between aluminium and indium, which he styled 'Eka-aluminium.' The properties of the first are found to correspond with those of the element scandium, since discovered by Nilson, while the second corresponds in most of its properties with the element gallium discovered later on by Lecoq de Boisbaudran. Again the element, discovered by Winkler early in 1886 appears to be identical with Mendelejeff's Eka-silicon and may fill up the blank space between titanium and zirconium. The existence of spaces in the table is to be explained by the supposition that these spaces indicate the position of elements as yet undiscovered.

The periodic law is useful also in the comparative study of the properties of elements already known.

§ 10.—Conclusion.

In all these speculations we must not lose sight of the fact that the general law upon which the relations between the properties of an element and its atomic weight depend, is still unknown. There are many of the groups of elements formed by the systematic arrangements of the atomic weights, the properties of the members of which would scarcely warrant their classification as a natural family, had not the relations of the numbers expressing their atomic weights indicated the necessity for such a classification. The uncertainty of our knowledge is also shown by the different results obtained when attempts have been made to classify the elements by means of any other properties save the fixed numerical values of their atomic weights.

There can no longer be any doubt that the system of the elements based upon the value of their atomic weights will form the basis of future doctrines of comparative affinity. We must advance with the aid of induction, and with special care. Hence the tendency to generalize must be kept in check as much as possible, and we must still depend upon experimental observations for our advance in this unexplored province. Hypothesis may certainly be used to aid our progress, but only when these are carefully distinguished from the results of observation. There is still much work for the hand and the mind; but it will be thoroughly rewarded. The prize is a systematic inorganic chemistry which will not fear comparison with the thoroughly developed system of organic chemistry.

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