LABORATORY INSTRUCTIONS

IN

GENERAL CHEMISTRY

ARRANGED BY

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WITH FIFTY-SIX ILLUSTRATIONS.

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PREFACE.

THIS little book consists of "Laboratory Instructions" to illustrate a course of study in General Chemistry.

Much of the material is original, having been developed in the course of ten years' experience in laboratory teaching, while those portions taken from other sources have been modified and added to, so that they might better meet the wants of students.

Particular attention is called to the Appendix, which contains demonstrations of a quantitative character, as from them the principles underlying the Science of Chemistry logically follow by induction. They also illustrate and exemplify the theories and laws on which it rests, and by virtue of which it remains to-day as one of the exact sciences.

These "Laboratory Instructions" can be used in connection with any standard text book, for example those by Richter, Newth, or Remsen.

I desire to express my thanks to President James MacAlister, LL. D., of the Drexel Institute; to Dr. John Marshall, Dean of the Medical Faculty of the University of Pennsylvania; and to Professor Albert P. Brubaker, M. D., of Jefferson Medical College, Philadelphia, whose favorable opinion of my MS. induced me to publish it.

Most of the illustrations in the book are taken from the laboratory notes of one of my students, Mr. D. Orren Hales, to whom I am indebted for the use of his drawings. The rest are taken from sketches of my own and from other sources.

I take pleasure in expressing my obligation to Mr. A. Henwood, Instructor in Chemistry in the Drexel Institute, who has given me much valuable aid and assistance in the preparation of the manual.

E. A. C.

PHILADELPHIA, *December*, 1900.

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LABORATORY DIRECTIONS.

1. Be cautious in the use of acids and alkalies. If spilled on person or clothing, neutralize at once; acids, with ammonia (dilute) or soda; alkalies, with weak acetic acid.

2. For burns, whether caused by fire or acids, use sodium bicarbonate, and afterwards dress with vaseline or an oil emulsion.

3. Be careful not to hold the hand under glass vessels, the contents of which are being heated. Always hold the Bunsen burner in such a way that the hand is at one side.

4. In heating test-tubes, always incline the tubes, and never hold them in an upright position. Always use a holder, and note the direction in which the tube is pointing.

5. Be careful in forcing glass tubes through corks and stoppers. Force the moistened tube gently, using a rotatory motion.

6. Never pour water into concentrated sulphuric acid. Pour the acid, which should be cold, gradually into the water. Mix in a thin glass vessel. When cold, transfer to the reagent bottle.

7. In case of fire, use the sand provided in the pails in the laboratory, or smother with a towel.

8. Never put away in the desk hot charcoal or metal apparatus which is hot. Cool first in water.

9. Keep the desk and apparatus scrupulously clean and neat.

10. Never throw anything solid into the sinks. Use the proper receptacles provided for this purpose.

11. Liquids can be emptied into the sink, taking the precaution to have plenty of water running from the spigot.

12. Keep neat and careful notes of all work done in the laboratory. Record all data and details of the work. Thoroughly understand each piece of work as it is done. Quality is more important than quantity.

EXPERIMENTS IN GENERAL CHEMISTRY.

Preliminary Work.

A.—Bending Glass Tubing.

Break the glass tubing into several pieces of the required lengths.

To break the glass tube to a desired length, make a scratch with the moistened edge of a triangular file, and holding firmly with both hands near the scratch, break by a quick downward movement.

Round the sharp edges of the ends by holding the tube in the blue flame of a Bunsen burner until soft and the inequalities are melted down.

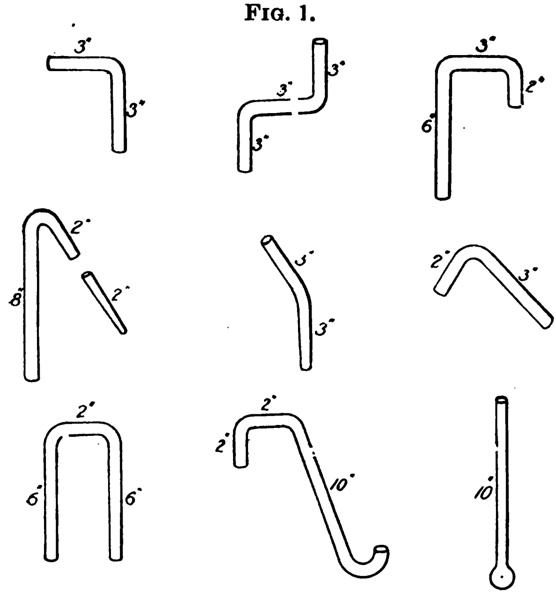
Place burner tube, giving a broad flame, on the Bunsen burner. To bend a tube, hold in the yellow flame and parallel to it; rotate constantly until soft. By a bending or pulling movement any desired angle may be given.

A tube that has been properly bent will have the same diameter of bore throughout, and will show no creases where it has been bent.

Bend several tubes of the shapes and lengths shown in the sketch. (See Fig. 1.)

To make a bulb on the end of a tube, heat the end until quite hot and until quite a lump of soft glass has collected; then by gently blowing from the open end of the tube, form a bulb of the desired size (*e. g.*, about one-half inch in diameter).

FIG. 1.



B.—Preparing Glass Rods.

Break the glass rods into pieces of six inches in length.

To break the rod, make a scratch with the edge of a triangular file and break as explained above.

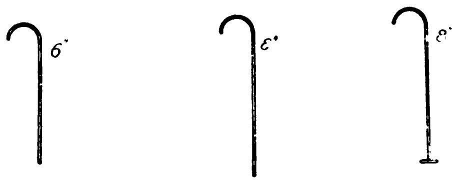
Round sharp edges in the blue flame of the Bunsen burner.

Bend the rods by heating in the blue flame.

To flatten the end of a rod, heat very hot until the glass is quite soft, and flatten by pressing the hot end on a cold surface, such as the soap-stone top of desk.

Bend several rods of the form shown in sketch. (See Fig. 2.)

FIG. 2.



(For manipulation of glass, read Newth, "Elementary

Practical Chemistry," Chap. IV, V. Also see: Shenstone, "Art of Glass Blowing.")

C.—Distillation.

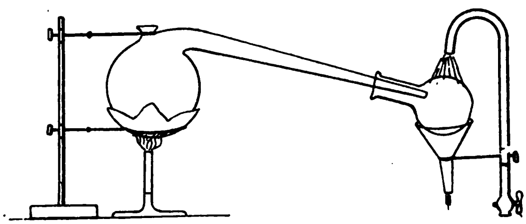
Weigh out two lots of 10 grams each of common salt on the horn scales.

Place one of these in an evaporating dish which has been accurately weighed. Heat this gently over a Bunsen burner to constant weight. Record all weights. Note the amount of hygroscopic moisture contained in the salt.

Introduce the second lot into a tubulated glass retort of 500 c.c. capacity. Dissolve in about 200 c.c. of distilled water.

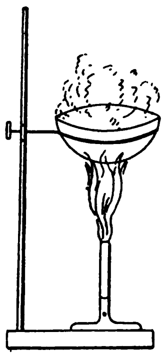
Connect the retort with a flask cooled with water from the

FIG. 3.



spigot (See Fig. 3). Distil down to small bulk. Save the distilled water obtained for future use. Pour what remains in the retort into a weighed evaporating dish; rinse the retort with distilled water, and add the rinsings to the contents of evaporating dish. (See Fig. 4.)

FIG. 4.



Evaporate to dryness. Ignite gently, being careful to avoid spattering, cool and weigh. Note the amount obtained. Compare with the amount of salt present in the sample first heated.

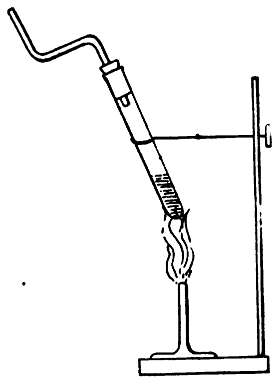
Note how nearly a correct figure has been obtained.

Explain the object of this experiment and what has been learned from it.

Physical Change.

Experiment 1.—Boil some water in a test-tube fitted with perforated cork and bent tube as shown in sketch. (See Fig. 5.) Explain the changes that take place and note the condensation of the steam in the bent angle of the tube. In what way has the water been changed?

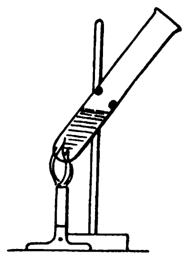
FIG. 5.



Experiment 2.—Rub a stick of roll sulphur briskly on the sleeve; approach to small bits of filter paper. Note the power it has now acquired over the paper. Explain. Has the sulphur itself been changed in any way?

Experiment 3.—Place a few lumps of roll sulphur in a test-tube. Warm very gently. Increase the heat gradually and heat strongly for some minutes. Note the various changes through which the sulphur passes. (See Fig. 6.) Remove the heat, and as the contents of the tube cool, note the changes that take place.

FIG. 6.



Experiment 4.—Heat a piece of platinum foil until red hot. Allow it to cool. Does it return to its original state? Has it in any way changed?

Experiment 5.—Dissolve a small lump of camphor in alcohol in a test-tube; add water; explain the result. Has the camphor become changed in substance?

Experiment 6.—Place a small grain of iodine in a piece of bent glass tubing sealed at one end. Seal the other end with the blue flame of a Bunsen burner, and after it has become cool, heat the end containing the iodine *very gently*. What

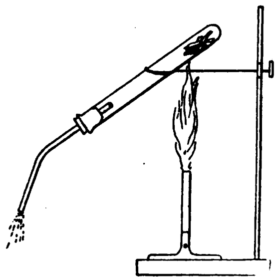
change has taken place in the iodine? In like manner, in a similar tube, *very gently* heat a small bit of camphor. Is a similar result obtained?

Note.—Describe, in general terms, what you understand by "Physical Change."

Chemical Change.

Experiment 7.—Place some bits of wood (match ends) in a test-tube fitted with a perforated cork and bent tube as shown in sketch. (See Fig. 7.) Heat gently in the flame of a Bunsen burner; light the gas that issues from the end of the tube. How does this experiment differ from Experiment 1?

FIG. 7.



Experiment 8.—Place a little granulated sugar in a test-tube. Heat very gently over a Bunsen burner, and finally strongly until no further action takes place. Note all changes. Apply a lighted match to the vapors that escape from the end of the tube and see if they are inflammable. Compare this experiment with the heating of the sulphur (Experiment 3). How are they alike and how do they differ?

Experiment 9.—Heat a small piece of the metal magnesium in the flame of a Bunsen burner. How does this result differ from the one obtained by heating the metal platinum? (See Experiment 4.)

Experiment 10.—In an evaporating dish half full of water, drop a small lump of gum-camphor; light with a match and observe the result. How does this experiment differ from Experiment 5?

Experiment 11.—Place some sugar in a small beaker and pour on it a few drops of concentrated sulphuric acid. What change takes place in the sugar?

Experiment 12.—Place a small filter paper in an evaporating dish and pour on it dilute sulphuric acid (made by adding one part concentrated acid to two parts of water); boil over a Bunsen burner until the solution is reduced to a very small bulk. What change has taken place in the paper? In what respect is this experiment similar to the last?

Experiment 13.—Mix about two grams of powdered sugar and an equal quantity of potassium chlorate on a sheet of glazed paper; pour on an iron sand-bath and let fall on the mixed powders a few drops of concentrated sulphuric acid from a long glass tube. (*Do this under the hood. Danger!*) Has a change of substance of the mixed bodies taken place?

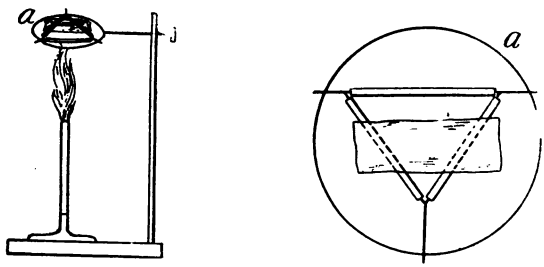
Note.—Describe, in general terms, what you understand by "Chemical Change."

Physical Forces Induce Chemical Change.

(a) Action of Heat.

Experiment 14.—Arrange a piece of platinum foil on a pipe-stem triangle supported on a ring stand as shown in sketch. (See Fig. 8.)

FIG. 8.



Place a crystal of "blue vitriol" (copper sulphate) on the foil and heat strongly with a Bunsen burner. Note the changes that take place.

Experiment 15.—Place a few crystals of microcosmic salt in a test-tube, heat gently, and note all the changes that take place. Note the odor given off and the formation of liquid in the cool part of the tube. Hold a piece of moistened red

litmus paper in the mouth of the tube while the contents are being heated, and note the effect. What does this show?

Experiment 16.—Place some sugar in a test-tube and heat over a Bunsen burner. How is the sugar changed? How does the residue differ from the sugar first taken?

Experiment 17.—Weigh on the horn scales:

Potassium nitrate	. . . . . .	3 grams.
" carbonate	. . . . .	2 "
Sulphur (powdered)	. . . .	1 "

Pulverize each separately before weighing, and after weighing, thoroughly mix on a sheet of glazed paper with the blending knife. Place a little of the mixture on a tin plate and heat over the Bunsen burner.

Caution.—Be careful not to have the face near, as an explosion takes place.

Query.—What experiments made under the heading "Chemical Change" showed the effect of *heat* in bringing about chemical changes?

(*b*) *Action of Light.*

Experiment 18.—Pour a solution of silver nitrate into a clean test-tube. Dissolve a little common salt in a test-tube, and pour into the silver solution. Note the immediate formation of a curdy, white precipitate; this is silver chloride. Allow it to stand exposed to bright sunlight for some time, and note the change in color that takes place.

Experiment 19.—Dip a piece of filter paper in a solution of silver nitrate. Then dip the paper into a solution of common salt in water, and place it in a dark place to dry—(inside the desk). When dry, or nearly so, place between two glasses and expose to bright light for some time, with part of the paper covered with letters cut from black glazed paper. Then remove the letters and note the results produced.

Experiment 20.—Gently warm a glass cylinder, pour into it a few drops of bromine water, and add two or three cubic centimetres of benzol. Stand in bright sunlight for a time and note the result.

Experiment 21.—Dissolve a clean crystal of ferrous sulphate in distilled water, and place in two test-tubes. To the one add a solution of potassium hydroxide (KOH), and to the other a solution of potassium ferro-cyanide ($K_4FeC_6N_6$). Stand in bright sunlight for some time and note the changes in color that take place.

What principles have been learned from the above experiments?

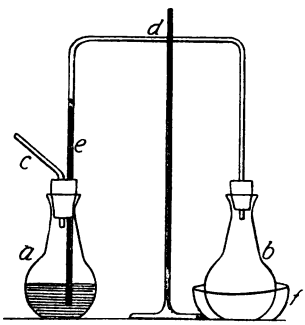
Chemical Changes Develop Physical Forces.

(a) Heat is Produced.

Fill two thermometers with a colored liquid as explained in lecture notes.

Experiment 22.—Pour into a beaker, or porcelain dish, a strong solution of caustic soda (sodium hydroxide). Place one of the thermometers in the beaker and support by a clamp. (See Fig. 9 A.) Add concentrated hydrochloric acid little by little, stirring the mixture with a glass rod. Add the acid until the solution has an acid reaction. How can this be ascertained? Note the heat developed. Use an air thermometer. Explain what has taken place and why heat was produced. Evaporate to dryness. What is the powder obtained?

FIG. 9 A.

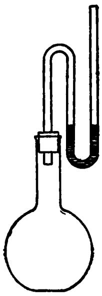


Caution.—Be very careful in using alkalies and acids, as they are extremely corrosive.

Experiment 23.—Place some pieces of quick-lime (calcium

oxide) in a beaker and about half cover them with water. Place a thermometer in the beaker. (See Fig. 9 B.) Note the increase in heat. Explain what has taken place and why heat was produced.

FIG. 9 B.



Experiment 24.—Place a few cubic centimeters of dilute sulphuric acid in a test-tube, drop into it a piece of granulated zinc. Note the chemical action that takes place and the heat that is developed.

Experiment 25.—In like manner drop a piece of copper into dilute nitric acid and note the results obtained and the production of heat.

(*b*) *Light is Produced.*

Experiment 26.—Hold a piece of magnesium in the flame of a Bunsen burner. Note the result.

Experiment 27.—Prepare the following colored fires:—

No. 1.—RED.

Strontium nitrate	4.	grams.
Sulphur	1.5	"
Potassium chlorate	2.5	"
Shellac	3.	"

No. 2.—GREEN.

Barium nitrate	3.2	grams.
Sulphur	.8	"
Potassium chlorate	3.2	"
Shellac	3.	"

Pulverize each ingredient separately and mix on a sheet of glazed paper with the blending knife. Place a little of each mixture in a tin dish and light. (*Do this under the hood.*)

Experiment 28.—Place a small piece of phosphorus, which has been dried by pressing for a moment between filter papers, in a clean iron dish. Drop into it by means of the forceps a small fragment of iodine. Note the chemical action and the production of light.

What principles have been learned from these experiments?

Chemical Reactions.

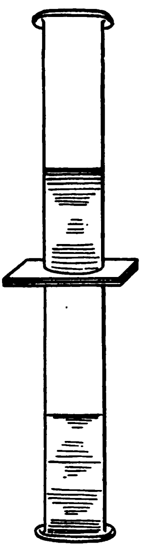
(1)—Synthesis.

Experiment 29.—Burn a small piece of magnesium ribbon by holding for an instant in the flame of a Bunsen burner. Explain the chemical action that takes place. Write the reaction.

Experiment 30.—Mix together some sulphur and some copper turnings. Heat in a test-tube and observe the changes that take place. What is the compound formed? Write the reaction.

Experiment 31.—Rub together in a porcelain mortar some sulphur and a little mercury. Note the change in color. What compound is formed? Write the reaction.

Fig. 10.



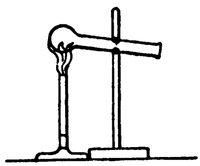
Experiment 32.—Boil a little ammonium hydroxide in a test-tube fitted with a perforated cork and bent tube, and pass the gas given off into a tall jar. In like manner in another tube, similarly fitted, boil a little hydrochloric acid and pass the gas evolved into another tall jar. Cover each jar with a piece of ground glass, and invert one over the other. Remove the covers of ground glass, and note the effect of the two gases mixing. (See Fig. 10.) What compound is formed by their union? Write all the reactions that have taken place during this experiment.

(2)—Analysis.

Experiment 33.—Arrange an apparatus as shown. (See Fig. 11.) Heat in the bulb tube of hard glass some red oxide of mercury (mercuric oxide), using the strong

blue flame of a Bunsen burner. Note the changes that take place. Examine the gas that is given off. See its effect on a lighted match. Note the silvery substance that drops into the jar on prolonged heating. What are the two substances formed, viz., the gas and the liquid silvery substance? Write the reaction that has taken place.

FIG. 11.



Experiment 34.—State how the action of the electric current on water is an example of analysis. Give a sketch of the apparatus. What are the relative amounts of the gases formed? What are the names given to the two gases that, chemically combined, form water and have been produced by its decomposition?

(3)—METATHESIS.

Experiment 35.—To a solution of ferric chloride, add ammonium hydroxide until alkaline. How may this be ascertained? Filter, examine the precipitate. Save the filtrate for examination.

Heat some of the precipitate on a piece of platinum foil. Dissolve some in hydrochloric acid and add ammonium sulpho-cyanide. What is the effect? What does this show?

Pour the filtrate into an evaporating dish, evaporate to dryness, and examine the powder that remains.

What was the composition of the precipitate?

What was in the solution?

Note.—By the filtrate is meant the liquid which passes through the paper when the precipitate is removed by filtration.

Write the reaction that has taken place.

Experiment 36—To a boiling solution of barium chloride, add dilute sulphuric acid until the solution is slightly acid, filter, and examine the precipitate.

Test its solubility in various solvents. Does it dissolve?

Heat a little on platinum foil. Does any change take place?

Examine the filtrate as follows:

Evaporate a portion to dryness in an evaporating dish, and hold a piece of blue litmus paper in the vapors given off. Note whether a residue remains. What is the effect on the paper? What does this show?

To another portion of the filtrate, add a few drops of a solution of silver nitrate. Compare the results with those obtained when a little of the sulphuric acid originally used is added to silver nitrate.

In a small tube add a few drops of hydrochloric acid to silver nitrate and compare the result.

What have you now ascertained the action of sulphuric acid on barium chloride to be? Write the reaction.

Experiment 37.—To a solution of copper sulphate ($CuSO_4$), add some caustic soda—sodium hydroxide ($NaOH$)—until a precipitate no longer forms. Filter.

Note its color, appearance, etc.

What is the composition of this precipitate?

What is contained in the filtrate?

Write the reaction that has taken place.

Mixtures.

Experiment 38.—Take 5 grams of powdered sulphur and an equal amount of finely divided iron; mix thoroughly with a steel spatula on a sheet of glazed paper. Place a little of the mixture on a watch-glass and examine with a hand magnifying glass.

Can you distinguish between the sulphur and iron? Draw a magnet over the mixture, and see if the iron can be separated by this means.

Experiment 39.—Pour some of the mixture into a beaker half full of water and stir with a glass rod. Can the constituents of the mixture be separated by their different gravities?

Experiment 40.—Take separate test-tubes, place a little iron in the one, a little sulphur in the other; attempt to dissolve each by pouring in a little carbon disulphide. Which one is soluble in this liquid? Place some of the mixture made in

Experiment 38 in a small beaker. Pour on it some carbon disulphide; stir with a glass rod, filter, and examine the residue to see if it is iron. (See if it is attracted by the magnet.) Evaporate the solution on the steam-bath, and see if sulphur is obtained.

Does it differ in any way from the sulphur first mixed with the iron?

Caution.—Carbon disulphide is an exceedingly inflammable liquid, and its vapor forms with air an explosive mixture. See that no flame is in the vicinity when it is being used. Use great care.

Experiment 41.—Make a mixture of chalk and pulverized sugar; heat in a beaker with water; filter; taste the solution; note the appearance of the residue.

Have the two substances been separated by this means?

Experiment 42.—Examine a small piece of "granite." Explain how this is an example of a "Mechanical Mixture."

Compounds.

Experiment 43.—Mix together on a sheet of glazed paper some iron filings and powdered sulphur, or use some of the mixture already made in Experiment 38. Place a little of the mixture in a small test-tube, heat over a Bunsen burner, and note what takes place.

When the tube is cool, place in a porcelain mortar, break, and examine the substance formed.

Attempt to separate the constituents in the manner performed in Exps. 38, 39, and 40.

Compare the results of these experiments with the experiments under the heading "Mechanical Mixture."

Can the iron and sulphur be separated in the same way that mixtures were separated?

Experiment 44.—Place about 50 c. c. of a solution of caustic soda (sodium hydroxide) in a beaker. Add dilute hydrochloric acid, stirring with a glass rod until the solution is acid. How can this be ascertained?

Pour the solution into an evaporating dish and evaporate to dryness over a Bunsen burner. Examine the powder obtained.

Has it the properties of the caustic soda or the hydrochloric acid? Explain what has taken place.

Experiment 45.—Dissolve a few scraps of copper in a test-tube with dilute nitric acid. Evaporate the solution to dryness in a porcelain dish. Is the copper again obtained?

Explain what has taken place.

Experiment 46.—Examine a crystal of calc spar. Prove experimentally that this is an example of a chemical compound.

Explain the main differences between a piece of granite and a piece of calc spar as illustrating a *Mixture* and a *Compound.*

Oxygen (O_2).

Experiment 47.—Introduce a little mercuric oxide ("red precipitate") into a tube of hard glass and heat strongly with a Bunsen burner. (See Fig. 11.)

Note the change that takes place.

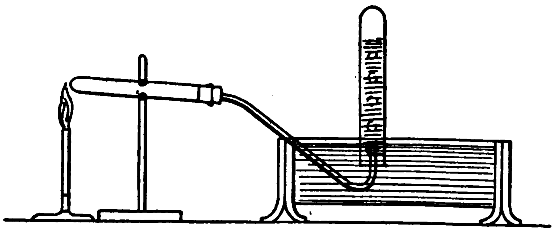
Introduce a glowing splinter of wood into the tube and note the effect. What gas has been given off?

See whether mercury can be distilled from the tube.

Write the reaction that has taken place.

Experiment 48.—Place about 5 grams of potassium chlorate in a test-tube fitted with a cork and bent delivery tube, heat

FIG. 12.



carefully with a Bunsen burner. (See Fig. 12.) Apply a

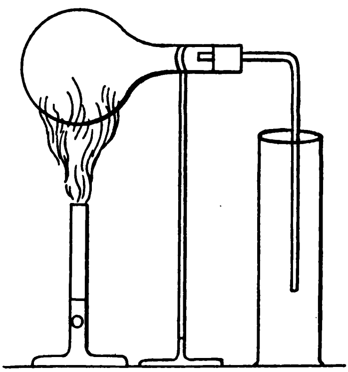
splinter of glowing wood at the delivery tube. Note that the substance after liquefying becomes solid again; heat more strongly and test the delivery tube again with a glowing splinter.

Collect some of the gas in a large test-tube over water. Use a pneumatic trough. Test the gas by introducing a lighted match. Examine the residue left in the tube, dissolve it in water, filter, and add silver nitrate. What does this reagent show?

Write all the reactions that have taken place.

Experiment 49.—Mix about 10 grams of potassium chlorate with an equal amount of manganese dioxide on a sheet of glazed paper. Introduce into a flask of about 250 c.c. capacity, fitted with a perforated cork and delivery tube, and heat gently with a Bunsen burner. . (See Fig. 13.) Note whether the gas is evolved more easily from this mixture than from the potassium chlorate alone. Fill several jars with the gas by downward displacement of the air.

FIG. 13.



Experiment 50.—Place a little barium dioxide (BaO_2) in a tube of hard glass and heat gently with a Bunsen burner. Test the gas given off with a glowing match end.

Experiment 51.—Place in a large test-tube a little sodium dioxide (Na_2O_2). Add a few drops of water, and test the gas given off as before. What remains in the test-tube? Write the reaction that has taken place.

Experiment 52.—Place a few grams of "nitre"—potassium nitrate—(KNO_3) in a porcelain crucible and heat over the Bunsen burner until it melts. Introduce a glowing match end into the molten mass and observe the effect. Explain what takes place.

Study the properties of Oxygen gas as follows:—

Experiment 53.—Hold slips of red and blue litmus paper in one of the jars of the gas and note whether there is any change in color. Has the gas an odor? Is it lighter or heavier than air?

Experiment 54.—Place a small piece of sulphur in a deflagrating spoon, light, and introduce into a jar of the gas. Note the result. Smell the contents of the jar. (*Caution.*) Introduce moistened slips of litmus paper.

Write the reaction that has taken place.

Experiment 55.—Burn a piece of wood in a jar of the gas. After the action is over, introduce a lighted match and note the result.

Pour into the jar a little lime-water; shake around and note change.

Pour the liquid into a small test-tube and add a few drops of hydrochloric acid. Note the effect.

Explain all these changes and write the reactions that have taken place.

Experiment 56.—Fasten a bit of candle into a deflagrating spoon, light, and introduce into a jar of gas. Note the result.

Does lime water now give similar results to those obtained in the last experiment?

Leave the candle in the jar for several minutes and note the change that takes place. Finally pour in a little lime-water.

See if a reaction takes place, similar to that in the last experiment. Explain. Write the reactions.

Experiment 57.—Cut off a small piece of phosphorus under water, (*Remember the caution given in regard to this substance.*) Dry for an instant between two pieces of filter paper and drop into a deflagrating spoon. Ignite by touching with a hot wire and at once introduce into a jar of the gas. Note the result.

After the action is over, hold in the jar slips of red and blue litmus paper.

Pour a little water into the jar and see if the substance formed is soluble.

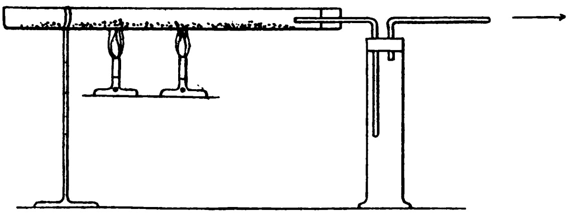
Write the reactions that have taken place.

Caution.—Phosphorus is an exceedingly dangerous substance and easily takes fire in the air. Keep under water and never touch with the fingers.

Preparation of Oxides.

Experiment 58.—Prepare copper oxide, CuO (cupric oxide), by passing a stream of oxygen or of air over copper filings placed in a combustion tube and heated with the Bunsen burner. (See Fig. 14.) Explain what has taken place chemically.

FIG. 14.



Experiment 59.—To a solution of copper sulphate, add a solution of sodium carbonate until the precipitation is complete. Stir with a glass rod and boil until the precipitate turns black. Filter, wash, and dry with a gentle heat. The powder obtained is copper oxide (CuO). Is this the substance obtained in the first experiment? Explain all that has taken place chemically.

Repeat, using NaOH instead of Na_2CO_3.

Are the same results obtained?

Experiment 60.—Place a few pieces of granulated zinc in a porcelain crucible. Heat strongly, stirring with a glass rod until the zinc is all converted into a powder. Explain what has taken place.

Experiment 61.—To a solution of a zinc salt, add a solution of sodium carbonate until the precipitation is complete.

Filter and wash thoroughly with water. Remove from the paper with the horn spatula and place the precipitate in a porcelain dish. Heat over the Bunsen burner. The powder obtained is zinc oxide (ZnO). Compare with the powder obtained by heating zinc in the air. Explain what has taken place chemically.

Repeat, using NaOH instead of Na_2CO_3.

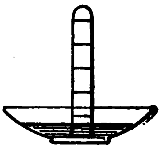
Are the same results obtained?

Volume Composition of the Atmosphere.

Experiment 62.—Divide a large-sized test-tube into five equal parts, by pasting on small rings of labels.

Place in an evaporating dish an alkaline solution of "pyrogallic acid." Invert the test-tube, slip over it a lead ring, and let it stand in the dish containing the "pyrogallic acid" for some time. (See Fig. 15.) Note that the liquid rises in the tube, and that in time it will rise to the first division on the paper scale in the tube.

FIG. 15.



What have you learned from this experiment?

Experiment 63.—Take a piece of combustion tubing about 15 inches long and close one end of it with a wooden cork. (See Fig. 16.)

FIG. 16.

Fill the tube with water, and then pour out into a measured cylinder. Read the number of c.c. This will give the capacity of the tube.

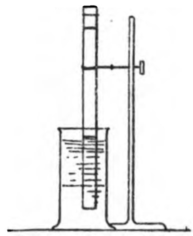
Dissolve about 1 gram of solid caustic soda or caustic potash in about 50 c.c. of water, and a like amount of crystallized ferrous sulphate in another 50 c.c. of water. Warm gently and stir each with a glass rod to aid solution.

From a small graduated cylinder pour 10 c.c. of the soda or potash into the combustion tube, and then add in like manner 10 c.c. of the ferrous-sulphate. Close the open end of the combustion tube with a cork (or hold the thumb over the open end), and shake vigorously for some little time.

Invert in a large beaker of water (See Fig. 17), remove the

cork (or the thumb) and allow to stand a few moments. The water will be seen to rise in the tube. Hold the tube so that the two levels are the same inside and outside of the tube, then place the thumb over the open end, remove from the beaker and invert.

FIG. 17.



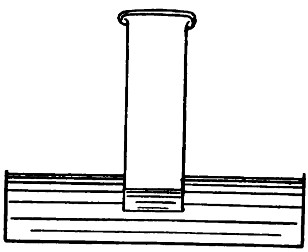
Pour the water that is now in the tube into a measured cylinder and ascertain the amount (*i. e.*, the number of c.c.). This amount, minus the twenty 20 c.c. of fluids added, will give the volume of oxygen which has been absorbed by the mixture introduced. Calculate the per cent. of oxygen in the atmosphere. Assuming the remainder to be nitrogen, about what is the ratio of these two gases in the atmosphere?

Experiment 64.—With the use of burettes and the Hempel apparatus containing small sticks of phosphorus ascertain the ratio of oxygen to nitrogen in the atmosphere.

(Consult the Instructor.)

Experiment 65.—*Modification of the above experiment.* Take the 12 inch cylinder and arrange as shown in sketch. (See Fig. 18.) Ascertain the capacity of the cylinder as in Experiment 63.

FIG. 18.



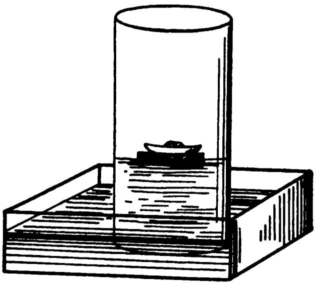
Attach a small piece of phosphorus to the bent wire. Allow to stand for some hours, gradually raising the phosphorus above the surface of the water. Level as before and calculate the percentage composition of the atmosphere.

Experiment 66.—Show the presence of water vapor and carbon dioxide in the air, by exposing a few pieces of calcium chloride in one beaker and a solution of calcium hydroxide in another. What results are observed? What do these results show?

Nitrogen.

Experiment 67.—Place the cover of a porcelain crucible on a cork and float in a trough full of water. Place a small piece of phosphorus on the porcelain cover, light it by touching with a piece of hot metal, and invert over it a large beaker. (See Fig. 19.) Note the result. Let it stand until the fumes are absorbed. Why does the water rise in the beaker? About how far does it rise? What conclusions can you draw from this? Explain the action that takes place. Put pieces of litmus paper in the water. What substance has been formed and dissolved in the water? Slip under the mouth of the beaker a square of ground glass. Quickly invert, and, partly removing the glass, introduce a lighted candle fixed in a deflagrating spoon. Note the result. Write all the reactions.

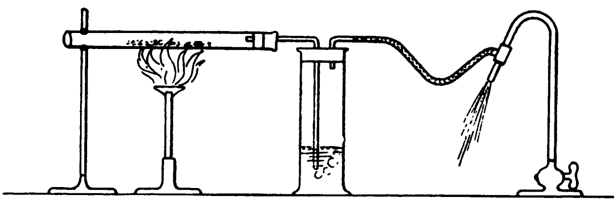
FIG. 19.



Bear in mind the caution given regarding the use of phosphorus.

Experiment 68.—Arrange an apparatus as shown in the sketch (See Fig. 20), fill the combustion tube with copper

FIG. 20.



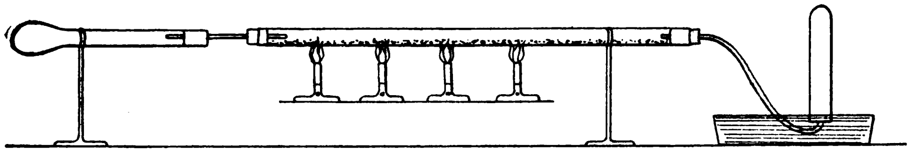
scraps. Heat strongly and aspirate a slow current of air through the apparatus. Examine the nitrogen that collects in the jar.

Experiment 69.—Mix a few grams of dried white of egg (albumen) with an excess of copper oxide. Place the mixture in an ignition-bulb tube, connect with a long ignition tube

filled with copper scraps. Heat the contents of the bulb tube strongly and also the long ignition tube, collect the gas evolved over water. (See Fig. 21.)

Examine the gas in the tube. Explain all that has taken place.

FIG. 21.

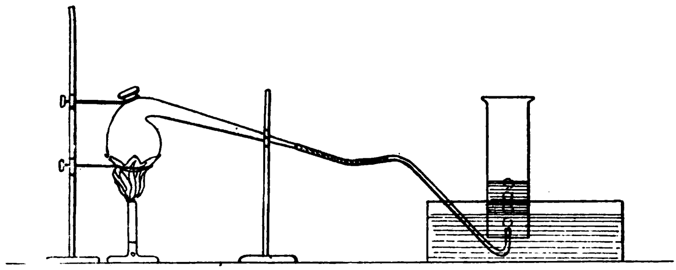


Experiment 70.—Weigh out:—

10 grams NH_4Cl (ammonium chloride),
10 " KNO_2 (potassium nitrite),
10 " $K_2Cr_2O_7$ (potassium dichromate).

Introduce into a retort and pour in 30 c.c. water. Arrange the apparatus as shown in the sketch (See Fig. 22), heat gently

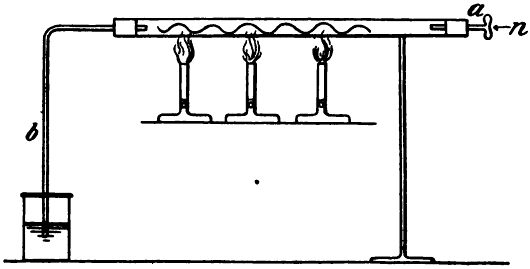
FIG. 22.



and gradually over a Bunsen burner, and as the gas comes off freely, lower or remove the flame. Collect several jars of the gas over water. Examine its physical and chemical properties. Does it act in the same way as the gas obtained above? Write the reactions by which the gas is prepared. Why was potassium dichromate added? Make a mixture of four parts of this gas with one part of oxygen gas in a tall glass cylinder. Does it behave like ordinary atmospheric air? What conclusions can you draw from these results?

Experiment 71.—Place a coil of magnesium ribbon, resting on a thin layer of asbestos in a piece of combustion tubing. Arrange the apparatus as shown in the sketch. (See Fig. 23.)

FIG. 23.



Pass a stream of nitrogen gas through the apparatus until all air is displaced. Close the end (*a*) of the ignition tube. Heat the contents of the tube strongly. Cool and observe the rising of the water in the tube (*b*). What causes this? Explain all that has taken place.

Water.

Experiment 72.—Heat a few fragments of wood gently in a test-tube and observe the moisture given off. Repeat the experiment, using a few small pieces of raw potato.

What do you learn from these experiments?

Experiment 73.—Repeat Experiment 72, using any organic substance (starch, sugar, wood, etc.).

Are similar results obtained?

What is meant by the term "destructive distillation?"

Experiment 74.—Burn a candle or a piece of wood. Hold over it a dry beaker and observe the condensation of water.

Query.—In all the above experiments how could it be proved that water was produced?

Experiment 75.—Introduce into an explosion pipette a mixture of hydrogen and oxygen gases. Explode by means of the electric spark. (Consult the Instructor.)

Also burn hydrogen obtained from a zinc and sulphuric acid generator and note the production of water. (Consult the Instructor.)

Also use the Hofmann apparatus for the decomposition of water. (Consult the Instructor.)

Experiment 76.—Pass a stream of dry hydrogen gas over a heated oxide. (Preferably copper oxide.)

Note the formation of water. (Consult the Instructor.) See Fig. —.

Experiment 77.—Place in a beaker a little dry barium hydroxide. *Add cautiously* a few drops of conc. sulphuric acid. Note the result.

Write the reaction that has taken place.

Experiment 78.—Repeat, using calcium oxide (quick lime.) Use great caution. Add only a drop or two at a time of the conc. sulphuric acid. Are similar results obtained?

Write the reaction that has taken place.

Experiment 79.—Place in a perfectly dry test-tube a little barium hydroxide. Heat gently. Note the formation of water vapor in the cooler part of the tube.

Experiment 80.—Bend a piece of glass tube, sealing one end. (See Fig. 1 second tube.) Prepare three of these tubes. Introduce into these tubes one or two crystals of "blue vitriol" ($CuSO_4,5H_2O$), "green vitriol" ($FeSO_4,7H_2O$), and "white vitriol" ($ZnSO_4,7H_2O$), respectively. Heat gently and observe the liquid that collects in the cool part of the tube.

What has taken place?

What name is given to water chemically combined in this manner?

Experiment 81.—Dissolve a small crystal of potassium permanganate in hydrant water, and prepare some distilled water, using the apparatus shown in Fig. 28. Put into a bottle and label.

Evaporate a little of the water on a piece of platinum foil, see if a residue remains. Also evaporate a little Schuylkill water and sea water (made by dissolving salt in water) in like manner. Compare the results.

Experiment 82.—Take two watch glasses; in one place some crystals of sodium sulphate ($Na_2SO_4.10H_2O$) and in the other a fragment of calcium chloride ($CaCl_2$); leave exposed for a few days.

What changes take place?

What two terms are used to express these changes?

Solution.

(*a*) *Gas in Liquid.*

Experiment 83. — Boil a little ammonium hydroxide (NH_4OH) in a test-tube. Quickly invert and place in a dish containing water and a little copper sulphate solution.

Explain what takes place.

(*b*) *Liquid in Liquid.*

Experiment 84.—Place in a test-tube a little water colored with a dyestuff, and in another put a little alcohol. Carefully add one to the other and observe whether they are miscible in all proportions.

(*c*) *Solid in Liquid.*

Experiment 85.—Place in beakers pulverized copper sulphate, potassium dichromate, and common salt. Add water to each. Stir with a glass rod. Boil and note whether they dissolve more easily in the cold or with heat.

Evaporate down to small bulk, and set aside to crystallize.

Examine the crystals formed and see whether or not they contain water of crystallization.

Query.—How could you show that the water had dissolved the various solids used?

Solubility of Various Chemical Compounds.

Experiment 86.—Determine the solubility of the following compounds :—

$NaCl$
$K_2Cr_2O_7$
$CaSO_4$
Na_2CO_3
KNO_3
NH_4NO_3.

Fill a large beaker with water. Place this over the Bunsen burner and heat to boiling. Fill a large sized test-tube about three-fourths with distilled water. Place in this enough of the compound under examination to form a saturated solution at 100° C., the temperature of the boiling water. Shake the contents of the test-tube thoroughly to insure complete saturation of the solution.

Keep the test-tube immersed in the beaker of water and place in it a thermometer.

Draw out 2 c.c. of the clear solution by means of a pipette. Place in a weighed 50 c.c. stoppered flask, and when cool ascertain the weight, and transfer to a weighed evaporating dish. Evaporate to dryness and ascertain the weight.

Remove the Bunsen burner and in like manner remove portions of 2 c.c. at varying temperatures as indicated by the thermometer.

Tabulate the result in the form of a plotted curve. (Consult the Instructor.)

Hydrogen.

Experiment 87.—Fill an evaporating dish with water acidulated with a little sulphuric acid.

Fill two small test-tubes with acidulated water, invert and place in the evaporating dish. Slip over them leaden rings to keep them in an upright position. Connect a wire from the electric light above the desk and pass a current of electricity through the water, having one wire inside of one tube and one in the other. Use a *direct* current.

What change takes place?

What are the respective volumes of the two gases obtained?

What is the behavior of each toward a lighted match?

Experiment 88.—Place water in two evaporating dishes. In the one throw a small fragment of sodium, and in the other a small fragment of potassium.

What takes place in each case?

Why is flame seen in one case?

Write the reactions that have taken place.

Caution.—Metallic sodium and potassium should be always *kept under naphtha or some other liquid which does not attack them. Never touch with the fingers, but use the forceps.*

Experiment 89.—Place some water in a large-sized evaporating dish. Fill a large test-tube with water, invert and place in the dish of water, slipping a leaden ring over the tube to keep it in an upright position.

Introduce a small piece of the metal magnesium into the tube.

Pour hydrochloric acid into the dish, mix with the tube, and observe the action on the metal.

Test the gas obtained with a lighted match. Does it burn?

Write the reaction that has taken place.

Experiment 90.—Arrange an apparatus as shown in Fig. 28. Introduce about 10 grams of zinc (granulated) into the flask. Cover with water and pour in dilute sulphuric acid through the thistle-bulb tube.

Collect several jars of the gas given off, using the pneumatic trough and collecting over water.

Write the reaction that has taken place.

When the action is over, pour out the contents of the flask into an evaporating dish and leave to crystallize.

What substance is obtained? Do the crystals contain water of crystallization? How could you show it?

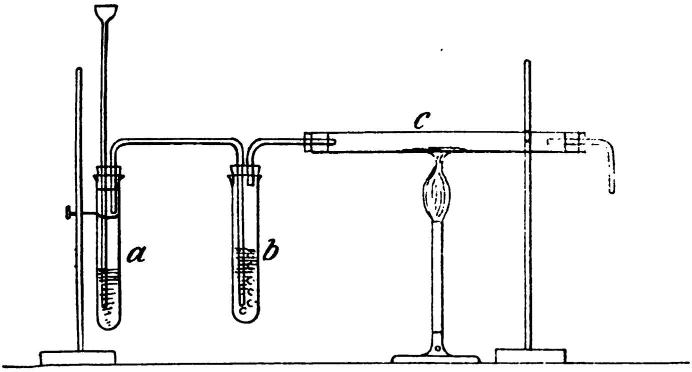
Experiment 91.—Examine a jar of the gas obtained. Note whether it is heavy or light; whether it has an odor or not, and if it reacts with litmus papers.

Experiment 92.—Ascertain if the gas burns and if it supports combustion. Mix a little air with a jar of the gas; light, holding the jar mouth downwards and observe the result.

Experiment 93.—Arrange an apparatus as shown in Fig. 24. Place in the piece of combustion tubing a little copper oxide (CuO). Pass a stream of dried hydrogen gas over the

powder, and heat at first gently and then strongly with a Bunsen burner. Explain what takes place.

FIG. 24.



What is meant by a reducing agent?

Explain how this experiment shows hydrogen to be a strong reducing agent. Write the reaction that has taken place.

Make a tabular statement of the physical and chemical properties of hydrogen.

Ammonia.

Experiment 94.—Ammonia can be produced by the destructive distillation of any organic substance containing nitrogen.

Query.—Recall what is meant by the term "destructive distillation."

Place in dry test-tubes a little egg albumin, or a small piece of meat. Heat strongly over a Bunsen burner. Test the gases given off by means of moistened red litmus paper. Note the change in color. What does this indicate?

Explain what has taken place.

Experiment 95.—Mix thoroughly equal parts of ammonium chloride and quicklime (NH_4Cl and CaO); place the mixture in a flask and heat gently. Collect the gas in a glass cylinder by upward displacement. (See Fig. 25.)

Write the reaction that has taken place.

Note the odor of the gas obtained.

Use *great caution* in smelling this gas!

What is left behind in the flask?

Pass some of the gas into water, using the device shown in Fig. 26.

FIG. 25.

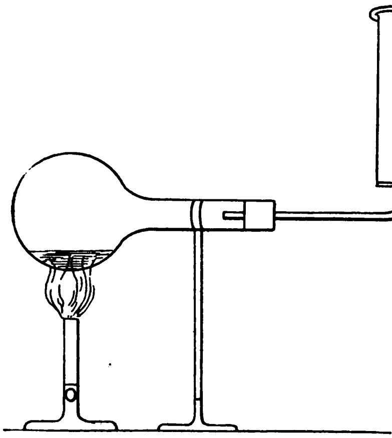
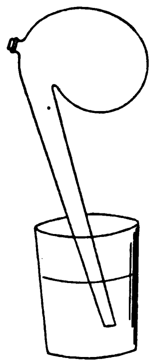


FIG. 26.



Pass the gas by means of a rubber tube into the retort through the tubulure. Remove the gas generator, close the rubber tube by means of a pinch-cock or a bit of glass rod, and observe the water rise in the retort as the gas dissolves.

Query.—Why is the gas not collected over water, as was done in experiments with other gases?

Experiment 96.—Try the action of the gas on a piece of wet litmus paper; on filter paper wet with a solution of copper sulphate; with mercurous nitrate.

What is the result in each case?

Hold a glass rod wet with hydrochloric acid in the gas. What is the result? Explain what takes place.

Experiment 97.—To a solution of ammonium chloride in a beaker, add a solution of sodium hydroxide or potassium hydroxide. Stir with a glass rod and warm gently.

Explain the reaction that takes place.

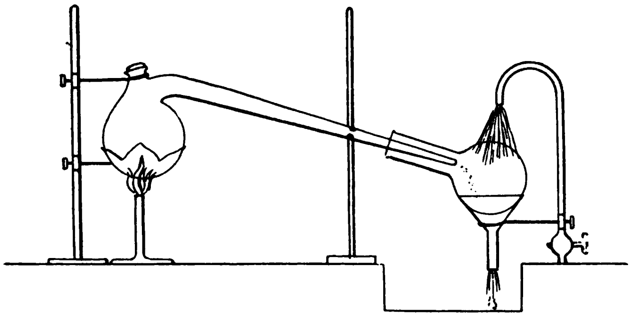
Test the gas that comes off as in the experiment above. Do you get the same results?

Query.—Could any salt of ammonium be used? Ascertain this experimentally.

Nitric Acid (HNO_3).

Experiment 98.—Arrange an apparatus as shown in the sketch. (See Fig. 27.)

FIG. 27.



In the retort place about 25 grams of sodium nitrate ($NaNO_3$) and pour in through a small funnel enough concentrated sulphuric acid (H_2SO_4) to cover it.

Replace the glass stopper and heat gently over a Bunsen burner and collect the distillate in a small flask cooled by a stream of water from the spigot.

Examine the liquid.

Dip bits of litmus paper into it; also bits of filter paper, cloth, etc.

What is the action?

Write the reaction of the sulphuric acid on the sodium nitrate.

Properties of Nitric Acid.

Experiment 99.—Try the action of the liquid on small pieces of copper, zinc, tin, and lead, placed in shallow evaporating dishes.

What takes place?

What are the products formed in each case?

How does the product obtained from the tin differ from the others?

Be very careful not to breathe the fumes given off.

Caution.—Nitric acid is an extremely corrosive and dangerous substance. Use great care. See that none of it touches the skin or the clothing.

Experiment 100.—Add to some nitric acid placed in an evaporating dish, ammonium hydroxide until the solution is alkaline. Boil down to small bulk and set aside to cool. Save the solid mass that forms for future work.

What is this solid?

Experiment 101.—Place a little concentrated nitric acid in an evaporating dish. Heat over a Bunsen burner, and when quite hot introduce a glowing match-end. Note the oxidizing effect of the nitric acid on the carbon.

Experiment 102.—Place a little dilute indigo solution in a test-tube. Add a few drops of nitric acid and note the result.

Experiment 103.—Place a small piece of phosphorus in a porcelain evaporating dish. Pour in a few c.c. of nitric acid. Warm gently on the water bath until the phosphorus has all disappeared.

Danger! Use great caution in doing this. Test the solution for phosphoric acid in the following manner:

Pour into a test-tube, add ammonium molybdate, and observe the formation of the yellow ammonium phospho-molybdate.

Experiment 104.—Place some nitric acid in a small flask and add sulphur. Boil for some time and test the liquid for sulphuric acid by means of barium chloride.

Observe the formation of $BaSO_4$ white insoluble precipitate.

Oxides of Nitrogen.

(N_2O. N_2O_2. N_2O_3. N_2O_4. N_2O_5.)

Experiment 105.—Preparation of N_2O.—Place in a retort the ammonium nitrate prepared from the last experiment. Connect with a bent delivery tube. (See Fig. 22.) Heat gently and collect the gas evolved over hot water.

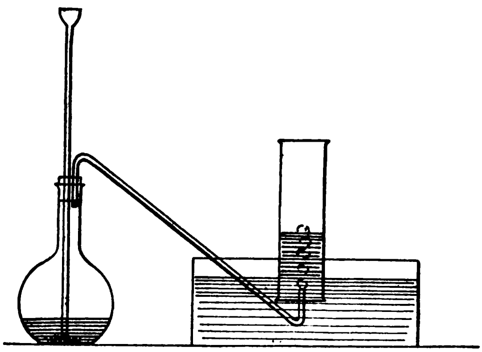
Why is it necessary to use hot water?

Collect several jars of the gas, and examine its properties.

Write the reaction that has taken place.

Experiment 106.—Preparation of N_2O_2 or NO.—Place some copper scraps in a flask fitted with a bent delivery tube. (See Fig. 28.) Pour some dilute nitric acid through the thistle-bulb tube, and collect the gas given off over water.

FIG. 28.



Write the reaction. Fill several jars with the gas and examine its properties.

What takes place when this gas comes in contact with the air?

What compound is formed?

What constituent of the air has combined with it? Prove this experimentally.

N_2O_3.—What is the name of this oxide?

What acid does it form when it combines with water?

Give the formula.

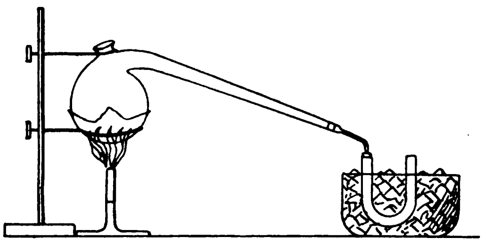
What name is given to salts of this acid?
Give examples with potassium and sodium.

Experiment 107.—Preparation of N_2O_4 or NO_2.—Place in a perfectly dry retort 25 grams of lead nitrate. (See Fig. 29.) Heat at first gently and finally rather strongly. Pass the gas evolved through a U-tube placed in a freezing mixture of cracked ice and salt. Pour the liquid obtained into a slightly warmed tall cylinder jar.

Examine the properties of the gas.

Write the reaction that has taken place.

FIG. 29.



N_2O_5.—What is the name of this oxide?
What acid does it form when it combines with water?
Give the formula.
What name is given to salts of this acid?
Give examples with potassium and sodium.

A Study of Neutralization.

Acids, Bases, and Salts.

Experiment 108.—Weigh 10 grams of solid caustic soda (NaOH) in a small beaker, and dissolve in about 200 c.c. of water. Add hydrochloric acid (HCl) slowly, stirring the solution with a glass rod and examining from time to time by means of a piece of blue litmus paper. So long as the solution is alkaline it will cause no change in the color of the paper, but the instant it passes the point of neutralization the

paper is turned red. When this point is reached transfer to an evaporating dish, evaporate to complete dryness and see what is left. Taste the substance.

Does it suggest any familiar substance?

Is it alkaline, acid, or neutral to test papers?

Heat a little of the powder in a test-tube with a few drops of sulphuric acid (H_2SO_4).

What gas is given off?

Write the reaction.

Dissolve another portion of the powder in a test-tube and add silver nitrate ($AgNO_3$).

What takes place?

Write the reaction.

Heat another portion of the powder in the loop of a platinum wire in the flame of a Bunsen burner.

What color is imparted to the flame?

The presence of what substance is indicated by this colored flame?

In consideration of all the above facts, what have you learned in regard to the action of hydrochloric acid (HCl) on sodium hydroxide (NaOH)?

Write the reaction that took place between the two substances, stating the relative weights that entered into chemical combination.

Query.—How much hydrochloric acid was needed to exactly neutralize the 10 grams of NaOH used?

Experiment 109.—Repeat the same experiment using nitric acid (HNO_3) instead of hydrochloric acid. Examine the product formed as in the experiment above. Also heat a little of it in a glass tube closed at one end.

What takes place?

Explain the action of nitric acid on sodium hydroxide and write the reaction, stating the relative weights of the two substances that have combined.

Experiment 110.—Similarly use sulphuric acid (H_2SO_4) and caustic soda (NaOH).

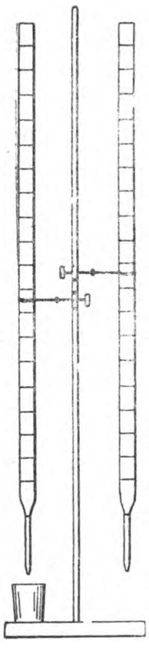
Note.—Use diute sulphuric acid. Write the reaction and state, as before, the relative weights that have chemically combined.

Query.—Do similar reactions take place when potassium hydroxide (KOH) is used instead of sodium hydroxide (NaOH)?

In all these cases of chemical combination of an acid with a base, note the heat given off. This is always an indication that chemical action has taken place.

What are the invariable products resulting from the action of an *acid* on an alkali (a *base*)?

FIG. 30.



Experiment III.—Place about 250 c.c. distilled water in each of four glass-stoppered bottles of convenient capacity. Add to the first about ten grams solid NaOH; to the second about the same amount of solid KOH; to the third about 10 c.c. conc. H_2SO_4 and to the last about the same amount of strong HNO_3. Shake the contents of each bottle thoroughly and label each. Thoroughly clean and rinse out a pair of burettes (50 c.c graduated in $\frac{1}{10}$th of 1 c.c.). Label the one "alkali" and the other "acid." (See Fig. 30.)

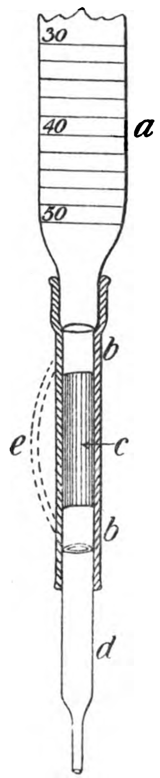
Pour some of the dilute NaOH previously prepared into the "alkali" burette. Rinse it out *with this solution* and then fill to the zero mark. In like manner fill the other burette with the dilute sulphuric acid.

Take a small perfectly clean beaker and pour into it about 100 c.c. distilled water. Now run in carefully from the burette a measured quantity of the dilute sodium hydroxide (say 10 c.c.). Add a few drops of an indicator (use litmus solution or phenol-pthalëin). Run in from the acid burette a sufficient quantity of the dilute sulphuric acid until neutralization is effected. Stir constantly with small clean glass rod, and note the point at which the color of the indicator changes, thereby showing the point

when neutralization is reached. (See Fig. 31 for the construction of the valve of the burette.)

Repeat this two or three times, taking the same quantity of the sodium hydroxide each time. Now repeat several times taking different quantities of the sodium hydroxide. Record all results and tabulate these results in the form of simple ratios.

FIG. 31.



Repeat all the above using the dilute solutions of potassium hydroxide and nitric acid. As before record all results and tabulate in the form of simple ratios.

If time permits repeat using sodium hydroxide and nitric acid, potassium hydroxide and sulphuric acid.

What important fact has been learned from these experiments?

Give a brief summary of all that you have learned from this Study of Neutralization.

Prepare a table showing the relative molecular quantities of bases and acids that neutralize one another on a basis of *one atom of replaceable hydrogen.* Use the following bases and acids.

NaOH	HCl
KOH	HNO_3
Na_2CO_3	H_2SO_4
$Na_2CO_3.10H_2O$	H_3PO_4
$NaHCO_3$	$HC_2H_3O_2$
NH_4OH	$H_2C_4H_4O_6$
NH_3	$H_2C_2O_4.2H_2O$
$Ba(OH)_2$	H_3BO_3

Chlorine.

Note.—Perform all the experiments under this heading under the "Hood."

Caution.—Be careful not to inhale any of the gas.

Inhale alcohol fumes as antidote.

Experiment 112.—Use the Hofmann apparatus for the electrolysis of hydrochloric acid and chlorides.

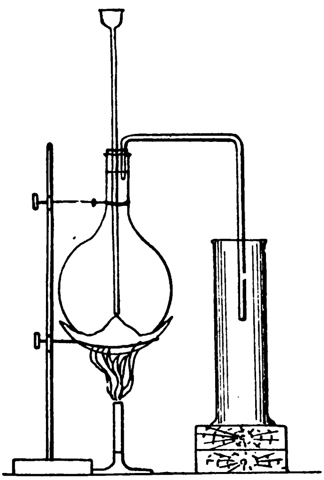
(Consult the Instructor.)

Experiment 113.—Place in a small test-tube a little common salt (NaCl); in another a little MnO_2 (manganese dioxide). Add some dilute H_2SO_4 (sulphuric acid) to each; heat in the flame of a Bunsen burner and note the results. Hold in the mouth of the test-tubes bits of litmus paper. Note the results. Now pour the contents of the test-tubes into another large test-tube, heat and observe the results.

Note color and odor of the gas evolved, and the effect of it on a piece of litmus paper.

Experiment 114.—Arrange an apparatus as shown. Use a flask of one litre capacity. (See Fig. 32.)

FIG. 32.



Place in the flask a mixture of common salt (NaCl) and MnO_2 (about 25 grams of each); pour through the thistle-bulb tube some dilute H_2SO_4. Heat gently and collect several jars of the gas by downward displacement. Write the reactions.

Note the color and odor of the gas. Does it support combustion? Note its effect on bits of moistened red and blue litmus paper. Is the gas soluble in water? Prepare about 200 c.c. of "Chlorine water."

Experiment 115.—Place a solution of sodium iodide in one test-tube and a solution of sodium bromide in another. Add to each a few drops of chloroform. Add now a few drops of the chlorine water prepared in the previous experiment; shake the contents of the tubes and note the results.

Explain the action of the chlorine.
Write the reactions.
Why is chloroform used?
Could anything else be used in its place?

Experiment 116.—Place some MnO_2 in a 250 c.c. flask arranged as in Experiment 114, add some HCl and heat gently.

Examine the gas evolved.
Is it the same as the gas obtained in the last experiment?
Write the reaction.
Collect several jars of the gas, as before.

Experiment 117.—Place some bleaching powder in a test-tube, add dilute HCl and heat gently.

Examine the gas obtained and compare results with the above.

Is the same gas obtained?

Properties of Chlorine.

Experiment 118.—Into a jar of the gas under examination, shake from a piece of paper a little powdered antimony (Sb).

Note the results.
What has taken place?
Attempt to dissolve some of the powder formed in pure water and in water acidulated with HCl.

Experiment 119.—Into a jar of the gas introduce a leaf of "Dutch metal" (an alloy of copper and zinc).

What takes place?
Explain the result.

Experiment 120.—Moisten a bit of filter paper with a few drops of hot turpentine and introduce by a pair of pincers into a jar of the gas.

Note the results.
Explain what takes place.

Experiment 121.—Into another jar of the gas introduce two pieces of red litmus paper, one dry and the other wet.

Compare the results.

Explain the chemical principle of the bleaching power of Chlorine.

Write the reactions.

Make a table of the physical and chemical properties of chlorine.

Hydrochloric Acid (HCl).

Experiment 122.—Arrange an apparatus as shown. (See Fig. 32.) (Use the Hood.) Place in the flask (of 1 litre capacity) about 30 grams of common salt (NaCl). Pour in through the funnel tube, gradually, about 50 grams of dilute sulphuric acid. Note the gas that is evolved. When the evolution of gas slackens, warm the flask gently, and heat until all the gas is expelled. Write the reaction that has taken place. Collect a few jars of the gas by downward displacement. Note the effect of the gas on pieces of litmus paper. Hold a glass rod wet with ammonium hydroxide in a jar of the gas. Explain what takes place. Invert a jar of the gas in the pneumatic trough filled with water.

Is the gas soluble in water?

Pass the gas into a small flask in which water has been placed.

Pour a little of this solution into a few fragments of zinc placed in an evaporating dish. Explain what takes place. What gas is evolved? How could you prove it?

When the contents of the large flask are cool, dissolve in the least possible amount of warm water, filter and set the clear solution aside to crystallize.

What substance is left in the flask after all the hydrochloric acid has been expelled?

What is the composition of the crystals obtained?

Experiment 123.—Precipitate the following insoluble chlorides:—

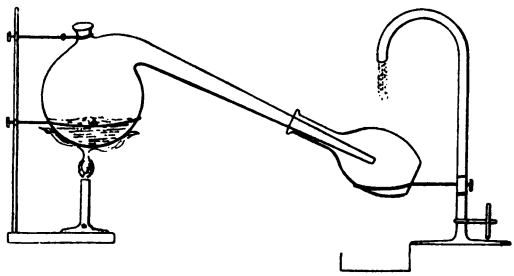
$AgCl$, Hg_2Cl_2, $PbCl_2$.

To solutions of silver nitrate, mercurous nirtate, and lead acetate, placed in test tubes, add the solution of hydrochloric acid prepared in the foregoing experiment. Note the color and appearance of each chloride formed.

Hypochlorous Acid (HClO).

Experiment 124.—Arrange an apparatus as shown. (See Fig. 33.)

FIG. 33.



Place in the retort a filtered solution of "bleaching powder" ($CaCl_2$, $Ca(ClO)_2$).

Add a little dilute nitric acid (one part of acid to three parts of water).

Heat gently and examine the liquid that distils over and collects in the small flask.

Note its action on pieces of litmus paper. Explain why it bleaches. Write the reaction by means of which this acid was prepared.

Preparation of Potassium Chlorate ($KClO_3$).

Experiment 125.—Dissolve 50 grams of potassium carbonate in the least possible amount of water. Pass a stream of chlorine gas through the solution until it no longer has an alkaline reaction.

Dilute to about 200 c.c. and set aside to crystallize.

Reaction :—

$$3K_2CO_3 + 6Cl = KClO_3 + 5KCl + 3CO_2.$$

Preparation of Sodium Chloride ($NaCl$).

Experiment 126.—Calculate the theoretical amount of crystallized sodium carbonate ($Na_2CO_3 10H_2O$) that it will be necessary to take to obtain 25 grams of sodium chloride ($NaCl$).

Weigh this amount as accurately as possible. Choose clear crystals. (Why?)

Introduce into a flask of 1 litre capacity and add hydrochloric acid (HCl), little by little, until no further action takes place and the contents of the flask have an acid reaction.

What gas is given off?

Introduce a lighted candle and note the result.

What compound is in the solution?

Write the reaction that has taken place.

Pour the contents of the flask into a large evaporating dish that has previously been carefully weighed. Rinse out the flask and add the rinsings. Evaporate to dryness, and gently ignite until the excess of acid has been expelled. Cool the dish and carefully weigh.

The weight of the dish subtracted from this weight will give the weight of the sodium chloride.

Record the weights, and explain the result of the experiment.

Dissolve the sodium chloride obtained in the least possible amount of hot water.

Set aside to crystallize spontaneously.

Note the curious "hopper-shaped" crystals obtained.

Bromine (Br).

Experiment 127.—Arrange a retort and flask as shown in sketch. (See Fig. 33.)

Introduce into the retort a few grams of sodium bromide ($NaBr$) and a little manganese dioxide. Add dilute sulphuric acid, pouring it into the retort by means of a small funnel, and warm the mixture gently.

Observe the red liquid that collects in the small flask.

Note.—Keep the flask as cold as possible with a constant stream of water from the hydrant.

Write the reaction that has taken place.

Experiment 128.—Pour a drop of the liquid into a jar which has been previously warmed. Note the fumes produced. Drop a little powdered antimony into the flask.

What takes place?

Experiment 129.—Into another jar that has been warmed pour a drop of the red liquid prepared in the first experiment.

Do the vapors support combustion?

Introduce pieces of moist litmus paper.

What takes place?

Experiment 130.—Pour the remainder of the red liquid in the flask into a jar; add water and see if it dissolves. (Save the result for future work.)

Note.—Make a table of the physical and chemical properties of bromine.

Preparation of Bromides.

Experiment 131.—Place the "bromine water" obtained in the last experiment in a porcelain evaporating dish and add a solution of sodium hydroxide, until the color has nearly disappeared. Stir constantly with a glass rod. Evaporate to dryness and then thoroughly ignite the residue.

What is this residue?

When cool, dissolve in water and keep for future ase.

Experiment 132.—Precipitate the following insoluble bromides:—

$$AgBr, Hg_2Br_2, PbBr_2.$$

To solutions of silver nitrate, mercurous nitrate, and lead acetate, placed in test tubes, add the solution prepared in the above experiment.

Note the color and appearance of each bromide formed.

Iodine (I).

Experiment 133.—Prepare iodine in the same way that bromine was prepared. (See Fig. 33.)

Note the appearance of the substance that collects in the small flask, and the color of the fumes.

Write the reaction that has taken place.

Experiment 134.—Dissolve a little of the solid obtained in the above experiment in alcohol. Prepare in a beaker a little boiled starch. Add a few drops of the solution and note the result.

Query.—Would chlorine or bromine give similar results?

Experiment 135.—Ascertain if iodine has any bleaching power.

How does it compare with bromine in this respect?

Note.—Make a table of the physical and chemical properties of iodine.

Preparation of Iodides.

Experiment 136.—Prepare sodium iodide in the same way that the bromide was prepared.

Dissolve in water, use the solution to prepare the following insoluble iodides. Note the color and appearance of each.

AgI,
Hg_2I_2, HgI_2.
PbI_2,

Precipitate these in the same way that the bromides were precipitated, using solutions of silver nitrate, mercurous nitrate, lead nitrate and mercuric chloride.

The Halogens.

(Chlorine, Bromine, Iodine.)

Comparative Study.

Experiment 137.—Place in three dry test-tubes a little $NaCl$, $NaBr$, and NaI, and add to each a few drops of concentrated sulphuric acid.

What results are obtained?

What do these results show?

Experiment 138.—Place in three dry test-tubes powdered manganese dioxide (MnO_2) mixed with NaCl, NaBr, and NaI respectively. Add to each a little dilute sulphuric acid, warm gently and note the results obtained in each case. Place in the mouth of each tube a bit of moistened red litmus paper; compare the results obtained.

Experiment 139.—Prepare a thin solution of boiled starch paste. Pour a little into three test-tubes and add to each a little chlorine water, bromine water, and a solution of iodine. Compare the results obtained.

Experiment 140.—Place a number of clean test-tubes in the test-tube rack, and precipitate in them the chlorides, bromides, and iodides of the metals silver, lead, and mercury. Make a comparative study of these compounds.

Query.—What is formed when a caustic alkali is added to chlorine, bromine, or iodine?

Make full tabular comparisons of the properties, both physical and chemical, of chlorine, bromine, and iodine; also of the results obtained in the above experiments.

Sulphur.

Physical Properties.

Experiment 141.—Examine a piece of roll sulphur, note its appearance, color, odor, hardness, taste, etc. Rub the piece of sulphur briskly on the coat sleeve, note the electricity developed and the power that it now has of attracting small bits of filter paper.

What change has taken place?

Is it physical or chemical?

Experiment 142.—Place a small fragment of sulphur in a test-tube, heat gently, and see if it can be made to assume the liquid and gaseous states. Heat more strongly and note the differences in viscidity of the sulphur at different stages of this experiment.

Be careful that the sulphur does not take fire and burn.

Experiment 143.—Determine the solvent action on sulphur of the following: water, acids, alkalis, alcohol, carbon disulphide.

To determine the solubilities, place small quantities of the sulphur in test-tubes, add the reagents named, shake well and note the effect.

Chemical Properties.

Experiment 144.—Attempt to light a small piece of sulphur and see if it burns.

Query.—Is the kindling point of sulphur high or low?

What is formed when sulphur burns?

Experiment 145.—Determine if sulphur can be made to combine directly with metals.

Heat together in test-tubes small quantities of sulphur with iron, copper, lead, and zinc.

Is a new compound formed in each case?

What are these compounds?

Allotropic Modifications of Sulphur.

Experiment 146.—Dissolve a small fragment of roll sulphur in carbon disulphide in a test-tube, pour the solution obtained into a large, clean, dry beaker and allow it to evaporate spontaneously. Note the form of the crystals obtained.

N. B.—Bear in mind the caution before given in regard to carbon disulphide. See that no flame is in the vicinity.

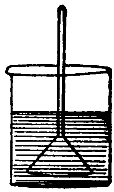
Experiment 147.—Fill a test-tube to the depth of about an inch with sulphur, warm until the sulphur is melted and limpid, remove the heat and as it cools watch the formation of crystals as they shoot out from the comparatively cold walls towards the centre of the liquid. Note the difference in form between these crystals and those obtained by evaporating the solution of sulphur in carbon disulphide in the previous experiment.

Experiment 148.—Place some sulphur in a test-tube, heat

until it turns dark colored and commences to boil; pour into a beaker full of cold water; note the plastic non-crystalline form of the sulphur obtained. (See Fig. 34.)

Dry some of this plastic sulphur; keep for a few days and note the changes that take place in it.

FIG. 34.

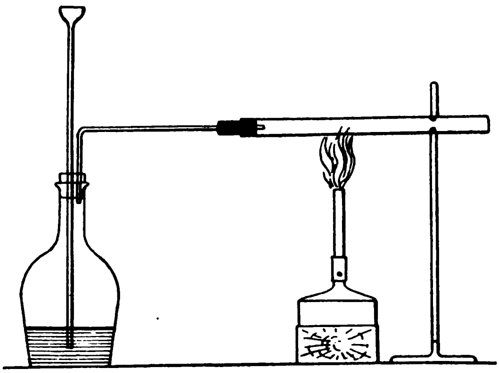


Caution.—Be careful that the sulphur does not take fire and burn in this experiment. The fumes formed by burning sulphur are very choking; be careful not to inhale them.

Hydrogen Sulphide (H_2S).

Experiment 149.—Pass dry hydrogen gas over a few fragments of sulphur placed in a combustion tube and heated with a Bunsen burner. Test the gas that escapes from the end of the tube with a paper moistened with $Pb(C_2H_3O_2)_2$. (See Fig. 35.)

FIG. 35.



Explain all that has taken place.

Experiment 150—Repeat the above, using a sulphide such as CuS in place of the sulphur.

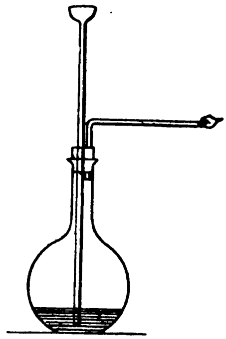
Are similar results obtained?

Experiment 151—Place a few fragments of FeS in a flask fitted with a thistle-bulb tube and delivery tube. Add dilute

H_2SO_4, or dilute HCl, and examine the gas given off. (See Fig. 36.)

Does it burn? Pass some of this gas over a heated oxide, such as CuO or PbO. What reaction takes place?

FIG. 36.



Sulphides (H_2S).

Experiment 152.—Arrange an apparatus as before and prepare H_2S by the action of dilute hydrochloric acid on FeS. Pass the H_2S gas into a bottle filled with water until the solution smells strongly.

Use this solution of H_2S in water to precipitate the sulphides of the following metals:—

Silver	(use $AgNO_3$).
Mercury	(" $Hg_2(NO_3)_2$).
Lead	(" $Pb(C_2H_3O_2)_2$).
Bismuth	(" $BiCl_3$).
Mercury	(" $HgCl_2$).
Copper	(" $CuSO_4$).
Cadmium	(" $CdSO_4$).
Arsenic	(" $AsCl_3$).
Antimony	(" $SbCl_3$).
Tin	(" $SnCl_2$).
Tin	(" $SnCl_4$).

Note the colors and properties of the precipitates obtained.

Write the reactions that have taken place.

Note.—Prepare the solution of H_2S *under the hood.*

Experiment 153.—Prepare a solution of ammonium sulphide by passing H_2S gas into a solution of ammonium hydroxide, placed in the bottle marked "Ammonium sulphide $(NH_4)_2S$."

With this solution so prepared, precipitate the following solutions:—

Iron	(use $FeSO_4$).
Aluminium	(" Al_2Cl_6).
Chromium	(" Cr_2Cl_6).
Nickel	(" $NiCl_2$).
Cobalt	(" $CoCl_2$).
Zinc	(" $ZnCl_2$).
Manganese	(" $MnCl_2$).

Query.—Do aluminium and chromium form sulphides? If not, what is the composition of the precipitates obtained by adding the $(NH_4)_2S$?

Note the colors and properties of the precipitates obtained.

Write the reactions that have taken place.

Note.—Prepare the solution of $(NH_4)_2S$ *under the hood.*

Sulphur Dioxide.

Experiment 154.—Place a small lump of sulphur in a deflagrating spoon, light it, and introduce into a glass jar; examine the gas formed. Note its bleaching action on moistened litmus paper or delicate tinted flowers.

Experiment 155.—Place in a flask a few scraps of copper, cover with a little concentrated H_2SO_4, and heat. Note the gas evolved. (See Fig. 32.)

Is it the same gas as that obtained in the experiment above?

Write reaction.

Experiment 156.—Place some powdered charcoal in a test-tube and cover with concentrated H_2SO_4; warm gently and note the gas evolved.

Is this gas the same as that obtained in the two experiments above?

Write the reaction.

Sulphur Trioxide.

Experiment 157.—Strongly heat a crystal of ferrous sulphate in a test-tube; after the water of crystallization is driven out, note the gas evolved; hold pieces of litmus paper in it; note the effect.

Write the reactions that have taken place.

Experiment 158.—Heat some *fuming sulphuric acid,* $H_2S_2O_7$, in a test-tube and note the gas evolved.

Is it the same as the gas obtained in the experiment above?

Write the reaction.

Query.—In what ways can you distinguish sulphur dioxide from sulphur trioxide?

Caution.—Be very careful not to inhale the fumes of SO_2 *nor of* SO_3. They are very suffocating.

Sulphurous Acid (H_2SO_3).

Experiment 159—Pass a stream of sulphur dioxide gas through water contained in a beaker; prepare the gas by acting on copper with sulphuric acid. (See Fig. 32.)

Examine the solution obtained; notice its odor, immediate action on blue litmus paper and the subsequent action.

What is the action of NaOH on this solution?

Is its odor changed?

Is its acid character changed?

Why are these changes produced by the NaOH?

Write all reactions involved in the above changes.

Experiment 160.—Pass a stream of sulphur dioxide gas through a solution of sodium or potassium hydroxide. Pass the gas until the solution smells strongly of SO_2; now add NaOH or KOH, drop by drop, until the solution is neutral to litmus paper. Concentrate the liquid so obtained and allow it to stand to crystallize.

What is the composition of these crystals?

Write all reactions involved in their formation.

What gas is liberated upon the addition of acids (HCl or H_2SO_4) to this salt?

Sulphuric Acid (H_2SO_4.)

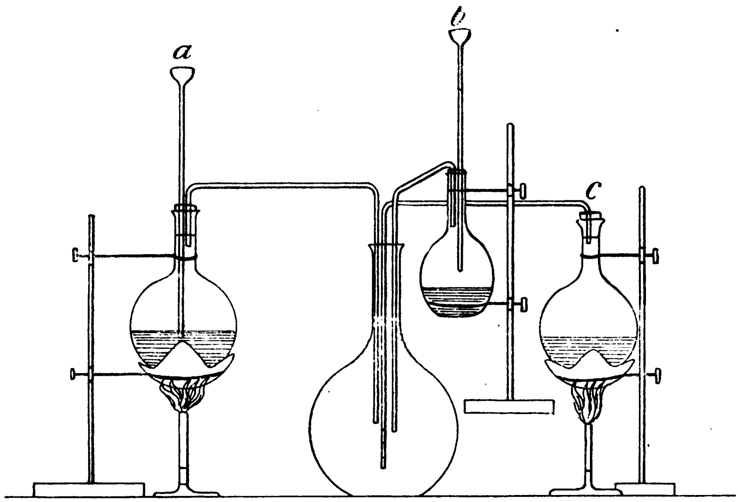
Experiment 161.—Place about 5 grams of flowers of sulphur in a flask and cover with concentrated nitric acid. Keep this mixture boiling gently for 15 or 30 minutes. (*Do this under the hood.*)

What is the action of nitric acid upon sulphur?

In answering this question, consider whether nitric acid is an oxidizing or reducing agent (see lecture notes on nitric acid); consider also the action of the resulting liquid on a solution of barium chloride, observing the facts that the precipitate is white, is insoluble in water and in acids.

Experiment 162.—Arrange an apparatus as shown in Fig. 37. In "a" place some copper scraps and cover with concen-

FIG. 37.



trated sulphuric acid; in "b" place a few scraps of copper turnings and nitric acid (1 to 1); and in "c" place water. Boil the water and cause the steam to pass into a litre flask; boil the Cu and H_2SO_4 and cause the gas to pass into the same flask; also pass in gas from Cu and HNO_3. When the brown fumes from the last are present in excess, discontinue their addition until the contents of the litre flask have again become colorless; then pass in the brown vapors once more, remove them when again in excess, and so continue alternately until an appreciable amount of liquid has collected at the bottom of the litre flask. Remove the three generators and pour the contents of the flask into a beaker. Pour one-half of the liquid into a clean evaporating dish, and evaporate to a syrupy consistency. Note the similarity between this and the concentrated sulphuric acid of commerce, "Oil of Vitriol."

Examine the behavior of the other half of the fluid toward litmus paper, sodium carbonate, and barium chloride. Record the result in each case and explain the action. Observe the action of the concentrated acid obtained above on a piece of filter paper.

In the light of the preceding experiment, give an explanation of the formation of H_2SO_4, by the method just employed.

Properties of Sulphuric Acid;

Experiment 163.—Observe the chief properties of sulphuric acid. Explain its action on organic substances.

Is it a bleaching agent?

Add dilute sulphuric acid to solutions of:—

Lead acetate ($Pb(C_2H_3O_2)_2$).
Barium chloride ($BaCl_2$).
Strontium chloride ($SrCl_2$).
Calcium chioride ($CaCl_2$).

What is the composition of the precipitates obtained?

Write all reactions.

Preparation of Vitriols.

Experiment 164.—*Green vitriol* (crystallized ferrous sulphate), $FeSO_4$, $7H_2O$.—Dissolve 25 grams of iron (nails) in about 200 c.c. dilute sulphuric acid (1–4), heat gently to aid the solution of the iron, and finally, when the action has almost ceased, pour into a clean evaporating dish. Filter if necessary. Allow to crystallize.

Write the reaction which occurs when iron dissolves in sulphuric acid.

Experiment 165.—*Blue vitriol* (crystallized copper sulphate), $CuSO_4$, $5H_2O$.—Dissolve 25 grams of copper in nitric acid (1–3), evaporate to remove the excess of the acid. Dilute with water, heat to boiling in a porcelain dish and add sodium carbonate solution until a precipitate no longer forms. Filter the black precipitate of oxide of copper and wash it with hot water. Finally dissolve this precipitate in hot dilute sulphuric acid, being careful to avoid an excessive quantity of the acid. Concentrate and allow to crystallize.

Write the reactions and state the reasons for the various operations performed in the preparation.

Experiment 166.—*White vitriol* (crystallized zinc sulphate), $ZnSO_4$, $7H_2O$.—Dissolve granulated zinc in dilute sulphuric acid, (1 : 4) keeping the zinc in excess so that the whole of the acid will be neutralized. Filter when the action has almost ceased and allow to crystallize.

Write the reaction involved in the solution of zinc in sulphuric acid.

Prove the presence of the water of crystallization in the vitriols.

Why do they bear the name "vitriols?"

Preparation of Alums.

Experiment 167.—*Common al'um*, $Al_2(SO_4)_3$, K_2SO_4, $24H_2O$. —Dissolve 20 grams of aluminium sulphate and 10 grams of potassium sulphate is not too large a quantity of boiling water. Filter the solution into a clean porcelain dish, protect from dust, and allow to stand until crystals form.

Experiment 168.—*Iron ammonium alum*, $Fe_2(SO_4)_3$, $(NH_4)_2$-SO_4, $24H_2O$.—Dissolve 20 grams of crystallized ferrous sulphate in water, add 3.5 grams concentrated sulphuric acid, stir and add a little concentrated HNO_3; heat to boiling, and continue the addition of the HNO_3 until the liquid is of a pure yellow or brown color and no further change seems to take place upon the addition of more acid. Evaporate the solution and add a filtered solution of ammonium sulphate (3.5 grams). Allow the liquid to crystallize.

Query.—Why was the nitric acid added?

Experiment 169.—*Chrome alum*, $Cr_2(SO_4)_3$, K_2SO_4, $24H_2O$.—Pass a stream of sulphur dioxide through a concentrated solution of potassium dichromate ($K_2Cr_2O_7$) containing H_2SO_4 until the deep red solution has changed color and smells strongly of SO_2. Allow the solution to stand in order to crystallize. If the solution is boiled or heated too hot, crystals cannot be obtained; hence make the solution sufficiently concentrated before passing the SO_2 gas, so as to do away with any necessity for evaporation after that operation. (Prepare the SO_2 from Cu and H_2SO_4.)

Why was the SO_2 employed?

Is its action in this case in accordance with its general character?

How could you show the presence of water of crystallization in the alums?

Write underneath one another the formulæ of these three alums, thus:—

$$Al_2(SO_4)_3,\ K_2SO_4,\ 24H_2O.$$
$$Fe_2(SO_4)_3,\ (NH_4)_2SO_4,\ 24H_2O.$$
$$Cr_2(SO_4)_3,\ K_2SO_4,\ 24H_2O.$$

Do you notice any relation between these formulæ; any reason why they should all be designated by the common term "alum?"

What is an alum?

Chromium. Acid Oxide (CrO_3).

Experiment 170.—Dissolve a few crystals of "chromic anhydride" (CrO_3) in water.

What acid is formed?

Add sodium or potassium hydroxide until the solution is neutral. Evaporate to small bulk and set aside to crystallize.

What compound has been formed?

Experiment 171.—Add some potassium hydroxide to a

solution of potassium dichromate, placed in a beaker, until the color changes from red to yellow.

Explain this change.

What compound has been formed?

Experiment 172.—To solutions of salts of silver, lead, mercury, bismuth, and barium placed in test-tubes, add respectively a solution of potassium chromate.

Note the characteristic chromate formed in each case.

Write all reactions.

Experiment 173.—To a solution of potassium dichromate ($K_2Cr_2O_7$) placed in an evaporating dish, add a little sulphuric acid (H_2SO_4) and then a few drops of alcohol. Boil gently and stir with a glass rod.

What changes take place?

Explain the action of the alcohol.

Write the reaction.

Chromium. Basic Oxide(Cr_2O_3).

Experiment 174.—Pour a solution of chromium sulphate into a beaker, boil and add a solution of sodium or potassium hydroxide as long as a precipitate forms. Filter and wash with hot water.

What is the composition of the precipitate?

What is contained in the solution?

How could you show this?

Experiment 175.—Heat a little of the precipitate on platinum foil.

What takes place?

What is the composition of the powder left on the foil?

Make a borax bead on a loop of platinum wire; when hot dip it into the powder, then hold in the flame again, fuse well, and note the color of the bead obtained.

Experiment 176.—Fuse some of the precipitate on platinum foil with a mixture of sodium carbonate and potassium nitrate. (*Use the dry substances.*) When well fused, cool and note the color. Place in the beaker, add water and acetic acid until the substance is dissolved; boil the liquid, and then add a solution of $Pb(C_2H_3O_2)_2$. What is obtained?

Explain all that has taken place chemically during this experiment.

Query.—Why was the acetic acid added?

Write the reaction.

Carbon (C.)

Experiment 177.—Heat some fragments of wood in a test-tube fitted with a perforated cork and bent tube. (See Fig. 7.) Observe the products formed.

Is the gas given off inflammable?

Examine the residue.

What is it?

Explain what you understand by the term "destructive distillation."

Experiment 178.—Repeat, using a piece of soft coal instead of the fragments of wood.

What is the residue left in the tube?

How do the products differ from those obtained in the preceding experiment?

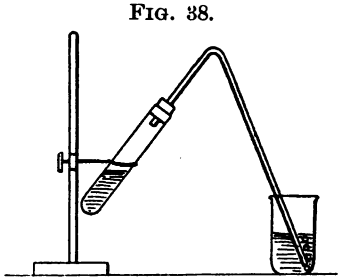
Experiment 179.—Place a little granulated sugar in a porcelain evaporating dish, pour on it some concentrated sulphuric acid (H_2SO_4) and note the result.

What is the black substance formed?

Explain the action of the sulphuric acid.

Experiment 180.—Mix together some copper oxide (CuO) and charcoal (C) and introduce into a test-tube fitted with a perforated cork and bent tube. Heat gently. (See Fig. 38.)

FIG. 38.



What gas is evolved?

Allow it to bubble through a solution of calcium hydroxide.

What takes place?

What does this experiment show?

Write all the reactions.

Experiment 181.—Examine a bit of charcoal (carbon, C).

Is it fusible?

Does it dissolve in any of the ordinary laboratory reagents?

Experiment 182.—Examine a piece of graphite (plumbago, black lead). Compare its properties with those of ordinary carbon (charcoal). Place a small piece in a combustion tube; heat strongly with a Bunsen burner; draw a current of air over it.

What is the product formed?

Experiment 183.—Repeat, using a piece of ordinary charcoal. Compare the results obtained.

What are the chief properties, physical and chemical, of the element carbon?

Carbon Dioxide.

(Carbonic anhydride) CO_2.

Experiment 184.—Arrange apparatus as in sketch. (See Fig. 39.)

FIG. 39.

Introduce a few pieces of marble ($CaCO_3$) into a flask and pour through the thistle bulb tube some dilute HCl.

Collect the gas evolved by downward displacement. Examine the gas obtained. Note color, odor, etc.

Introduce a lighted candle. Does it support combustion?

Pour a jar of the gas on a lighted candle placed in a beaker.

Is the gas heavier or lighter than air?

Experiment 185.—Pour a little lime-water ($Ca(OH)_2$) into a jar filled with gas, shake about and note the difference in appearance of the liquid.

Write the reaction.

Experiment 186.—Place some lime-water ($Ca(OH)_2$) in a beaker, blow through the solution for several minutes with a glass tube. (See Fig. 40.)

FIG. 40.

What takes place?

What gas is shown to be contained in the breath by this experiment?

Experiment 187.—Allow a small beaker containing a solution of calcium hydroxide ($Ca(OH)_2$) to stand for some time exposed to the air.

Explain the change that takes place.

What gas is shown to be contained in the air by this experiment?

Write the reaction that has taken place?

Experiment 188.—Pass a stream of CO_2 gas through a solution of sodium hydroxide until the solution is saturated. Boil the solution down to a small bulk and set aside to crystallize.

What are the crystals formed?

How could you show that the CO_2 has chemically combined?
Do they contain water of crystallization?
How could you show it?

Experiment 189.—Place some dry sodium carbonate in a tall glass cylinder. Pour on to it a little hydrochloric acid and observe the result. Introduce a lighted taper. Evaporate the residue in the jar to dryness, in a dish.

What is the powder obtained?

Properties of CO_2.

Experiment 190.—Collect a test-tube full of the gas over water. Place the thumb over the mouth of the tube and stand in a porcelain evaporating dish containing a solution of sodium hydroxide. Note the effect. Explain what has taken place.

Experiment 191.—Collect a jar of the gas by downward displacement, and introduce into it a strip of burning magnesium ribbon. Explain what has taken place.

Carbon Monoxide (CO).

Experiment 192.—Arrange an apparatus as shown using a flask instead of the first test-tube. (See Fig. 43.) Introduce into the flask 5 grams crystallized oxalic acid ($H_2C_2O_4$, $2H_2O$) and place in the test-tube a solution of caustic potash (KOH) or caustic soda (NaOH); pour in through the thistle-bulb tube 25 grams concentrated sulphuric acid (H_2SO_4), heat gently and gradually, and after the action has gone on for some time, light the gas as it issues from the small tube.

Caution.—A mixture of CO and air is explosive, therefore allow the gas to flow for some time before applying the flame

Test the gas with a small test-tube before lighting it.
Why was the alkaline liquid placed in the test-tube?
Could any other substance have been used?
Explain the experiment and write all reactions.

Experiment 193.—Place in a test-tube a little formic acid ($HCHO_2$). Pour on it a little concentrated H_2SO_4; shake and apply a match to the gas given off.

Does it burn?
Explain what has taken place.
Write the reaction.

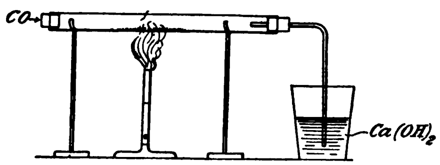
Experiment 194.—Place about 30 grams conc. H_2SO_4 in a 250 c.c. flask, add 10 grams $K_4Fe(CN)_6$ broken into small lumps. Warm gently and collect the gas evolved over water. Examine its properties.

Experiment 195.—Cork one end of the combustion tube, fill with water and stand upright in the pneumatic trough. Pass into this tube the gases evolved by the action of oxalic acid and H_2SO_4. When the tube is full of the mixed gases, remove and stand in a beaker containing NaOH. Observe the rise of the liquid in the tube.

What are the respective volumes of the two gases obtained by the action of the sulphuric acid on the oxalic acid?

Experiment 196.—Place a quantity of CuO in a piece of combustion tube. Arrange the apparatus as shown in the sketch. (See Fig. 41.) Pass a stream of CO over the CuO, heating the latter with a Bunsen burner with broad flame.

FIG. 41.



Place in a small beaker a solution of $Ca(OH)_2$, and as the action proceeds note the formation of a white precipitate.

What does this show?
Explain what has taken place.
Query: How could you tell that metallic copper is formed?

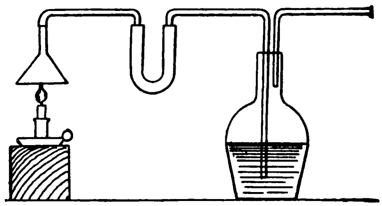
Properties of CO.

What have you ascertained the properties of CO to be?
How is it like to and does it differ from CO_2?

Chemistry of Combustion.

Experiment 197.—Arrange an apparatus as shown in the sketch. (See Fig. 42.) In the tube between the funnel and the flask place some pieces of calcium chloride; in the flask

FIG. 42.



place some "lime-water," $Ca(OH)_2$. Attach a rubber tube from the Chapman filter-pump, light the candle placed below the funnel, and slowly draw a current of air through the apparatus. Note the effect on the lime-water in the flask.

Pour a little lime-water into a test-tube; blow through it with a glass tube and note that the same result is produced.

What are the chief products of combustion?

How can the presence of each be shown?

Combustible and Incombustible.

Experiment 198.—Burn a piece of magnesium ribbon by holding it in the flame of a Bunsen burner.

Carefully collect the white powder formed.

Experiment 199.—Place the powder obtained in the last experiment in a little heap and apply the flame of a Bunsen burner to it.

What is the result?

Explain these two simple experiments, showing how they illustrate the terms "*Combustible*" and "*Incombustible.*"

Combustible and Supporter of Combustion.

Experiment 200.—

(*a*) Burn illuminating gas in air.

(*b*) Burn air in illuminating gas.

(*c*) Burn hydrogen in oxygen.

(*d*) Burn oxygen in hydrogen.
(Consult the Instructor.)

Ignition Point of Inflammable Substances.

Experiment 201.—
(*a*) Illuminating gas. (Use the Davy lamp.)
(*b*) Liquids: carbon bisulphide, ether, benzene, etc.
(*c*) Solids: gun-cotton, gun-powder, yellow and red phosphorus.
(*d*) Metals: magnesium, zinc, tin, platinum.
(Consult the Instructor.)

Meaning of the Term Combustion.

Experiment 202.—
(*a*) Broad sense of term. (Burn Sb in Cl; place P and I together.)
(*b*) Restricted sense of term. (P + O; Zn + O, etc.)
(*c*) Special sense of term. (C + O; carbonaceous material + O.)
(Consult the Instructor.)

Products of the Combustion of Carbonaceous Materials.

Experiment 203.—
Burn illuminating gas, wood, candle, oil, etc.
Examine products for:
(*a*) Carbon dioxide (CO_2).
(*b*) Water (H_2O).
(*c*) Carbon (C).
(Consult the Instructor.)

Query.—Why is the weight of the products greater than the original material?

What is the difference between combustion and destructive distillation?

Structure of Flame.

Experiment 204.—
(*a*) Flame of illuminating gas.
(*b*) Flame of oil from wick.
(*c*) Flame of candle.
(Consult the Instructor.)

Calorific Value of Fuels.

Experiment 205.—

(*a*) Gaseous fuels.

(*b*) Liquid fuels.

(*c*) Solid fuels.

(Consult the Instructor.)

Preparation of Carbonates.

Experiment 206.—"Zinc White."

Prepare by adding sodium carbonate to zinc sulphate, filtering, washing and drying the precipitate obtained. Keep in a box; label. Reaction:—

$$5ZnSO_4 + 5Na_2CO_3 + 3H_2O =$$
$$2ZnCO_3, 3Zn(OH)_2 + 5Na_2SO_4 + 3CO_2.$$

Experiment 207.—"White Lead."

Prepare by adding sodium carbonate to lead acetate, filtering, washing and drying the precipitate obtained.

Reaction:—

$$3Pb(C_2H_3O_2)_2 + 3Na_2CO_3 + H_2O =$$
$$2PbCO_3, Pb(OH)_2 + 6NaC_2H_3O_2 + CO_2.$$

Experiment 208.—Precipitate in test-tubes the carbonates of:

(*a*) Mg, Ca, Sr, Ba.

(*b*) Fe,* Al,* Cr.*

(*c*) Zn, Mn, Ni, Co.

Note the appearance of each.

Arsenic (As). Antimony (Sb). Bismuth (Bi).

Experiment 209.—Preparation of AsH_3.—Arrange two test-tubes as shown in sketch. (See Fig. 43.) In the first, place a few pieces of granulated zinc and a few drops of dilute sulphuric acid. In the second place a little water. After all the air has been expelled, light the gas that issues from the small tube. Add a few drops of an arsenic solution. Note the color

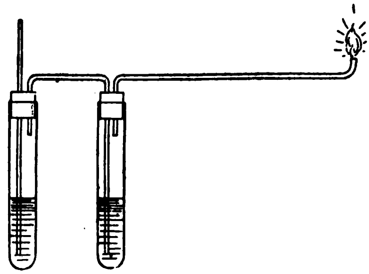
* These metals form hydroxides.

of the flame. Hold a piece of cold porcelain in the flame and note the deposit formed.

What is it?

Write the reactions.

FIG. 43.



Experiment 210.—Preparation of SbH_3.—Repeat the experiment, using a solution of an antimony compound instead of the arsenic.

Compare the results obtained.

Experiment 211.—To solutions of salts of arsenic, antimony and bismuth, placed in test-tubes, add respectively the following reagents and note the results produced in each case:—

Hydrogen sulphide	H_2S
Ammonium sulphide	$(NH_4)_2S$
Sodium hydroxide	$NaOH$
Ammonium hydroxide . . .	NH_4OH

Experiment 212.—Place a little bismuth chloride in a test-tube. Boil down to a small bulk and then fill the tube up with water.

What compound has the water precipitated?

Write the reaction that has taken place.

Silicon (Si).

Experiment 213.—Place some sodium silicate ("soluble glass") in a beaker. Dilute somewhat, and all hydrochloric acid little by little.

Note the effect.

What is the precipitate formed?

Decant off the liquid and heat some of the jelly-like substance in a porcelain dish, or on a sheet of platinum.

Note the result.

What is the white powder that remains?

Tin (Sn).

Experiment 214.—Examine a few fragments of granulated tin. Note the chief physical and chemical properties.

Experiment 215.—Prepare SnO_2 by adding strong nitric acid to some pieces of tin placed in a porcelain dish. Heat gently until the tin has all been converted to the oxide. Evaporate to dryness and examine the residue left. Test its solubility in H_2O, HCl, HNO_3.

Lead (Pb).

Experiment 216.—Examine a small piece of lead. Note its chief physical and chemical properties.

Experiment 217.—Heat a mixture of PbO and charcoal in a glass tube closed at one end. Explain what takes place chemically. Cool; break the tube and examine the metallic lead formed.

Experiment 218.—Precipitate, in test tubes, the following compounds of lead:—

PbS, ...	$Pb(C_2H_3O_2)_2 + H_2S$,
$PbSO_4$, ...	" $+ H_2SO_4$,
$PbCl_2$, ...	" $+ HCl$,
PbI_2, ...	" $+ KI$,
$PbCO_3$, ...	" $+ Na_2CO_3$,
$PbCrO_4$, ...	" $+ K_2Cr_2O_7$.

Write the reactions that have taken place.

Experiment 219.—Show that lead salts are a very delicate test for H_2S by holding a paper wet with a lead salt over a bottle containing H_2S.

Mercury. (Hg.)

Experiment 220.—Examine some mercury. Note its physical state. Ascertain some of its physical properties. Heat a little in a bit of glass tubing closed at one end and note the result. Examine some of the chemical properties of the metal. Test its solubility in nitric acid, hydrochloric acid, sulphuric acid, aqua regia, alkalies. Drop a fragment of zinc in a little mercury in a watch glass.

What change takes place in the zinc?

Define an amalgam.

Experiment 221.—Mix together equal parts powdered cinnabar and quicklime (mercury sulphide and calcium oxide). Place the mixture in a tube of hard glass, and heat for some time over a Bunsen burner. Note the result. Explain the reaction.

What compound remains in the tube?

Add a little hydrochloric acid to this residue.

What gas is given off?

Experiment 222.—Heat a little "red precipitate" (mercuric oxide) in a tube of hard glass. Note the results. Explain the action that takes place.

When was this experiment previously performed and for what purpose?

Distil off the mercury.

Experiment 223.—Pour a solution of mercuric chloride into a beaker. Suspend in it a small stick of zinc, let stand for some time and note the result.

How can you prove that the deposit formed is metallic mercury?

Repeat the experiment, using a piece of copper foil, and compare the results obtained.

Experiment 224.—To a solution of $HgCl_2$, add a solution of $SnCl_2$. Explain all that takes place.

Experiment 225.—To a solution of mercurous nitrate, add

dilute HCl. Note the precipitate obtained. Write the reaction that has taken place.

How does this compound differ from the compound of mercury used in the preceding experiment?

Boil a little of this precipitate with HCl and a crystal of $KClO_3$.

What takes place?

Experiment 226.—Pass a stream of H_2S gas through solutions of $Hg_2(NO_3)_2$, and $HgCl_2$, contained in separate beakers. Note the change of color that takes place as the precipitates form.

Can you explain the reason for this?

Pass the gas for some time until the precipitation is complete.

How can you tell whether the precipitation is complete or not?

Have you any reason to believe the two precipitates to be the same in composition?

Explain the reactions that have taken place. Filter the precipitates and wash with water. Treat portions of the precipitate with the following reagents: HCl, HNO_3, $(NH_4)_2S$.

What results are obtained?

Treat portions of both precipitates with "aqua regia" in evaporating dishes and boil.

Do they dissolve?

What compound is formed in both cases?

Prove this by adding $SnCl_2$ to each of the solutions.

Is the same compound formed in both cases?

Boron (Bo).

Experiment 227.—Hold a stick of meta-boric acid (HBO_2) in the flame of a Bunsen burner. Note the color.

Experiment 228.—Place in a porcelain dish a little powdered "borax" ($Na_2B_4O_7$). Cover with conc. H_2SO_4, add a little alcohol and light. Observe the color of the flame. Explain all that has taken place. Write the reaction.

Experiment 229.—Prepare borax beads on a loop of platinum wire, and try the "bead test" with salts of: Cu, Co, Ni, Fe, Cr. Note the characteristic color of each.

Query.—Why does the borax swell up when heated?

Aluminium (Al).

Experiment 230.—Examine some metallic aluminium. Note its chief physical and chemical properties. Try the action of KOH, HCl, and HNO_3.

Experiment 231.—To a solution of aluminium sulphate, add KOH carefully, stirring with a glass rod until a precipitate is obtained. Place a little of this in a test-tube, and see if it dissolves in an excess of KOH. Filter the precipitate, wash with water, and heat a little on platinum foil.

What is the powder obtained?

Write all reactions.

Experiment 232.—Dissolve a crystal of the "potash alum," obtained from a previous experiment, in water. Test the solution for aluminium as described above.

Query.—How could you show that the alum contains potassium; that it is a sulphate, and that it contains water of crystallization? Try these various tests.

Note.—Potash alum K_2SO_4, $Al_2(SO_4)_3$, $24H_2O$.

Magnesium (Mg).

Experiment 233.—Add to a solution of magnesium sulphate, ammonium hydroxide and ammonium chloride.

To clear the solution obtained, add a solution of Na_2HPO_4, sodium hydrogen phosphate, until a precipitate no longer forms.

Filter and wash.

What is the precipitate obtained?

What is contained in the solution?

Write the reaction.

Zinc (Zn).

Experiment 234.—Dissolve 10 grams of granulated zinc, placed in an evaporating dish, in dilute sulphuric acid.

When the zinc is nearly all dissolved, filter into a beaker, and add sodium carbonate as long as a precipitate forms.

Filter and wash with hot water.

Dry the precipitate on a steam bath.

What is its composition?

Write all the reactions that have taken place.

Query.—How could you show that the powder contains zinc; that it is a carbonate; that it is a hydroxide?

Cadmium (Cd).

Experiment 235.—Dissolve about 5 grams of cadmium chloride in water. Pass a stream of hydrogen sulphide gas through it until a precipitate ceases to form and the solution smells strongly of the gas. (*Do this under the "hood."*) Filter off the precipitate and wash with water. Dry on the steam bath and keep for future use. Note the color of the compound obtained and its appearance.

What is its composition?

Heat a little of it on charcoal with the blow-pipe and observe the metallic-looking tarnish formed on the coal.

(*Note.—This is a characteristic reaction for cadmium compounds.*)

Fuse a little of it on a stick of charcoal with dry sodium carbonate. Place a bit of the fused mass on a bright silver coin and moisten with a drop of water.

How does the result show that sulphur was present in the compound?

Dissolve some of the powder obtained in the preceding experiment in nitric acid.

What compound of cadmium is formed?

What gas is given off?

Evaporate until a concentrated solution is obtained and crystallize the compound.

Dissolve some of the crystals in a little water, and add a few drops successively to solutions of potassium hydroxide, ammonium hydroxide, sodium carbonate, and sodium phosphate. Note the precipitate obtained in each case.

What does the solution contain in each of these tests?

Calcium. (Ca.) Strontium. (Sr.) Barium. (Ba.)

Experiment 236.—To solutions of $CaCl_2$, $SrCl_2$, and $BaCl_2$, placed in test-tubes, add a solution of NaOH and observe the precipitates formed in each case.

What are the compositions of these precipitates?

How could oxides be prepared from them?

What is formed in the solution in each case?

How could you prove it?

Experiment 237.—In like manner add a solution of $(NH_4)_2CO_3$.

What are the precipitates formed in each case?

How could you prove it?

What change in composition takes place when these precipitates are strongly heated?

Try it on a piece of platinum foil. Try the effect of hydrochloric acid (HCl) on these precipitates.

Experiment 238.—Preparation of barium chromate. To a solution of $BaCl_2$ add a solution of K_2CrO_4 as long as a precipitate forms. Filter, wash, and dry.

What is the composition of the powder obtained?

Write the reaction.

Experiment 239.—Try the flame test with compounds of Ca, Sr, and Ba. Use a platinum wire.

Sodium. (Na.) Potassium. (K.) Ammonium. (NH_4.)

Experiment 240.—Examine a solution of caustic soda (NaOH). Note its action on litmus paper. Pour some into an evaporating dish and add, while stirring with a glass rod, dilute hydrochloric acid until the solution has an acid reaction. Evaporate to dryness.

What is the powder obtained?

Is it caustic soda?

If not, why?

Experiment 241.—Repeat the above experiment, using caustic potash (KOH) instead of caustic soda.

How does the result differ?

Experiment 242.—Try the flame tests on sodium and potassium compounds. Use a platinum wire.

What characteristic colors do they impart to the flame?

Experiment 243.—Place a solution of ammonium chloride in a test-tube. Add a solution of NaOH or KOH until alkaline. Warm gently.

What takes place?

What gas is given off?

How could you show it?

Write the reaction.

Experiment 244.—To solutions of NaCl, KCl, and NH_4Cl, placed in test-tubes, add respectively a few drops of $PtCl_4$.

What takes place in each case?

Write the reactions.

Manganese. (Mn.)

Experiment 245.—To a solution of manganese salt, such as $MnSO_4$, add KOH and filter. Note the color of the precipitate obtained and the action of the air on this compound on standing.

Fuse a little of this precipitate on platinum foil with a mixture of Na_2CO_3 (dry), and KNO_3 (dry). Note the color of the result.

What has taken place?

Drop the platinum foil and fused mass into a beaker of water, and boil. Note the color of the solution obtained.

What has taken place?

Study MnO_2. See previous work on Cl and O.

Iron. (Fe.)

(a) *Ferrous Compounds.*

Experiment 246.—Examine a crystal of ferrous sulphate ("green vitriol").

Heat a fragment in a test-tube, at first gently, and then strongly.

What takes place?

Dissolve about 5 grams of ferrous sulphate in boiling water. Filter, and use the solution in the following experiments:—

Experiment 247.—To the solution of ferrous sulphate placed in test-tubes, add the following reagents:—

1. NaOH (or KOH).
2. Na_2CO_3.
3. $(NH_4)_2S$.
4. Na_2HPO_4.
5. $K_4FeC_6N_6$.
6. $K_3FeC_6N_6$.

What is formed in each case?

Make a table of the precipitates, giving their compositions, color, etc.

Experiment 248.—Make a bead test, using borax on a loop of platinum wire, with the first of the above precipitates.

What is the color of the bead obtained?

(b) *Ferric Compounds.*

Experiment 249.—Examine a crystal of "iron ammonium alum" (ferric ammonium sulphate).

Heat a fragment in a test-tube, first gently, then strongly in the flame of a Bunsen burner.

What takes place?

What is the powder that remains?

Save some of this powder for the bead test.

Experiment 250.—Dissolve about 5 grams of the "iron ammonium alum" in boiling water, filter, and use the solution in the following experiments:—

Experiment 251.—To the solution from the last experiment, placed in test tubes, add the following reagents:—

1. NaOH (or KOH).
2. NH_4OH.
3. Na_2CO_3.
4. $(NH_4)_2S$.
5. NH_4CNS.
6. Na_2HPO_4.
7. $K_4FeC_6N_6$.
8. $K_3FeC_6N_6$.

What is formed in each case?

Make a table of the precipitates, giving their composition, color, etc.

Experiment 252.—Make a bead test, using borax on a loop of platinum wire, with the powder obtained in first experiment. (Experiment 249.)

What is the color of the bead obtained?

Query.—In what ways can ferrous and ferric compounds be distinguished from one another?

What is the most delicate test for a ferrous compound? for a ferric?

Make a comparative table of these reactions, showing in parallel columns the results obtained from ferrous and ferric compounds.

Cobalt. (Co.) Nickel. (Ni.)

Experiment 253.—To solutions of cobalt nitrate and of nickel nitrate placed in test-tubes, add respectively the following reagents:—

1. KOH (or NaOH).
2. NH_4OH and NH_4Cl.
3. Na_2CO_3.
4. $(NH_4)_2S$.
5. Na_2HPO_4.
6. KNO_2 and $HC_2H_3O_2$.

Note the precipitates formed in each case, their compositions, colors, etc.

Experiment 254—Make two borax beads on a loop of platinum wire.

Make bead tests, using cobalt hydroxide and nickel hydroxide, respectively.

Note the characteristic colors of the beads obtained.

Copper. (Cu.)

Experiment 255.—Examine some pieces of copper. Note color, hardness, etc. Try the solubility of copper in the common mineral acids. Hold a steel knife blade in the solution obtained.

What takes place?

What is the best solvent for this metal?

Heat a small piece of copper on a stick of charcoal.

Is any coating formed?

What takes place?

Experiment 256.—To a solution of a copper salt, add sodium carbonate as long as a precipitate forms. Boil the contents of the beaker until the precipitate turns black.

What takes place chemically?

Filter, dry, and save the powder. Make a borax bead and from this powder make a bead test.

What is the color of the bead obtained?

Experiment 257.—To a solution of copper sulphate placed in a beaker, add a few crystals of tartaric acid. Stir until dissolved, add an excess of NaOH, and then add a small piece of glucose. Boil.

What takes place?

What is the composition of the precipitate formed?

What are the formulæ and colors of the oxides of copper?

Experiment 258.—To a solution of copper sulphate placed in test-tubes, add respectively the following reagents:—

1. NaOH (or KOH).
2. NH_4OH.
3. $NaCO_3$.
4. $(NH_4)_2S$—(also H_2S).
5. Na_2HPO_4.
6. $K_4FeC_6N_6$.

Note the precipitate formed in each case; color, composition, etc.

Do copper compounds impart any characteristic color to a flame?

Silver (Ag).

Experiment 259.—To a solution of $AgNO_3$ add dilute HCl. Note the precipitate formed. Note the reaction that has

taken place. Note action of NH_4OH and of HNO_3 on this precipitate.

Heat some of this precipitate on a stick of charcoal with the blow-pipe.

What takes place?

Experiment 260.—Suspend a stick of zinc in a solution of $AgNO_3$.

What takes place?

Examine the deposit formed.

Is it soluble in HNO_3?

What is the best solvent for silver.

Gold (Au).

Experiment 261.—Note the action of HCl and of HNO_3 separately on leaf gold. (Use concentrated acids.)

Mix these together.

Is the gold dissolved? Why?

What is the best solvent for gold? Why?

Evaporate the solution in an evaporating dish until most of the free acid is expelled. Dilute with water and divide in two portions. To one add a freshly prepared solution of $FeSO_4$, to the other a solution of oxalic acid ($H_2C_2O_4$).

What takes place in each case?

Fuse the precipitates obtained on a stick of charcoal with the blow-pipe.

What takes place?

Platinum (Pt).

Experiment 262.—Note the action of HCl (conc.) and HNO_3 (conc.) on scraps of platinum.

Mix them together, warm, and note the result.

What is the best solvent for platinum? Why?

To a solution of $PtCl_4$ add some KCl. Note the precipitate. Write the reaction. Repeat using NH_4Cl instead of KCl. Note the similar precipitate obtained. Write the reactions. Evaporate each of these precipitates and ignite.

What is left behind in each case?

Note.—**Save all gold and platinum residues.**

APPENDIX.

APPENDIX.

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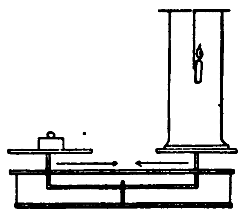
APPENDIX.

I.

LAW OF THE CONSERVATION OF MATTER.

(1.) Weigh a cylinder with ground top, containing a candle fastened in a deflagrating spoon and covered with a square of ground glass. (See Fig. 45.)

FIG. 45.



Record the weight obtained.

Remove ground glass; light the candle and immediately replace the glass. Allow to stand until the flame of the candle is extinguished by the accumulation of gases produced by the combustion. When apparatus is cold, weigh again.

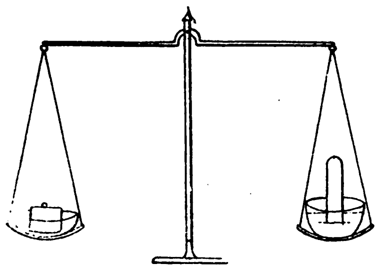
Record the weight obtained.

Note the diminution in size of the candle.

(2.) Place on the scale pan of a balance, a porcelain dish containing dilute sulphuric acid. In this dish invert a test-tube which is also filled with dilute sulphuric acid, and is kept in an upright position by a leaden ring slipped over it.

(See Fig. 46.) Also place on the scale pan a small fragment of zinc (about 0.1 gram).

FIG. 46.



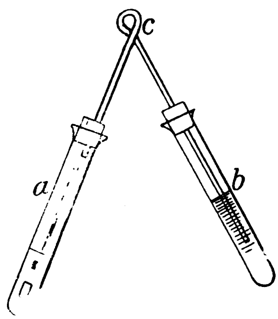
Weigh the whole and record the weight.

Drop the zinc into the dish of acid and *immediately* place the inverted test-tube directly over it. Note the disappearance of the zinc and the formation of a gas.

After the action is over weigh again and record the weight.

(3.) Arrange an apparatus as shown in Fig. 47. The tube "*b*" should have an outlet. The tube "*a*" contains hydrochloric acid, the tube "*b*" potassium hydroxide. Counterbalance the apparatus, and weigh with it a small fragment of calcite.

FIG. 47.



Remove the tube "*a*," drop into it the fragment of calcite and *immediately* replace the cork.

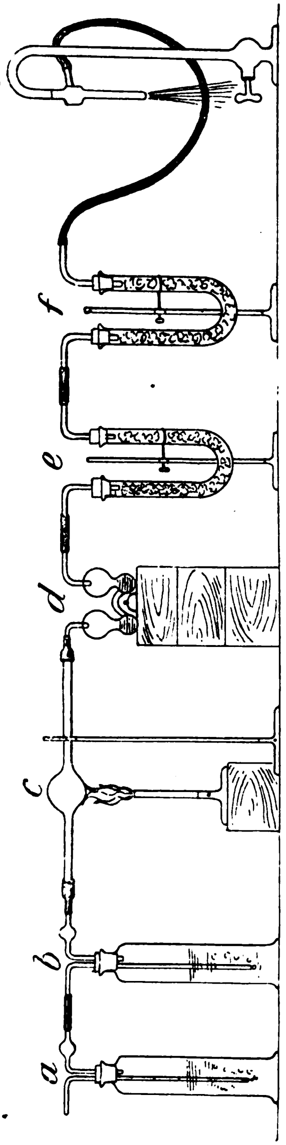
Note the action that takes place.

Is there any loss or gain in weight?

(4.) Arrange an apparatus as shown in Fig. 48.

Perform the experiment as described in Mendeléeff's "Principles of Chemistry," Vol. I, *Introduction*.

FIG. 48.



a. Contains Conc. H_2SO_4.
b. Contains KOH solution.
c. Contains "Azurite" (Hydrated copper carbonate).
d. Contains Conc. H_2SO_4.
e. Contains KOH, solid.
f. Contains $CaCl_2$.
g. Aspirator.
Record all weights as directed.

Tabulate the results of the above four experiments in the following form:—

Experiment.	Total weight before.	Total weight after.	Loss.	Gain.
(1.)	. . . .	. . . .	. . . .	. . . .
(2.)	. . . .	. . . .	. . . .	. . . .
(3.)	. . . .	. . . .	. . . .	. . . .
(4.)	. . . .	. . . .	. . . .	. . . .

What conclusion do you draw from the above results?

State this conclusion in the form of a general law.

How do you explain the small loss or gain that may have been obtained in performing the above experiments?

Why do these not modify the general conclusion?

II.

LAW OF DEFINITE PROPORTIONS.

(Law of Lavoisier.)

(*a*) Formation of Magnesium Oxide:

(1) By burning Mg in air or oxygen.

Clean a piece of magnesium ribbon by rubbing with a piece of charcoal. Accurately weigh 0.5 gram, or less, of this magnesium ribbon cut into small pieces. Introduce the weighed metal into an ignition bulb tube which has been previously weighed together with a plug of glass wool, which should slightly stopper the end of the tube. Heat at first gently and afterwards strongly with a Bunsen burner until the metal is completely converted into oxide. Cool and weigh. Again heat the contents of the tube, cool and weigh it. (A layer of asbestos will keep the tube from breaking.)

Repeat until constant weight is obtained, thereby insuring the complete conversion of the metal into the oxide.

(See Newth's "Elementary Practical Chemistry," Chap. xi, pp. 104–108.)

(2) By dissolving Mg in dilute nitric acid, and igniting the nitrate obtained.

Weigh an ignition bulb tube and plug of glass wool as in the previous experiment. Introduce into the tube an accurately weighed quantity of clean, bright magnesium ribbon. Take about 0.2 grams. Remove the glass wool plug and add nitric acid a few drops at a time until the metal is all dissolved. Heat gently with a Bunsen burner and evaporate to dryness. Now replace the plug of glass wool and heat strongly until the nitrate is entirely decomposed and brown fumes are no longer given off. Cool and weigh.

To insure complete decomposition of the nitrate into the oxide, again heat, cool and weigh.

Repeat this until constant weight is obtained.

Record all weights in the above experiments.

(*b*) Formation of Silver Chloride.

(1) By precipitation of silver nitrate with hydrochloric acid.

Weigh accurately about 0.5 gram of pure silver. Place in a small covered porcelain crucible which has been weighed, and dissolve in a little dilute nitric acid. Add 2–3 c.c. dilute HCl. Evaporate to dryness on the water-bath. Fuse dry residue over a Bunsen burner. Allow to become cold and weigh. Record all weights.

(2) By precipitation of silver nitrate with sodium chloride.

Accurately weigh about 0.5 gram pure silver. Place in an Erlenmeyer flask and add enough dilute nitric acid to dissolve it. Add enough solution of sodium chloride completely to precipitate the silver chloride. Add water. Cork the flask and shake thoroughly. Allow the silver chloride to settle, and pour off the clear supernatant liquid. Add water again and shake as before. Repeat several times, until the supernatant liquid poured off shows the absence of sodium chloride.

How can this be ascertained?

Remove the cork from the flask and place over the mouth of the flask a porcelain crucible that has been weighed. Invert and place in a vessel half-filled with water. Gently tap the sides of the flask until all the silver chloride has settled into the crucible. Remove the flask carefully and lift out the porcelain crucible containing the silver chloride. Carefully pour off as much water as possible and evaporate to dryness on the water-bath. Ignite as before, cool and record weight. Always heat to constant weight.

(*c*) Formation of Copper Sulphide.

Accurately weigh an ignition bulb tube. Introduce about 1 gram (also accurately weighed) of pure copper wire cut into small pieces.

Introduce 1 grm. of flowers of sulphur into the tube, and shake so that the sulphur is well mixed with the copper.

Heat strongly over the Bunsen burner until all action is over and the excess of the sulphur is all expelled.

Arrange the tube as shown in Fig. 11. When the tube has cooled, weigh. Record all weights.

Repeat the experiment using various amounts of sulphur (*e. g.*, 2, 3 or 4 gms.) and the same amount of copper.

Record all weights. Always heat to constant weight.

Tabulate the results of experiments (*a*), (*b*) and (*c*) in the following form :—

Experiment.	Compounds Formed.	Weight of Compounds Formed.	A. Weight of Metal Used.	B. Weight of Substance Combined with Metal.	Ratios A.: B.
(*a*) 1 . . .	Magnesium oxide.				
(*a*) 2 . . .	Magnesium oxide.				
(*b*) 1 . . .	Silver chloride.				
(*b*) 2 . . .	Silver chloride.				
(*c*) 1 . . .	Copper sulphide.				
(*c*) 2 . . .	Copper sulphide.				
(*c*) 3 . . .	Copper sulphide.				
(*c*) 4 . . .	Copper sulphide.				

(*d*) Formation of Sodium Chloride.

Weigh about 10 grams of sodium hydroxide and dissolve in about 200 c.c. of water and place in a cork-stoppered bottle. Into another glass-stoppered bottle pour 20 c.c. strong HCl and add about 200 c.c. water. Label each bottle.

By means of a graduated glass cylinder pour about 10 c.c. of the sodium hydroxide solution in a weighed porcelain dish. Add about 20 c.c. of the hydrochloric acid solution.

This will give an excess of HCl above the amount necessary to form sodium chloride, and this acid being volatile the excess will be expelled in the subsequent ignition.

Evaporate, gently, the solution in the dish to dryness. Use great care to avoid any possible spattering of the substance. Now heat strongly to expel any moisture that remains. Ignite to constant weight.

Repeat, using the same amount of the NaOH and also different amounts of the acid (e. g., 15, 20 or 30 c.c.). Record all figures and tabulate the results obtained in the form of the ratio of the quantity of sodium chloride produced to the volume of sodium hydroxide solution used.

Show that the ratio of the weight of the sodium chloride obtained to a given quantity of solution of sodium hydroxide from which it was found, is always a *constant*—C.

$$\text{Thus: } \frac{\text{Wgt. NaCl}}{\text{Vol. NaOH solution}} = \text{C.}$$

What conclusion do you draw from the results of the preceding four experiments?

State this conclusion in the form of a general law.

How do you explain any apparent inconsistencies?

Why do not these modify the general conclusions?

III.

LAW OF MULTIPLE PROPORTIONS.

(*Law of Dalton.*)

(*a*) 1 and 2.—Two compounds of copper and oxygen are known, one red and the other black, which have been found on chemical analysis to have different compositions.

Accurately weigh about .5 gram of each in weighed porcelain boats. Carefully introduce the boats into a long glass combustion-tube and pass over them a stream of dry hydrogen gas. Heat the tube gently at first and afterwards strongly until the reduction is complete.

How may this be ascertained?

Allow the tube to cool in a stream of hydrogen gas. Why?

When cold carefully remove the porcelain boats and again weigh. Record all weights.

(*b*) 1 and 2.—Repeat the above experiment, using two different oxides of lead. Use the yellow and the brown oxides. Record all weights.

Tabulate the results of the above two experiments in the following form:—

Experiment.	Name Oxide.	Weight Oxide.	Name Metal.	A. Weight Metal.	B. Weight Oxygen Combined + A.	Ratio A. : B.
(*a*) 1 . .						
(*a*) 2 . .						
(*b*) 1 .						
(*b*) 2						

What ratios exist between the amounts of B which combine with a fixed amount of A?

What conclusion do you draw from these results?

State these conclusions in the form of a general law.

How do you explain any apparent inconsistencies, and why do not these modify the general conclusion?

AN EXPERIMENT FROM WHICH CAN BE OBTAINED BY INDUCTION BOTH THE LAWS OF DEFINITE AND OF MULTIPLE PROPORTIONS.

Action of an Acid on a Solution of Sodium Carbonate.

Dissolve about 10 grams of dry sodium carbonate in about 500 c.c. of water. Place in a glass stoppered bottle. Label "Sodium carbonate solution."

Add 50 c.c. of conc. sulphuric acid to 500 c.c. of water. Stir thoroughly and when cool place in a glass-stoppered bottle and label "Dilute sulphuric acid."

Thoroughly clean two burettes. Label the one "sodium carbonate" and the other "sulphuric acid."

Rinse out each with the respective solutions prepared.

Fill the burettes with the prepared solution.

Titrate the one against the other, using:

1. Cochineal or methyl orange as an *indicator*.

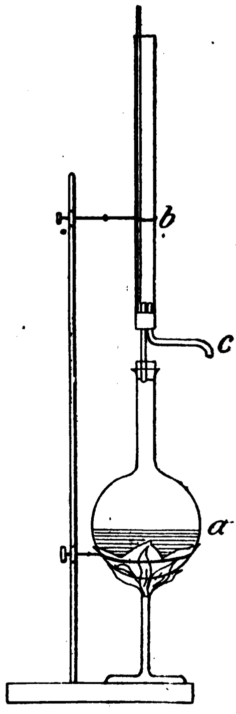
2. Phenol phthalein as an *indicator*.

Record all results.

(Consult the instructor for details of this work.)

Tabulate results in columns, showing that in the first case the compound formed may be represented by AB, and in the second case by A_2B, thereby demonstrating the Laws of Definite and of Multiple Proportions.

FIG. 49.



IV.

LAWS GOVERNING THE VOLUMES OF GASES.

(*a*) Influence of Temperature on Gas Volumes.

(*Law of Charles.*)

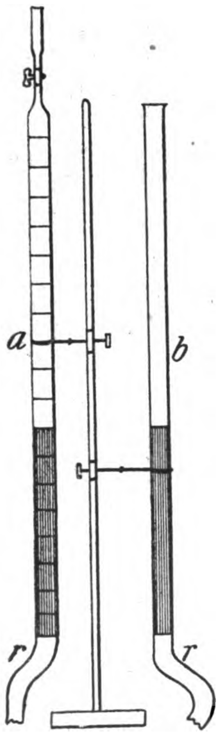
Arrange an apparatus as shown in Fig. 49.

The small tube should be of an internal bore of about 2 mm. diameter and perfectly dry. A little cylinder of mercury is introduced into this tube and one end sealed in the flame of a Bunsen burner. The outer tube is the ordinary piece of combustion tubing about 45 c.m. long and 1.5 c.m. diameter. Lay the small tube on the table and accurately measure with a metre stick the space occupied by the enclosed volume of air between the index of mercury and the sealed end of the tube. Now place the small tube within the upright combustion-tube. Boil the water in the flask beneath and after the steam has passed through the tube for some length of time, again measure the volume of air enclosed in the small tube.

Repeat this several times, using different volumes of air.

Record all results; note the temperature of the room.

FIG. 50.



Express results according to the proportion—

$$V : V' :: T : T'.$$

To ascertain accuracy of results divide V′ by V and T′ by T.

State the Law of Charles.

Subject a certain volume of air to different temperatures by means of the vapors of boiling liquids, other than water, *e. g.*, chloroform, benzene, alcohol, etc. Also subject this volume, by immersion in a vessel containing melting ice, to a temperature of 0° C.

Plot these results to a convenient scale on cross-section paper, the absolute temperatures being placed on the vertical axis and the volumes on the horizontal axis.

From the form of the curve so obtained, state the Law of Charles.

(*b*) Influence of Pressure on Gas Volumes.

(*Law of Boyle.*)

Arrange an apparatus as shown in Fig. 50.

Open stop-cock at top of tube "*a*." Lower the tube "*b*" until the mercury is at the same level in both tubes and the tube "*a*" contains a convenient volume of air. Read this volume. Now close the stop-cock. Read the barometric pressure of the room in terms of mercury. Raise or lower the tube "*b*" and measure the distance between the mercury levels. This figure added to or subtracted from the atmospheric pressure gives the increased or diminished pressure to which the volume of air is now subjected. Read the volume of the enclosed gas under this increased or diminished pressure.

Repeat several times, using different volumes of air.

Record all results.

Express results according to the proportions—

$$V : V' :: P' : P.$$

To ascertain accuracy of results divide V by V′ and P by P′, or multiply V by P.

State the Law of Boyle.

Assuming Boyle's Law and given one litre of a gas at 1 atmosphere pressure, plot the results on cross-section paper, which would be obtained if the pressure were made successively 2, 3 and 4 atmospheres. Join the points so obtained by a smooth curve. This curve is the graphical expression of the Law of Boyle.

Note that the area enclosed by any volume and its corresponding pressure is always the same.

$$PV = C \ (a\ constant.)$$

(*c*) Combination of Gas Volumes by Simple Ratios.

(*Laws of Gay Lussac.*)

(1) *Gases combine in simple proportions by volume.*

(2) *The volume of the gaseous compound obtained always bears a simple proportion to the volumes of gases from which it has been formed.*

Examples:

(1) 2 vols. Hydrogen + 1 vol. Oxygen = 2 vols. Water (as gas).

(2) 1 vol. Hydrogen + 1 vol. Chlorine = 2 vols. Hydrochloric acid.

(3) 2 vols. Carbon monoxide + 1 vol. Oxygen = 2 vols. Carbon Dioxide.

(4) 1 vol. Nitrogen + 3 vols. Hydrogen = 2 vols. Ammonia.

The reverse of these statements is also true.

(For details of laboratory work on the above see Richter, "Inorganic Chemistry," p. 78.)

HYPOTHESIS OF AVOGADRO.

In proposing a hypothesis to explain empirical facts, the one is to be given preference that most simply explains the facts. Then, if further facts require a modification it is not necessary to reject the hypothesis but simply to add to it.

Avogadro following this idea proposed in 1811 a hypothesis to explain and account for the very simple laws which govern gases. His hypothesis was as follows:

"*Equal volumes of gases measured under like conditions of temperature and pressure contain an equal number of ultimate particles or molecules.*"

(For fuller details concerning this hypothesis read Richter, "Inorganic Chemistry," p. 73; Ramsey, "Chemical Theory," chap. VII.)

V.

DETERMINATION OF MOLECULAR WEIGHTS.

(*a*) Vapor Density Method.

Law of Avogadro.—Equal volumes of all gases under like conditions of temperature and pressure contain an equal number of molecules.

The specific gravity of a substance in the form of a gas, when referred to hydrogen as unity, termed its vapor density, is one-half its molecular weight. Thus,

$$H (1) : \text{Vapor Density} :: H_2(2) : \text{Molecular Weight},$$

or,

$$\text{Molecular Weight} = 2 \times \text{Vapor Density}.$$

Determine the molecular weight of water by means of its

vapor density, using the method of Victor Meyer. (See Whiteley, "Organic Chemistry," page 26.)

(*b*) Lowering of Freezing Point.

(Raoult's Cryoscopic Method.)

Law of Raoult.—Molecular quantities of different substances, non-electrolytes, dissolved in the same amount of a solvent, cause the same depression in the freezing point of that solvent.

Let t° = depression produced by "p" grams of substance in 100 grams of the solvent; then the coefficient of depression $\frac{t^\circ}{p}$ will be the depression for 1 gram of substance in 100 grams of solution.

The molecular depression is the product obtained by multiplying $\frac{t^\circ}{p}$ by the molecular weight of the substance. This is constant for all substances having the same solvent:—

$$M \times \frac{t^\circ}{p} = C \therefore$$

$$M = C \times \frac{p}{t^\circ}$$

M = molecular weight,
C = constant.

Table of constants:

Water	19
C_6H_6	4.9
Glac. Acet. Acid	39
C_2H_5OH	76
Naphthalene	70

Determine the molecular weight of cane sugar by Raoult's Cryoscopic method.

(See Perkins' "Organic Chemistry," page 50, and Whitely, "Organic Chemistry," p. 30.)

(*c*) Chemical Analysis.

Analyze Silver Acetate, ($AgC_2H_3O_2$). Note the weight of Ag obtained from a given weight of $AgC_2H_3O_2$.

As a rule, the chemical analysis of a substance will give only its empirical formula, *i. e.*, the simplest formula.

Physical methods establish the true molecular formula.

VI.

DETERMINATION OF EQUIVALENT WEIGHTS.

The equivalent weight of a metal is that weight which replaces unit weight of hydrogen.

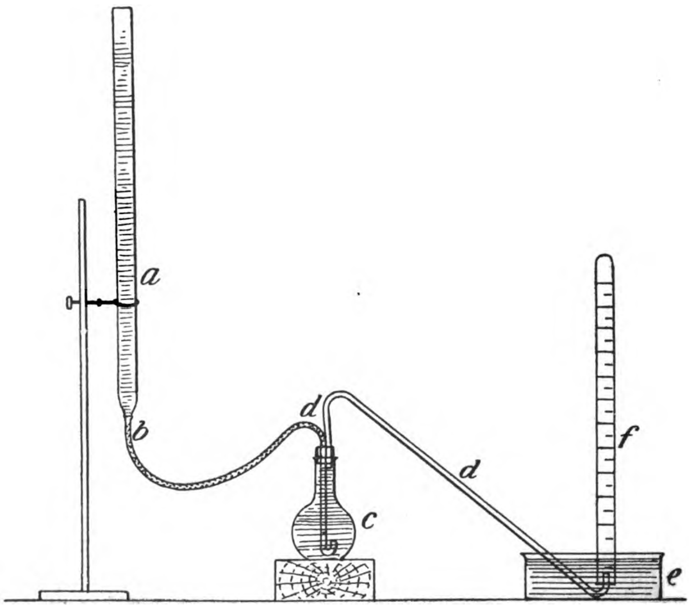
It is obtained by dividing the quantity of the metal taken by the weight of the hydrogen liberated by its action on an excess of acid, or from the weight of oxygen with which a given quantity of the metal combines.

The equivalent weight is either the atomic weight or a multiple of it.

(*a*) Amount of Hydrogen Liberated.

Determine the equivalent weight of zinc by determining the weight of hydrogen liberated from an acid by a known weight of the metal. Arrange an apparatus as shown in Fig. 51. (See Remsen, "Introduction to Chemistry," Chap. XIII.)

FIG. 51.

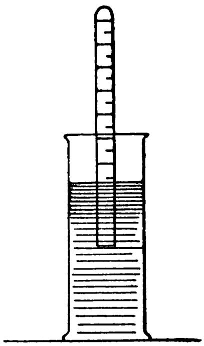


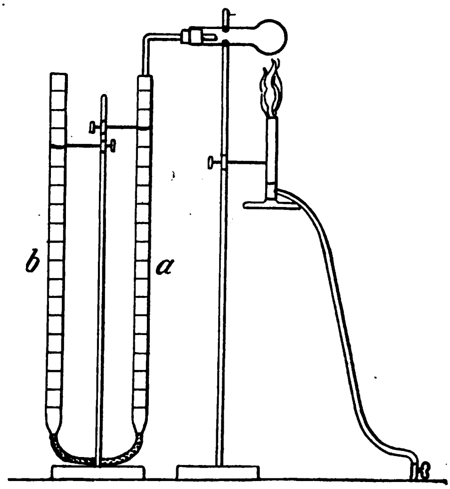
13

Read the volume of the liberated gas as shown in Fig. 52.

FIG. 53.

FIG. 52.





In like manner determine the equivalent weights of:—

Magnesium.

Cadmium.

Aluminium.

Iron.

(*b*) Amount of Oxygen combined.

Determine the equivalent weight of tin from the formation of its oxide. Heat a weighed quantity of tin in air.

VII.

DETERMINATION OF ATOMIC WEIGHTS.

(*a*) By analysis of a compound, given its formula.

1. Weigh accurately about .5 gram of mercury oxide in a weighed bulb tube of hard glass. Arrange an apparatus as shown. (See fig. 53.) Heat until all the oxygen is expelled and the volume remains constant.

Read this volume and reduce it to N. T. P. (0° C. and 760 mm.). When bulb tube is cold, weigh and determine weight of metallic mercury.

Determine the weight of the oxygen obtained.

Observe the relative volumes of hydrogen and oxygen which result on the decomposition of water with the electric current. (Use Hofmann apparatus.)

Tabulate results as follows:—

Observe weight of bulb tube		
Observe weight of bulb tube + mercury oxide		
Observe weight of mercury oxide		
Observe weight of bulb tube + mercury		
Observe weight of mercury		
Compute weight of oxygen		
Observe volume of oxygen		
Observe barometric pressure		
Observe temperature		
Compute corrected volume of oxygen		
Compute weight of one (1) litre of oxygen		
Given weight of one (1) litre of hydrogen (N. T. P.)	.0896 gms.	
Observe volume composition of water		
Compute weight composition of water, *per cent.*		
Determine molecular weight of water		
Compute molecular formula for water		
Assume atomic weight of hydrogen	1	
Compute atomic weight of oxygen		
Assume molecular formula of mercury oxide	HgO	
Compute atomic weight of mercury		

2. In like manner repeat, using potassium chlorate ($KClO_3$), and ascertain the weight of 1 litre of Oxygen gas (N. T. P.) using a similar apparatus.

(*b*) Method of Specific Heat.

If to equal masses of different substances, a given quantity of heat be added, they will in general assume different temperatures; or, if equal masses of different substances be heated through the same range of temperature, there will be in general different quantities of heat absorbed.

The number of calories required to raise the temperature of a body 1° C. is called the heat capacity of that body. The ratio of the heat capacity of a body to the heat capacity of an equal weight of water is called the specific heat of that body. It follows from this definition that the specific heat of water is unity.

Law of Dulong and Petit: The product specific heat × *atomic weight* = 6.4 *for all elements in the solid state.*

This constant, 6.4, is called the *atomic heat.*

The product (molecular weight × specific heat) is called the *molecular heat.* It has been found to be equal to the sum of the atomic heats of the constituent elements.

From the Law of Dulong and Petit it follows:

$$\text{Atomic weight} = \frac{6.4}{\text{Specific Heat}}$$

Hence the atomic weight of an element may be obtained if the specific heat in the solid state is known.

Arrange an apparatus as shown in Figs. 53 and 54. (Fig. 55 shows the gas regulator.)

FIG. 54.

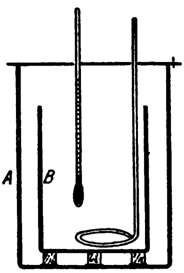


FIG. 55.

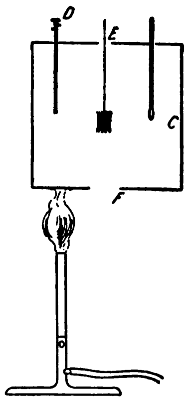
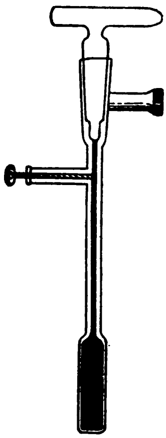


FIG. 56.



Gas Regulator.
(Reichert.)

Description of the Apparatus.

A large beaker or jar, *A*, with three cork supports, fastened with sealing-wax.

A beaker, *B*, of 250 c.c. capacity.

2 delicate thermometers.

A glass stirrer, bent into a ring at the base.

A heavy card-board cover.

A tin can *C*, carrying a gas regulator *D*, and having two holes, *E* and *F*; *F* being large enough to allow the bundle of zinc sticks to be lowered through it. The can, *C*, should be supported at such a height that *A*, can be brought beneath it.

Determination of Water-equivalent of the Material of the Calorimeter.

Weigh the beaker *B*, the stirrer and the thermometer to decigrams and multiply this weight by 0.198 the specific heat of glass. Call this value "*H.*" It is approximately equal to the number of grams of water having the same heat capacity as the material of the calorimeter.

Determination of the Specific Heat of Zinc.

Take about 25 grams of pure zinc in the form of sticks 6 or 7 cm. long. Weigh them to decigrams; tie them into a bundle with a long delicate silk thread and suspend them in the air-bath *C*, the temperature of which is accurately adjusted at 100° C. Call this "W."

Place 200 grams (weighed to decigrams) of distilled water, whose temperature is 5 or 6 degrees below that of the room, in the calorimeter. Put on the cover, stir thoroughly, and when the zinc has been heated for ten or fifteen minutes, note the temperature of the water to tenths of a degree. Call this "t."

Bring the calorimeter under the air-bath and *quickly* lower the bundle of zinc sticks into the water. Replace the cover, remove the calorimeter from the vicinity of the air-bath, stir thoroughly and note the maximum temperature attained by the water to tenths of a degree. Call this "θ."

Let "c" = specific heat of zinc. Then since the heat given out by the zinc in cooling from 100° to θ is taken up by the water and material of the calorimeter, the same being raised in temperature from t to θ, we have

$$W\,c\,(100 - \theta) = (200 + H)\,(\theta - t)$$

or

$$c = \frac{200 + H}{W}\;\frac{\theta - t}{100 - \theta}$$

W being the weight of zinc and c the specific heat of zinc.

Then by the Law of Dulong and Petit

$$\textit{Atom. wt. of zinc} = \frac{6.4}{c}$$

In like manner determine the atomic weights, from their specific heats, of other metals, *e. g.*, Cu, Ag, Al, etc.

(For fuller details concerning atomic weights obtained from specific heats, see Richter, "Inorganic Chemistry," p. 73, and Ramsey, "Chemical Theory," Chap. X.)

TABLES.

TABLES.

Table of Atomic Weights.

Atomic Mass.

Prepared by Prof. F. W. Clarke (1900).*

	Clarke. H = 1.	Clarke. O = 16.	Richards.	German.
Aluminum	26.9	27.1	27.1	27.1
Antimony	119.5	120.4	120.0	120.
Argon	?	?	39.9?	40.
Arsenic	74.45	75.0	75.0	75.
Barium	136.4	137.40	137.43	137.4
Bismuth	206.5	208.1	208.0	208.5
Boron	10.9	11.0	10.95	11.
Bromine	79.34	79.95	79.955	79.96
Cadmium	111.55	112.4	112.3	112.
Caesium	131.9	132.9	132.9	133.
Calcium	39.8	40.1	40.1	40.
Carbon	11.9	12.0	12.001	12.00
Cerium	138.0	139.0	140.	140.
Chlorine	35.18	35.45	35.455	35.45
Chromium	51.7	52.1	52.14	52.1
Cobalt	58.55	59.00	59.00	59.
Columbium	93.0	93.7	94.	94.
Copper	63.1	63.6	63.60	63.6
Erbium	164.7	166.0	166.	166.
Fluorine	18.9	19.05	19.05	19.
Gadolinium	155.8	157.0	156. ?	....
Gallium	69.5	70.0	70.0	70.
Germanium	71.9	72.5	72.5	72.
Glucinum	9.0	9.1	9.1	9.1
Gold	195.7	197.2	197.3	197.2
Helium	?	?	4.0?	4.
Hydrogen	1.000	1.008	1.0075	1.01
Indium	113.1	114.0	114.	114.
Iodine	125.89	126.85	126.85	126.85
Iridium	191.7	193.1	193.0	193.
Iron	55.46	55.88	56.0	56.
Lanthanum	137.6	138.6	138.5	138.
Lead	205.36	206.92	206.92	206.9
Lithium	6.97	7.03	7.03	7.03

*Journal of the American Chemical Society, Vol. XXI, No. 2, February, 1900.

	Clarke. H = 1.	Clarke. O = 16.	Richards.	German.
Magnesium	24.1	24.3	24.36	24.36
Manganese	54.6	55.0	55.02	55.
Mercury	198.50	290.0	200.0	200.3
Molybdenum	95.3	96.0	96.0	96.
Neodymium	142.5	143.6	143.6	144.
Nickel	58.25	58.70	58.70	58.7
Nitrogen	13.93	14.04	14.045	14.04
Osmium	189.6	191.0	190.8	191.
Oxygen	15.88	16.000	16.0000	16.00
Palladium	106.2	107.0	106.5	106.
Phosphorus	30.75	31.0	31.0	31.
Platinum	193.4	194.9	195.2	194.8
Potassium	38.82	39.11	39.140	39.15
Praseodymium	139.4	140.5	140.5	140.
Rhodium	102.2	103.0	103.0	103.
Rubidium	84.75	85.4	85 44	85.4
Ruthenium	100.9	101.7	101.7	101.7
Samarium	149.2	150.3	150.0	150.
Scandium	43 8	44.1	44.	44.1
Selenium	78.6	79.2	79.2	79.1
Silicon	28.2	28.4	28.4	28.4
Silver	107.11	107.92	107.930	107.93
Sodium	22.88	23.05	23.050	23.05
Strontium	86.95	87.60	87.68	87.6
Sulphur	31.83	32.07	32.065	32.06
Tantalum	181.5	182.8	183.	183.
Tellurium	126.5	127.5?	127.5?	127.
Terbium	158.8	160.	160.	....
Thallium	202.61	204.15	204.15	204.1
Thorium	230.8	232.6	233.	232.
Thulium	169.4	170.7	170.?	....
Tin	118.1	119.0	119.0	118.5
Titanium	47.8	48.15	48.17	48.1
Tungsten	182.6	184.	184.4	184.
Uranium	237.8	239.6	240.	239.5
Vanadium	51.0	51.4	51.4	51.2
Ytterbium	171.9	173.2	173.	173.
Yttrium	88.3	89.0	89.0	89.
Zinc	64.9	65.4	65.40	65.4
Zirconium	89.7	90.4	90.5	90.6

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Comparison of Thermometers.

From Gould's Dictionary of Medicine.

FAHR.	CENT.	REAU.	FAHR.	CENT.	REAU.	FAHR.	CENT.	REAU.
212	100	80	122	50	40	32	0	0
210	98.9	79.1	120	48.9	39.1	30	-1.1	-0.9
208	97.8	78.2	118	47.8	38.2	28	-2.2	-1.8
206	96.7	77.3	116	46.7	37.3	26	-3.3	-2.7
204	95.6	76.4	114	45.6	36.4	24	-4.4	-3.6
202	94.4	75.6	112	44.4	35.6	22	-5.6	-4.4
200	93.3	74.7	110	43.3	34.7	20	-6.7	-5.3
198	92.2	73.8	108	42.2	33.8	18	-7.8	-6.2
196	91.1	72.9	106	41.1	32.9	16	-8.9	-7.1
194	90	72	104	40	32	14	-10	-8
192	88.9	71.1	102	38.9	31.1	12	-11.1	-8.9
190	87 8	70.2	100	37.8	30.2	10	-12.2	-9.8
188	86.7	69.3	98	36.7	29.3	8	-13.3	-10.7
186	85.6	68.4	96	35.6	28.4	6	-14.4	-11.6
184	84.4	67.6	94	34.4	27.6	4	-15.6	-12.4
182	83.3	66.7	92	33.3	26.7	2	-16 7	-13.3
180	82.2	65.8	90	32.2	25.8	0	-17.8	-14.2
178	81.1	64.9	88	31.1	24.9	-2	-18.9	-15.1
176	80	64	86	30	24	-4	-20	-16
174	78.9	63.1	84	28.9	23.1	-6	-21 1	-16.9
172	77.8	62.2	82	27.8	22.2	-8	-22.2	-17.8
170	76.7	61.3	80	26.7	21.3	-10	-23.3	-18.7
168	75.6	60.4	78	25.6	20.4	-12	-24.4	-19.6
166	74.4	59.5	76	24.4	19.6	-14	-25.6	-20.4
164	73.3	58.7	74	23.3	18.7	-16	-26.7	-21.3
162	72.2	57.8	72	22.2	17.8	-18	-27.8	-22.2
160	71.1	56.9	70	21.1	16.9	-20	-28 9	-23.1
158	70	56	68	20	15	-22	-30	-24
156	68.9	55.1	66	18.9	15.1	-24	-31.1	-24.9
154	67.8	54.1	64	17.8	14.2	-26	-32.2	-25.8
152	66.7	53.3	62	16.7	13.3	-28	-33.3	-26.7
150	65.6	52.4	60	15.6	12.4	-30	-34.4	-27.6
148	64.4	51.6	58	14.4	11.6	-32	-35.6	-28.4
146	63.3	50.7	56	13.3	10.7	-34	-36.7	-29.3
144	62.2	49.8	54	12.2	9.8	-36	-37.8	-30.2
142	61.1	48.9	52	11.1	8.9	-38	-38.9	-31.1
140	60	48	50	10	8	-40	-40	-32
138	58.9	47.1	48	8.9	7.1	-42	-41.1	-32.9
136	57.8	46.2	46	7.8	6.2	-44	-42.2	-33.8
134	56.7	45.3	44	6.7	5.3	-46	-43.3	-34.7
132	55.6	44.4	42	5.6	4.4	-48	-44.4	-35.6
130	54.4	43.6	40	4.4	3.6	-50	-45.6	-36.4
128	53.3	42.7	38	3.3	2.7	-52	-46.7	-37.3
126	52.2	41.8	36	2.2	1.8	-54	-47.8	-38.2
124	51.1	40.9	34	1.1	0.9	-56	-48.9	-39.1

T., **Centigrade,** the freezing point is 0° and the boiling point at 100°. T., **Fahrenheit's,** the interval between freezing and boiling is divided into 180 equal parts, each called a degree, the zero point being 32 degrees or divisions below the freezing of water. T., **Reaumur,** the freezing point is 0°, and the boiling point 80°. To convert the registration of one thermometer into that of another the following formulæ are useful:—

Let $F =$ No. of degrees Fahrenheit.
Let $C =$ No. of degrees Centigrade.
Let $R =$ No. of degrees Reaumur.

Then to convert

Fahr. to Cent.

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

Cent. to Fahr.

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

Fahr. to Reaum.

$$\frac{4\,(F-32)}{9} = R.$$

Reaum. to Fahr.

$$\frac{9\,R}{4} + 32 = F.$$

Table of Metric Weights and Measures.

MEASURES OF LENGTH.

1 metre = 10 decimetres = 100 centimetres = 1000 millimetres.
1 metre = 1.09363 yards = 3.2809 feet = 39.3709 inches.

MEASURES OF CAPACITY.

1 cubic metre = 1000 litres = 1,000,000 cubic centimetres = 1,000,000,000 cubic millimetres.

1 litre = 61.02705 cubic inches = .035317 cubic foot = 1.76077 pints = .22097 gallon.

MEASURES OF WEIGHT.

1 gram = weight of 1 c.c. of water at 4° C.
1 kilogram = 1000 grams = 100,000 centigrams = 1,000,000 milligrams.
1 kilogram = 2.20462 lbs. = 35.2739 ounces = 15432.35 grains.

Tension of Aqueous Vapor in Millimetres (*Regnault*).

Temp.	Tension.	Temp.	Tension.	Temp.	Tension.
0°	4.6	11°	9.8	21°	18.5
1	4.9	12	10.4	22	19.7
2	5.3	13	11.1	23	20.9
3	5.7	14	11.9	24	22.2
4	6.1	15	12.7	25	23.6
5	6.5	56	13.5	26	25.0
6	7.0	17	14.4	27	26.5
7	7.5	18	15.4	28	28.1
8	8.0	19	16.3	29	29.8
9	8.5	20	17.4	30	31.6
10	9.1				