Version 2.0

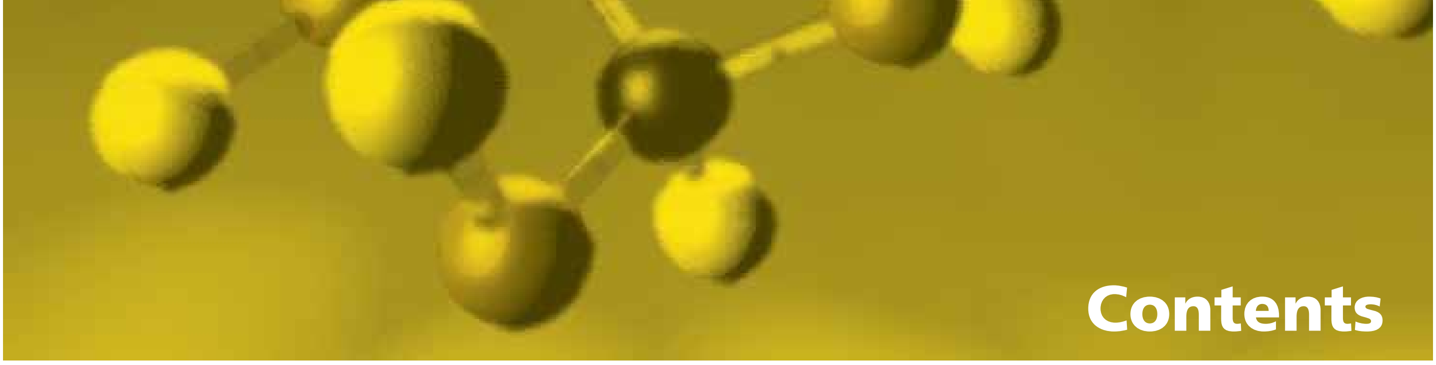
CHEM C3000

Thames & Kosmos

CHEMISTRY SET

WARNING — This set contains chemicals and parts that may be harmful if misused. Read cautions on individual containers and in the manual carefully. Not to be used by children except under adult supervision.

Experiment Manual



Contents

Safety and Precautions

The Experiments

Appendix

The CHEM C3000 contains the following parts:

Component Tray 1

No.	Description	Item No.
1	Two dropper pipettes	232134
2	Rubber stopper without hole	071078
3	Rubber stopper with a hole	071028
4	Cork stopper with a hole	071118
5	Test tube brush	000036
6	Test tube holder	000026
7	Safety glasses	052347
8	Magnesium strip	771761
9	Lid opener tool	070177
10	Test tube stand	070187
11	Copper wire	703059
12	Clip for 9-volt square battery	042106
13	Funnel	086228
14	Two large graduated beakers	087077
15	Two lids for large graduated beakers	087087
16	Immersion heater	065458
17	Angled tube	065378
18	Pointed glass tube	065308
19	Sodium hydrogen sulfate (also known as sodium bisulfate)	033402
20	Calcium hydroxide	033432
21	Potassium hexacyanoferrate(II)	033422
22	Sodium carbonate	033412
23	Ammonium chloride	033452
24	Potassium permanganate mixture (Potassium permanganate-sodium sulfate mixture 1:2 m/m)	771530
25	Sulfur	033262
26	Copper(II) sulfate	033242
27	Litmus powder	771500
28	Five test tubes	062118
29	Vial for litmus solution	771501
30	Safety cap with dropper insert for litmus vial	704092
31	Double-headed measuring spoon	035017

CAUTION! Some parts in this kit have pointed corners, sharp corners, or sharp edges required by their function. There is a risk of injury!

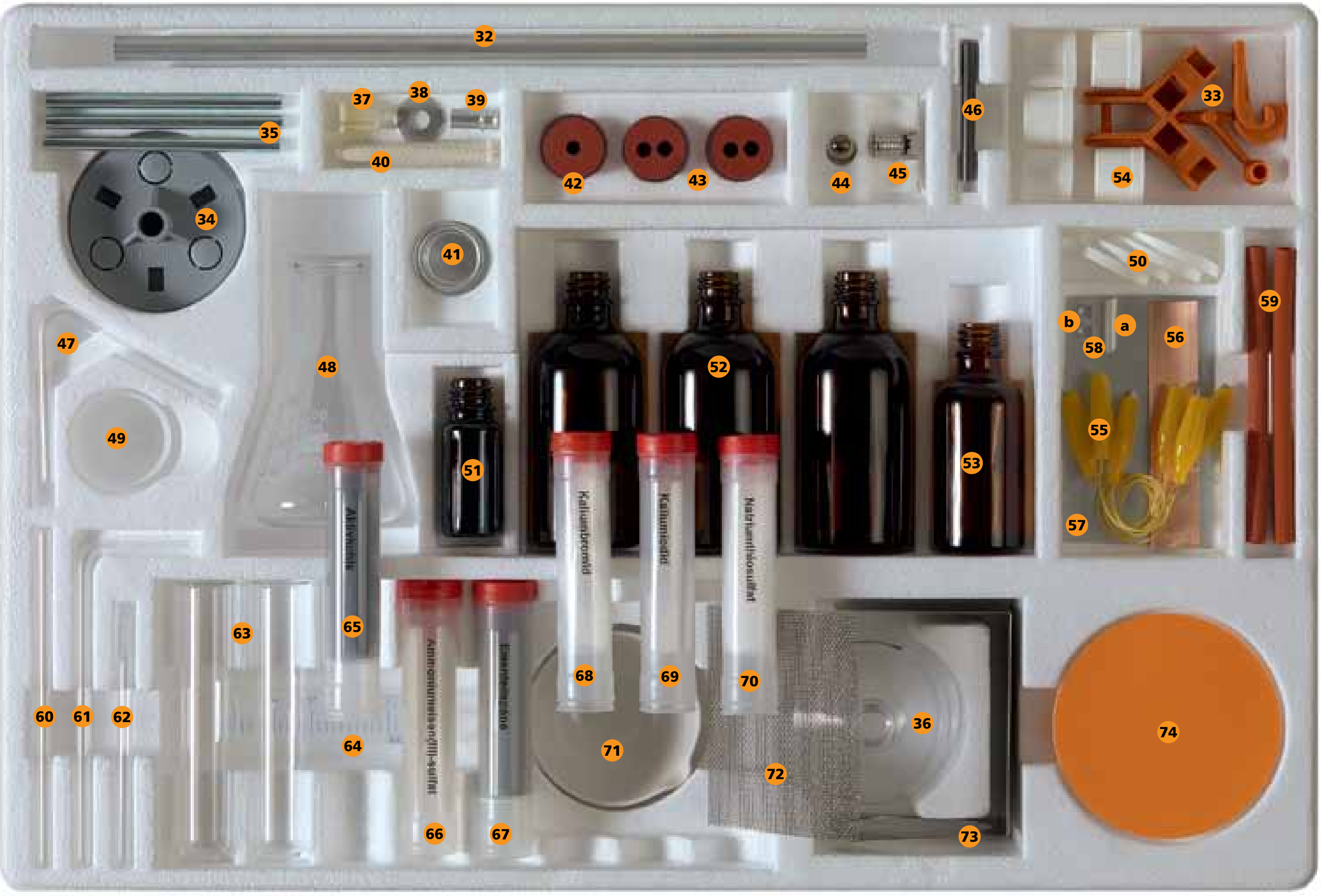
Save the packaging and instructions, since they contain important information.

We reserve the right to make technical changes.

Please check to make sure that all of the parts and chemicals listed in the parts list are contained in the kit.

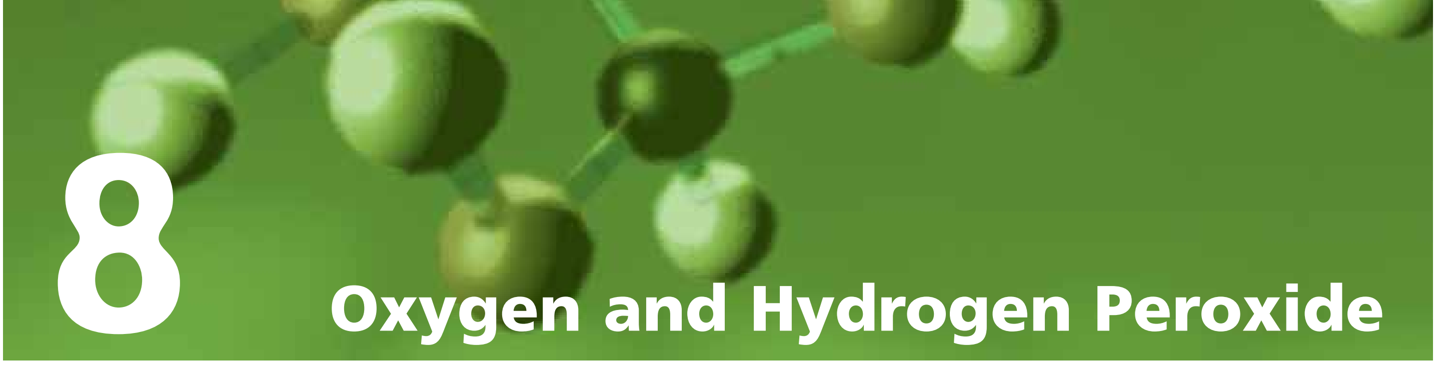
How can individual parts be reordered?
Contact Thames & Kosmos at 800-587-2872 or visit our website at www.thamesandkosmos.com to inquire about an order.

Additional materials required
On page 16, we have made a list of the additional materials required for a number of experiments.



Component Tray 2

No.	Description	Item No.
	Tripod stand consisting of	
32	Tripod pipe	035057
33	Tripod collar	035056
34	Tripod base	083247
35	Three rods for tripod base	011307
	Alcohol burner consisting of	
36	Burner base	061117
37	Insulating piece	048067
38	Aluminum disk	021787
39	Wick holder	021777
40	Wick	051056
41	Burner cap	021797
42	Rubber stopper with a hole	071028
43	Two rubber stoppers with two holes	071038
44	Light bulb (6 V; 50 mA)	704094
45	Bulb socket	702218
46	Carbon electrode	026217
47	Acute-angle glass tube	065268
48	Erlenmeyer flask	062138
49	Four small graduated beakers	061150
50	Four lids for small graduated beakers	061160
51	Bottle, 10 ml (for silver nitrate solution)	701883
52	Three bottles, 100 ml (for caustic soda, hydrochloric acid and hydrogen peroxide)	each 703853
53	Bottle 50 ml (for ammonia solution)	701413
54	Five safety lids for bottles each	075088
55	Three wires, double-ended with alligator clips each	000267
56	Copper sheet	703858
57	Zinc sheet	771431
58	Bag with silicone hose coupler (a) and two glass balls (b)	771432
59	Two rubber hoses each	044473
60	Straight glass tube	065188
61	Angled tube	065378
62	Pointed glass tube	065308
63	Two test tubes each	062118
64	Plastic syringe	086258
65	Activated charcoal	033202
66	Ammonium iron(III) sulfate	033442
67	Iron filings	033512
68	Potassium bromide	033332
69	Potassium iodide	033352
70	Sodium thiosulfate	033252
71	Evaporating dish	063057
72	Wire netting	100187
73	Burner stand	703859
74	Filter paper (round filter)	080156
75	Label sheet (not pictured)	703856



8 Oxygen and Hydrogen Peroxide

If you travel under water, you have to take oxygen along with you.

You are constantly breathing in oxygen. Have you ever noticed a sour taste as you did so? Of course not. But the word oxygen is composed of Greek roots meaning "acid producer." So what sense are we to make of this name? It is actually based on an error. The French chemist Antoine Lavoisier (1743–1797) thought that oxygen was the characteristic component of acids, which as you know isn't true. Hydrogen, not oxygen, is the common feature of acids. Out of respect for the significant achievements of the French chemist, though, the old name has been retained: French *oxygène*, German *Sauerstoff* (= "acid material"), English *oxygen*.

In the gas mixture of the air, oxygen is "diluted" with four times its quantity of nitrogen. In the following experiments, you will be producing somewhat larger quantities of undiluted oxygen in order to study the combustion-supporting effect of this gas.

For the hobby chemist, the oxygen-rich compounds **potassium permanganate** and **hydrogen peroxide** are the handiest things for making oxygen.

Strongly colored — potassium permanganate

You already used potassium permanganate for the slow oxidation of sugar (Experiment 68). Your experiment kit contains a potassium permanganate mixture consisting of one part potassium permanganate to two parts sodium sulfate.

! For **potassium permanganate**, note the "Hazardous materials and mixtures" information on p. 7–9.

EXPERIMENT 72

Place 1 spoon tip of the potassium permanganate mixture in the Erlenmeyer flask and fill the flask with water up to the 100-ml mark. Close the flask with the stopper and shake. You will get a deep purple solution with an intensity of color that will only deepen as you continue to shake. At the beginning, you will see undissolved crystals at the bottom of the flask releasing even darker clouds of color.

Intense in color though potassium permanganate may be, it is nevertheless a rather sensitive compound.

A paper towel can rob the purple potassium permanganate solution of its charm.

TIP

A simple wood stain

If you ever want to color a model made of light wood with a brown stain that won't hide the grain, potassium permanganate would be a good choice. Dissolve 1 small spoonful of the mixture in half a test tube of water and paint the wood with the purple solution. The wood will take on a brown color tone.
Leftover solution: A7

Ethanol — main component of denatured alcohol — converts potassium permanganate into manganese dioxide.

You can stain light-colored wood with the potassium permanganate mixture.

EXPERIMENT 73

Place a few drops of the purple solution from the previous experiment on a piece of paper towel. The purple color will disappear in the blink of an eye, leaving yellowish-brown stains on the paper.

> **!** For **denatured alcohol**, note the "Hazardous materials and mixtures" information on p. 7–9.
> Have an adult pour the required amount of alcohol for you.

EXPERIMENT 74

Measure 10 ml of the purple solution into a test tube and add 1 ml of denatured alcohol. Insert the boiling rod and heat. The purple will gradually turn yellowish-brown and then brown. Set the test tube in the test tube stand to cool. There will be brown flakes that gather at the bottom of the test tube. Precipitate: A6, leftover potassium permanganate solution: A7

In Experiments 73 and 74, as well as with the wood stain, the potassium permanganate decomposes and leaves a deposit of manganese dioxide. Potassium permanganate, which has the formula $KMnO_4$, consists of the elements potassium, K, manganese, Mn, and oxygen, O. When the oxygen is given off, it creates manganese dioxide, with the formula MnO_2.

Making oxygen

> **!** Be careful when handling glass tubes. Note the information on p. 15. In case of injury: **First Aid 5** (back cover).

EXPERIMENT 75

Measure 5 spoonfuls of the potassium permanganate mixture into a dry test tube and assemble the experimental apparatus shown on p. 52. This is the same setup that you used for the production of hydrogen in Experiment 51. Place two test tubes in the basin and have the stoppers ready to seal them (plug the two-hole stopper openings with the little glass balls). Heat the potassium permanganate mixture. Let the first few gas bubbles escape as they come out of the tube. Then collect the gas in two test tubes, one after the other. After each one is filled, seal it under water with a stopper and place it in the test tube stand. Save the heated test tube with its contents for Experiment 78.

> **!** Move aside the test tube stand with the heated test tube, so that the angled tube no longer dips into the water. Otherwise, the cold water could rise back up into the hot test tube, which would probably cause it to shatter. Now you can extinguish the burner flame.

Oxygen

8 O
Oxygen
16.00

Properties:

- odorless, colorless, combustion-supporting gas
- density 1.4291 g/L at 0 °C and 1013 hPa; atomic mass 16.00 u
- in addition to diatomic molecules, O_2, there is also a triatomic form, O_3 (ozone)

Production:

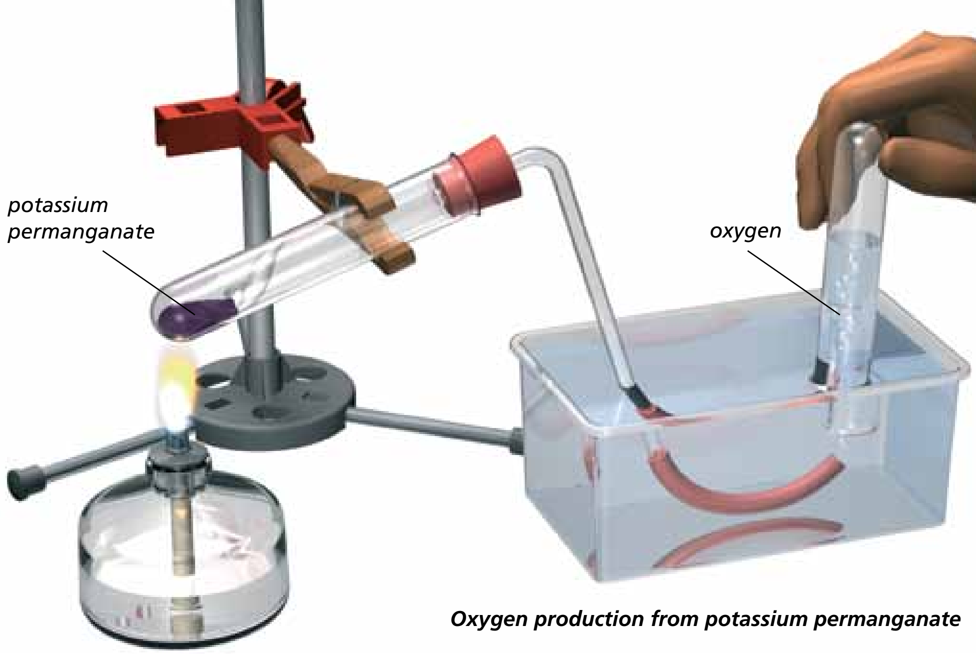
- through distillation of liquefied air (separation of oxygen from the other components)
- in the lab, from oxygen-rich substances or through electrolysis of water

Use:

- multiplicity of uses in place of air in industrial processes (such as metal production and processing, chemical industry, glass industry)
- liquid oxygen for explosives and rocket fuel
- energy production in fuel cells

Liquid oxygen is used to power rockets, among other things.

Testing for oxygen (glowing splint test)



Oxygen production from potassium permanganate

EXPERIMENT 76

Additional material: *Tealight candle*

Light a wooden splint in the tealight flame, and then blow out the burning splint so it's just glowing. Open the first of the filled test tubes and lower the glowing splint into it. It will sizzle and then burst back into flame. The **glowing splint test** serves as a test for oxygen.

Oxygen wakes up a drowsy flame

The double-headed measuring spoon used as a combustion spoon

! For **copper sulfate**, **ammonium iron(III) sulfate**, and **potassium hexacyanoferrate(II)**, note the warnings in "Hazardous substances and mixtures" on pp. 7–8, and for **copper sulfate** follow the instructions for disposal on p. 77.

EXPERIMENT 77

Perform this experiment outside or near an open window. Ventilate well afterwards! Bend the double-headed measuring spoon as shown in the illustration. Now, it will serve as a combustion spoon. Fill the small end with sulfur. Light the sulfur in the burner flame and lower the spoon into the second oxygen-filled test tube. The previously weak little blue flame will get larger and brighter. The sulfur has combined with the oxygen to form sulfur dioxide, SO_2.

Cleaning the combustion spoon: Working outside, grab the large end of the spoon with the test tube holder and hold the small end in the burner flame long enough for the sulfur residues to burn completely away.

Even an iron wire will burn in pure oxygen.

EXPERIMENT 78

Additional material: *Small candle (like a birthday candle)*

Use melted wax to stick a 1 cm-long candle piece to the small hollow of the combustion spoon. Now add another 2 spoonfuls of the potassium permanganate mixture to the test tube from Experiment 75 and use the same apparatus that you used to produce oxygen. Fill a test tube with oxygen, seal it shut under water, and set it in the test tube stand. Light the candle, open the test tube, and lower the combustion spoon with the burning candle into it. The yellow candle flame will turn bright white.

Hold the combustion spoon in the burner flame as in the previous experiment, until the paraffin residues have burned off.

Save the test tube in which you heated the potassium permanganate mixture for the following experiments.

Chem Facts

All combustion processes unfold more vigorously in pure oxygen than in air, which only contains one fifth oxygen.

Manganese, the quick-change artist

You already separated brown manganese dioxide out of the purple potassium permanganate solution. But manganese can also take other colors in its various compounds.

EXPERIMENT 79

Add 1 small spoon tip of the reaction residue from Experiment 78 to the water-filled evaporating dish. The water will immediately turn a deep green color. The color will soon return to purple, though, which is the color of potassium permanganate. A7

The green potassium manganate decomposes when you add acid, and the purple permanganate returns.

! For **sodium hydroxide**, note the “Hazardous materials and mixtures” information on p. 7–9.

EXPERIMENT 80

Fill the cleaned evaporating dish with water and add 10 drops of sodium hydroxide. Stir well with the boiling rod and add 1 small spoon tip of the residue from the oxygen production experiment. Stir again. This time, the green color will hold longer. If you acidify the solution, for example with vinegar (5% acetic acid), the color will change to purple. A7

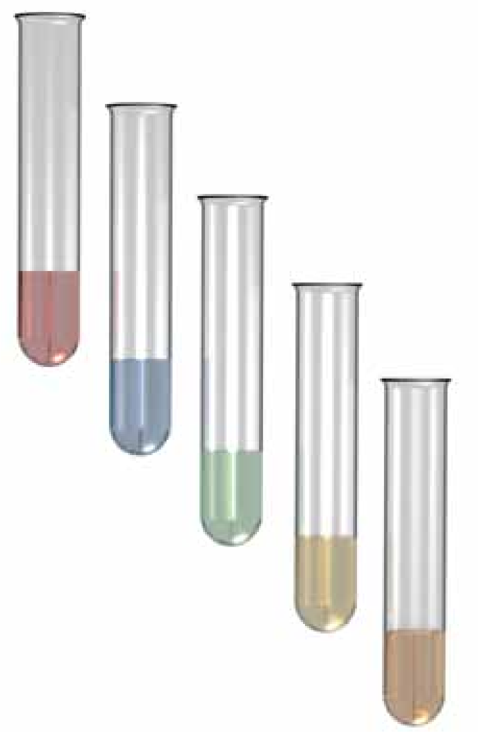
The way a chemical reaction proceeds will often depend on whether it takes place in an **acidic** or **alkaline** solution. For an acidic reaction, **acids** will do the trick, while **alkalis** or **bases** work for alkaline reactions, such as lye or sodium carbonate solution. You will learn more about acids and bases in Chapter 14.

The unstable green potassium manganate also appears as an intermediate stage in the following “play of colors.”

EXPERIMENT 81

Prepare a strongly diluted, but still clearly purple, solution from 1 small spoon tip of the potassium permanganate mixture. Add a few drop of sodium hydroxide and 1 spoon tip of finely-powdered sugar. Seal the test tube with the stopper and shake. The test tube contents will turn from bluish-purple through blue and green to yellow and finally brown. A1

As in Experiment 68, what took place here was a slow combustion of the sugar. The oxygen required for this was given off in stages, resulting in various-colored intermediate stages. The last stage is the brown manganese dioxide.



From potassium permanganate to manganese dioxide.

What is a peroxide?

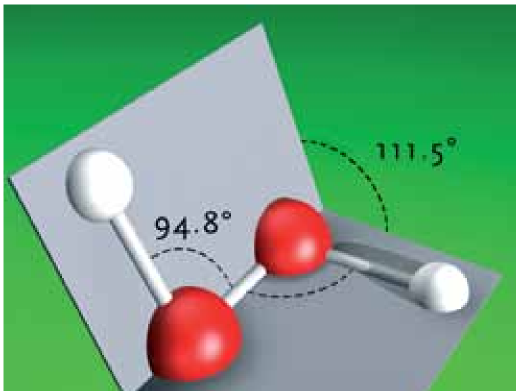
When hydrogen is combusted, water is created. Water is the oxide of hydrogen, and really should be called hydrogen oxide, technically speaking. But even chemists don’t say that. **Hydrogen peroxide**, on the other hand, is a common term in technical circles, and a frequently used chemical in the lab. The Latin root per means, among other things, “over,” “more,” and it relates here to oxygen. Hydrogen peroxide contains twice as much oxygen as water, and its formula is H_2O_2.

A compound that decomposes easily

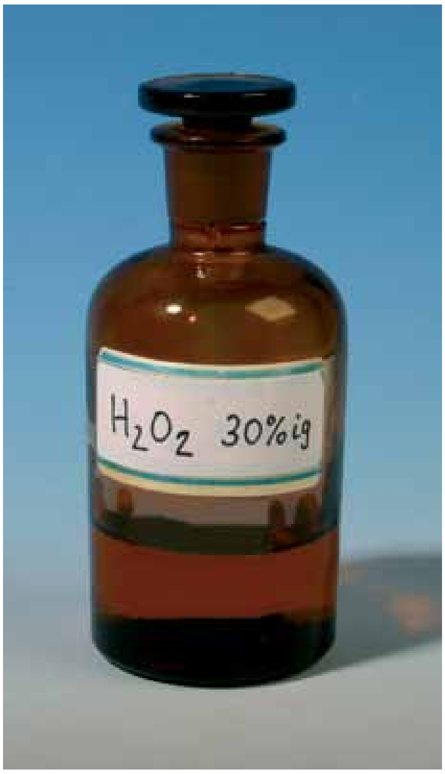
Hydrogen peroxide decomposes when exposed to heat, alkalis, heavy metal compounds, and a lot of other substances, with oxygen released in the process:

$$2\ H_2O_2 \rightarrow 2\ H_2O + O_2\uparrow$$

Get a few wooden splints (shish kebab skewers) ready for the following experiments.



Three-dimensional structure of the hydrogen peroxide molecule



Hydrogen peroxide — an important reagent in the lab

TECHNOLOGY AND ENVIRONMENT

Versatile hydrogen peroxide

Hydrogen peroxide was first produced in 1818 by the chemist Louis-Jacques Thenard (1777–1857) from barium peroxide and sulfuric acid. Pure hydrogen peroxide is a colorless liquid that can be mixed with water in any proportion, and is just under 1.5 times heavier than water: Its density at 20 °C is 1.45 g/cm^3, while that of water is 0.998 g/cm^3. A curiosity: When mixed with water, the solution becomes more viscous (thicker). The reason: The forces of attraction between the H_2O_2 and the H_2O molecules are stronger than those between the molecules of the pure substance.

You will be able to see for yourself in a range of experiments that hydrogen peroxide is an unstable substance that decomposes readily into water and oxygen. So you have to stabilize the product available in the store (usually a 30% solution or the 3% solution you are using) through additives that prevent or at least slow down the process of decomposition.

Hydrogen peroxide is most often used as a bleaching agent in the textile, paper, and laundry detergent industries, as well as for cosmetics and hair bleaches. It is also increasingly used instead of chlorine for disinfecting and deodorizing water for drinking and swimming. Hydrogen peroxide, while created in the body by metabolic processes, is nevertheless harmful and is therefore broken down by the body's own enzymes (see Chapter 23).

!

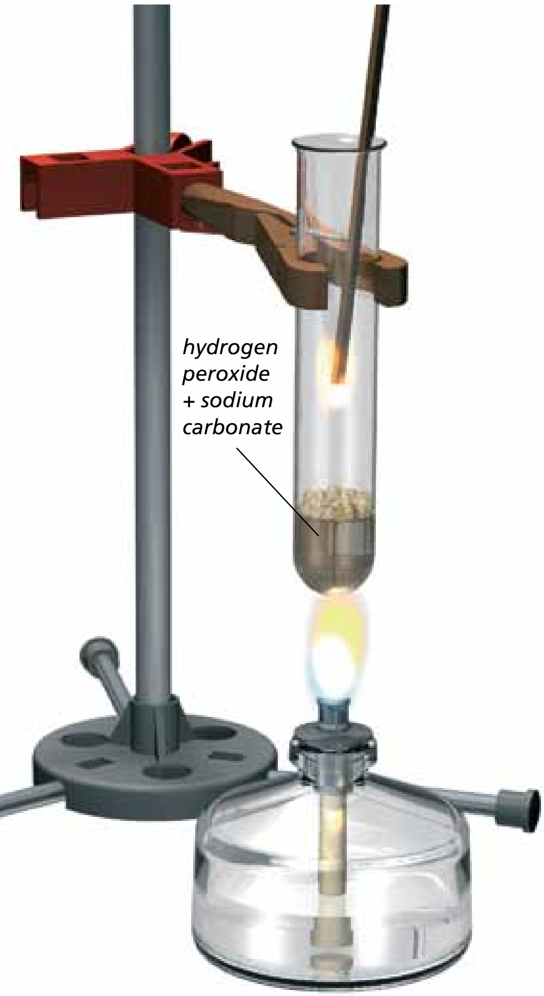
For **copper sulfate**, **ammonium iron(III) sulfate**, and **potassium hexacyanoferrate(II)**, note the warnings in "Hazardous substances and mixtures" on pp. 7–8, and for **copper sulfate** follow the instructions for disposal on p. 77.

EXPERIMENT 82

Clamp a test tube straight upright in the tripod and add 5 ml of hydrogen peroxide and 1 spoonful of sodium carbonate. Heat lightly to get the reaction going. Then you can pull away the burner. Perform the glowing splint test! A2

EXPERIMENT 83

The glowing splint test often won't work in an open reaction vessel, especially if too much water vapor is created in the heating process. In that case, adjust the previous experiment as follows: Close the test tube with the stopper with the hole in it and the angled tube lengthened with the rubber hose. Once again, let the first few bubbles of gas (air!) escape. Then collect the oxygen in test tubes using water as in Experiment 75. Now the glowing splint test should work. A2



Decomposition of hydrogen peroxide with sodium carbonate and the glowing splint test

Activated charcoal — active as a catalyst, too

The word "catalyst" is probably known to you from the catalytic converters used on cars. For the chemist, catalysts are things that accelerate reactions. Catalytic converters owe their name to the fact that this kind of reaction accelerator is used in the exhaust detoxification process. The word catalyst derives from the Greek word *katalyein*, translated as "dissolve," "cancel," "release." To set a reaction in motion, the existing bonds have to be released. That can be helped along the use of catalysts. What's interesting is that the catalyst takes part in the reaction in a hidden manner, and doesn't even show up in the reaction product.

Chem Facts

A catalyst is a material that accelerates a reaction without showing up in the final product.

Chlorine

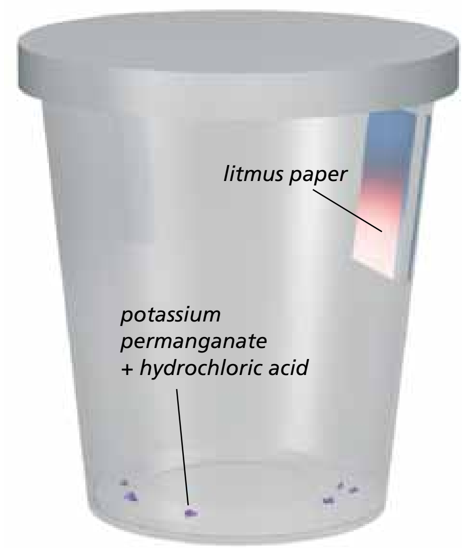
17 Cl
Chlorine
35.45

Properties:
- sharp-smelling, greenish-yellow, toxic gas
- density 3.214 g/L at 0 °C and 1013 hPa; atomic mass 35.45 u
- very reactive, strong oxidizing agent

Production:
- industrially through chloralkali electrolysis from table salt solution
- in the laboratory from hydrochloric acid and oxidizing agents (like potassium permanganate)

Use:
- for reactive intermediate products during chemical synthesis
- for the manufacture of plastics
- for solvents, crop protection agents, medicines

Chlorine disinfects water in swimming pools.



Chlorine bleaches litmus paper.

Hydrogen chloride has a great passion: water. About 450 L of hydrogen chloride will dissolve in 1 L of water at 20 °C. That yields concentrated hydrochloric acid. In this experiment, only a little hydrogen chloride dissolves in the narrow glass tube at first. The outer air pressure slowly presses the water column up in the glass tube. But when the first drops emerge from the tip of the glass tube, a large portion of the hydrogen chloride gas suddenly dissolves and the air pressure presses the water into the mostly gas-depleted test tube. That's what makes the fountain work.

Chem Facts

About 450 liters of hydrogen chloride dissolve in 1 L of water at 20 °C.

Toxic gas and disinfectants

Now you're going to meet a gas whose smell you might know from swimming pools. In some places, it's also added to drinking water. In both cases, its purpose is disinfection. While chloride ions, Cl^-, are quite harmless in table salt solution (and in your soup!), **chlorine** (composed of Cl_2 molecules) is an aggressive, toxic gas. You see, it makes a huge difference whether or not a chlorine atom has an additional electron from a sodium atom or from another chlorine atom to fill out the eight-electron shell.

For safety's sake, we will only be experimenting with very small quantities of the gas, and we'll render leftover chlorine harmless using a "chlorine killer." By the way, chlorine gets it name from its color (Greek *chloros* = yellowish-green).

Carry out the experiments with **chlorine** outside or near an open window. Ventilate well after the experiment. Be sure to keep to the indicated quantities.

DANGER

Chlorine is toxic when inhaled, causes severe eye and skin irritation and can irritate the respiratory tract. — Do not inhale gas. — IF INHALED: Take to fresh air and place in a comfortable position that makes it easiest to breath. Call POISON CONTROL CENTER or doctor immediately. — IN CASE OF CONTACT WITH EYES: Rinse carefully with water for several minutes. If possible, remove any contact lenses. Continue rinsing. If eye irritation persists, seek medical attention.
For **potassium permanganate** and **hydrochloric acid**, note the "Hazardous substances and mixtures" information on p. 7–9.

EXPERIMENT 117

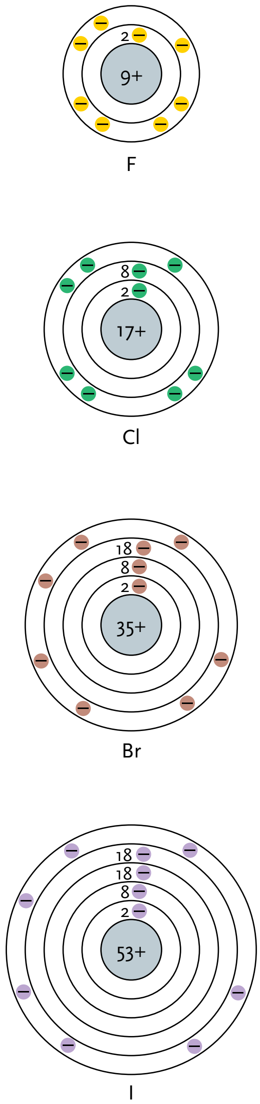
Affix a 1 cm-long moistened piece of blue litmus paper to the wall of a small graduated beaker as shown in the illustration. Place 1 small (!) spoon tip of potassium permanganate in the beaker and add 1 pipette of hydrochloric acid to it. Seal the beaker. The litmus paper first turns a reddish color (because of the hydrochloric acid vapors), then gradually fades. Keep the sealed graduated beaker for the next experiment.

Potassium permanganate gave off oxygen, which released chlorine from the hydrochloric acid. This can be expressed in simplified form as follows:

$$2\ HCl + O \rightarrow H_2O + Cl_2$$

The chlorine bleached the litmus dye.

11 The Halogens: A Family of Elements



Atomic models of the halogens fluorine, chlorine, bromine and iodine. The elements have 7 electrons in their outer shell.

You already know one element family: the noble gases. Their common structural feature is an outer shell fully occupied by 8 electrons. You might suppose that other element families have the same structural features — and you'd be right. The **halogens** are a great example of this.

The name comes from the Greek and means "salt-former." What this refers to is that the halogens are able to combine directly with metals to form salts, the **halogenides** — thus skipping the detours through the acids, bases and oxides (we'll look at those more in depth in Chapter 14). What the halogens have to do with halogen lamps is revealed by the info box on p. 78.

You've already worked intensively with one halogen: **chlorine**, Cl. In this chapter, we'll be adding **bromine**, Br, and **iodine**, I. Other halogens are **fluorine**, F, which you'll get to know shortly, and the radioactive element **astatine**, At, which is the rarest element occurring in nature of which only tiny quantities exist.

Unlike the noble gases, the halogens are extremely reaction-happy. Like chlorine, all of the members of the family have 7 electrons in their outer shell, so they're desperate to fill their outer shell by taking on an electron. This happens either by bonding with an ion or through covalent bonds, for example in the double-atom molecules F_2, Cl_2, Br_2, I_2. The existence of At_2 molecules hasn't been confirmed so far.

A versatile reagent

Silver nitrate, which you've used to detect chloride, is also an indicator for bromide and iodide. Your kit contains **potassium bromide**, KBr, and **potassium iodide**, KI, two typical salts that are similar to sodium chloride.

> **!** For **silver nitrate solution** and **potassium bromide**, note the "Hazardous substances and mixtures" information on p. 7–9.

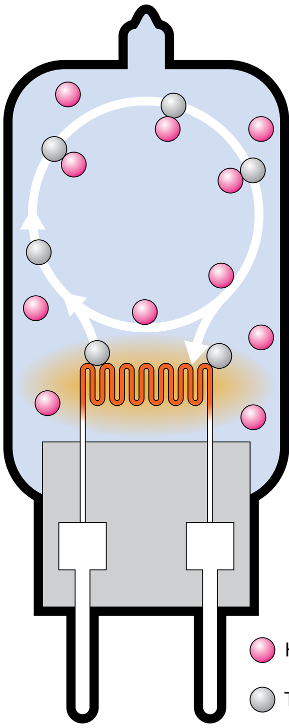
EXPERIMENT 124 Dissolve 1 spoon tip of potassium bromide in 2 ml of water and add 3–4 drops of silver nitrate solution to it. Keep the precipitate for Experiment 126.

EXPERIMENT 125 Repeat the experiment with potassium iodide. Put the precipitate in a test tube and keep it for Experiment 127.

The Ag^+ ions in the silver nitrate and the halogenide ions Cl^- and I^- produce similar precipitates, although the silver iodide has a strong yellow hue:

$$Ag^+ + Cl^- \rightarrow AgCl \text{ (silver chloride)}$$
$$Ag^+ + Br^- \rightarrow AgBr \text{ (silver bromide)}$$
$$Ag^+ + I^- \rightarrow AgI \text{ (silver iodide)}$$

Silver nitrate, $AgNO_3$, produces white to yellowish precipitates with halogenides that are readily soluble in ammonia and sodium thiosulfate solutions.



Functional diagram of a halogen lamp. Tungsten atoms leave the filament, bond fleetingly with the halogen, and return to the filament after the bond is broken.

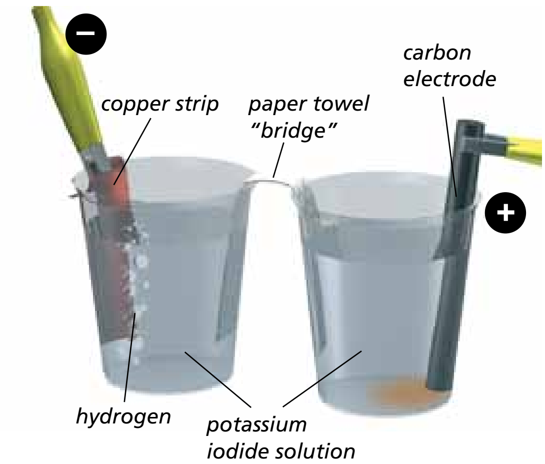
TECHNOLOGY AND ENVIRONMENT

Traffic circle in the halogen lamp

Incandescent light bulbs contain a metal thread that is usually made of the element tungsten, W (it gets its symbol from its other name, wolfram), which glows when electrical current passes through it, thus emitting light. Due to the high temperature at which it operates (2500–3000 °C), a portion of the tungsten — that is of the thread, also called a coiled filament — becomes thinner and thinner until it eventually breaks. Before that happens, vaporizing tungsten condenses on the inner wall of the light bulb, settling there as a dark coating and reducing the light output.

These disadvantages are avoided for the most part in halogen lamps. They contain small amounts of halogen compounds such as methyl bromide or methyl iodide. The tungsten that would otherwise settle on the wall of the bulb temporarily bonds with the halogen and returns to the coiled filament as a result of thermal flow. There the compound breaks down again into tungsten, which settles on the metal thread, and into the halogen components, which bond again with vaporized tungsten near the wall of the bulb.

Success! The coiled filament ages less quickly and the reduction of the light output caused by blackening of the wall of the bulb is avoided. But even halogen lamps don't last forever: The problem is that the tungsten doesn't settle on the thin parts of the coil but only on the thicker parts, where it is a little cooler.



Electrolysis of potassium iodide solution. In the photo below, a porous diaphragm separated the cathode and anode compartments. This does not prevent the current from passing through — just as a paper bridge does not.

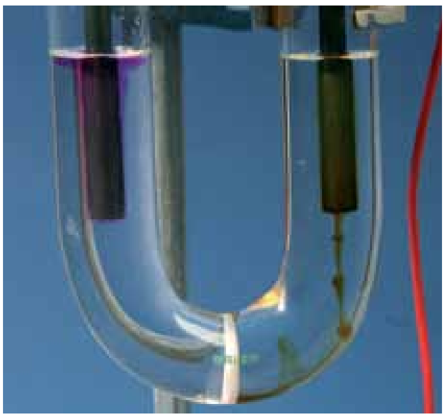
EXPERIMENT 135

Assemble the experimental setup as shown. Dissolve 2 spoonfuls of potassium iodide in 30 ml of water and evenly divide the solution between two graduated beakers. Also soak the paper towel "bridge" with the solution. Close the electrical circuit and observe what happens in the two beakers. A colorless gas is produced in the cathode beaker. Dip red litmus paper into the solution: it turns blue. The solution in the anode beaker turns yellow; keep it for Experiment 138. Cathode beaker: A1

As in Experiment 122, hydrogen is released at the cathode. The simultaneously-produced caustic potash (potassium hydroxide) solution, KOH, colors the red litmus strip blue. In the anode beaker, the released iodine dissolves in the potassium iodide solution.

Chem Facts

Oxidizing agents (like potassium permanganate) or electrical current oxidize iodide to form iodine.



When iodine solution turns pale

! For **iodine**, heed the hazard warnings and safety guidelines on p. 77.

EXPERIMENT 136

Take just enough of the dark-brown iodine solution you prepared in Experiment 134 to cover the rounded bottom part of a test tube, and dilute it with 5 ml of water. You will get a yellowish-brown solution that you will need for Experiments 137 and 139. If brown cloudiness occurs during dilution, add 1 spoon tip of potassium iodide to it. The cloudiness will dissolve and disappear.

Iodine doesn't dissolve very well in water (1 g iodine in 3.5 L water). But iodine dissolves well in ethanol (the main component of denatured alcohol) as well as in potassium iodide solution. When you subjected the potassium iodide solution to electrolysis (Experiment 135), the precipitated iodine also dissolved in the iodide solution, or the electrolyte.