

# Chem C1000



Please observe the safety information, the advice for supervising adults on page 5, the safety rules on page 6, the information about hazardous substances and mixtures (chemicals) on pages 7-8 and their environmentally sound disposal on page 75, the safety for experiments with batteries on page 7, and the first aid information on the inside front cover.

**WARNING.** Not suitable for children under 10 years. For use under adult supervision. Contains some chemicals which present a hazard to health. Read the instructions before use, follow them and keep them for reference. Do not allow chemicals to come into contact with any part of the body, particularly the mouth and eyes. Keep small children and animals away from experiments. Keep the experimental set out of reach of children under 10 years old. Eye protection for supervising adults is not included.

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



# First Aid Information

When conducting experiments with chemicals and in case any accidents should happen during experimentation:

**In case of injury, seek immediate medical help.**

1. **In case of eye contact:** Wash out eye with plenty of water, holding eye open if necessary. Rinse from the nose outward. Seek immediate medical advice.
2. **If swallowed:** Wash out mouth with water, drink some fresh water. Do not induce vomiting. Seek immediate medical advice.
3. **In case of inhalation:** Remove person to fresh air. For example, move person into another room with open windows or outside.
4. **In case of skin contact and burns:** Wash affected area with plenty of water for at least 10 minutes. Cover burns with a bandage. Never apply oil, powder, or flour to the wound. Do not lance blisters. For larger burns, seek immediate medical help.
5. **In case of doubt, seek medical advice without delay.** Take the chemical and its container with you.
6. **In case of injury always seek medical advice.**
7. **In case of cuts:** Do not touch or rinse with water. Do not apply any ointments, powders or the like. Dress the wound with a germ-free, dry first-aid bandage. Foreign objects such as glass splinters should only be removed from the wound by a doctor. Seek medical advice if you feel a sharp or throbbing pain.

Warning! Contains functional sharp points or edges that pose a risk of injury.

Keep the packaging and instructions, as they contain important information.

## Poison Control Centers (United States)

In case of emergency, your nearest poison control center can be reached everywhere in the United States by dialing the number:

**1-800-222-1222**

## Local Hospital or Poison Centre (Europe)

Record the telephone number of your local hospital or poison centre here:

Write the number down now so you do not have to search for it in an emergency.

# **Chem C1000**

## **Experiment Manual**

**Kurt Waselowsky**

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Thames & Kosmos, LLC, Providence, RI, USA

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### Chem C1000 contains the following parts:

No.	Description	Item No.
1	Safety glasses	717019
2	Three dropper pipettes	232134
3	Clip for 9-volt battery	712310
4	Safety cap with dropper insert for litmus bottle	704092
5	Copper wire	000063
6	Two large graduated beakers	087077
7	Two lids for graduated beakers	087087
8	Four test tubes	062118
9	Test tube brush	000036
10	Rubber stopper with hole	071028
11	Rubber stopper without hole	071078
12	Funnel	086228
13	Sodium carbonate, 12g	772504
14	Potassium hexacyanoferrate(II), 4g	772505
15	Calcium hydroxide, 8.5g	772506
16	Ammonium iron(III) sulfate, 5g	772507
17	Copper(II) sulfate, 8g	772522
18	Citric acid, 10g	772528
19	Litmus powder, 1g	772502
20	Small bottle for litmus solution	771510
21	Lid opener	070177
22	Double-headed measuring spoon	035017
23	Angled tube	065378
24	Experiment station (part of the polystyrene insert)	711461
25	Filter paper sheets (not pictured)	702842

The experiment station (for more info, see p. 10) can be divided here using a sharp knife. An adult must do this step.

Please note: The actual design of your experiment station and component storage tray may vary from what is pictured here.

Keep the packaging and instructions, as they contain important information.

**Please check** whether 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 customer service to inquire about an order.

#### Additional materials required

On page 13, we have made a list of the additional materials required for a number of experiments.

# Safety Information

## Information for Parents and Adults

### What you need to know about chemistry experiment kits

With Chem C1000, children can take their first steps into a field of study that may not yet be part of their regular curriculum in school, but, as experience shows, fascinates children who are curious, have fun discovering new things, and want to know what causes the phenomena they encounter in their everyday environment. Chem C1000 provides an engaging introduction to chemistry, bridging the realms of play and learning. The experiments can be done without previous knowledge of chemistry or lab experience because the experimental procedures are described in detail in this manual.

The experiment kit makes it possible to perform 125 simple, exciting, and instructive experiments ranging from creative experimentation with color on paper to tracking down a notorious climate-changing molecule in a number of different places.

It is normal to have questions about the safety of a chemistry experiment kit. You know that improper use of chemicals can lead to injuries or other health risks. This kit complies with all applicable U.S. consumer product safety regulations, including those for chemistry sets. Moreover, this kit complies with the more rigorous European safety standard EN 71-4, in which the safety requirements for chemistry experimental kits are established, to reduce risks to a minimum. This standard forms the reliable basis of all Thames & Kosmos chemistry experiment kits. The standard contains requirements for the manufacturer — for example, the requirement that no particularly hazardous substances can be used. They also require the manufacturer to carefully inform the parents or adult supervisors of the possible hazards and to require them to accompany their children in their new hobby with a helping hand. Therefore, please read all of the safety information prescribed by EN 71-4 listed in the box on the right as well as the advice below. Emphasize to your child the importance of following all of this information, and the importance of carrying out only the experiments that are described in this manual.

The **safety rules** (pages 6-7), the **information about hazardous substances and mixtures** (pages 7-8), and the safety information accompanying each experiment inform you of the risks and help you evaluate the suitability of each experiment for your child. In case anything should happen, please find the **first aid information** and the **poison control center contact information** on the inside front cover. To avoid injuries please inform your child how to handle the angled glass tube on page 12 (place it into the rubber stopper and remove it from the stopper).

Instructions for **setting up the work area** are described on page 10. Proper **waste disposal** techniques are on page 75. A list of **additional materials** required can be found on page 13.

We advise you to carry out the experiments in the prescribed order, since the knowledge of working techniques described in the earlier experiments is a prerequisite for the later ones.

We wish your young chemist, as well as you, a lot of fun and success with the experiments!

Tips and information for you and the child performing experiments.

- Safety Rules (p. 6 – 7)
- Hazardous Substances and Mixtures/Chemicals (p. 7 – 8)
- Setting up the Work Area (p. 10)
- Handling the Angled Glass Tube (p. 12 – 13)
- Additional Materials Required (p. 13)
- Preparation of the Litmus Solution (p. 14 – 15)
- Proper Waste Disposal (p. 75)
- Information about the Safety Glasses (inside back cover)
- First Aid Information (inside front cover)
- Poison Control Centers (inside front cover)

#### Advice for Supervising Adults

A. Read and follow these instructions, the safety rules and the first aid information, and keep them for reference.

B. The incorrect use of chemicals can cause injury and damage to health. Only carry out those experiments which are listed in the instructions.

C. This experimental set is for use only by children over 10 years.

D. Because children's abilities vary so much, even within age groups, supervising adults should exercise discretion as to which experiments are suitable and safe for them. The instructions should enable supervisors to assess any experiment to establish its suitability for a particular child.

E. The supervising adult should discuss the warnings and safety information with the child or children before commencing the experiments. Particular attention should be paid to the safe handling of acids (e.g. citric acid), alkalis (bases, e.g. sodium carbonate) and flammable liquids (denatured alcohol).

F. The area surrounding the experiment should be kept clear of any obstructions and away from the storage of food. It should be well lit and ventilated and close to a water supply. A solid table with a heat resistant top should be provided (also see page 10).

## Safety Rules

All of the experiments that are described in this manual can be performed without danger if you carefully follow the tips and rules summarized below.

### Safety rules for chemical experiments

1. **Read these instructions before use, follow them and keep them for reference.** Pay special attention to the quantity specifications and the sequence of the individual tasks. Only perform experiments that are described in this instruction manual.
2. **Keep young children, animals and those not wearing eye protection away from the experimental area.**
3. **Always wear eye protection.** If you wear corrective eyeglasses, you will need protective goggles for those who wear eyeglasses. When working, wear appropriate protective clothing (old smock and smooth fingered gloves).
4. **Store this experimental set out of reach of children under 10 years of age** (for example, in a lockable cabinet). This includes the additional materials.
5. **Clean all equipment after use.**
6. **Make sure that all containers are fully closed and properly stored after use.**
7. **Ensure that all empty containers are disposed of properly.**
8. **Wash hands after carrying out experiments.** Chemicals that accidentally get onto your skin must be rinsed off immediately under running water.
9. **Do not use any equipment which has not been supplied with the set or recommended in the instructions for use.**
10. **Do not eat or drink in the experimental area.** Do not use any eating, drinking or other kitchen utensils for your experiments unless it is specifically recommended. Any containers or equipment you use in your work should not be used in the kitchen afterwards. Dispose of used drinking straws in the garbage immediately after the experiment. Do not save and reuse them.
11. **Do not allow chemicals to come into contact with the eyes or mouth.**
12. **Do not replace foodstuffs in original container. Dispose of immediately** (in the household trash or the sink). If you are investigating food products (e.g., sugar, flour, table salt), fill the required amount into one of the graduated beakers (do not use the double-headed measuring spoon for this). Make note on the beaker of what it contains and the date it was filled.
13. The tealight candle required for some experiments has to be placed on a fire-resistant surface (such as an old plate). Extinguish the flame at the end of the experiment at the latest as well as when you leave the experiment area even if just for a moment.
14. During the experiments with open flame, be sure that there are no flammable objects or liquids nearby. Have a bucket or box with sand ready for extinguishing any fires. If the fire can't be extinguished right away, notify the fire department immediately.
15. Any filled container or experimental apparatus that is to remain standing for a longer period of time (e.g., for the precipitation of substances) has to be labeled and stored out of reach of young children and animals.
16. Get any additionally required materials ready before starting an experiment.
17. Handle breakable materials (e.g., the glass test tubes or angled tube) carefully.

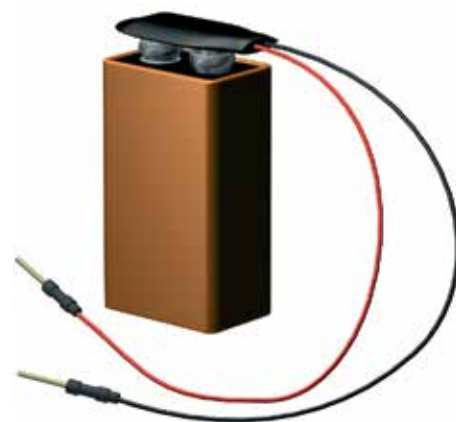
Also note the information on the chemical vial labels, the information about "Hazardous substances and mixtures" on p. 7 – 8 as well as the safety and waste disposal instructions for the individual experiments (for example, regarding how to properly handle the angled tube). If additional products are required, also take note of the warnings on their packaging (e.g., for denatured alcohol).



## Safety for experiments with batteries

**WARNING!** Only for use by children aged 10 years and older. Instructions for parents or other supervising adults are included and have to be observed. Keep the packaging and instructions as they contain important information.

- For some experiments, you will need a 9-V type 6LR61 square battery, which could not be included in the kit due to its limited shelf life. Press the battery clip equipped with red and black wires onto the battery.
- Have an adult check your experimental setup before performing the experiment.
- Non-rechargeable batteries are not to be charged. They could explode.
- Rechargeable batteries are only to be charged under adult supervision.
- Rechargeable batteries are to be removed from the toy before being charged.
- Exhausted batteries are to be removed from the toy.
- Dispose of used batteries in accordance with environmental provisions, not in the household trash.
- The supply terminals of the battery clip are not to be short-circuited: Neither the battery wire contacts nor the wires connected to them should touch each other. Make sure there is no unintended short circuit due to conductive metal objects, such as coins or a keychain. A short circuit can cause the wires to overheat and the battery to explode.
- Don't throw batteries into the fire and don't store them near heat sources.
- Avoid deforming the batteries.
- Never perform experiments using household current. You know that you should never insert any objects into the wall socket holes. The high voltage can be extremely dangerous or fatal!
- Don't use any voltage source other than the specified battery, including a power supply unit.
- After you are done experimenting, remove the battery clip from the battery.



## Hazardous substances and mixtures (chemicals)

### How they are labeled and how to properly handle them

In the following, we provide you with a list of the chemicals contained in this chemistry set as well as the chemicals you will be obtaining in addition that are classified as hazardous substances or hazardous mixtures. For each substance, the list shows **hazard statements** (in blue) and precautionary statements for avoiding the hazards. In the margin, you will find the corresponding **pictograms** and a **signal word** for the scope of the danger: **WARNING** means limited risk, and **DANGER** means elevated or significant risk.

The figure to the right shows the pictograms that appear in the following list of substances. The text under each one cites dangers associated with substances that are designated with the pictogram. A substance designated with a pictogram may be associated with one or more of the hazards listed under that pictogram. For example, copper sulfate (exclamation point) is harmful and irritant. You can find out exactly what hazards are associated with a certain substance by referring to the following list.

The pictograms are a component of Regulation (EC) No 1272/2008, also called the GHS Regulation. GHS stands for **G**lobally **H**armonized **S**ystem, a system whose aim is to achieve a classification and designation of hazardous substances and mixtures that is uniform throughout the world.

Some of the chemicals listed below only cause limited risks. Therefore they are not labelled with a pictogram or signal word. Nevertheless the given precautionary statements should be observed.



Substance/mixture causes

- corrosion (destruction) of metals
- burns on skin
- severe eye damage



Substance/mixture is

- acutely harmful if swallowed, in contact with skin, or inhaled
- an irritant to skin, eyes, and respiratory tract
- sensitizing to skin
- narcotic



Substance/mixture is

- acutely very toxic
- very toxic or toxic with long lasting effects to aquatic life

### Ammonium iron(III) sulfate

Not a hazardous substance

Obtain special instructions before use. May cause slight eye irritation. – Prolonged or repeated contact may dry skin and cause irritation.



### Calcium hydroxide

Causes skin irritation. – Causes serious eye damage. – May cause respiratory irritation.

Do not breathe dust. – Wear eye protection. – IF ON SKIN: Wash with plenty of soap and water. – IF INHALED: Remove victim to fresh air and keep at rest in a position comfortable for breathing. – IF IN EYES: Rinse cautiously with water for several minutes. Remove contact lenses, if present and easy to do. Continue rinsing. – Get medical advice/attention.

These warnings also apply to its aqueous solution, called lime water (prepared in Experiments 9 and 10).

### Citric acid

Not a hazardous substance

Avoid contact with skin and eyes. Avoid formation of dust and aerosols. Further processing of solid materials may result in the formation of combustible dusts.



### Copper(II) sulfate

Harmful if swallowed. – Toxic in contact with skin. – Causes skin irritation. – Causes serious eye irritation.

Wash face, hands and any exposed skin thoroughly after handling. – Do not eat, drink or smoke when using this product. – Wear protective gloves/protective clothing/eye protection/face protection. – Specific treatment (See first aid instructions). – IF IN EYES: Rinse cautiously with water for several minutes. Remove contact lenses, if present and easy to do. Continue rinsing. If eye irritation persists: Get medical advice/attention. – IF ON SKIN: Wash with plenty of soap and water. Call a POISON CENTER or doctor/physician if you feel unwell. Remove/Take off immediately all contaminated clothing. Wash contaminated clothing before reuse. If skin irritation occurs: Get medical advice/attention. – IF SWALLOWED: Call a POISON CENTER or doctor/physician if you feel unwell. Rinse mouth.

Very toxic to aquatic life with long lasting effects.

Avoid release to the environment. Comply with the instructions for disposal on p. 75.

### Litmus powder

Not a hazardous substance

### Potassium hexacyanoferrate(II)

Not a hazardous substance

Obtain special instructions before use. May be harmful if swallowed. – Harmful to aquatic life with long lasting effects. – May cause slight eye irritation.



### Sodium carbonate

Harmful if inhaled. – Causes serious eye irritation.

Avoid breathing dust/fume/gas/mist/vapors/spray. – Use only outdoors or in a well-ventilated area. – IF IN EYES: Rinse cautiously with water for several minutes. Remove contact lenses, if present and easy to do. Continue rinsing. If eye irritation persists: Get medical advice/attention. – IF INHALED: Remove victim to fresh air and keep at rest in a position comfortable for breathing. Call a POISON CENTER or doctor/physician if you feel unwell.

### Hydrogen peroxide (3% aqueous solution) (Not included)

Do not get in eyes or on skin.

**Warning!** The following applies to all chemicals: **Store locked up. Keep out of reach of children.**

This applies to all children except for the experimenting child who is being instructed and supervised by an adult.

Also follow this precautionary statement:

**IF SWALLOWED: Get immediate medical advice/attention and have product container or label of chemical substance at hand.**



For the sake of environmental protection:

**Dispose of contents/ containers (of no-longer-needed chemicals) to a hazardous waste disposal location.**

For information on the protective equipment required, see "Work area and equipment" on p. 10.

# 1

## What is chemistry?

You are probably familiar with tailors, people who work with fabrics. They find, examine, dye, cut, sew, and stitch various fabrics together into all sorts of different styles of clothing. Well, chemists are people who work with the “fabrics” of our world. They explore the properties and transformations of materials. Of course, you know that we’re not talking about actual fabrics or clothing materials here, but the stuff out of which everything that exists is composed.

All of these materials are what chemistry, the science of materials and transforming them, is all about. With this kit, you too can perform transformations of materials and chemical reactions like a professional in the laboratory. This kit includes exciting experiments: sudden changes in color, hissing and bubbling gases, and magical physics tricks with water and air. You will learn all about the seven chemical substances contained in the cylindrical vials in the kit. Don’t be intimidated by their complicated sounding names, like potassium hexacyanoferrate(II) — by the time you are finished with this kit, you will know them all very well!

The scope of your explorations will by no means be limited to these seven chemicals. No, your experiments integrate materials from your surroundings and your environment as well. Accordingly, the informational boxes “Side Notes” and “Technology and Environment” are provided in addition to the instructions for the experiments in order to broaden your perspective.

One more note on the experiments: Oftentimes, several experiments are collected into an “experiment section” with an explanation at the end of the section. It is often best to do these experiments together in one session.

Chem C1000 is nothing more and nothing less than a key that opens the gate to the world of materials. Of course, to conquer this world, considerable effort as well as a number of special tools are required. For more than a few kids, the explorations undertaken with an experiment kit have led them to study chemistry later on and work in one of the many fields of chemistry.

But even if you end up doing something completely different — participating in some way in the quest for solutions to our environmental problems, for example — it can be useful to have had a peek at the world of materials. It will then be easier for you to form your own scientific opinions and back them up too.



*“Oh dear, these aren’t the right fabrics for sewing!”*



# 2

## Work area and equipment



*"Let's conduct some research!"*

You should have an appropriate work area for your experiments. Although you can't really wreak too much havoc with this chemistry set, the living room table or a new desktop could get some stains on them. A sturdy old table or workbench are the most suitable. The work surface should be washable and heat-resistant. The kitchen or any place where foods are kept is not a suitable place for chemical experiments. Your work area needs to be brightly lit and freely accessible; it should also be possible to ventilate it well in case it gets stinky. Unneeded objects and flammable materials, especially fabrics, have no business being in the work area. When performing experiments, do not wear a scarf, neckerchiefs, or loose sleeves and if you have long hair, tie it back.

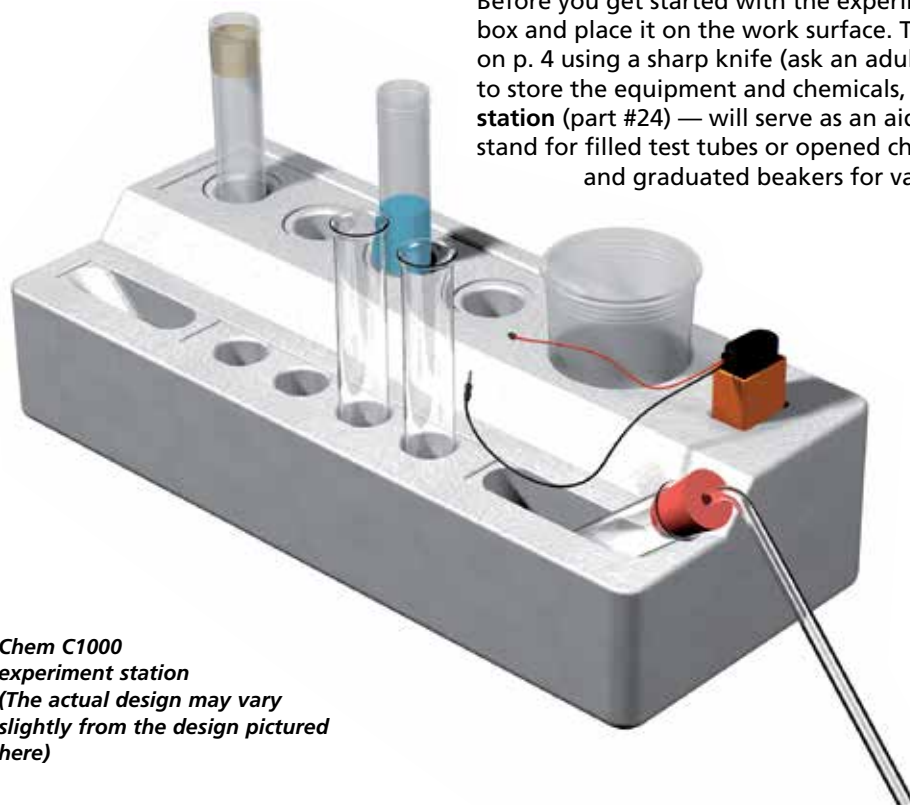
You will need water to perform your experiments and to clean the equipment. A nearby water faucet and drain would of course be ideal. Otherwise, ready a dispensing bottle with water (you can use a carefully cleaned dish-washing liquid bottle for this, for example). Discard reaction products containing copper residues as indicated in Chapter 11. You can pour all other reaction products into the drain with plenty of water. Or you can toss the waste into a small plastic jug half-filled with water labeled "chemical waste" that you empty and clean right after the experiments. To immediately wipe up spilled liquids, use a rag. After using it, rinse it out thoroughly.

The most important part of your protective equipment is the **safety glasses** (part #1). An old smock serves as **protective clothing**, and fingered gloves with a smooth surface are suitable as **protective gloves**. Keep the frequently used additionally required materials in your work area as well (See p. 13 for a list).

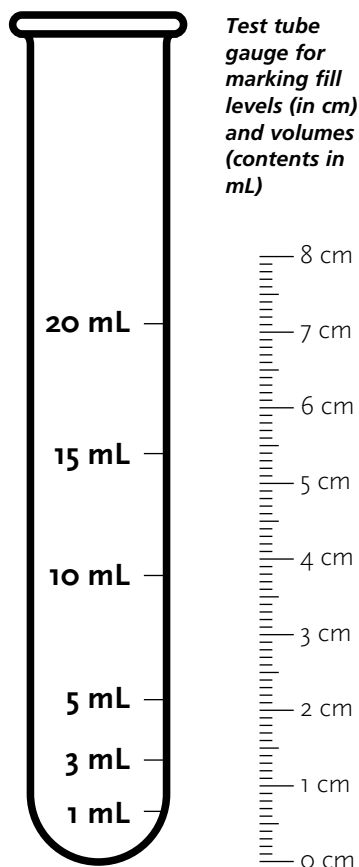
If other people have access to your work area, you will need to put all of the equipment and chemicals away somewhere that they cannot be reached by young children, preferably locked away in a cabinet.

### The equipment

Before you get started with the experiments, pull the polystyrene insert out of the box and place it on the work surface. The insert can be divided at the place shown on p. 4 using a sharp knife (ask an adult to do that for you). The larger part is used to store the equipment and chemicals, and the smaller part — the **experiment station** (part #24) — will serve as an aid to you while you're working. It is used as a stand for filled test tubes or opened chemical containers or as a holder for test tubes and graduated beakers for various reactions.



*Chem C1000  
experiment station  
(The actual design may vary  
slightly from the design pictured  
here)*



The most important device when performing chemical experiments is the **test tube** (part #8). It is used to hold liquid or solid substances, particularly to make solutions of solid substances. To clean the test tubes, use the **test tube brush** (part #9). But don't press too hard or you might break the thin glass. Cleaning is best done under running water. Place the cleaned tubes to drip dry in a clean container (e.g., a screw-top jar, see "Additional materials") with the opening facing downward.

Place filled test tubes into the small holes in the experiment station. The slanted shafts are for holding the test tubes that are set up with a **stopper with hole** (part #10) and **angled tube** (part #23) for certain experiments (more on this in the next section).

You will use the **stopper without hole** (part #11) to seal test tubes in which you dissolve substances in water with a shaking action. When shaking a test tube, always press on the stopper as shown in the illustration.



*When shaking a test tube, always press on the stopper with your thumb!*

For some experiments, you will need somewhat larger containers: the two **graduated beakers** (part #6) with **lids** (part #7). When performing experiments, a graduated beaker fits securely in the large recess in the experiment station, for example in Experiments 38 and 88. The small compartment to the side is for accommodating the square 9-volt battery. Using the beakers, you can measure 10, 25, 50, 75 and 100 milliliters. One milliliter (abbreviated as mL) is one-thousandth of a liter and corresponds to the cubic centimeter used for solids (abbreviated as cm<sup>3</sup>, cube having equal edges of 1 cm). In the test tube experiments, however, we indicate the quantity of liquid according to the approximate height of the column of liquid in the test tube, for example: "3 cm of water." To open a filled and sealed graduated beaker, place it on the table, hold it firmly, take a hold of the tab and pull the lid upward.

The **chemical vials** (parts #13 – 19) have two chambers: a large one for larger quantities and a small one for chemicals of which only a little is needed. The amounts added are as needed and are in line with the safety standards applicable to experiment kits. The uniform size of the vials is a result of the size of the labels, which contain information prescribed by law. The illustration shows how to open the **safety closures** using the **lid opener** (part #21). When opening a vial, sometimes a small amount of the chemical that is stuck to the lid will fall on your hand or on the work surface. This can be prevented by forcefully tapping the vial on the working surface several times before opening it. You can place opened chemical vials into the large recesses in the experiment station, since they easily tip over, which would make a mess and waste material. Once you have removed the amount you need, close the vial immediately. See p. 14 for an explanation of how to open and close the safety closure with **dropper insert** (part #20) of the bottle for the **litmus solution** (part #20).

If you have trouble opening or closing the safety closures, ask an adult for help.



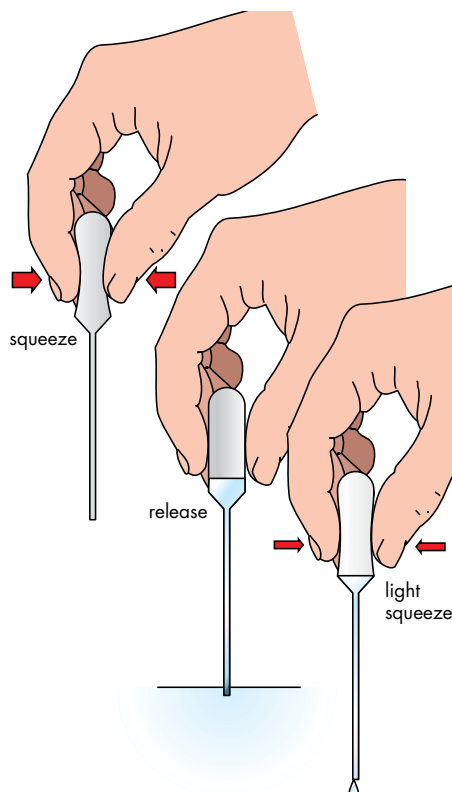
*This is how to open the chemical vials using the lid opener.*



### The double-headed measuring spoon

1 small spoon

1 large spoon



How the dropper pipettes work.

To remove the chemicals from the supply bottles, use the **double-headed measuring spoon** (part #22). When the instructions refer to "1 spoonful" or "1 large spoonful," that means one level scoop from the larger end, and "1 small spoonful" means a level scoop from the small end.

"1 spoon tip" means about half of the smaller end. You must wash and dry the double-headed measuring spoon after each use, since otherwise you will introduce trace amounts of the chemicals into other vials.

When you want to add liquids drop by drop, use the **dropper pipette** (part #13). The illustration to the left shows how to use it. Squeeze the top bulb of the pipette between your thumb and forefinger and dip the end into the liquid. As soon as you release pressure on the bulb, the liquid rises into the pipe. Then, you can release individual drops one by one by applying gentle pressure on the bulb. After use, clean the pipette by filling it repeatedly with water, shaking, and emptying.

You won't need the **funnel** (part #12) just for filling liquids (e.g., the litmus solution), but above all for filtering. You will learn how that works right off the bat in the first experiment. You will need the filter paper sheets included in your kit. If you run out of filter paper sheets, you can also use round **white coffee filters** for this, or you can trim yourself some round filters from **white filter bags**.

## Working with the angled glass tube

Placing the angled tube into the rubber stopper with hole and removing the tube from the stopper is not as easy as it looks. Ask an adult to help you with it.

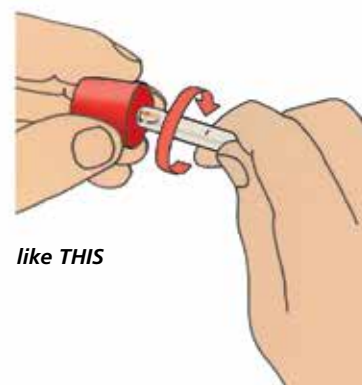
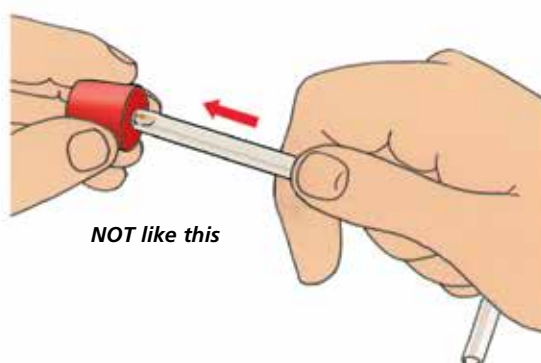
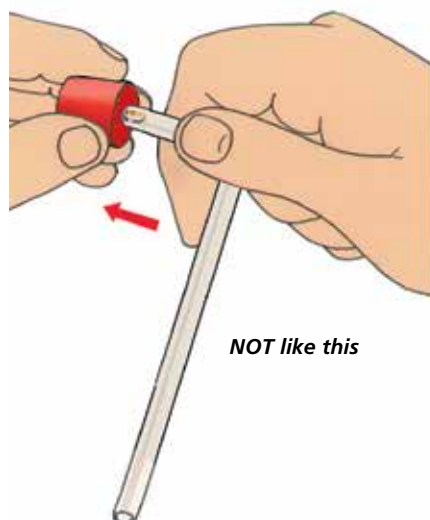


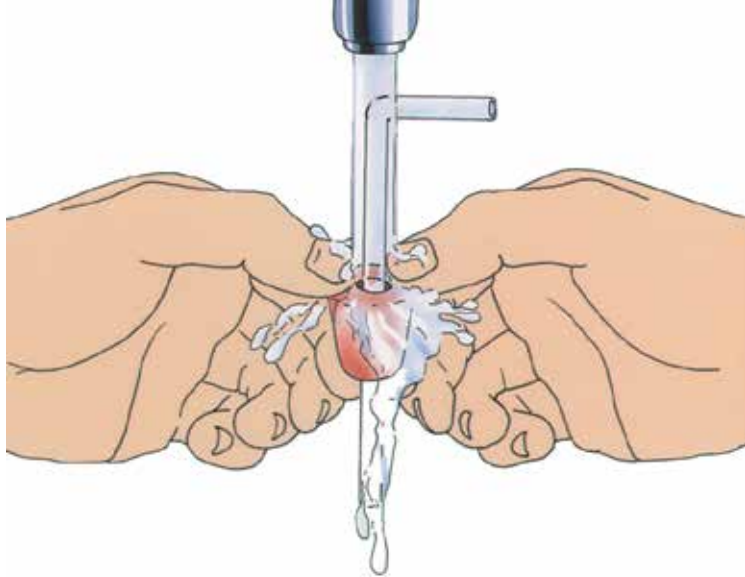
Broken glass tubes can cause unpleasant cuts and wounds. If an injury occurs: **First Aid 7** (inside front cover).

First hold the hole in the stopper and the angled tube under running water. Then twist the tube into the hole, preferably back and forth, and do not force it by any means. Hold the angled tube as shown in the illustration. You can give it extra protection by using a thick cloth to hold onto the glass tube.

It is often more difficult to take the angled tube out of the stopper than it is to put it in. To do this, gently pull the glass tube to both sides to form a small gap between the stopper and glass tube so that water can penetrate into it. Do the same thing on the opposite side of the stopper. Hold the glass tube right above the stopper and pull it out of the hole while constantly twisting it back and forth. **It is essential that you follow these instructions!**

The angled tube is not pushed into the stopper (a and b) but twisted in (c). Also, hold the long legs of the angled tube close to the stopper.





*A trick to make it easier to loosen the angled tube in the rubber stopper.*

## Additional materials

### Containers that don't cost anything and otherwise end up in the trash:

- Two or three transparent, clean jelly or jam jars with lids (called glass jars with lids in the manual)
- Plastic bowl or container like the ones ice cream or ground beef come in (called basins in the manual)
- Small, clean glass bottles

### Items that can be found in the household or easily obtained in the supermarket:

- Aluminum foil, paper towels, tealight candle, coffee filter (round filter or filter bags), cotton pads or balls, matches
- Demineralized (distilled) water, sodium bicarbonate (like Arm & Hammer® baking soda)
- White vinegar (wine vinegar)

### From the drug store or pharmacy:

- Hydrogen peroxide (100 mL of 3% aqueous solution)  
This substance is needed starting in Chapter 8 and should be given to the experimenter only as required.

Materials that are only needed in isolated cases (e.g., nails or paperclips), will be listed for the respective experiment.



*Materials that would otherwise end up in the trash.*

## Cleaning the containers

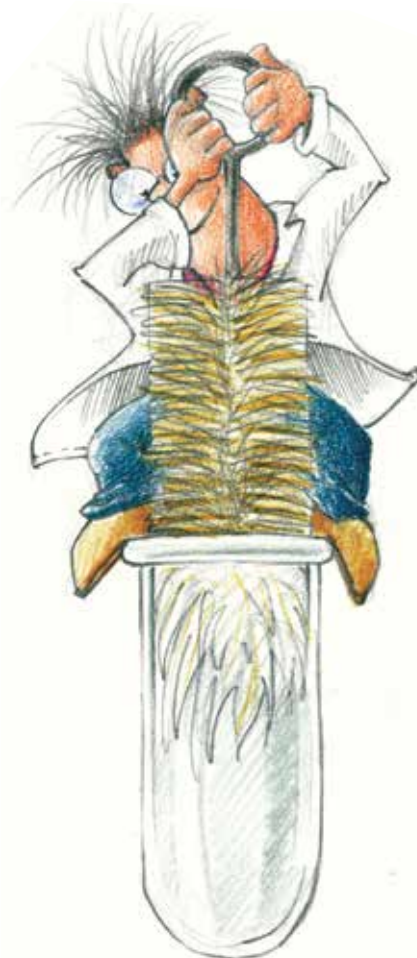
In chemistry, cleanliness is an essential prerequisite for the success of the experiments, since contaminants can trigger undesired secondary reactions. So make it a habit right from the beginning to carefully clean the used containers when an experiment is completed.

In most cases, cold water and a test tube brush are enough. If that doesn't do the trick, try warm water and a little dish-washing detergent. That will help especially with greasy or oily grime. Then rinse thoroughly!

Two tips:

**Lime residues** — white films after experiments with lime water — disappear after treatment with household vinegar.

**Berlin blue** (or Prussian blue) softens alkaline detergents. In both cases, rinse well afterward, of course.



# 3

## Magic blue and secret inks



Round filters and filter bags from which you can trim round filters.

For now, the magic blue is still hidden in the chemical vial labeled "Litmus powder." When you open the vial, you will discover a dark, fine-grained substance inside. To perform experiments with it, you will need to prepare a litmus solution, which lasts one day. You know: sugar and salt dissolve in water so easily that it seems as if it disappears. It's not quite so easy with the litmus powder. First you need to become acquainted with one of the most important laboratory techniques: **filtering**.

### Super-sieves in action

You will need the filter paper sheets from your kit for these experiments. If you run out of filter paper sheets, you can always use white coffee filters for filtering: either the round ones or the larger filter bags out of which you can cut round filters (diameter approx. 9 cm). In the experiments, we will call these filters "filter paper."

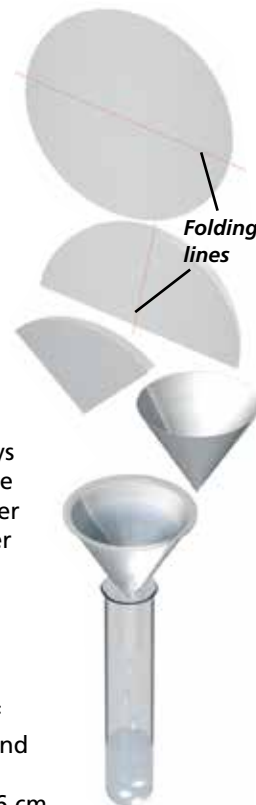


The solution prepared from litmus powder is filtered.

#### EXPERIMENT 01

##### Additional material: Sand

Fold a round filter paper as shown in the illustration. You will end up with a cone consisting of one layer of filter paper on one side and of three layers on the other. Place the filter cone into the funnel and moisten it with a little water. This will help it stick better to the wall of the funnel. In a sealed test tube, shake some sand with 6 cm of water (remember, keep your thumb on the stopper!) and pour the mixture into the filter cone. The sand remains in the filter and a nearly clear liquid, the filtrate, drips into the test tube below.



### What's happening here?



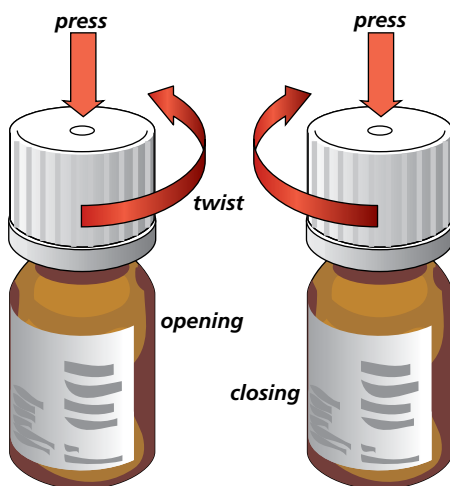
Sand is made of small quartz crystals. The particles are difficult to dissolve or not soluble at all in water and too large to pass through the tiny pores of the filter paper. In contrast, the particles of water and the soluble substances are so small that they overcome the "filter blockade" with no trouble. By using super-fine sieves, you can separate the soluble from the insoluble components of a mixture.

#### EXPERIMENT 02

**Preparing the litmus solution.** Place 3 cm of water in a test tube and add 3 small spoonfuls of litmus powder to it (level scoop). Close the tube with the stopper, shake

vigorously and allow the closed tube to stand for one day somewhere that is out of the reach of young children.

Now set up the funnel and filter for filtering like in the previous experiment. Place the funnel on the vial provided for the litmus solution and pour the deep-blue mixture into the filter. You can dispose of the insoluble leftovers in the trash. If denatured alcohol (careful, fire hazard!) is available, an adult should add a half pipette of it to the vial.



How the safety closure of the vial for the litmus solution works.



Litmus is a plant product with a limited shelf life.  
The denatured alcohol acts as a preservative.

DANGER



**Denatured alcohol:** Highly flammable liquid and vapor. Keep away from heat/sparks/open flames/hot surfaces. – No smoking. – Keep container tightly closed.

Now place the safety closure with dropper insert onto the vial and close it (by turning the closure clockwise). The illustration on the previous page shows how to close and open the safety closure by turning and pressing it at the same time.

When refilling (recommended due to the limited shelf life), the dropper insert must first be removed. Have an adult help you with this if necessary. Okay, now the litmus solution is ready to use.



**Question 1.** The safety closures are designed to prevent small children from opening the vials. What features make this a safety closure, and why?\*

## Blue here, red there

### EXPERIMENT

03

**Additional material:** White vinegar (wine vinegar)

To a test tube with 3 cm of water, add 3 drops of litmus solution. The solution is light blue. Pour a little vinegar into a test tube and drip it into the litmus solution using a pipette. The first drops already cause the color to turn bright red. You will need this solution for the next experiment.

Regarding Experiments 4 – 8:



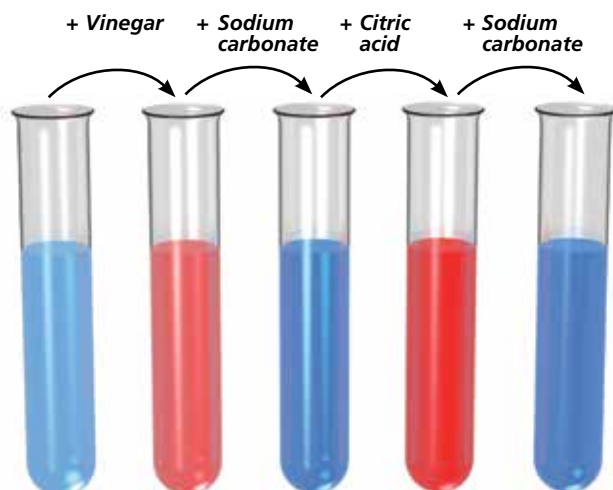
For **sodium carbonate** and **citric acid**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

### EXPERIMENT

04

Add a spoon tip of sodium carbonate to the red solution from the previous experiment. Rock the tube gently back and forth: The solution turns blue again, this time

a bit darker than the aqueous (watery) solution in Experiment 3. Don't throw this solution away, either!



Blue here,  
red there

## Side Notes

### The optical brightener of the Netherlands

Litmus is a mixture of plant materials that is obtained above all from lichens such as *Rocella montagnei* shown below. The Dutch once used the dye for “bluing” their laundry. The blue covers the yellow tint in textiles and thus produces the “whitest of whites.” Today, the “optical brighteners” contained in detergents — products of chemical ingenuity — do the trick.

The name comes from the Netherlands, too: lakmoes, a Dutch word whose origin linguists are not completely in agreement about. There is pretty much a consensus about the second syllable, moes, meaning puree, pulp, mush. Whereas some think that lak means paint (hence “colored pulp”), others say that hidden in lak is the Old Dutch word lêken, meaning to drip. Both interpretations are consistent with the preparation process. The lichens were ground with water between millstones into a colored pulp that was allowed to ferment and drain. The dried deep-blue mass was then ground into powder.



The lichen *Rocella montagnei*, from which litmus powder is obtained (photo: Prof. Dr. V. Wirth, Karlsruhe).



Pressed litmus (© and photo: chemie-master.de)

## EXPERIMENT 05

Add 1 spoon tip of citric acid (written as C in technical language) to the blue solution. Even without rocking it back and forth, the color changes to red. Save the solution for the next experiment!

You can probably guess what comes next.

## EXPERIMENT 06

Add some sodium carbonate to the red solution. If it remains red, you put too much citric acid in before. You will then need to add a bit more sodium carbonate to make the color change back.

It will probably come as no surprise to you anymore that lemon juice (chemists abbreviate it with L) turns blue litmus solution red. Try it out!



Acids and bases in the laboratory

## What's happening here?



There are two types of substances that battle with each other and react differently with the magic blue: the **acids** and the **alkalis**. The acids include the acetic acid contained in vinegar and the citric acid contained in lemons and lemon juice. The alkalis include sodium carbonate, which is important for hobby chemists. The aqueous solutions of the alkalis are also called **lyes** in everyday language or **bases** in more technical language. Sodium hydroxide and caustic potash solution play an important role in professional laboratories. When sodium carbonate is dissolved, some sodium hydroxide is formed.

Your experiments have shown:

- **Acids** turn blue litmus solution red (acidic reaction)
- **Bases** (solutions of alkalis) turn red litmus solution blue (alkaline reaction)

The litmus solution also shows whether a solution has an acidic or alkaline reaction. Such sleuths among the substances are called **indicators** (from Latin *indicare* = to indicate).

The words "acid" and "acidic" also come from Latin (*acidus* "sour, sharp"). The words "alkali" and "alkaline," on the other hand, can be traced back to the Arabic word *alqali*, which means something like "plant ash." The alkaline effect of plant ashes, which were used to make soap, was known even in antiquity. You will become acquainted with alkaline-reacting soaps in Chapter 9.

## Working magic with magic blue

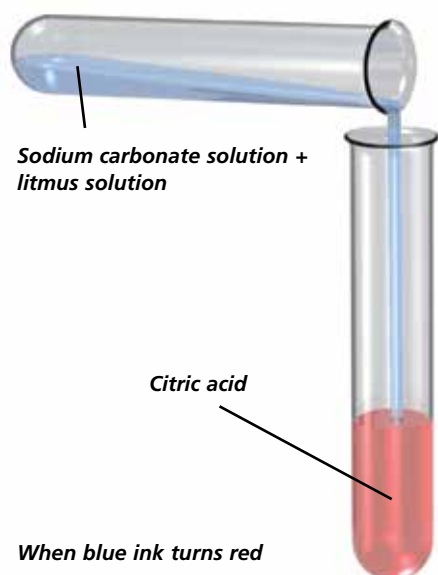
You've probably already asked yourself why we've been calling litmus dye magic blue. It's because you can do some scientific "magic" with it, of course — for example **by turning blue ink red**. However, a good magician can't allow the audience to get too good a look at the cards or, in this case, in the test tube.

## EXPERIMENT 07

Prepare a blue "ink" in a test tube composed of 5 cm of water and 10 drops of litmus solution. You can deepen the hue using a few specks (!) of sodium carbonate. Add 1 spoon tip of citric acid to a dry test tube. If you hold the tube in such a way that the round base is at eye level with your spectators, they won't be able to see the crystals in it. Now you chant the proper spell as you pour the "blue ink" into the "empty" test tube and presto! Amazement all around.



**Question 2.** How can you turn red ink blue — that is, the other way around from the experiment? Indicate the formula.





## The gas from the bottle of mineral water

When you open a bottle of mineral water, lots of little bubbles of gas rise up. This gas is called **carbon dioxide**. There are mineral waters that naturally contain carbon dioxide. In most cases, the gas is pressed into the mineral water. When the bottle is opened, a portion of the “cooped-up” gas escapes in the form of the observed bubbles. You will perform more experiments with this gas in Chapter 6.

### EXPERIMENT 08

**Additional material:** Sparkling mineral water that is as fresh as possible

Fill a graduated beaker up to the 25 mL mark with water and add 1 spoon tip of sodium carbonate and 10 drops of litmus solution to it. Now add a shot of mineral water to the beaker. You won't be surprised about the red coloration; after all, you're very familiar with the sour taste of mineral water.



### What's happening here?



Most of the carbon dioxide pressed into mineral water dissolves into the water just like salt and sugar do. A small portion of the gas combines with the water to form **carbonic acid**. If there is a note on the bottle of mineral water that says “Laced with carbonic acid,” then it's not completely correct, chemically speaking. The water was laced with carbon dioxide gas, a small portion of which then formed carbonic acid with the water. By the way, the carbonic acid didn't just redden the litmus dye but had to first “get rid of” the added sodium carbonate so that the color change could happen. Instead of “get rid of,” we say “neutralize” in this case (more on this in Chapter 8: Water and its elements).

So you can detect carbon dioxide and carbonic acid using blue litmus solution. But the red coloration is not characteristic — in technical language, we say, “not specific” — since other acids have the same effect, as you found out. You will now prepare a specific detection agent for carbon dioxide and get to know another substance from your assortment of chemicals in the process: calcium hydroxide.



If you have a mineral water maker at home, you can prepare your mineral water yourself. The device contains a cartridge that contains pressurized, liquefied carbon dioxide that is pressed into the screwed-on bottle when the valve is opened.

## From lime milk to lime water



For **calcium hydroxide** and **lime water**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

### EXPERIMENT 09

Add 1 spoonful of calcium hydroxide to a test tube that is half-filled with water, close the tube with the stopper and shake vigorously (thumb on the stopper, wearing safety glasses!). A white liquid is formed that is called “lime milk” because of its appearance but of course cannot be drunk. Let the closed tube stand for a while out of the reach of young children. You will then see that the white powder has settled on the bottom of the tube but that the liquid above it is still cloudy. Will the cloudiness be held back (cleared up) if we filter it?



**Calcium hydroxide** — important for detecting carbon dioxide



Lime milk, lime water, and lime water with litmus solution

## EXPERIMENT 10

Place a filter in the funnel and place the funnel onto an empty test tube. Now pour the lime milk bit by bit into the filter. If the filtrate is still cloudy — because the filter's pores were a little too large — filter the cloudy liquid through a “double filter” (place two filter papers on top of each other and fold or cut a double filter from a filter bag). Keep the filtrate for the following experiments.

Have the water and calcium hydroxide separated again? The following experiment will provide the answer.

## EXPERIMENT 11

**Additional material:** White vinegar

Prepare a red, that is, acid-reacting solution in a test tube composed of 5 cm of water, 5 drops of litmus solution and several drops of vinegar. Add half of the clear filtrate from Experiment 10 to this solution. You will need the other half of the filtrate for Experiment 12.



**Question 3.** What do you observe? What can you conclude from this?

## What's happening here?

If you shake calcium hydroxide with water, a small portion of the substance dissolves but the majority sinks slowly to the bottom or can be filtered off. The clear, alkaline-reacting substance is called lime water, not to be confused with hard tap water (you will deal with that in Chapter 9). So calcium hydroxide and lime water are among the opponents of the acids. The white cloudiness in Experiment 12 — called precipitate in technical language — consists of hard-to-dissolve calcium carbonate. Calcium carbonate is formed when carbon dioxide acts on lime water. In Chapter 9, you will also find out why the precipitate disappears again when excess carbon dioxide is present.

## Turning water into milk:

Magic blue isn't the only thing you can do magic tricks with: as unremarkable as it may look, lime water is great for magic tricks too. To do this, you pitch the experiment last performed as a transformation of water into milk. An important note, though: You must not drink your magic drinks or give them to anyone else to drink!

**TIP**

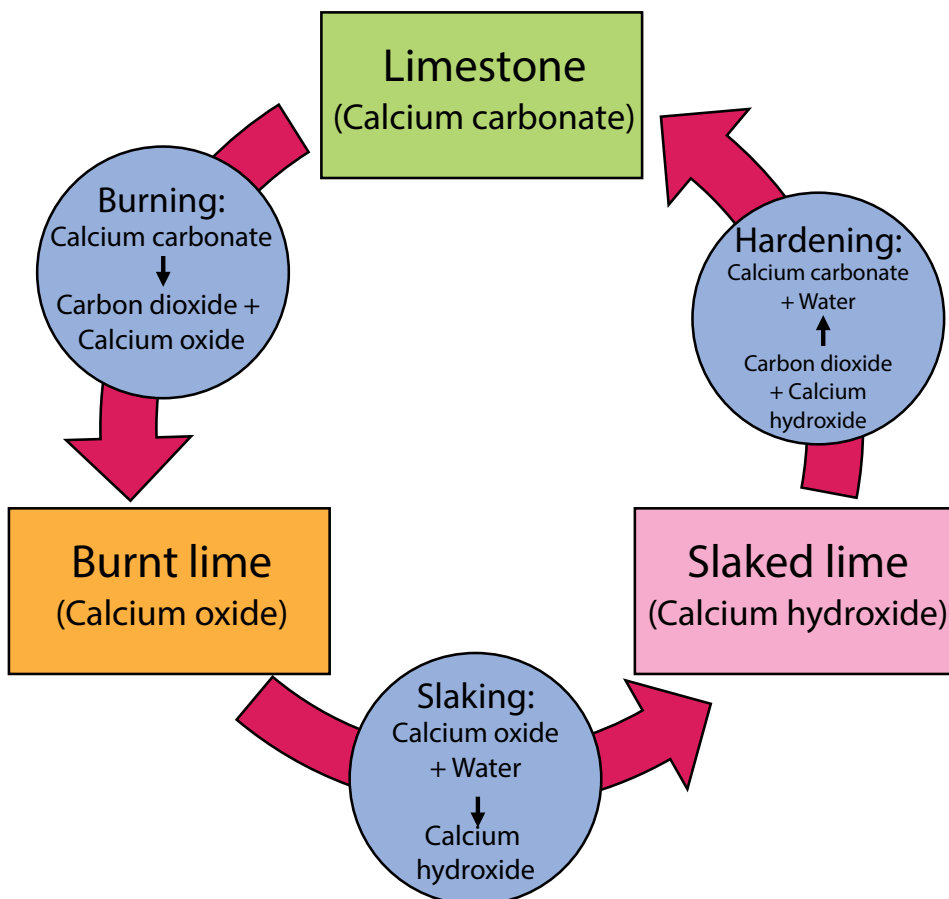
## A precipitate indicates the presence of carbon dioxide

Now you will use the alkaline-reacting filtrate from Experiment 11 as a means for detecting carbon dioxide.

## EXPERIMENT 12

**Additional material:** Mineral water that is as fresh as possible

Fill a test tube halfway up with mineral water and pour the mineral water into the remaining lime water in small portions. More and more cloudiness forms that looks very similar to lime milk. If you add more fresh mineral water, the cloudiness may disappear, meaning the contents of the test tube may become clear again. In all cases, the cloudiness disappears when you add vinegar or citric acid. Try it!





## From limestone to lime

The detection of carbon dioxide using lime water plays an important role in chemical investigations, since the gas — as you will soon see — is involved in a wide variety of processes. But by far the largest quantity of calcium hydroxide is used for quite another purpose: building houses. In regions with abundant lime, limestone — calcium carbonate, chemically speaking — is broken down and often “burned” right then and there in tall shaft furnaces. The burning breaks the carbonate down into calcium oxide and carbon dioxide. The calcium oxide, “quicklime” or burnt lime, is “slaked” through the addition of water (as in the expression “to slake your thirst”). Slaking is the term used by those working in construction to refer to the conversion of the calcium oxide into calcium hydroxide, which they therefore call slaked lime.

In earlier times, the slaking was done exclusively at the construction site, but nowadays it takes place in lime plants for the most part. Using slaked lime, sand, and water, the bricklayers then prepared the mortar that holds the building stones together as a binder. In this step, exactly the opposite happens as during the burning of the lime. The calcium hydroxide component of the mortar absorbs carbon dioxide from the air and hardens again into calcium carbonate, into limestone. This process is referred to as the “carbonation” or “hardening” of the mortar. Chemically speaking, the same reaction takes place as when detecting carbon dioxide using lime water: the formation of calcium carbonate — in one case as the hardening of the mortar and in the other case as the formation of a precipitate.



Shaft furnaces in which slaked lime is manufactured from limestone (photo: Dyckerhoff AG, Lengerich plant)

## Another blue — this time from Berlin

In the following experiments, you will be introduced to two new chemicals: ammonium iron(III) sulfate and potassium hexacyanoferrate(II). You must use the latter of the two substances sparingly. It is harmful to aquatic organisms but is indispensable to you. So prepare a potassium hexacyanoferrate(II) solution that you will need for a number of experiments.



For **ammonium iron(III) sulfate** and **potassium hexacyanoferrate(II)**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.



### Preparing Potassium

#### hexacyanoferrate(II) solution:

Dissolve 2 spoonfuls of potassium hexacyanoferrate(II) in a test tube in 6 cm of water and store the solution in a carefully cleaned, labeled vial. Remove only the amount needed in each case (usually only a few drops). You can throw the reaction products (most notably Berlin blue) into the drain with a little detergent. Rinse well afterward.

First let's allow the solutions of the two substances react with each other in a test tube.

### EXPERIMENT 13

Add 1 spoon tip of ammonium iron(III) sulfate to a test tube filled halfway up with water, seal the tube and shake. The light-violet crystals dissolve and you obtain a faintly yellowish-brown solution. If you now add a few drops of the potassium hexacyanoferrate(II) solution you just prepared, you get a splendid blue: **Berlin blue**, to be precise.

If the ammonium iron(III) sulfate has melted and solidified again, you can loosen the crystal plug from the storage vial, dry it with absorbent paper, and grind it between two layers of clean paper using a heavy object (for example, a hammer, but don't pound, just press). Then place the substance back into the dried storage jar.



## EXPERIMENT 14

Add 1 drop of ammonium iron(III) sulfate to a small piece of filter paper. Rinse the pipette out well (or use the other pipette) and now put 1 drop of potassium hexacyanoferrate(II) solution in the same place. A deep Berlin blue forms on the paper.



### What's happening here?

Calcium hydroxide (pronounced KAL-see-um hi-DROX-ide), potassium hexacyanoferrate(II) (pronounced po-TASS-ee-um hex-a-sigh-ann-o-FARE-ate 2), ammonium iron(III) sulfate (pronounced ah-MOAN-ee-um EYE-urn 3 SUL-fate)... why do chemical names have to be so complicated and such tongue-twisters? Because to those in the know — and that means you, too, sooner or later — they say something about the composition of substances.

Over time, scientists found out that there are certain substances that cannot be broken down into simpler ones. They are called **elements** or “basic materials,” and there are less than 120 of them. They include the light metals calcium, sodium, and potassium. If two or more elements combine with each other, they form **chemical compounds**. Calcium hydroxide, sodium carbonate, and potassium hexacyanoferrate(II) are examples of chemical compounds. You can see immediately from the names that they contain the elements calcium, sodium, and potassium, respectively.

The other elements in the compounds above are a bit more concealed. If a compound contains “carbonate” in its name, then we know it contains carbon. If its name contains “hydroxide,” we know that it contains hydrogen and oxygen. When components occur more than once in a compound, chemical names also include Greek numerals, for example *di* for two and *hexa* for six.

We will reveal some more about the Roman numerals (II and III) in chemical names at the end of Chapter 7. Take a look at the chemical elements table below. There you will find the elements that you'll encounter in your experiments, some as stand-alone elements (in copper wire, for example) and some as compounds (in copper sulfate, for example).

### Chemical elements

Name	Latin or Greek name (chemistry symbol)	Type
Aluminium	Aluminium (Al)	Light metal
Calcium	Calcium (Ca)	Light metal
Carbon	Carboneum (C)	Non-metal
Chlorine	Chlorum (Cl)	Gas
Copper	Cuprum (Cu)	Heavy metal
Hydrogen	Hydrogenium (H)	Gas
Iodine	Iodum (I)	Non-metal
Iron	Ferrum (Fe)	Heavy metal
Nitrogen	Nitrogenium (N)	Gas
Oxygen	Oxygenium (O)	Gas
Potassium	Kalium (K)	Light metal
Sodium	Natrium (Na)	Light metal
Sulfur	Sulphurium (S)	Non-metal

*In 1661, Robert Boyle (1627 – 1691) named substances that cannot be broken down into simpler ones “elements” (photo: Deutsches Museum, Munich).*



**Question 4.** Potassium hexacyanoferrate(II) is a compound that contains a total of six elements. One is of course potassium. Using the chemical elements table above and a little of your feel for language, can you identify another of them based on the name?





## Blue secret ink

On some special occasion, you may want to send a secret message that only the addressee can decode. You can do it using the following secret ink.

*Chemical elements at play in your experiments: Copper (1), iron (2), carbon in the form of activated charcoal (3), iodine in crystalline form (4), sodium under petroleum (prevents oxygen from getting in) and cut on the glass plate (5), aluminum as foil and grit (coarse powder) (6), sulfur as powder and in pieces (roll sulfur) (7), calcium in grainy form and in larger chips*



For **ammonium iron(III) sulfate**, **sodium carbonate**, and **potassium hexacyanoferrate(II)**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

Filter paper or cotton balls that are soaked or moistened with chemical solutions should only be touched with protective gloves.

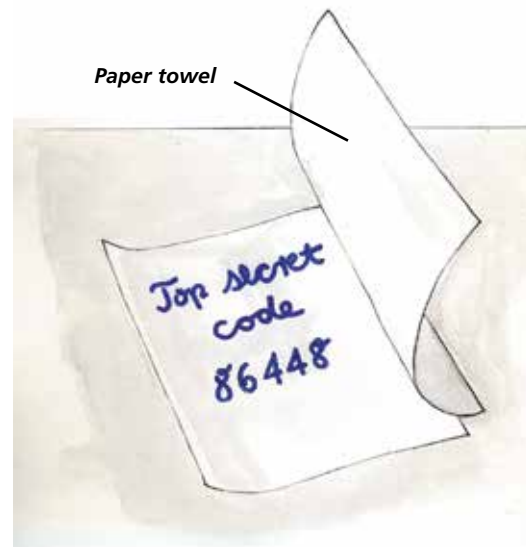
For the following experiments, you will need a **steel fountain pen** or a **fine hair paint brush** from a painting kit.

### EXPERIMENT 15

**Additional material:** Paper to write on, paper towels

The ammonium iron(III) sulfate (1 small spoonful in 2 cm of water) is the ink with which you write your secret message onto paper (yellowish paper works best). After it's dry, your writing is invisible.

Paper towel



**Question 5.** How can the addressee render your message visible again? Keep the message for the next experiment.

### EXPERIMENT 16

**Additional material:** Cotton

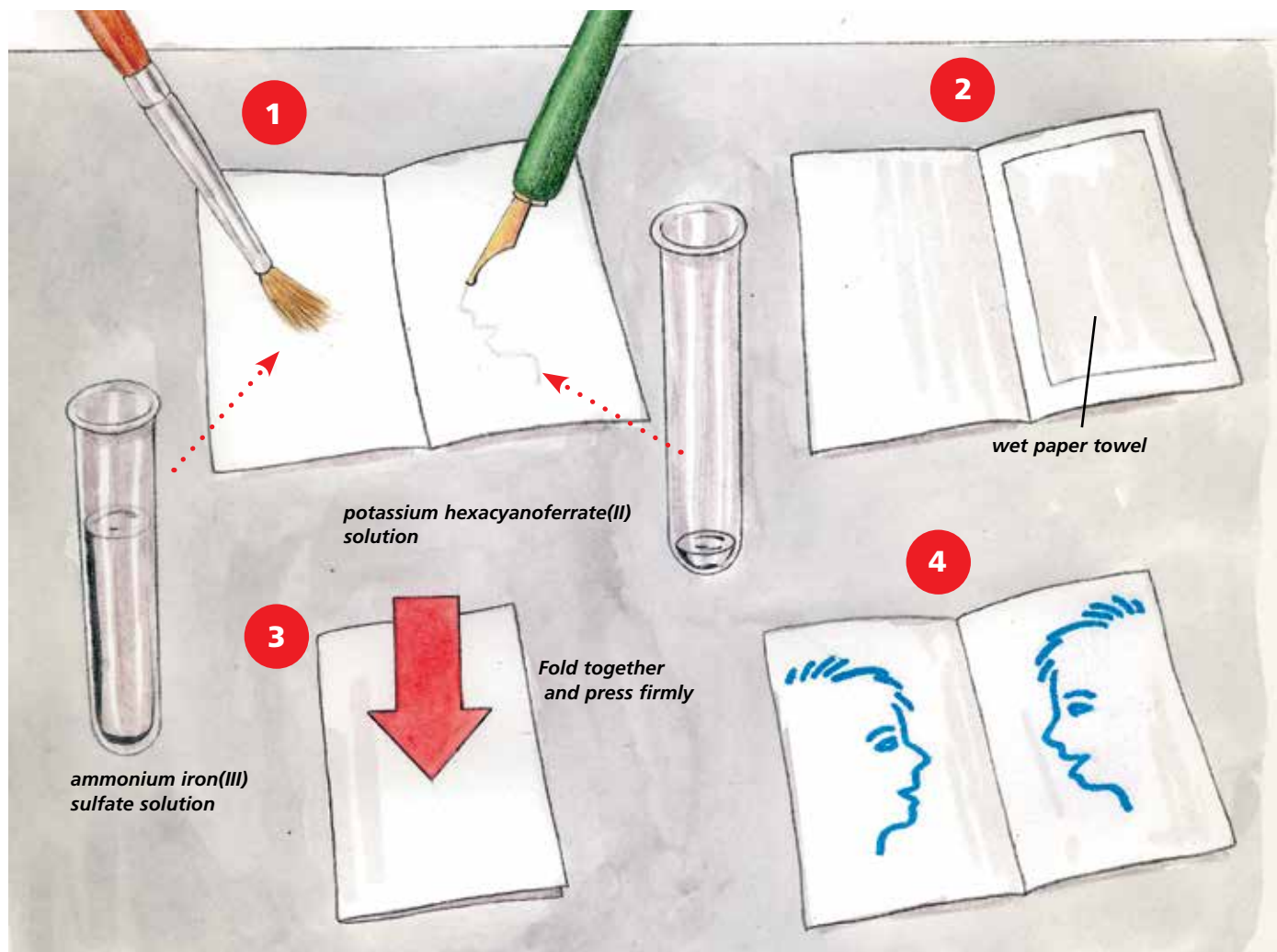
You can make the Berlin blue writing disappear again, too. Add 1 cm of water to each graduated beaker. In one beaker, dissolve 1 small spoonful of sodium carbonate, and dissolve 1 small spoonful of citric acid in the other. Stir each with the spoon, rinsing the spoon off thoroughly before stirring the second beaker.

Dab the writing using a cotton ball that you have previously dipped into the sodium carbonate solution. Wait until the paper is dry and now dab it with a cotton ball that you have moistened with the citric acid solution.



**Question 6.** What can you conclude from the dabbing experiments?





## Image and mirror image

With ammonium iron(III) sulfate and potassium hexacyanoferrate(II), you can make other “printing products” too — a folded image, for example.

### EXPERIMENT 17

**Additional material:** White paper to write on, paper towels

1 Fold a white sheet of writing paper (half of a standard 8½ by 11 inch sheet) as shown in the illustration. Using the paint brush, moisten the inside completely with ammonium iron(III) sulfate solution (1 small spoonful in 2 cm of water). On the other side, draw or paint a picture (a profile, for example) using the potassium hexacyanoferrate(II) solution. Let the whole thing dry. 2 Then moisten a paper towel with water and press it briefly onto your drawing. Remove the moist paper towel. 3 Press the two sides of the folded sheet firmly together. 4 When you open the folded sheet, the drawing and its mirror-inverted impression emerge.

You can also send secret messages this way. This way, the addressee doesn’t need any special chemicals to “develop” it — just a paper towel and water.

## Secret inks from the kitchen

But you don’t even necessarily need special chemicals for the secret ink itself. It also works with lemon juice, household vinegar, and onion juice.



Secret inks that are available in any household



In the following experiments, be careful not to burn yourself. Ask an adult to help you.

For the following experiments, you will also need **white writing paper** and **barbecue tongs**.

### EXPERIMENT 18

Squeeze the juice out of a lemon and use the juice as "ink." The pen or brush should not be too thin. After it's dry, you won't be able to see a trace of this ink, either. To make the ink visible, you will need to heat up the paper a lot without allowing it to go up in flames. Using the barbecue tongs, hold it over (not on!) an electric heating plate. What do you see?

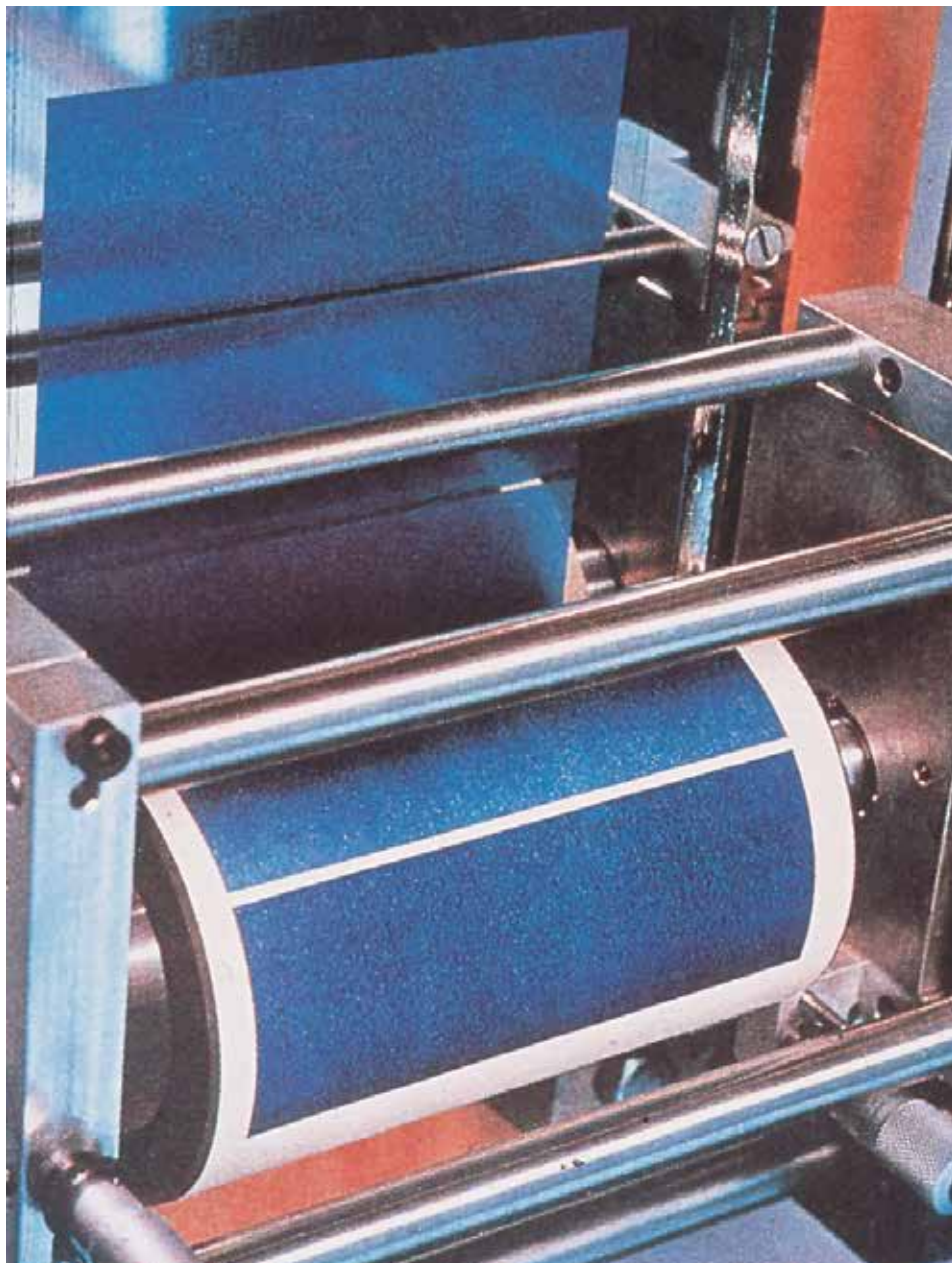
### EXPERIMENT 19

Repeat the experiment, this time using household vinegar. Be patient, though — it takes a bit longer for the writing to appear.

### EXPERIMENT 20

Preparing onion juice will probably bring you to tears. The strong-smelling ingredients in onions irritate the eyes like tear gas. It's best to prepare the juice using a garlic press. Now you can test out the juice as a secret ink just like in the previous experiments (heat it up well again).

Unlike the Berlin blue ink, these secret inks from the kitchen are made visible by charring (or "carbonization") of the carbon-containing compounds contained in the "inks." As a result, the carbon becomes visible (brown to black coloration).



## Side Notes

### Berlin blue — an accidental discovery

In 1706, a paint manufacturer from Berlin named Diesbach was mixing a scale insect brew (yes, you read that correctly) with alum, iron sulfate, and an "alkali" that had been discovered not long before for making a carmine red paint. The alkali was the stinky animal oil prepared by Johann Conrad Dippel (1673 – 1734) from animal blood. Diesbach was extremely surprised when, instead of the red he wanted, he got a deep blue product.

We now know that "Dippel's oil" (as it was called) contained a number of substances, some of which were very hazardous. Highly toxic cyanide may have been present in it, or yellow blood-lye salt, an old name for the potassium hexacyanoferrate(II) you're using. Which name do you like more? In both cases, Berlin blue can be produced together with iron compounds.

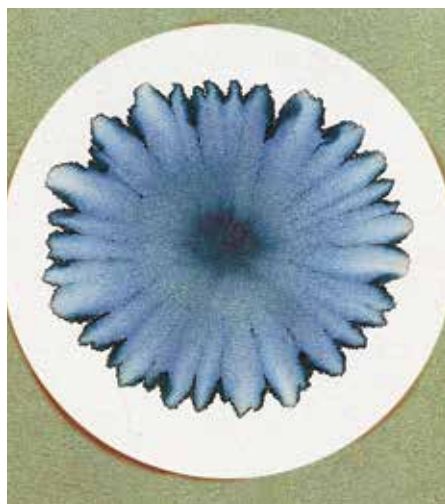
Great success came from this accidental discovery. The merits of the new dye were praised: it was suitable for painting with water or oil or for printing wallpaper, it didn't fade when exposed to light, it was resistant to acid and posed no health hazards (unlike others, for example poisonous paints containing arsenic). In the 19th century, expansive factories for the production of Berlin blue sprung up.

Even after the discovery of a number of new dyes, Berliner blue is still used, especially in the printing industry (photo: Degussa AG, Wolfgang).



# 4

## Test papers and changing colors



*Blooming ink stain*

By now you've noticed: Colors play a large role in chemistry — for example the blue and red of litmus dye or the striking Berlin blue. You've also seen that chemical reactions need not necessarily take place in the test tube or even larger containers, but that colorful reactions can make their mark quite nicely on filter paper, too. And filter paper will be home to more surprising plays of color in this chapter. What's so special about filter paper, anyway?

### Two kinds of paper

#### EXPERIMENT 21

**Additional material:** *White writing paper*

Cut out two circles of equal size from the filter paper and writing paper you're using (diameter approx. 9 cm). Then place the filter made of filter paper into the funnel. Pour 6 cm of water into a test tube and pour the water into the filter. Note the time (simply by counting) that is needed until all of the water has passed through the filter.

Repeat the experiment with the writing paper filter. Moisten it, too, so that the paper sticks to the wall of the funnel.



**Question 7.** What did you observe in comparing the two filters?

#### EXPERIMENT 22

**Additional material:** *White writing paper, wooden skewers or knitting needles*

Fill a screw-top jar with 5 cm of water and place the wooden skewer or knitting needle over the opening. Cut 2 cm wide strips out of writing paper and filter paper that are long enough that they dip into the water when you hang them next to each other over your "curtain rod."



*True papyrus (photo: Liné1, wikipedia, CC-BY-2.0)*



*Writing paper and filter paper when dipped into a liquid*



**Question 8.** What do you observe? You can make it easier to see if you use blue or red litmus solution instead of water (see Exp. 2 and 3).

## What's happening here?



Experiment 21 shows that writing paper isn't suitable for filtering. On the other hand, you can't write very well on filter paper, not even on the best kinds from professional laboratories. At least not with ink. The writing would run. Writing paper is fully glued and made to be extra smooth. Filter paper isn't coated, so liquids that are put on it have unimpeded access to the inside of the paper. The inside consists of tiny **capillaries**, or tiny hair-like tubes. Due to their small diameter, the capillaries exert so much attractive force on the water that it is able to climb up the paper against gravity (Experiment 22).



*Modern paper manufacturing. Highly diluted with water, a mixture of cellulose, wood pulp, and other substances runs on the continuous paper sieve, on which sheets are formed through fiber mingling and dewatering (photo: Verband Deutscher Papierfabrikanten e. V., Bonn).*

## Test paper for copper and iron

In many cases, it's easier and faster working with test papers than with reactions in test tubes.

Your acquaintance potassium hexacyanoferrate(II) is a widely used, versatile detection agent. Here you'll make a test paper out of it with which you can detect not only iron compounds but copper compounds, too.



For **potassium hexacyanoferrate(II)**, **ammonium iron(III) sulfate**, and **copper sulfate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8 as well as the instructions for disposal for **copper sulfate** on p. 75.

When performing experiments with soaked papers, wear your protective gloves.

### EXPERIMENT 23

Here again, cut out several 2 cm wide strips of filter paper. Soak them in potassium hexacyanoferrate(II) solution using the technique according to the tip on p.

19. Place the test strips on paper towel(s) and allow them to dry. You will need them for the following experiments.

### EXPERIMENT 24

In a test tube, dissolve 1 spoon tip of ammonium iron(III) sulfate in 2 cm of water; in another tube, dissolve 1 spoon tip of copper sulfate in 2 cm of water as well. Place 1 drop

of each solution onto separate areas of the test papers you made in the previous experiment. The solution of the iron salt makes a blue spot, as expected: Berlin blue. The copper salt brings about a reddish-brown color. Keep the solutions for Experiment 25.

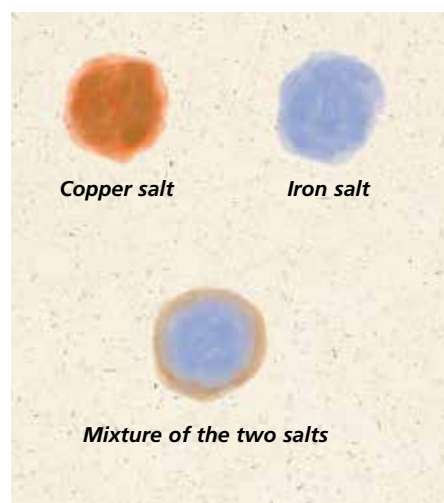
## Side Notes

### The story of paper

Paper has a shrub to thank for its name — the papyrus shrub. Around 4000 – 3000 B.C., the Egyptians had a need for material to write on. They produced it from the pulp of the shrub, which they cut into narrow strips and processed into thin sheets. Sheets that were written on, called papyri (plural of papyrus) for short, have been found in Egypt in dry layers of soil.

The inventor of modern paper manufacturing was a Chinese man who had the idea of suspending plant fibers in water in 105 A.D. Individual sheets were then formed by hand from the suspension with the aid of sieves.

Like other inventions from the Far East, the method was brought to Europe by the Arabs. But it was quite a long time until the first paper mills were constructed in Europe in the 13th century. Paper mills are water-powered stamping mills (driven by "mill wheels") for breaking down the raw material, which consisted of rags for the most part. A good many inventions and improvements to the papermaking process were necessary before paper could become the high-quality mass product it is today.



*How the solutions of ammonium iron(III) sulfate and copper sulfate react with the test paper (upper illustration). And this is how the mixture of solutions reacts: Experiment 25 (lower illustration).*





## What's happening here?

You're probably puzzled by the terms "iron salt" and "copper salt." Why "iron salt"? Iron is iron and salt is salt, right? — In chemistry, though, there's not just table salt that you sprinkle on your breakfast eggs, but a whole group of compounds that are referred to as **salts**. Common characteristics are their regular structure (crystalline form) and the conductivity of their solutions for electric current (Experiments 75 and 76).

Iron salts contain the element iron, and copper salts contain the element copper. Table salt is a sodium salt. It contains the elements sodium and chlorine and is therefore called "**sodium chloride**" by scientists. Another sodium salt you already know is **sodium carbonate**.

In the last experiment, the blue iron hexacyanoferrate(II) and the reddish-brown copper hexacyanoferrate(II) were produced. These are also salts — as is potassium hexacyanoferrate(II).



Not just table salt is a salt. From left to right: Sodium chloride, copper sulfate, sodium carbonate, ammonium iron(III) sulfate

### EXPERIMENT 25

Mix the two solutions from the previous experiment and add 1 – 2 drops of the mixture to a new test strip prepared in Experiment 23. A dark stain appears, the color of which is hard to determine. But after a few seconds, a brownish edge should appear. To achieve better separation, let's alter our method: Save the solution mixture for the next experiment.

## Let the salts climb

### EXPERIMENT 26

**Additional material:** Wooden skewer or knitting needle, paper towels

Fill a screw-top jar with 1 cm of water. Cut out a 3 cm wide strip of filter paper so that it dips a few millimeters into the water when it is hung over the skewer (the illustration shows the experimental setup). Take the dipped strip out — it will serve merely as a "pattern" — and then cut out a strip of the same length. In the middle of the strip, draw a dot about 1.5 cm from the bottom edge using a pencil. Now add 1 drop of the solution mixture from the previous experiment onto the dot and allow the drop to be absorbed. If a blow-dryer is available, use it to dry the strip. Add a second drop to the dot. The moist spot should not reach the edge of the strip. Hang the strip over the skewer/knitting needle and observe how the water rises in the filter paper. When the leading edge of the water is only 2 cm from the upper edge, remove the strip from the glass and allow it to dry (it's best to hang it somewhere).

Place the dried strip onto paper towel. Soak an approximately 2 × 10 cm piece of filter paper with potassium hexacyanoferrate(II) solution using the dripping technique (see tip on p. 19) and press it onto the dry strip. When you lift the hexacyanoferrate paper off, you get an "image" that should look something like that shown to the left.

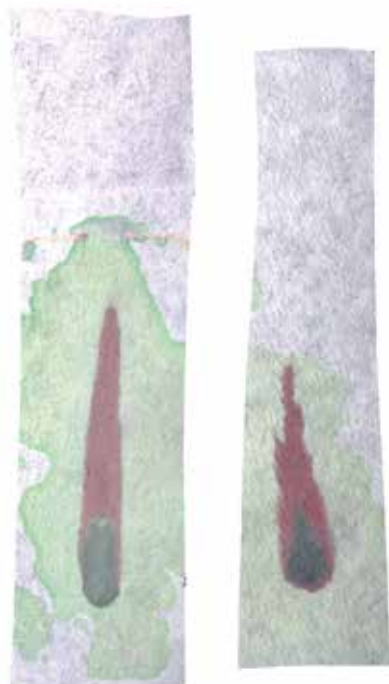
Even though the experiment was a bit complicated, the result is easy to understand. The copper salt and iron salt "climbed up" the paper to different heights and were located accordingly by the detection agent. You've separated the copper salt and iron salt present as a solution mixture from each other on the filter paper.



**Question 9.** Did the two runners in the race climb not only to different heights but at different speeds, too?



The "climbing garden" of the salts



The copper salt (reddish) is several nose-lengths ahead of the iron salt (dark gray).



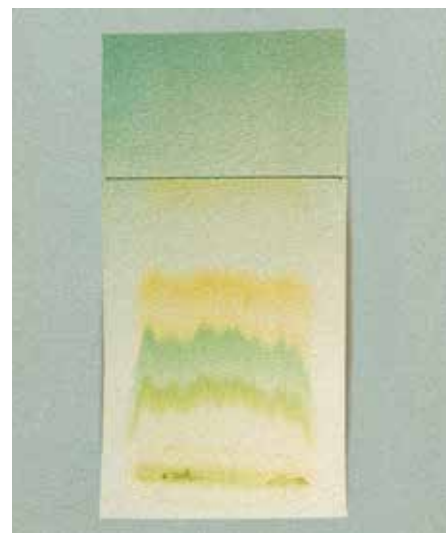
# TECHNOLOGY AND ENVIRONMENT



## Chromatography or writing with colors

The experiment you just performed already has a touch of professionalism to it: you performed ascending chromatography. What? Well, ascending means climbing or rising, and the water climbed up in the strip of filter paper. And chromatography can be translated into common language as “color recording” (Greek *chromos* = color, *graphein* = to write, to record).

The components of mixtures of substances travel different distances on paper or other absorbent materials and are recorded and documented by detection agents as color spots. Such a recording is called a chromatogram. Professionals don't usually use water when performing chromatography but fancy solutions that enable especially good separation.



Even leaf green (or chlorophyll) isn't a uniform dye; it is composed of several components (photo: Dr. W. Botsch, Schwäbisch Gmünd).

## Markers show their true colors

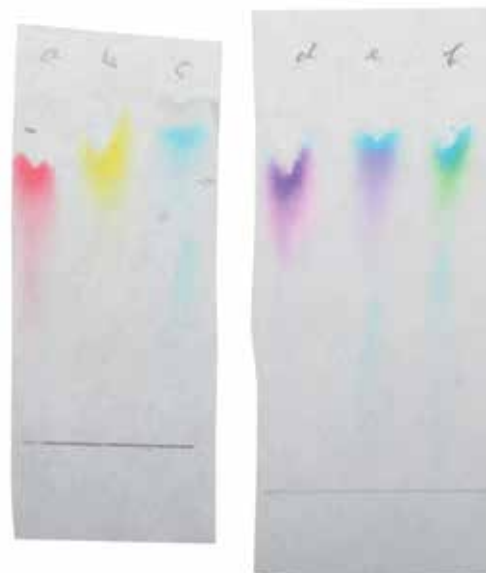
Surely you have a whole array of colorful felt-tipped pens or markers. Markers designed for children contain water-soluble dyes. They are odorless or smell a little like vinegar at most. Markers or pens with the imprint “water-resistant” or “permanent” contain quickly evaporating and intoxicatingly smelling solvents, which is why they also say “Close after using.” You will only need the water-soluble kind for the following experiments.

You will soon see why dyes from felt-tipped markers are well-suited to chromatographic experiments. One advantage is that you don't have to “develop” the chromatogram like you did with the copper-iron separation by pressing the potassium hexacyanoferrate(II) paper on it. Cut out a few 4 – 5 cm wide pieces of filter paper. Here, too, select the length so that the hanging strips dip only a few millimeters into the water.

### EXPERIMENT 27

**Additional material:** Wooden skewer or knitting needle

Using a pencil, draw a starting line on the strip 1.5 cm above the lower edge. Make a light-red, a yellow, and a light-blue dot on the line (each 2 – 3 mm in diameter). These will be the starting points. You can assign letters to them as shown and keep an index on another sheet of paper showing which original color belongs to which letter. Hang the strip in the jar. End the experiment when the leading edge of the water is 1 – 2 cm from the skewer.



Chromatography of single-dye and multiple-dye felt-tipped markers

### EXPERIMENT 28

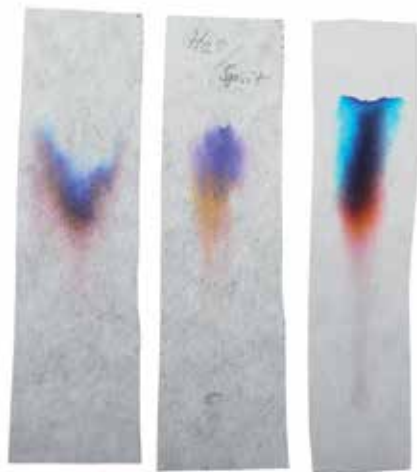
Repeat the experiment with a blue, a dark-blue and a violet marker.



## What's happening here?

Even though the different assortments of felt-tipped pens and markers differ, they do have certain things in common, so your experimental results shouldn't be all that different from the illustrations to the right. The red, yellow, and light-green markers usually contain a single dye, whereas the blue and violet ones contain dye mixtures that are separated through chromatography. Did the red components move along more slowly in your experiments, too?

You won't be particularly surprised about some of the results. You know from working with paints that blue plus red makes violet, and blue plus yellow makes green (test out green markers). But you might be surprised about all that's hidden in a dark colored marker, for example a black one. We'll try it out in Experiment 29.



Three black felt-tipped markers show their true colors.



Dispersion of dye without and with stumbling blocks

## EXPERIMENT 29

If possible, chromatograph three different black felt-tipped markers or pens next to each other. You may soon discover that there is more than one black. Indeed, when it comes down to chromatography, markers are forced to show their true colors.

## Now the markers make their rounds

In the following experiments, you won't learn anything new about felt-tipped marker dyes, but you will use a new method: You will chromatograph horizontally using a colored circle as the starting line.

## EXPERIMENT 30

Cut out several square filter papers with sides measuring approx. 7 cm. Using a felt-tipped marker or pen, draw a circle with a diameter of about 2 cm on one of the papers and place it on an open glass jar or graduated beaker. Now drip water into the center of the circle. The water reaches the colored circle with the second or third drop. How things go from here depends on whether your marker contains only one dye or a dye mixture. Continue to add drops until the leading edge of the water is 1 – 2 cm from the edge of the filter paper and allow the paper to dry.

In the following experiment, you will put some stumbling blocks or, better yet, some potholes in the way.

## EXPERIMENT 31

Fold one of the square pieces of filter paper as shown in the illustration and make the small diamond-shaped holes. Unfold the filter paper and place it under several heavy books so that it is pressed somewhat flat again. Proceed as in the previous experiment. By changing the distance between the midpoint of the paper and the cuts, you can make different images in different experiments.

How the dyes flow depends a lot on the filter paper used as well as on the "mobile solvent." Up to now, you've used water. The next experiment will show how saline solution affects the result.

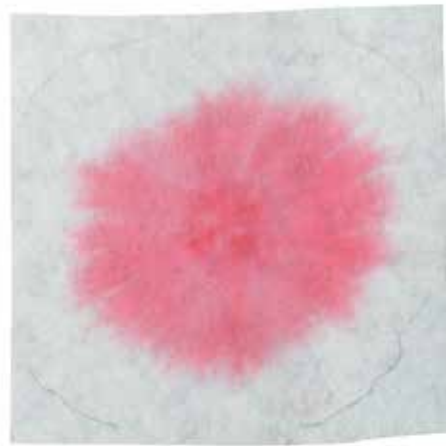
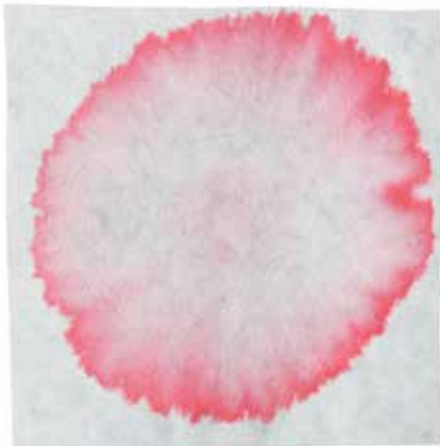
## EXPERIMENT 32

**Additional material:** Table salt

This time, in the middle of two filter papers, do not draw circles but round colored dots with a diameter of about 1 cm. Use a single-dye felt-tipped marker or pen, red for example.

Place the papers onto the graduated beakers and drip water on one red dot and saline solution on the other (dissolve 1 – 2 large spoonfuls of table salt in a test tube with 6 cm of water).

On the paper treated with water, the dye travels outward and the middle around the drip point becomes almost colorless. The saline solution you drip on delays the dispersion of the dye. Unlike in the control experiment with water, the dye remains quite a bit behind the leading edge of the water; there is no brightening of the middle. You will need the saline solution again for the next experiment.



Dispersion of dye with water (left) and with saline solution

## Mysteriously appearing blue ring

And finally, a not-so-simple but very impressive experiment.



For **ammonium iron(III) sulfate** and **potassium hexacyanoferrate(II)**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

### EXPERIMENT 33

Place a square or round filter paper (approx. 7 – 8 cm sides or diameter) onto a sufficiently large screw-top lid turned over. Place 1 drop of ammonium iron(III) sulfate solution (1 spoon tip in 2 cm of water) into the middle of the filter paper. Now allow the spot to grow to 5 – 6 cm diameter by adding saline solution drop by drop. Wait until the saline solution is fully absorbed. Then, in the middle of the spot, place 1 drop of sodium carbonate solution (1 small spoonful in 2 cm of water) and then 1 drop of the potassium hexacyanoferrate(II) solution prepared according to the tip on p. 19. Wait again until the solutions have soaked in. Now add saline solution again drop by drop to the starting point. After the second or third drop, a deep-blue, jagged ring suddenly appears — something like in the illustration.

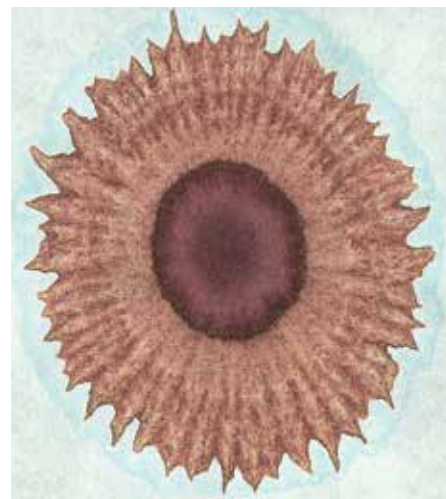
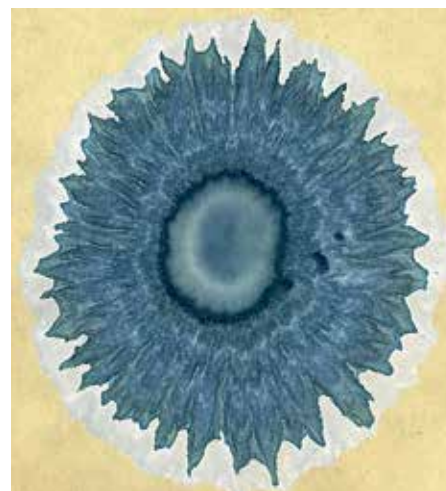


A reaction on filter paper:  
Berlin blue ring

### Side Notes

Pattern images for painters and fabric designers

In the experiments in this chapter, the interplay of colors and shapes has been emphasized, while the physical and chemical processes took a back seat for now. In so doing, you are tracing the footsteps of a man who took a similar approach. The chemist F. F. Runge (1794 – 1867) experimented with dyes and color reactions on filter paper and so became the father of paper chromatography — a technique that only became an integral part of chemical practices much later, during World War II (1939 – 1945). Runge published a book about his "self-grown pictures." He wanted to give drawers, painters, and fabric designers inspiration for their work.



Produced according to the original formula of Friedlieb Ferdinand Runge, the inventor of this painting technique.

## What's happening here?



You already know how Berlin blue is formed; and you also know that Berlin blue is broken down by sodium carbonate solution (Experiment 16). While the majority of the ammonium iron(III) sulfate is transported by the saline solution to the "outskirts" of the paper, the rest remaining in the center would produce a distinct blue coloration. That is prevented by the sodium carbonate. Through the addition of saline solution again, the potassium hexacyanoferrate(II) also travels to the outskirts — even farther than the sodium carbonate, in fact. In the area where there is no longer a noteworthy amount of sodium carbonate, there is nothing left to stop the formation of the Berlin blue.

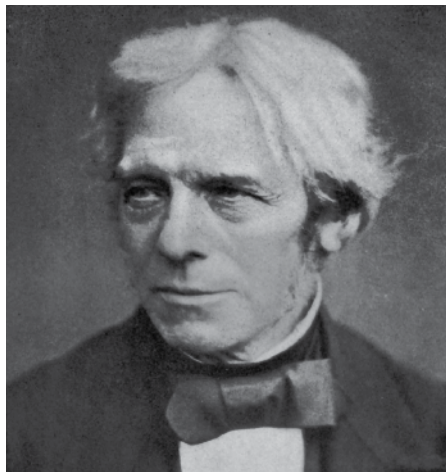


**Question 10.** This experiment requires a little flair and patience. Sometimes parts of the blue ring fade again, especially if more water or saline solution is dripped on. Why do you think that is?



# 5

## Some physics for hobby chemists



Michael Faraday (1791 – 1867) made important discoveries both as a physicist and as a chemist (photo from Milikan & Gale, *A First Course in Physics*, 1906).

Chemistry and physics — what is the difference between the two sciences, anyway? A simple example will make the difference clear: take air, for example.

Chemistry asks about the material composition of air. Is air a simple or a compound substance? What properties does air have and what properties do its components have? For example in terms of flammability, suitability for breathing, impact on climate?

The physicist pursues the question of how a certain quantity of air — five liters, for example — behaves under changing conditions: under changing pressure, for instance, or when the temperature rises or falls.

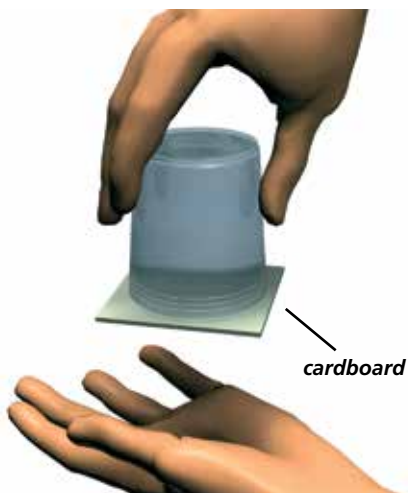
Practically speaking, chemistry and physics can't be easily separated. A chemist needs to know about the effects of air pressure if he wants to produce and collect gases. He is confronted with the surface tension of water if he is developing a means for eliminating this disruptive effect when washing clothes. We will deal with both of these topics in this chapter.

### Air pressure seals jars

#### EXPERIMENT 34

**Additional material:** Piece of smooth cardboard

Fill a test tube — you had best do this over the sink — to the brim with water and cover the opening with a piece of smooth cardboard or strong paper. Press the “lid” against the opening and turn the tube over. Let the lid go! You can also repeat the experiment with one of the graduated beakers.



The air pressure seals the beaker.

### What's happening here?

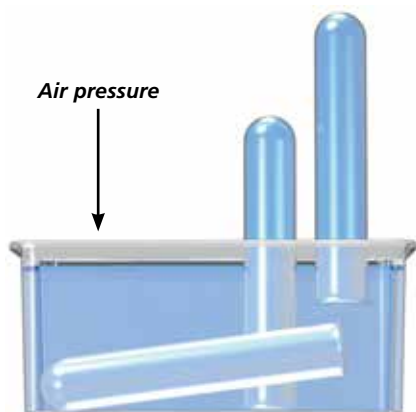


There is no magical force at work here keeping the beaker sealed, but air pressure.

Air pressure presses the lid firmly against the water-filled container so that no water can flow out. In fact, the air pressure is able to hold up a 10 meter- (almost 33 foot-) high column of water. In other words, your test tube or beaker would need to be taller than 10 meters for the experiment not to work.

#### EXPERIMENT 35

Place about 3 cm of water into your plastic basin and lay a test tube down flat in it. If you lift the test tube a little on the open side, it will fill completely with water. All of the air will bubble out. Hold the test tube up as shown in the illustration. No water flows out now, either. You will need these materials in the next experiment as well.



The air pressure ensures that the water stays in the test tube.



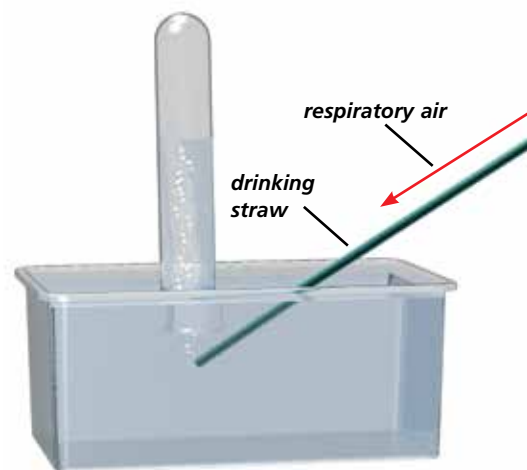
## EXPERIMENT 36

**Additional material:** Drinking straw

Now push the end of the straw under the opening of a test tube that is upright and completely filled with water and carefully blow in the other end of the straw. Air bubbles rise in the test tube which force the water out of the tube. Dispose of used drinking straw in the garbage.

### What's happening here?

In Experiment 35, too, air pressure ensures that the water doesn't flow out of the test tube. In that case, the air pressure doesn't act directly on the opening of the test tube but indirectly via the water in the basin. In Experiment 36, you became acquainted with a method chemists can use to trap a gas — respiratory air in this case — through water. The water serves as a confining liquid; it ensures that only the desired gas is trapped.



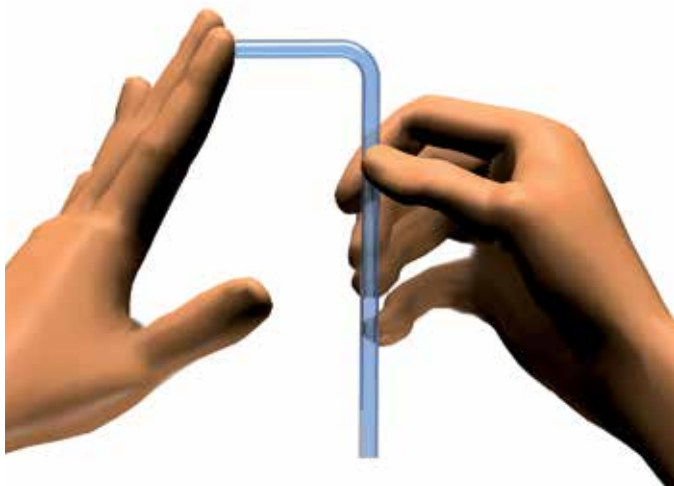
*The air forces the water out of the test tube.*

## Water flows over the mountain

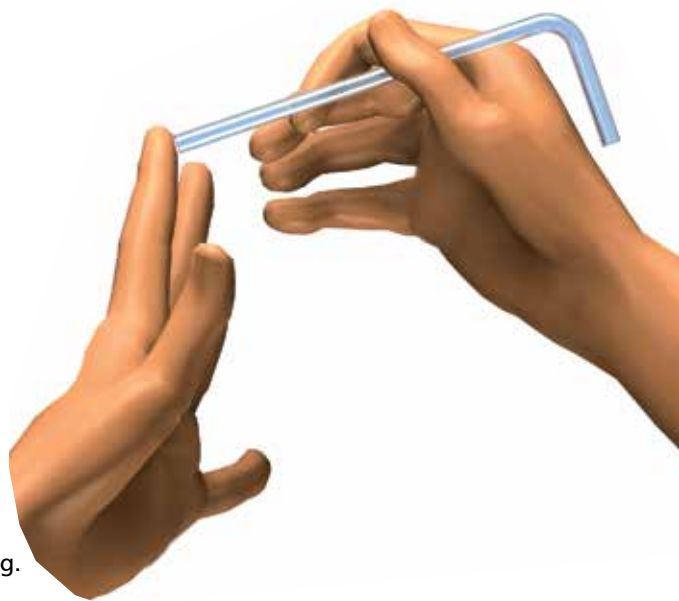
In general, water flows downhill unless it is placed under pressure, for example in a water line. Then it can spray into the air, too, like it does through a garden hose. But water climbing up over the edge of a cup and running away does sound somewhat mysterious. Let's do an experiment, but first some preliminary tests.

## EXPERIMENT 37

Fill your basin with water and place the angled tube into it; it fills up with water. Remove small air bubbles by lifting the tube under water a little on one side and tapping against it with a finger. Hold the long leg of the tube with one hand and seal the opening of the short leg using the index finger of the other hand. Lift the tube out of the water and hold it as shown in the illustration on the left. The water remains in the tube. Aha! It's the air pressure. Remove your finger from the opening, and the water flows out.



Fill the angled tube with water like before, seal one opening with your finger again and hold the tube as shown in the illustration on the right. What will happen if you now release the opening? Will the water flow out of both ends or only out of one? And if only out of one, which one? You can set up this experiment as a series of tests and record the results. You will then realize that the water always flows out of the lower opening. In any case, some water flows up over the hill. And if the opening of the short leg is lower, the water will even crawl up the long leg in order to drop down the short one.



*The index finger on one opening of the tube prevents the water from flowing out regardless of how you hold the tube. The second opening is sealed by the air pressure.*



*With a sure eye and a steady hand, you can make the water “hover” in the angled tube.*



*Water flows over the mountain.*

Here’s another trick. Fill up the tube again under water and hold the closed tube with your index finger as shown in the illustration. Try to put the tube in such a position that both openings are at the same level. If you’re successful, the water will stay in the tube even after releasing the opening. You’ll need to practice this a little before you get it.

## EXPERIMENT 38

Up to now, not much water has flowed over the mountain yet. You just need to make sure that the water in the tube doesn’t just flow out but that a further water supply is there to keep things going. Assemble the illustrated experimental setup, first with an empty tube. Place the upper beaker on the lower level or in the large opening of the experiment station. Then fill the angled tube with water again, seal the opening of the long leg with your finger, and hang the tube over the edge of the filled beaker. Release your finger! The water flows into the lower beaker and flows and flows, but only as long as the short leg is still dipped in water. And even that only works if the outlet opening is lower than the water level in the supply beaker. Lift the long leg slowly before your little “well” dries up. When the outlet opening approaches the level of the water in the upper beaker, it slows to a drip; if you lift the tube further, the water in it flows back into the upper beaker.

## What’s happening here?



This device is known as a siphon and, believe it or not, the force at work here isn’t suction. So how can we explain the fact that water is able to climb over the mountain? The answer has to do with air pressure, naturally:

It has less effect on the lower opening of the tube than on the upper one (or, more precisely, on the surface of the water in the upper beaker). So it’s easier for the water to exit the tube at the lower opening. And why does the air pressure have less effect at the lower end of the angled tube? The reason is that the water column in the tube exerts a stronger counter force there.

The whole thing works quite nicely with a hose, too, of course. You can transfer water from a filled rain barrel into a watering can standing on the ground. You just have to fill the hose with water first.

## Heat expands bodies

When you think of “bodies,” you probably think of human or animal bodies. In physics, we use this term in a much broader sense. For physicists, an essential characteristic of a body is its extension, regardless of whether it is a solid, liquid, or gas. Yes, air is a gaseous body. Experiment 36 already showed that it has an extension, that is, that it fills up space, because it forced the water out of the test tube. And once it’s inside, it doesn’t let the water in.



**Question 11.** How can you demonstrate this through a simple experiment?

Air demands space. And it becomes larger when heated.



Be careful when twisting the angled tube into the stopper. Follow the instructions on p. 12/13.  
In case of injury: **First Aid 7** (inside front cover).

## EXPERIMENT 39

Pour enough water into a test tube so that the lower end of the angled tube dips about 1 cm into it. Instead of water, you can also use red or blue litmus solution.

Make note of how high the water column in the angled tube is. Now hold the upper part of the test tube with your hand. The water rises in the angled tube. In a magic show, you can of course attribute the rising of the water column to mysterious forces that you control at will.

After releasing it from your heating grip, hold the test tube under cold running water. The column of liquid sinks below where it was originally. If the angled tube doesn't dip too deeply, some air bubbles will rise up through the water if cooled off a lot.



Simple air thermometer

## What's happening here?



There are no magical forces at play in this experiment. It's the heat from your hand that makes the water in the glass tube rise. In simple thermometers, colored alcohol (or, in old ones, the liquid metal mercury) expands when heated. But it's only able to expand by rising in a tube. Your experimental apparatus is an air thermometer. The air in the upper part of the test tube expanded and pressed the water in the angled tube upward. You can try out with friends to see who gives off the most heat. But don't squeeze too hard, because the glass test tube might break.



**Question 12.** How do you explain the behavior of your thermometer when it's cooled off?

## The dry coin

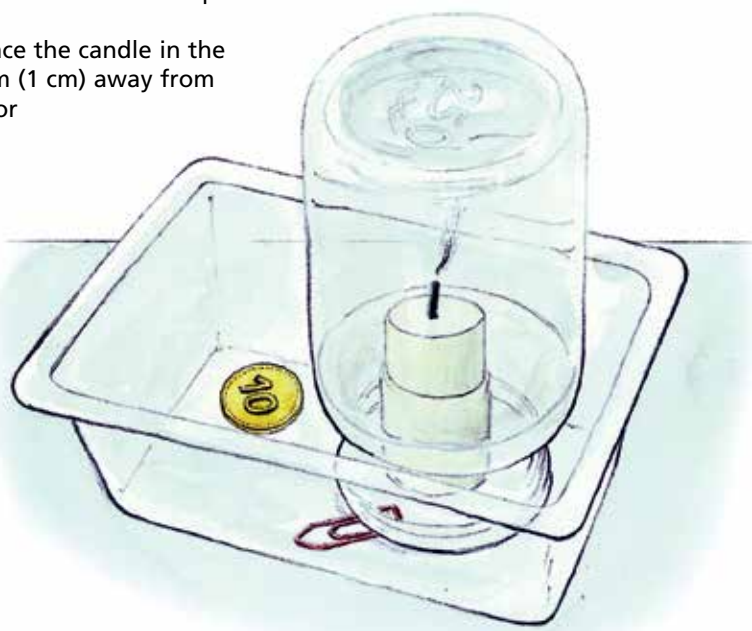
## EXPERIMENT 40

You can do this classic experiment in front of an audience. For this experiment, you will need a **self-standing candle 3 to 4 cm (1.5 in)** in height, a **paper clip**, the **plastic basin**, and **screw-top jar** — and, of course, a **coin**. Place the coin in the basin and pour in just enough water to cover the coin. Then tell your audience that anyone can have the coin if they just figure out how to get it out of the basin without dipping their fingers in the water. It's against the rules to pour the water into another container, lift one side of the basin, or use a tool to scoop the coin out.

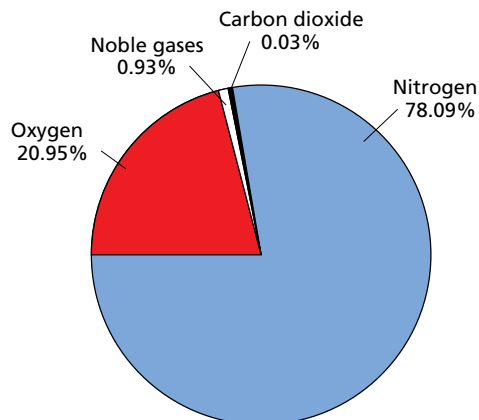
While everyone's scratching their heads, place the candle in the basin and put the paper clip in the water about 2 cm (1 cm) away from the candle. Light the candle and hold the glass jar for 2 – 3 seconds over the flame. Then place the jar over it, taking care to place the jar so that its edge is on the paper clip.

Isn't that cool? The water is sucked into the jar, the candle goes out, and the coin is lying in a dry basin. You're the lucky winner.

The dry coin







## What's happening here?



Both chemistry and physics play a role in this neat little experiment. As you will learn later, air is a mixture of different gases, mostly **oxygen** and **nitrogen**. The graph shows a proportional breakdown of the components in air. Most of the oxygen under the jar is used up by the candle, because fire is the combustion of oxygen. In other words, a portion of the air in the jar "disappears" because of this chemical reaction. The external air pressure becomes greater than the pressure under the jar and the water is pushed from the basin into the jar. In addition, the cooling of the warmed "leftover air" after the candle goes out plays a role in the suction effect as well.

## TECHNOLOGY AND ENVIRONMENT



### When thermal expansion becomes a hazard

Sometimes a hot summer is enough to warp railroad tracks so much that trains are derailed. To prevent that, a small gap used to be left between successive rails, the so-called rail joint. Older folks will still be familiar with the seemingly incessant melody "da-dam-da-dam" that it caused that not only got on travelers' nerves but also damaged trains' wheels. Today, the rails are welded and joined to solidly with the concrete ties that it's no longer possible for the rails to move sideways when exposed to

extreme heat. The expansive forces are diverted into the subsoil.

In steel bridges, there's another way to manage this problem. Only one side of the bridge is solidly connected with the bridge pillars. On the other side, between the pillars and the bridge, there is a roller bearing that can "roll along" without a problem when the roadway lengthens due to heating.

In piping that is exposed to large fluctuations in temperature — for example, pipelines or steam pipes — so-called "expansion bends" are integrated to absorb the expansion of the pipes.



When thermal expansion becomes a hazard (photo: ulrich-rapp.de)





**Question 13.** Why do you think that the cables in high-voltage lines (overland lines) droop more in summer than in winter?

## Abrupt end of a mountain of water

Now you're going to create a mountain of water and then make it collapse. This is a great way to become acquainted with a phenomenon — a fancy way of saying event or observable fact — that is awe-inspiring time and time again.

### EXPERIMENT 41

Place a beaker in the basin and fill it to the very top with water. Now, slowly pour even more water into the beaker from a test tube. You will be astounded how much more water the jar can take without overflowing. Take a look at the beaker from the side. The surface of the water has bulged upward. It looks like the water is covered with skin that holds it together and keeps it from overflowing.

### EXPERIMENT 42

**Additional Material:** Dishwashing detergent

Now, while watching the jar from the side, add a drop of dishwashing liquid to the water's surface. The mountain of water collapses and water flows over the edge of the jar.



*Does water have a skin?*

## What's happening here?



Maybe you've already heard or read that all substances are composed of very tiny particles. We will get into that a bit more in the next chapter. Water and the substances contained in detergents and soaps — the **surfactants** — are no different. There are particularly strong attractive forces between the water particles on the water's surface that does in fact act like a skin holding the water together. The technical term for this is **surface tension**. Surface tension is also responsible for the fact that small quantities of water form into drops.

When surfactants are added to the water, the surfactant particles gather on the water's surface, getting between the water particles and therefore greatly reducing the surface tension. The "skin" tears, and that spells the end of the mountain of water. Soaps and dishwashing liquids reduce surface tension, an ability that is very important for washing. We will deal more with surfactants and their effects in Chapter 9.



*Waterskaters depend on the surface tension of water, too (photo: Diffené, Neustadt/W.).*



The needle floats but it isn't unsinkable.

## Downfall of a sewing needle

### EXPERIMENT 43

Fill your thoroughly washed basin with water and carefully place a sewing needle on the water's surface. Don't give up if it does not work the first time you try. The sewing needle will float on top of the water, even though the steel that it is made from is much heavier than water. If you look closely, you will see that the needle is slightly sunk into the water. It really looks like the water has an elastic skin that prevents the needle from sinking.

### EXPERIMENT 44

**Additional material:** Dishwashing detergent

Add one drop of soap solution to the water. The needle goes down abruptly.

### EXPERIMENT 45

**Additional material:** Razor blade, paper clip, dishwashing detergent

An old-fashioned razor blade (admittedly not so easy to come by) or small utility knife blade can float on water too, and can even carry cargo (e.g. a paper clip). If you add dishwashing soap to the water, you can watch how the water loses its surface tension. The indentation of the water surface will gradually subside, the water will begin to creep over the razor blade in a few spots, and finally it will sink.



The floating razor blade is even able to carry a load.

## What's happening here?



The "skin" of the water is stretched under the weight of the needle or razor blade, but it holds firm. Like with the collapse of the mountain of water, the surfactant particles cause the attractive forces between the water particles to be blocked, causing the skin to tear here as well and making it incapable of holding things up.

The next two experiments also deal with the deployment of surfactant particles on the water's surface. In the transition between the tensed to the relaxed state, forces are released that can be converted into forces of propulsion depending on the setup.

### EXPERIMENT 46

**Additional material:** Matchsticks, dishwashing detergent

Fill the basin with water and launch a small fleet of paper matchsticks in it. If you dab the angled tube with a little dishwashing soap and dip it into the water in the middle of the fleet, the "matchstick ships" will shoot away from it in all directions.

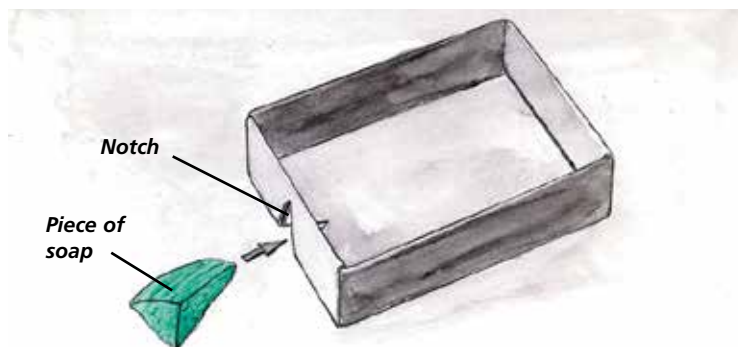
### EXPERIMENT 47

**Additional material:** Matchbox, piece of soap

Make a notch in a matchbox, and then jam a small piece of soap into the notch. Set the matchbox in a large basin of water. It will be propelled forward for a while by its "soap motor."



The little match boats scoot away without delay.



Assembling your matchbox yacht

# 6

## Carbon dioxide — a familiar old friend

You've already been introduced to the gas we'll be exploring more in-depth in this chapter. At the outset of your experiments, you got to know it as the "active ingredient" in mineral water. You know that it is responsible for the sour taste of mineral water because it bonds with water to form carbonic acid. And you have practiced a technique for specifically detecting the gas using lime water.

### Shake the gas out of the mineral water

Now you're going to change the carbon dioxide detection technique a little: You're not going to add the dissolved gas in the form of mineral water to the detection agents like in Experiments 8 and 12, but will feed only the gas through the litmus solution and lime water.



Be careful when twisting the angled tube into the stopper. Follow the instructions on p. 12/13.  
In case of injury: **First Aid 7** (inside front cover).

#### EXPERIMENT 48

**Additional material:** Fresh sparkling mineral water  
Add 3 cm of water and 3 drops of litmus solution to a test tube. Fill a second test tube halfway up with sparkling mineral water that is as fresh as possible. On this tube, place the stopper with the angled tube as shown in the illustration. Dip the free end of the angled tube into the blue litmus solution. If you shake the tube with the mineral water slightly, gas will bubble up through the litmus solution, which turns red.

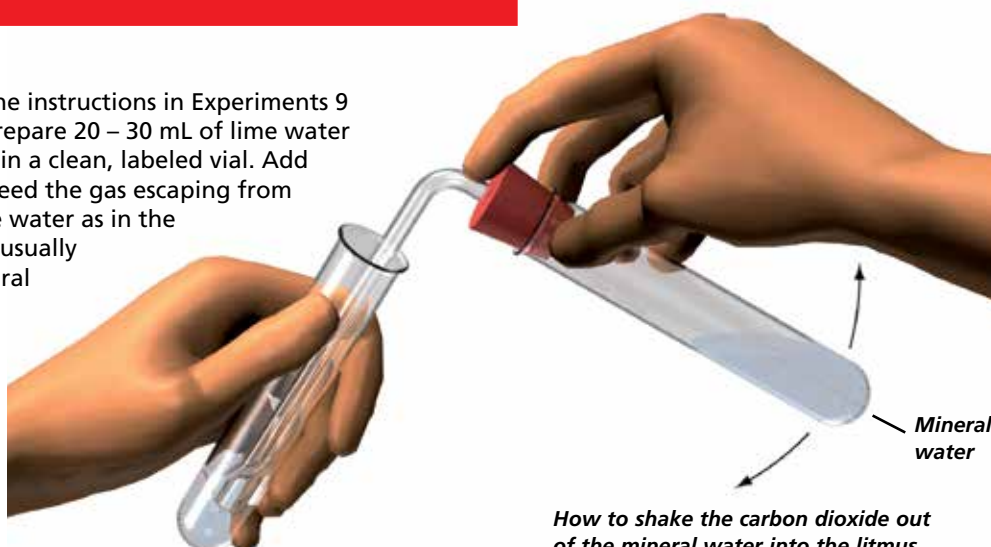


For **calcium hydroxide** and **lime water**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

#### EXPERIMENT 49

Following the instructions in Experiments 9 and 10 to prepare 20 – 30 mL of lime water and store it in a clean, labeled vial. Add

3 cm of lime water to a test tube and feed the gas escaping from the shaken mineral water into the lime water as in the previous experiment. The cloudiness is usually not as pronounced as in a shot of mineral water but is clearly visible.



*How to shake the carbon dioxide out of the mineral water into the litmus solution or lime water (index finger on the stopper!).*



*In sweet lemonade, too, carbon dioxide provides the refreshing tingling sensation.*





Acidic spring at Kronthal im Taunus, Germany (photo: Taunustipp, [www.taunustipp.de](http://www.taunustipp.de))

## Side Notes

### Where carbon dioxide trickles out of the rock

As you've heard, carbon dioxide is usually added to mineral water. But there are mineral water springs that naturally contain carbon dioxide and thus carbonic acid in such a quantity that the gas pressure makes the water gush from the source. If natural mineral water contains more than 250 mg of natural carbon dioxide per liter, then the spring is referred to as an acidic spring. Acidic springs have been known since the Middle Ages. They were held in very high regard due to their favorable influence on health and well-being and became the focus of many popular spas.



The household chemicals soda ash and baking soda

## What's happening here?



### About atoms and molecules

The formula  $\text{CO}_2$  has become a household name known even by people who have nothing to do with chemistry. The main reason for this is the hot topic of climate change, which we'll get back to. But you'd probably like to know what the letters and the number two mean. To explain, let's go back a bit and talk about atoms and molecules.

You've already heard about elements and compounds and that all substances are made of very small particles. The smallest particles of the elements are called **atoms**. One of these "smallest of particles" can't be broken down any further, which is why the ancient Greeks called them *atomos* (= indivisible).

The smallest particles of the element carbon are carbon atoms, and the smallest particles of the element oxygen are oxygen atoms. For chemists, the symbols listed on p. 20 are not just abbreviations of the element names, but each also represents one atom of the respective element:



When the elements carbon and oxygen combine to form carbon dioxide, one carbon atom (symbol C) and two oxygen atoms (symbol O) always join to form one carbon dioxide particle (formula  $\text{COO}$  or  $\text{CO}_2$ ):



The smallest particles of chemical compounds are called **molecules** (Latin *molecula* = small mass). In this case, the molecule consists of three atoms, but giant molecules with thousands of atoms even exist. The formula  $\text{CO}_2$  says briefly and precisely what otherwise takes a lot of words to say.



**Question 14.** The formula for water is  $\text{H}_2\text{O}$ . What does this formula say about the composition of the water molecule?

## Soda ash and baking soda — two household chemicals

Soda ash — known as washing soda — is still used by homemakers even in the era of specialized and heavy-duty detergents. Baking soda is helpful in the kitchen, for example to neutralize excess acid when bottling acidic fruit or to delay the curdling of milk.

Both compounds naturally have precise scientific names: soda ash is sodium carbonate — hey, we have that in our kit — and baking soda is **sodium hydrogen carbonate** or **sodium bicarbonate**. (The same compound is also commonly known as bicarbonate of soda.) You can see immediately that both compounds contain the element sodium, as well as the element carbon, which has hidden itself away in "carbonate."



**Question 15.** How does sodium carbonate differ from sodium hydrogen carbonate? Consult the table on p. 20.

Get yourself some baking soda at the supermarket. You can often use baking soda instead of sodium carbonate, which should be used sparingly.

## Carbon dioxide gets expelled

Up to now, you've only gotten to know carbon dioxide as a component of mineral water. You you'll expel it out of soda ash and baking soda, too.



For **sodium carbonate**, **citric acid** and **lime water**, note the "Hazardous substances and mixtures" information on pp. 7 – 8. Be careful when twisting the angled tube into the stopper. Follow the instructions on p. 12/13. In case of injury: **First Aid 7** (inside front cover).

### EXPERIMENT 50

Prepare a test tube with 3 cm of lime water and arrange the stopper with the angled tube. Now add 2 spoonfuls of sodium carbonate and 2 spoonfuls of citric acid to a second test tube. Add 2 – 3 cm of water to that and seal the tube as quickly as possible with the prepared stopper. Assemble the illustrated experimental apparatus. There will be hissing in one test tube and emphatic bubbling in the other. No doubt: The cloudiness is indicative of carbon dioxide. That can only be coming from the sodium carbonate.



*Citric acid expels the carbon dioxide from sodium carbonate.*

### EXPERIMENT 51

This time, ready a test tube with diluted blue litmus solution (3 drops from the vial in 3 cm of water). Proceed as in the previous experiment and feed the escaping gas into the litmus solution. It'll turn red in no time.

### EXPERIMENT 52

Repeat Experiment 50 with baking soda instead of sodium carbonate.

### EXPERIMENT 53

Repeat Experiment 51 with baking soda instead of sodium carbonate.

So citric acid also releases carbon dioxide from baking soda.



*The acid contained in wine releases carbon dioxide from baking soda.*

## What's happening here?



While carbon dioxide escapes quickly from mineral water, the gas in the two salts sodium carbonate and sodium hydrogen carbonate is well preserved — well stored, so to speak. But only as long as these salts don't come into contact with acids. Acids expel the carbon dioxide in an effervescent frenzy. Consequently, you can use sodium carbonate or sodium hydrogen carbonate to detect acid in places litmus solution won't work very well, for example with red wine. If you add 1 spoonful of baking soda to 3 cm of wine (have an adult pour it for you), you will observe a more or less lively development of gas depending on how much acid the wine contains. It goes without saying that you'll discard the wine.



## All kinds of effervescent powders

For the following experiments and some later ones, you will need effervescent powder. Get yourself some pouches of effervescent, or fizzing, drink mixes. You can also use a pulverized effervescent tablet (a vitamin tablet, for example). You will also need a packet of baking powder.



For **calcium hydroxide**, **lime water** and **citric acid**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

### EXPERIMENT 54

Ready 3 cm of lime water again and set up the stopper with the angled tube (heed the warnings on p. 12/13!). Into a second test tube, place 1 – 2 cm of effervescent powder and 3 cm of water. Put the stopper on quickly and feed the escaping gas into the lime water. Aha:  $\text{CO}_2$ .

### EXPERIMENT 55

Fill a graduated beaker up to the 50 mL mark with water and add 5 drops of litmus solution to it. Now sprinkle some effervescent powder into the blue solution. The red coloration is indicative of acid.

### EXPERIMENT 56

Repeat Experiment 54 with baking powder instead of the effervescent powder: add 4 cm of water to 2 cm of baking powder. Feed the escaping gas into lime water. Here, too, the familiar cloudiness indicates carbon dioxide.



## What's happening here?



**Effervescent powder** consists mostly of sodium hydrogen carbonate (baking soda) and citric acid or tartaric acid. Other ingredients: sugar, flavorings and some food coloring (red for raspberry, yellow for lemon). In a sealed, dry state, the mixture keeps for a pretty long time. When water is added, the baking soda and the acid react and release carbon dioxide (Experiment 54). You detected the acid (which is usually present in abundance to give it a sour taste) in Experiment 55.

**Baking powders** have a very similar composition. Here the idea is for the carbon dioxide released through the combination of acid and baking soda to lighten the dough for cakes and pastries. Since air always contains some moisture, the manufacturer adds flour or other substances to the baking powder to absorb moisture. After all, the carbon dioxide isn't supposed to be released right in the bag.



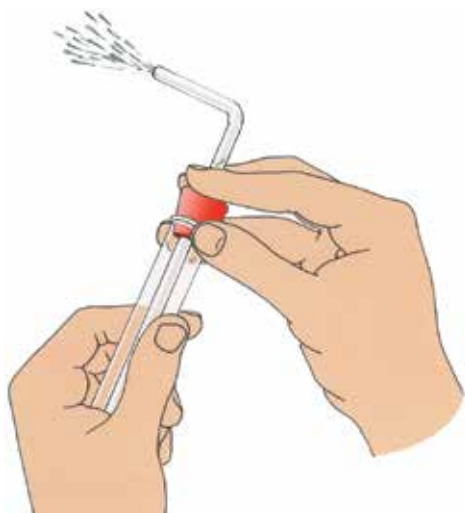
**Question 16.** If you use too much or exceedingly fresh effervescent powder in Experiment 54, carbon dioxide detection may not be possible given the amount of foam formed. Why?



Be careful when twisting the angled tube into the stopper. Follow the instructions on p. 12/13.  
In case of injury: **First Aid 7** (inside front cover).

### EXPERIMENT 57

Set up the angled tube and stopper as shown in the illustration. Add 3 spoonfuls of citric acid and 4 spoonfuls of baking soda to a dry test tube. The next maneuvers have





to be performed with lightening-quick speed. Fill the test tube to the top with water — this is best done over the sink — and place the stopper with the glass tube onto it. When doing this, hold the test tube with one hand and the stopper with the other. The liquid fizzes out of the test tube. The pressure is so strong that the stopper would fly out if you weren't holding it firmly. This would not only ruin your experiment but your angled tube, too. Take care not to spray the liquid into anyone's face; spray it in the sink and rinse the sink out thoroughly after the experiment.

## Tendency to settle

Although the gases are all airy types, there are considerable differences between them. The lightest gas, hydrogen, takes off in a twinkling of an eye. The poisonous, green gas chlorine hangs around a long time in photographic processing basins if it is not driven out. Have a look at the table with the weights by liter of the gases on page 42. One liter of chlorine weighs only 3.21 g, but chlorine is 35 times heavier than hydrogen. Carbon dioxide weighs in at almost 2 g. It's 1.5 times as heavy as air and shows a tendency to settle, as we'll soon see.

### EXPERIMENT 58

#### Additional material: Matchsticks

Insert the angled tube into the stopper with a hole (follow the instructions on p. 12/13!). To a dry test tube, add 2 spoonfuls of baking soda, 2 spoonfuls of citric acid, and 3 cm of water. Close the test tube as quickly as possible with the prepared stopper and feed the escaping carbon dioxide — as shown in the illustration — into a graduated beaker.

When the escaping gas subsides, lower a burning matchstick into the beaker. It goes out! Do the match test two or three more times at intervals of 6 – 7 seconds (counting them out). The match's flame should be extinguished in the third or even fourth test. The carbon dioxide apparently isn't in any big hurry to leave. The experiment also works if you lift the test tube with the angled tube out of the beaker before the matchstick test and place it in the experiment station.



*The matchstick is extinguished in the beaker filled with carbon dioxide. Usually you can repeat the test 2 or 3 times at an interval of 6 to 7 seconds.*

*The outside and the inside of the portable fire extinguisher by Minimax. When the striker knob was hit, it destroyed the acid tube. The acid released carbon dioxide from 6 liters of sodium hydrogen carbonate solution which drove the extinguishing liquid through the riser and out of the nozzle. (Photo: Minimax Mobile Services GmbH & Co. KG, Bad Urach, [www.minimax-mobile.com](http://www.minimax-mobile.com))*

## Side Notes

### The cornet bag and its successors

Your "wet fire extinguisher" is based on the same principle that led to the invention of the first portable fire extinguisher at the beginning of the 20th century. Due to its shape, it was also called "cornet bag." You can see how it works in the illustration below. Today, devices using chemical means to create pressure are considered outdated. There are still wet extinguishers, however: They contain a cartridge with liquefied carbon dioxide (similar to the cartridge in the mineral water maker, see p. 17). When the lever is pressed, the valve opens and the "ready-to-use" carbon dioxide that is released drives the extinguishing liquid out of the device. Besides wet extinguishers, most of the portable fire extinguishers used today don't extinguish with water or aqueous solutions but with carbon dioxide itself. They take advantage of the fact that carbon dioxide is "heavier" than air (see Experiments 58 – 60) and therefore have a suffocating effect on fires by cutting off their supply of oxygen.





For calcium hydroxide, lime water, and citric acid, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

## EXPERIMENT 59



Litmus solution    Lime Water

Carbon dioxide can be poured into the beaker like a liquid.

To one graduated beaker, add blue litmus solution (3 drops in 3 cm of water), and add lime water to the other, each 1 cm. Ready a small screw-top jar. Now proceed like in the previous experiment, adding 3 spoonfuls of baking soda, 3 spoonfuls of citric acid, and a bit more water to the test tube. This should be done very quickly again. Feed the carbon dioxide into the screw-top jar. When the escaping gas subsides, place the test tube back in the experiment station.

Take the screw-top jar and pour the carbon dioxide into the beaker as if it were an invisible liquid — half in one beaker and the "rest" in the other. Distributing the invisible gas "fairly" between the two beakers isn't so easy. Seal the beakers with the lids and shake. The familiar reactions prove that the carbon dioxide found its way into the beakers.

## Extinguishing with CO<sub>2</sub> — the right way

In Experiment 58, you lowered the burning match into the gas-filled graduated beaker. The flame had to go out in that case. The following experiment shows how you can extinguish a candle that is burning in the open and how you have to set it up so that the experiment fails.

## EXPERIMENT 60

**Additional material:** Tea candle

Place the tealight candle onto an old plate and light it. Create some carbon dioxide in a test tube by mixing 1 spoonful of baking soda, 1 spoonful of citric acid, and 2 cm of water. When the column of foam sinks back, bring the test tube closer to the flame from the side as in the illustration. Naturally, you have to keep the solution from getting on the tealight candle.



Weights per liter of some gases that you've already become acquainted with in your experiments.

Carbon dioxide extinguishes the tealight candle flame.



**Question 17.** Repeat the experiment but pour the gas over the flame. It will usually continue to burn. Can you think of why?

## What's happening here?



Experiments 58 – 60 confirm it: carbon dioxide is “heavier” than air. Take another look at the table with the weights per liter on p. 42. Carbon dioxide therefore sinks to the bottom of containers, displacing the lighter air and thus displacing the air component oxygen, which is responsible for maintaining fires. If the oxygen is used up (Experiment 40) or displaced by carbon dioxide (Experiment 60), the candle goes out. This latter fact explains why carbon dioxide is suitable as an extinguishing agent. In Experiment 60, you spread out a little “carpet of  $\text{CO}_2$ ” on the surface of the tealight candle which robbed the flame of the oxygen it needed to keep burning.

## Burning — from a chemist's perspective

When a candle burns, it becomes smaller. Where does the burnt candle mass go, anyway? It doesn't disintegrate into nothing. It breaks down into combustion products, which you will study more closely now.



When the candle flame is cooled off by the knife blade, carbon dioxide is precipitated in the form of soot.



For Experiments 61 – 64, use a **tealight candle on a fire-resistant surface**, for example on a plate.

### EXPERIMENT 61

**Additional material:** Knife

Place the tealight candle on an old plate, light it, and hold the bare knife blade a few millimeters over the wick in the flame. Black soot forms.

### EXPERIMENT 62

Hold a screw-top jar over the tealight candle flame as shown in the illustration. Put the lid on the jar while it's still upside down and stand it right side up. Open the jar, add enough lime water so that the bottom is covered, close the jar again, and shake it. The clearly visible cloudiness indicates carbon dioxide.



Combustion gases of the tealight candle are trapped.



**Question 18.** How did the “heavy” carbon dioxide get into the screw-top jar?

### EXPERIMENT 63

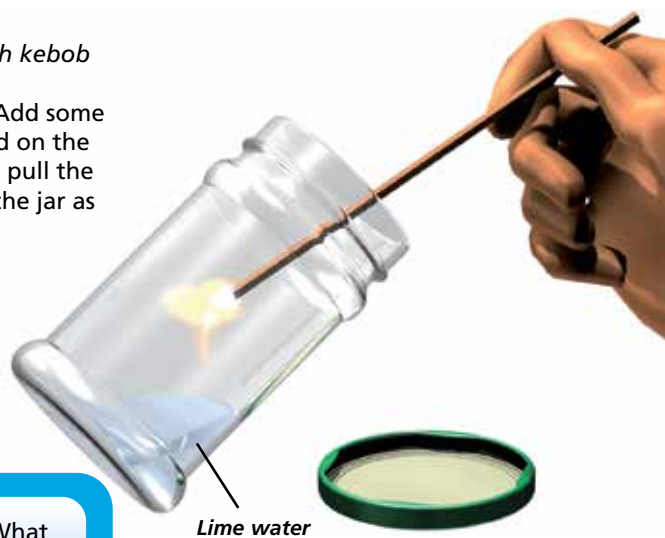
**Additional material:** Small sticks of wood or shish kebob skewers

For this experiment, you will need an assistant. Add some lime water again to the rinsed-out screw-top jar. Light a small stick of wood on the tealight candle and hold the burning stick for 5 – 6 seconds in the jar. Then pull the stick out, carefully give it to your assistant so they can put it out and close the jar as quickly as possible. Shake! Now observe the familiar cloudiness.

### EXPERIMENT 64

**Additional material:** Stick lighter

Repeat Experiment 62 by holding the flame of the lighter under the jar instead of the candle. Check the combustion gases for carbon dioxide using lime water.



Carbon dioxide is produced when wood burns, too.



**Question 19.** Let's assume that you've run out of lime water. What could you use to check for carbon dioxide in Experiments 62 – 64?



You may have noticed that, in this experiment, the wall of the jar “steams up” (like a mirror in the bathroom after a shower). If not, try it again.



## What's happening here?

The black coating on the knife in Experiment 61 is made of **carbon**. It is produced during incomplete burning. The knife dissipates heat so that the required combustion temperature is not reached.

The fuels you used — candle mass (paraffin), wood, liquefied gas — contain carbon. When burned, carbon bonds with the oxygen from the air to form carbon dioxide. The water that is produced indicates that the fuels also contain the element **hydrogen** in addition to carbon. When hydrogen and oxygen are united, water is formed.

Petroleum and petroleum products — which include the paraffin in the candle and the lighter gas — are **hydrocarbons**, that is, compounds that only contain hydrogen and carbon. The main products of petroleum processing are gasoline, diesel, and heating oil. Their increased consumption is responsible for more and more CO<sub>2</sub> being produced.



*In the realm of the hydrocarbons: At this refinery, crude oil (petroleum) is processed into gasoline, Diesel and heating oil (photo: Lurgi, Frankfurt am Main).*

## Living CO<sub>2</sub> factories



For **calcium hydroxide** and **lime water**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

### EXPERIMENT 65

Add some lime water to a graduated beaker so that the bottom is just covered. After only a short time, the appearance of the liquid's surface takes on a matte appearance. A thin “skin” forms. If you rock the beaker gently back and forth, the skin will break at several places and you will be able to see “channels” between the parts of the skin. The skin is composed of calcium carbonate. You have detected the carbon dioxide present in the air even though it makes only 0.03% of the air (see graphic on p. 34).

### EXPERIMENT 66

Pour some lime water into a sealed screw-top jar. You won't see any noticeable clouding. In this experiment, the low quantity of calcium carbonate is too finely distributed to see with the naked eye.

### EXPERIMENT 67

Like in Experiment 35, fill a small screw-top jar completely with water in a deep basin or in the sink. Now push a drinking straw under the mouth of the jar like in Experiment 36 and blow air into the jar until the air has completely displaced the water from the jar.



*When lime water is exposed to air for an extended period, a layer of calcium carbonate forms on top.*



**Question 20.** How can you tell that the jar is full?

Close the jar under water with the lid and remove it from the water. Open the jar, add at least as much lime water to it as there is still water inside and close it again immediately. If you now shake it vigorously, clouding will become clearly visible.

Unlike the “fresh air” that you tested in Experiment 66, exhaled air contains approximately 4% carbon dioxide, which is 130 times more. The foods that people and animals eat also contain carbon and hydrogen, among other things, and yield the “combustion products” carbon dioxide and water after conversion in the body — a process called metabolism. Dispose of the used drinking straw in the garbage.



The exhaled air is collected in the screw-top jar.

## TECHNOLOGY AND ENVIRONMENT



### A disrupted cycle

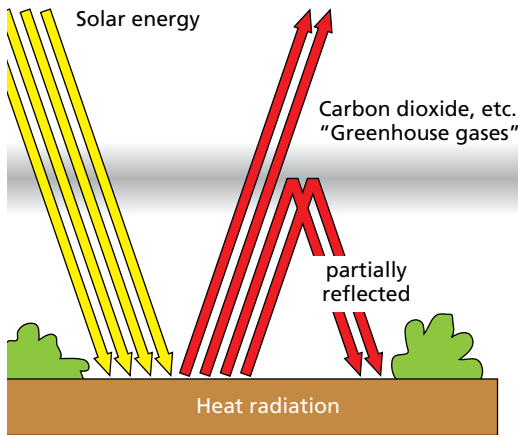
Carbon dioxide and some other gases (methane and water vapor, for example) ensure that Earth doesn't become bitterly cold but remains instead within tolerable temperature ranges for the most part. These gases allow the sun's rays to pass through the atmosphere to Earth's surface, but then absorb some of the heat radiation reflected back to it. (see graphic on the right). They act in a way that is similar to the glass roof of a greenhouse and are therefore called “greenhouse gases.”

The greenhouse gases only become a problem when there is excess production of energy from coal and petroleum, since too much carbon dioxide is formed as a result of these processes. Surely you've heard that the resulting temperature

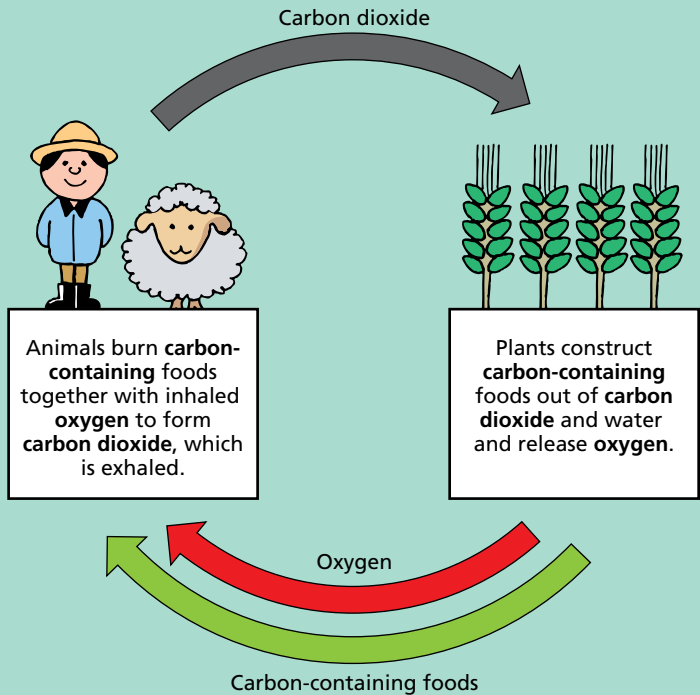
increase can lead to the melting of polar ice and thus to widespread flooding. When you hear talk about the **greenhouse effect**, it is this global (worldwide) heating-up of Earth's atmosphere that is being described.

But the carbon dioxide exhaled by people and animals is not responsible for the increase in temperature. There is a balance in nature between production and consumption of carbon dioxide and oxygen (see graphic below). Through the over-production of carbon dioxide, the cycle loses its balance, especially if the carbon dioxide-consuming rain forests are overexploited at the same time. Carbon dioxide is not a pollutant; it is a basic substance of life.

The crucial thing is how people treat the aspects of nature with which they are entrusted.



Carbon dioxide, water vapor, and some other gases absorb a portion of the heat radiation reflected by Earth's surface.



People and animals give off what plants need; plants give off what people and animals need: an ideal interplay that humans should not throw out of balance.

# 7

## Iron and copper — two heavy metals



*Copper and iron as elements and in the compounds copper sulfate and iron chloride*

In this chapter, experiments with the elements copper and iron and some of their compounds are on the agenda. You have already become acquainted with ammonium iron(III) sulfate and copper sulfate, and you may have asked yourself: What do these compounds, these salts, have in common with the bare metals copper and iron? Maybe the copper has only hidden itself away in the copper sulfate?

### Bring the copper out of the solution

It's best to begin the following experiment in the evening and let the test tube stand overnight out of the reach of pets and young children.



For **copper sulfate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8 and the instructions for disposal on p. 75.

#### EXPERIMENT

### 68

**Additional material:** Bare iron nail

In a test tube, dissolve 1 small spoonful of copper sulfate in 2 cm of water and place the nail in the solution. Did anything happen? As mentioned before, have some patience. Good night!



*The iron nail collects the red copper out of the blue solution.*

### What's happening here?



Good morning! Did you sleep well? You've probably hurried to your work area even before breakfast to find out how the nail in the blue solution is coming along. A dark coating has settled on the nail, and the solution is no longer light blue but greenish. Save it for Experiment 73.

If you rock the jar gently back and forth, brownish glitter will detach from the coating and fall quickly to the bottom of the jar. This is the copper. You can filter it off and dry it and then it will appear reddish brown. You have in fact fetched it out of the copper sulfate solution.

Yes, compounds often look quite different and have very different properties than the elements from which they are constructed — a peculiarity that you will encounter time and time again in your chemical forays.

### Potassium hexacyanoferrate(II) — versatile detection agent

Did you prepare a small supply of potassium hexacyanoferrate(II) solution as recommended in the tip on p. 19? If you don't have any left, in a test tube dissolve 1 spoonful of the substance in 3 cm of water.





Aluminum and gold — not the lightest and not the heaviest metal: Aluminum scrap, gold bars and nuggets.



For citric acid and potassium hexacyanoferrate(II), note the “Hazardous substances and mixtures” information on pp. 7 – 8.

## EXPERIMENT 69

**Additional material:** Iron nail

In a test tube, add 10 drops of potassium hexacyanoferrate(II) solution to 2 cm of water and place the nail in it. Even after a long wait, you won’t notice any changes. You will need the test tube with its contents for the next experiment.

## EXPERIMENT 70

Add 1 spoonful of citric acid to the test tube with the nail. Slowly, the solution around the nail turns blue. Naturally, you already suspect you know what’s going on.

## EXPERIMENT 71

Dissolve 1 small spoonful of copper sulfate in 3 cm of water and add 2 – 3 drops of potassium hexacyanoferrate(II) to it. A reddish-brown precipitate forms (you already performed the reaction on test paper in Experiment 24). You can identify copper compounds in this way. But how do things stand with the copper wire?

## EXPERIMENT 72

Prepare a potassium hexacyanoferrate(II) solution like in Experiment 69 and place the copper wire in it. Nothing stirs. Add 1 spoonful of citric acid to it. The reddish-brown precipitate can’t be conjured up; not even a reddish-brown coloration. The copper wire looks a bit more “matte” at best. Allow the test tube to stand overnight out of the reach of pets and young children and check in the morning to see if there are any changes.



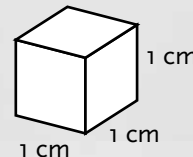
## What’s happening here?

The experiments show that iron and copper don’t react to potassium hexacyanoferrate(II) in their elemental form. The metals have to be present in soluble form as compounds, for example as copper sulfate or as iron citrate. Iron citrate was formed in Experiment 70 when citric acid acted on the iron. In Experiment 72, even citric acid fails to get the job done. This is because copper is a nobler metal than iron. Nothing can be accomplished here with diluted, weak acids unless the acid is given a lot of time. If you look at the test tube from Experiment 72 on the next day, a brown cloudiness will have formed. The oxygen in the air has helped in this case.

## Side Notes

### About light and heavy metals

In this chapter, you are dealing with two heavy metals: iron and copper. The table on p. 20 differentiates between light and heavy metals. When does a metal belong to one group, and when does it belong to the other?



A limit of 4.5 grams per cubic centimeter has been set. Do you remember? A cubic centimeter is a cube whose sides each measure 1 cm (see p. 11). Potassium and sodium are among the lightest metals and weigh in at 0.82 g and 0.97 g per cubic centimeter. These two metals are so light that they float on water (1 g per cubic centimeter). Other light metals are calcium and aluminum, which weigh 1.55 g and 2.70 g per cubic centimeter. While potassium, sodium, and calcium react violently with water, aluminum is not only weather-resistant but also offers great technical advantages when “lightness” is essential — in the construction of airplanes, for example.

A cubic centimeter of the heavy metals iron and copper dealt with in this chapter weighs 7.87 g and 8.92 g, respectively. The noble metals silver, gold and platinum are among the heaviest chunks of matter around, weighing in at 10.5 g, 19.32 g, and 21.45 g per cubic centimeter, as is the only metal that is a liquid at normal atmospheric pressure: mercury (13.54 g).



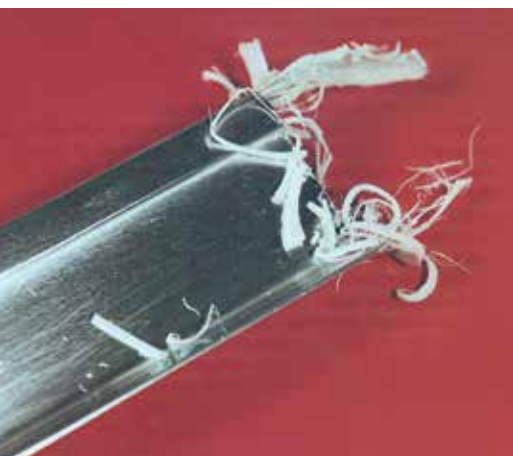
A green protective layer forms on copper domes as a result of the gases contained in the air (carbon dioxide, sulfur dioxide) that protects the metal from being destroyed further: the patina.

Though iron isn't a noble metal, technology wouldn't get very far without iron and steel.



## About noble and ignoble metals

How noble or ignoble a metal is can be seen not only using diluted acids but in how the metals interact with each other, too.



The ignoble metal aluminum is weather-resistant because it gets covered with an oxide layer that denies the oxygen in the air further access. A few drops of mercury chloride solution are enough to destroy the protective layer. The aluminum is oxidized and the oxide grows out of the metal.



For **sodium carbonate**, **ammonium iron(III) sulfate** and **potassium hexacyanoferrate(II)**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

### EXPERIMENT 73

Add 2 drops of the solution set aside from Experiment 68 to 3 cm of water and add 1 drop of potassium hexacyanoferrate(II) to it. It turns blue! Add 1 small spoonful of sodium carbonate to it and shake the tube sealed with the stopper. The blue coloration disappears (see Experiment 16). No doubt about it: it's Berlin blue.



**Question 21.** What conclusion do you draw based on this result?

Now let's turn the tables. Can we dissolve copper using the same method?

### EXPERIMENT 74

**Additional material:** Small iron nail, steel wool

Prepare a copper sulfate solution as in Experiment 68. Put an iron nail or some steel wool into the solution. Let the test tube sit for several hours, keeping it away from young children. Pour the resulting faintly green solution into another test tube and add a piece of copper wire to it. Does the solution turn blue?



**Question 22.** What conclusion would one draw from blue coloration?



## What's happening here?



The blue coloration in Experiment 73 proves that iron from Experiment 68 "has gone into solution" (the chemistry expression is "dissolved"). Copper sulfate solution and iron yielded copper and iron sulfate solution.

As Experiment 74 shows, a reversal of this reaction isn't possible. The self-made iron sulfate solution and the copper yield neither iron nor copper sulfate.

In the table to the right, you will find a series of metals listed, most of which will be familiar to you. The higher in the list a metal is, the less noble it is, which is to say the easier it goes into solution. The noble character increases from top to bottom. The important commercial metals aluminum, zinc, and iron belong to the ignoble metals. Notice that the noble metals silver, gold, and platinum are at the bottom of the list; they go into solution only very unwillingly but, for the same reason, are all the easier to precipitate out of their solutions.

**A rule of thumb:** A metal listed in the series can precipitate the metals below it in the series from their solutions. For example, iron can precipitate copper out of copper sulfate solution, since copper is below iron.

Desire to go into solution decreases



Potassium  
Calcium  
Sodium  
Magnesium  
Aluminium  
Manganese  
Zinc  
Iron  
Tin  
Lead  
Copper  
Silver  
Gold  
Platinum

**Topsy-turvy world?** The ignoble metals are at the top, while the noble metals are at the bottom.



**Question 23.** Can noble gold precipitate ignoble zinc out of zinc citrate solution? (Zinc citrate is formed when citric acid acts on zinc.)

## Battery current dissolves metals

As the table shows, red copper comes just before the noble metals, and in fact it's referred to as a semi-noble metal. It proved to be rather resistant to the citric acid. Professionals use caustic nitric acid to dissolve copper. You can do it using a 9-volt battery. You will also need paper towel, which you will soak with saline solution (2 spoonfuls of table salt in 25 mL of water). Other additional materials are indicated in the experiments.



Before performing the experiments, take another look at the "Safety for experiments with batteries" (p. 7).



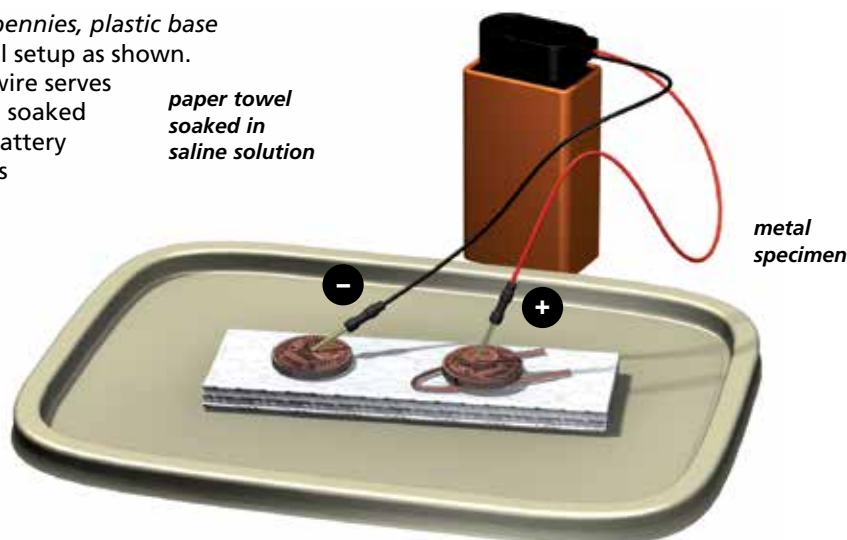
**Nitric acid, also known as aqua fortis** (Latin for "strong water"), dissolves copper and silver but not gold. It therefore used to be used to determine whether gold was "diluted" with less precious material.

## EXPERIMENT 75

**Additional material:** Two pennies, plastic base  
Assemble the experimental setup as shown.  
The coin over the copper wire serves

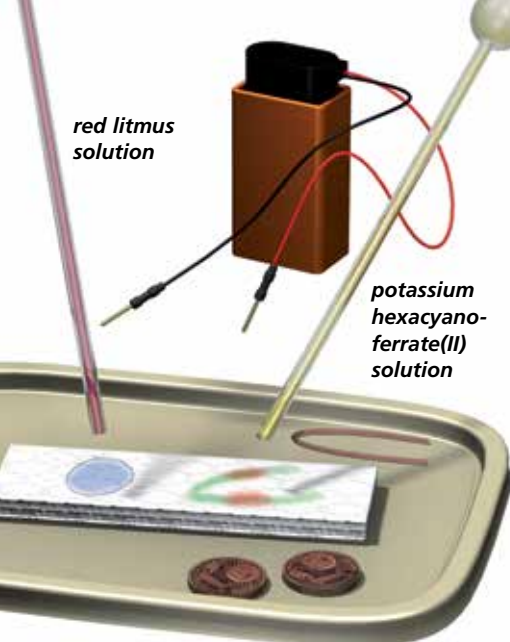
only as a contact; it should not touch the paper towel soaked in saline solution. Now press the contact pins of the battery cables to the two coins and be sure that the polarity is correct: The red cable (positive pole) leads to the coin on the copper wire bow. Allow the current to flow about 10 seconds; you might notice a faint bubbling underneath the coin with the black cable. Have a look at the paper under the copper bow. A greenish imprint has been formed. Add 1 – 2 drops of potassium hexacyanoferrate(II) solution to where the imprint was made. The brownish-red coloration demonstrates that some copper has come loose. Mark the place where the coin connected to the negative pole using a pencil line. Save the paper towel for Experiment 77.

*paper towel soaked in saline solution*

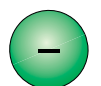
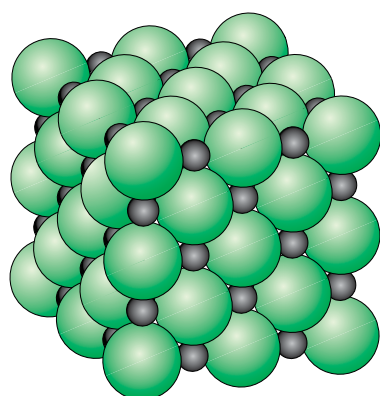


**The battery current dissolves metals.**





Detection of the reaction products after dissolution of copper

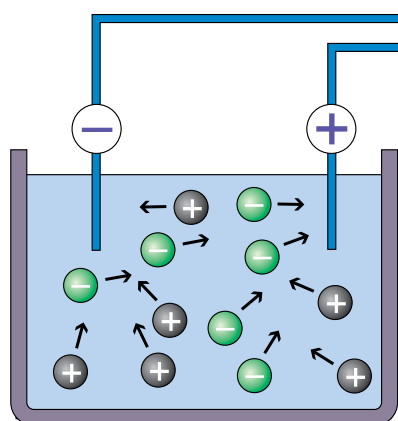


Chlorine ions, negatively charged



Sodium ions, positively charged

The smallest particles of sodium chloride — sodium ions and chloride ions — form a crystal lattice with a regular structure.



Migration of the ions through the sodium chloride solution

## EXPERIMENT 76

**Additional material:** Besides the coins, 2 iron nails

Repeat Experiment 75 with a new paper towel soaked in saline solution and replace the copper wire bow with the two nails (lay them next to each other so that the coin is steady. Run current again for about 10 seconds. Did the bubbling happen this time? It shouldn't be hard for you to prove that some iron went into solution. Keep the paper towel for the next experiment.

Metals are dissolved at the positive pole. So what happens at the negative pole? If you take that coin away, you won't notice any changes. Mark the place again with a pencil line.



For **citric acid**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

## EXPERIMENT 77

To a test tube with 2 cm of water, add 5 drops of litmus solution and 2 – 3 small (!) crystals of citric acid. Let a little of the red solution drip onto the marked place on the two paper towels. You will notice a faint blue coloration. Apparently, an alkaline-reacting substance has formed. Throw the paper towels into the trash.



**Question 24.** Sometimes the blue coloration doesn't occur. Why could that be?

## What's happening here?



### When electrical current gets in on the act

The latest experiments have demonstrated that, when battery current (direct current) passes through a saline solution (sodium chloride solution), chemical reactions take place: Metals go into solution at the positive pole and an alkaline-reacting substance is formed at the negative pole.

To understand what's happening here, you need to know a little about how salts are constructed. The smallest particles of sodium chloride — and other salts as well — are not molecules like with  $H_2O$  and  $CO_2$ , but electrically charged atoms or atom groups. These charged particles are called **ions** since they — as you will see in a moment — are able to migrate (Greek *ion* = going, migrating). Sodium chloride consists of positively charged sodium ions and negatively charged chloride ions that alternate regularly with each other in the salt crystal and are held together by their opposite charges. When dissolved in water, the crystals break up and the ions move around freely in the solution. They are the reason for the **electrical conductivity** of salt solutions.

If direct current from the battery flows through the sodium chloride solution absorbed by the paper, the negatively charged chloride ions migrate to the positive pole and the positively charged sodium ions to the negative pole according to the maxim: opposites attract.

Experiments 75 to 77 are a variation of the **chloralkali electrolysis** (Greek *lysis* = loosening) that is performed on an industrial scale. In this process, the chloride ions are discharged as chlorine gas at the positive pole and hydrogen gas and soda lye (sodium hydroxide solution) are produced at the negative pole. In your experiments, the hydrogen slipped away unseen; however, it becomes apparent sometimes in the form of gas bubbles. You were able to detect an alkaline-reacting substance, soda lye. Instead of the chlorine being released, the metals associated with the positive pole went into solution as positively charged ions. But that was of course the purpose of the exercise. You can learn more about chloralkali electrolysis in the "Technology and environment" info box on p. 52.

## Interplay of colors with sodium carbonate



For **sodium carbonate**, **ammonium iron(III) sulfate**, **citric acid**, and **copper sulfate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8, and for **copper sulfate** follow the instructions for disposal on p. 75.

Phew! After that excursion into theory, let's perform some experiments with some color-happy compounds.

### EXPERIMENT 78

In a test tube, dissolve 2 small spoonfuls of ammonium iron(III) sulfate in 6 cm of water. Pour

half of the solution into a second test tube and save it for Experiment 79.

In a third test tube with 4 cm of water in it, add 2 spoonfuls of sodium carbonate. You will need to seal the test tube with the stopper and vigorously shake it in order for everything to dissolve.

Now add a little less than half of the sodium carbonate solution to the ammonium iron(III) sulfate solution. A yellow-brown precipitate of iron(III) hydroxide forms. You will need the rest of the sodium carbonate solution in the next experiment.

### EXPERIMENT 79

Add 1 small spoonful of citric acid to the ammonium iron(III) sulfate solution you saved. The

solution turns a light yellow. If you now add the rest of the sodium carbonate solution, the solution will remain yellow and clear. There is no formation of precipitate. Apparently, the citric acid is responsible for this.

### EXPERIMENT 80

In a test tube, dissolve 2 small spoonfuls of copper sulfate in 6 cm of water and set aside half of the solution for the next experiment again.

Prepare sodium carbonate solution like in Experiment 78 and add about one third of it to the copper sulfate solution. A light-blue precipitate is formed that contains copper hydroxide and copper carbonate.

### EXPERIMENT 81

Add 1 small spoonful of citric acid to the copper sulfate solution you saved and shake it.

You won't notice any changes. Gradually add the rest of the sodium carbonate solution to it. Again, no precipitate is formed, but the blue color of the solution grows deeper into an appealing dark blue.

The diagram illustrates the experiments using test tubes and arrows indicating the addition of substances:

- Experiment 78:** A test tube containing a pale yellow solution of ammonium iron(III) sulfate. An arrow labeled "+ Sodium carbonate solution" points to a second test tube containing a yellow-brown precipitate. Text to the right states: "Sodium carbonate solution precipitates yellow-brown iron(III) hydroxide from ammonium iron(III) sulfate solution (Experiment 78), and citric acid prevents the precipitation (Experiment 79). A yellow solution is formed."
- Experiment 79:** A test tube containing a pale yellow solution of ammonium iron(III) sulfate. An arrow labeled "+ Citric acid" points to a second test tube containing a light yellow solution. A second arrow labeled "+ Sodium carbonate solution" points to a third test tube containing a yellow solution.
- Experiment 80:** A test tube containing a light blue solution of copper sulfate. An arrow labeled "+ Sodium carbonate solution" points to a second test tube containing a dark blue solution. Text to the right states: "Sodium carbonate solution precipitates blue-green copper carbonate/copper hydroxide out of copper sulfate solution (Experiment 80). If the copper sulfate solution is previously laced with citric acid, there is no precipitate and a dark-blue solution is formed (Experiment 81)."
- Experiment 81:** A test tube containing a light blue solution of copper sulfate. An arrow labeled "+ Citric acid" points to a second test tube containing a light blue solution. A second arrow labeled "+ Sodium carbonate solution" points to a third test tube containing a dark blue solution.



Industrial chloralkali electrolysis (photo: Lurgi, Frankfurt am Main)

## What's happening here?



Sodium carbonate produces characteristically colored precipitates with many metal compounds. Some acids — including citric acid and tartaric acid — are able to prevent the formation of precipitate when their molecules “hug” the iron or copper ions, thus denying access to the precipitating agent. Precipitating agents (sodium hydroxide in this case) are substances that produce precipitates from solutions. The dark-blue solution from the last experiment can be used to detect fructose. You will be performing just such an experiment with honey, in fact (Experiment 110). The oftentimes intensely colored compounds produced in this way are called complexes (Latin *complexus* = embrace, hug). You wanted to know what the deal is with the Roman numerals in compounds, right? Well, the Roman numerals refer to multiple charges in metal ions. Sodium ions have one charge, so they are said to be singly charged. Calcium ions have two charges, so they are doubly charged. In ammonium iron(III) sulfate, the iron is triply charged; in potassium hexacyanoferrate(II), the iron has a double charge. The Roman **charge numbers** are only used for metals with different charges, such as iron.



**Question 25.** What part of the name tells you that potassium hexacyanoferrate(II) contains iron? The table on p. 20 will help you.

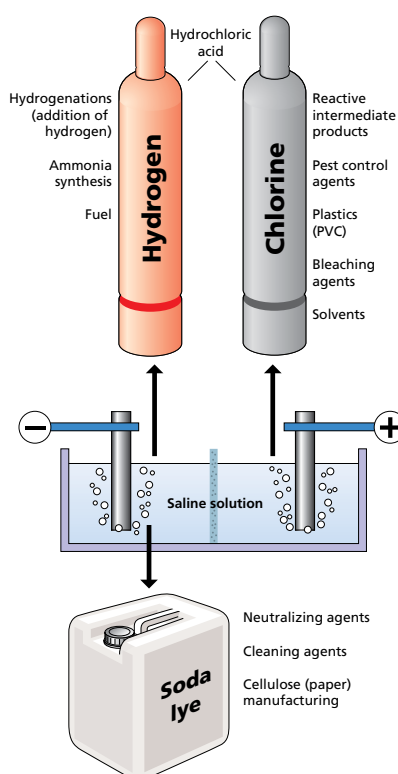
## TECHNOLOGY AND ENVIRONMENT



### Professional chloralkali electrolysis

Chloralkali electrolysis is among the most important processes in the basic materials industry. As already mentioned, it involves the recovery of hydrogen, soda lye, and chlorine. The process starts with saline solution in sealed containers through which direct current is fed. But, unlike in your experiments, the precipitation of the chlorine will only work if the current is fed not over disintegrating metals but instead over resistant electrodes made of graphite or titanium. Electrodes are conductive components that dip into the saline solution during electrolysis to feed electrical current in and out of the solution.

The hydrogen produced during electrolysis is used as a clean fuel as well as for the production of other chemical and energy products. The soda lye and sodium hydroxide that is recovered from it through evaporation are used in numerous technical processes and as a component of cleaning agents. Chlorine is used for the manufacture of pest control agents, plastics, and solvents. Solvents containing chlorine frequently pose less of a fire hazard than hydrocarbons (gasoline) but can be very harmful to health. In concentrations that are harmless to humans, chlorine is used to disinfect swimming pools. You might therefore be familiar with the typical smell. But chlorinated drinking water — yes, it exists — isn't recommended for those who enjoy an aromatic tea.





# 8

## Water and its elements

It would be no exaggeration to call water the most important foodstuff. Nowhere does this become clear than during natural disasters when the water supply breaks down and thousands of people have to suffer from thirst. Water is the solvent for the majority of the substances that control the life-supporting processes in plant and animal organisms.

### About saturated and unsaturated solutions

Water plays an important role in the laboratory and in your chemical experiments, too. You have already prepared numerous aqueous solutions, for example, of copper and iron salts, of acids and alkalis, of detection agents such as potassium hexacyanoferrate(II) or litmus dye.

There are readily soluble and poorly soluble substances. But even with readily soluble substances, water isn't able to dissolve unlimited amounts.

#### EXPERIMENT

### 82

**Additional material:** Table salt

To a test tube with 5 cm of water, add table salt by the spoonful. After each spoonful, close the tube with the stopper and shake! Continue adding salt until there is undissolved, leftover salt at the bottom even after shaking quite a while.



Two faces of water: Refreshing drink and solvent for the preparation of diluted acids.

### What's happening here?



The given amount of water — 5 cm in the test tube in this case (approx. 13 mL) — is not able to hold more material; you have prepared a **saturated solution**. As long as a solution continues to dissolve material, it is an **unsaturated solution**.

A saturated saline solution is "heavier" than the solvent water. This statement only makes sense, of course, when we compare the same volumes (spatial parts). One liter of saturated saline solution weighs about 1.2 kg, and 1 liter of water weighs 1 kg. The difference becomes clear in the following experiment.

#### EXPERIMENT

### 83

**Additional material:** Small chicken egg (discard after the experiment), table salt, spoon

Fill a graduated beaker up to the 100 mL mark with water and add 3 teaspoons of table salt to it. Carefully close the beaker with the lid, press against the bottom of the beaker with one hand and against the lid with the other hand, and shake thoroughly. If there is still some undissolved salt, it won't affect the experiment. Fill the other graduated beaker up to the 100 mL mark with water and carefully let the egg slide in (use a spoon). The egg will sink to the bottom of the beaker. Fish the egg out of the water and let it slide into the saline solution: It floats! Save the saline solution for Experiment 84.

**Egg in water and in saturated saline solution**



## Side Notes

### Bathing in saturated saline solution

If you've swum in the ocean, you know that it tastes salty. The salt content is 3.5%, that is, 100 g of sea water contains about 3.5 g of salt. The Baltic Sea is not quite as salty: about 1.5 – 2%.

The Dead Sea, which lies between Israel and Jordan (take a look in an atlas), is particularly salty. It is fed by the Jordan River but has no outlet. The inflow from the Jordan and the evaporation by the heat there more or less balance each other out. The salt content is said to be between 28 and 33%. The excess salt is deposited along the banks of the inland sea.

The salts contained in the Dead Sea have a healing effect on certain skin diseases, for example on psoriasis. People bathing in the Dead Sea are carried by the concentrated solution. You simply lay back and can read the newspaper if you want.

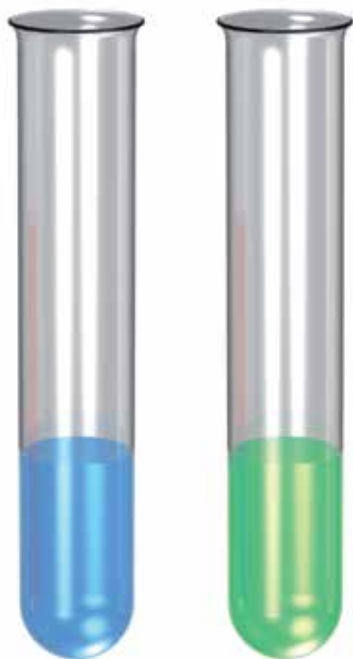


The water in the Dead Sea contains so much salt that you can lay down on the water without sinking.

Of course, the experiment only works with objects that are somewhat heavier than water but lighter than the saline solution. A huge egg with a volume of 1 liter (imagine an egg-shaped liter bottle) would weigh about 1.05 kg, so it would be heavier than 1 liter of water but lighter than 1 liter of saturated saline solution.

## Full, but there's still room for more

The following experiment shows that a saturated saline solution, that is, sodium chloride solution, is perfectly capable of still having an appetite for other salts.



The color of copper sulfate solution depends on the solvent.



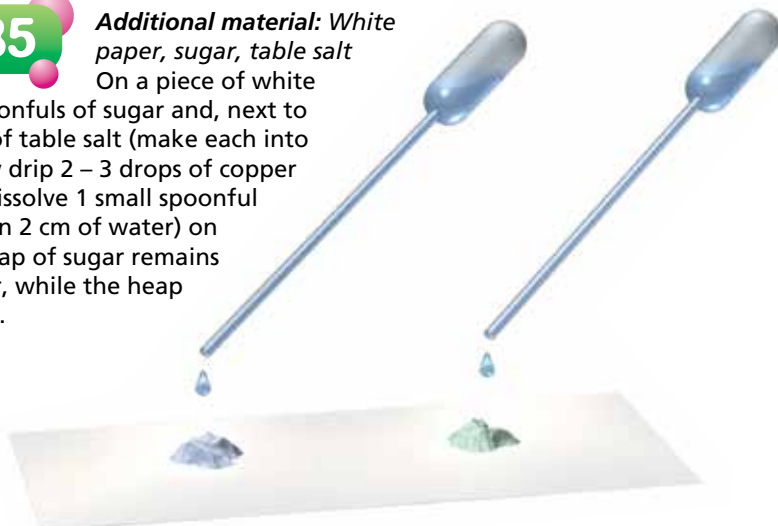
For **copper sulfate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8 and the instructions for disposal on p. 75.

### EXPERIMENT 84

To a test tube, add 3 cm of the saturated solution from Experiment 83, add 3 cm of water to a second test tube and add 1 small spoonful of copper sulfate to each. Seal the tubes successively with the stopper and shake them thoroughly. The blue crystals have dissolved in the saline solution, too, but the solution is no longer blue like in the control tube, but green. You can use this color change to differentiate between sugar and table salt without a taste test.

### EXPERIMENT 85

**Additional material:** White paper, sugar, table salt  
On a piece of white paper, place 3 spoonfuls of sugar and, next to that, 3 spoonfuls of table salt (make each into a little heap). Now drip 2 – 3 drops of copper sulfate solution (dissolve 1 small spoonful of copper sulfate in 2 cm of water) on each heap. The heap of sugar remains a light bluish color, while the heap of salt turns green.



How you can differentiate sugar from table salt — without a taste test

## What's happening here?



Experiment 84 makes it clear that saying that a solution is saturated only makes sense if one also states the substance with which it's saturated. Lime water, saturated calcium hydroxide solution, is also able to dissolve other materials — sodium chloride, for example. The green coloration of the copper sulfate solution is a result of the chloride ions present in such abundant quantities in the solution. You already know the dark-blue solution that is formed when citric acid is added to alkaline copper sulfate solution. Like all metal ions, copper ions don't swim around "naked" in the solution but dress up in different outfits: in a light-blue one made of water molecules, in a green one made of chloride ions, in a dark-blue one made of citrate ions (resulting from citric acid and sodium carbonate).



Copper ions dress up in different-colored clothing made of: water molecules, chloride ions, and citrate ions.

## Cubes and parallelograms

Up to now, you've experimented with solutions first and foremost and have seen that reactions often only occur with substances in the dissolved state. Many times, though, it's necessary in chemical operations to work the other way and remove the solvent from a solution. In so doing, the dissolved substance settles, crystallizing out. The crystals of the different salts have characteristic shapes. Let's take a closer look at crystals of sodium chloride (table salt).

### EXPERIMENT 86

**Additional material:** Lid from a canning jar

Prepare a saturated saline solution and add a few milliliters of the solution to a bowl. Label the bowl and put it in a warm place — in the sun or near a heater, depending on the season. This procedure requires some patience, but after some time the water will have evaporated. Little crystals are left behind in the bowl. The table salt has crystallized out. If you have a magnifying glass handy, you should take a closer look at them. You might be able to make out tiny cubes in some places. This crystal shape is characteristic for sodium chloride; it fits with sodium chloride.



The cubic structure of the crystals is typical for sodium chloride, for example in the mineral halite (photo: R. Bode, Salzhemmendorf).



**Question 26.** Why does the cubic shape "fit" with sodium chloride? Leaf through the previous chapter again.

### EXPERIMENT 87

**Additional material:** Lid from a canning jar

In a test tube with a little water in it, dissolve 1 small spoonful of copper sulfate and allow the water to evaporate like in the previous experiment. This time, you will get crystals in the shape of little parallelograms. This is what mathematicians call rectangles with two sets of parallel sides but with pointed (acute) and blunt (obtuse) angles.

Here, the ions in the crystal lattice are arranged in a more complicated structure but, as with sodium chloride, this structure is what determines the shape of the copper sulfate crystals.



Copper sulfate crystals — here on the mineral barite — exhibit the characteristic parallelogram shape.



If the lid is going to be used again for canning again, it needs to be washed thoroughly beforehand — preferably in a dishwasher.





Extraction of rock salt using modern drum milling machine (photo: CVA Kirschner, Flein) and salt production from sea water (photo: Lutz Travira, Portugal).



## TECHNOLOGY AND ENVIRONMENT

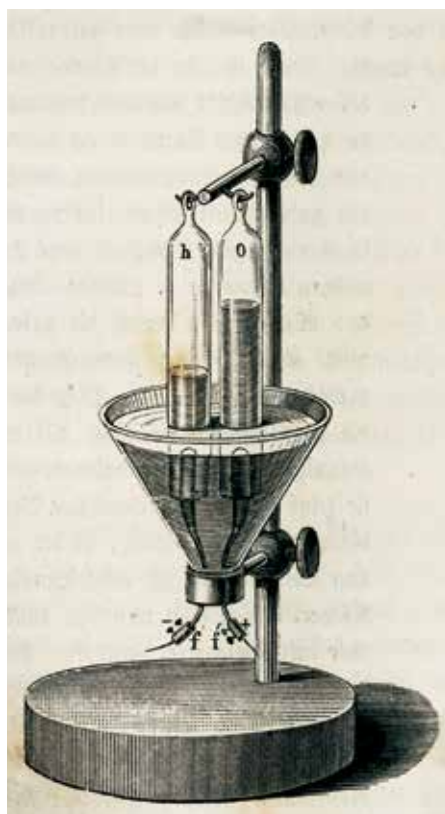


### Table salt — food product and raw material

For ages, “salt” has been an indispensable semi-luxury food product in the kitchen. People and animals need it to maintain essential life functions. **Table salt** — which, as you now know, is composed primarily of sodium chloride — is obtained from a variety of sources. Concentrated solutions, also called brine, are boiled down. Brine occurs naturally in part, or sometimes it is made by dissolving underground salt deposits. These salt deposits — the

salt from ancient seas that are now dried up — are also mined, leading to the term **rock salt**.

Large quantities of salt are extracted from the sea. In hot, dry regions, sea water is fed into shallow basins where the water evaporates and leaves the salt behind (similar to Experiment 86). You can buy sea salt obtained in this way at the supermarket. You already know in connection with the chloralkali electrolysis process that table salt is also an important raw material for the chemical industry.



As early as the 19th century, special devices were made in order to recover hydrogen and oxygen through electrolysis. From Stöckhardt, School of Chemistry (1863)

## Water — connection of two gases

You already know that water isn't an element like the Greek philosopher Empedocles assumed, but a chemical compound. We have also revealed the chemical formula of water —  $\text{H}_2\text{O}$  — to you. You also answered question 14, of course: Water molecules consist of 2 atoms of hydrogen (symbol H) and 1 atom of oxygen (symbol O). This is another case of chemical hide and seek: Two gases join together and a liquid is formed. At room temperature, that is. At temperatures of over  $100^\circ\text{C}$ , the boiling point, and at normal air pressure, water is a gas. In Experiments 86 and 87, water made off below the boiling point in the form of a gas: It evaporated.

In the following experiment — the electrolysis of water — you may not produce the gases hydrogen and oxygen in very large amounts, but you will at least be able to see them bubble a little.



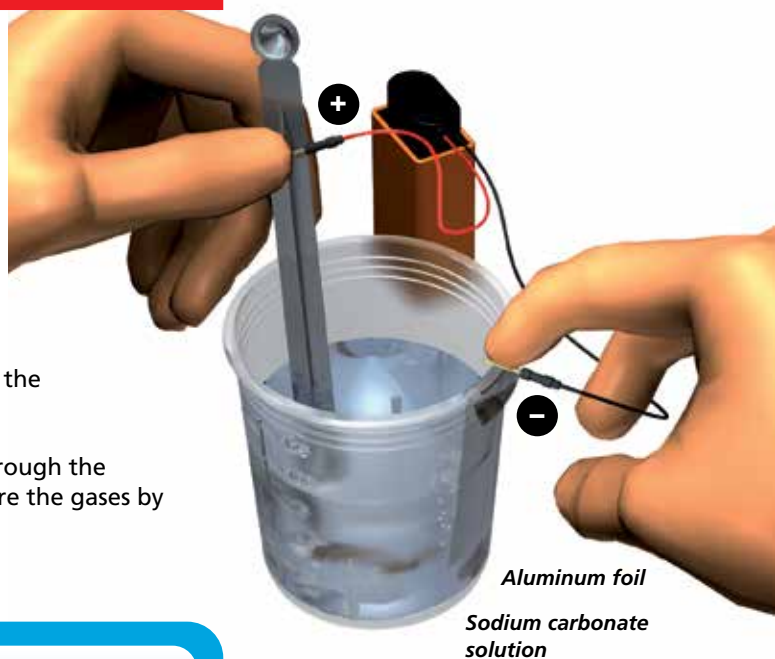
For **sodium carbonate**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

## EXPERIMENT 88

To a graduated beaker half filled with water, add 2 spoonfuls of sodium carbonate. Close the

beaker with the lid and shake it so that the sodium carbonate dissolves. After removing the lid, hang two strips of aluminum foil into the beaker. These will be the electrodes. Using your hands, complete the experimental setup as shown. Ensure that the contact pin of the red cable (positive pole) has contact with the spoon. The beaker will be especially stable if you place it into the large hole in the experiment station. A fine stream of gas rises at the negative pole: the hydrogen; at the positive pole, the spoon gets covered with larger-sized bubbles: the oxygen.

As mentioned before, you won't get enough of the two gases through the electrolysis of water to study their properties. It's better to prepare the gases by chemical means.



Aluminum foil

Sodium carbonate solution

*Electrical current breaks down water (that has been made conductive with sodium carbonate): The spoon (positive pole) gets covered with little oxygen bubbles, and hydrogen rises in the form of even smaller bubbles at the aluminum foil (negative pole).*



**Question 27.** Why does the addition of sodium carbonate make water conductive?

## Oxygen from hydrogen peroxide

You're going to use hydrogen peroxide as your supplier of oxygen. Ask your parents to get you 100 mL of 3% hydrogen peroxide solution from the pharmacy if they do not already have some available at home.

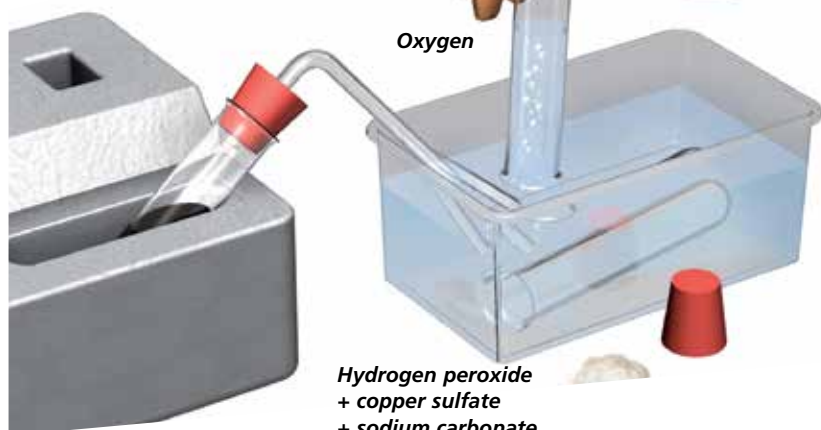


For **hydrogen peroxide** and **copper sulfate**, note the “Hazardous substances and mixtures” information on pp. 7 – 8 as well as the instructions for disposal for **copper sulfate** on p. 75. Be careful when twisting the angled tube into the stopper. Follow the instructions on p. 12/13. In case of injury: **First Aid 7** (inside front cover).

## EXPERIMENT 89

**Additional material:** Hydrogen peroxide, wad of cotton. Have an adult help you with this experiment. Prepare a basin filled with water. Place two clean test tubes in the basin so they fill with water. Carefully twist

the angled glass tube into the stopper. Place 2 spoonfuls of sodium carbonate and 1 small spoonful of copper sulfate into a third test tube. Add 5 cm of hydrogen peroxide to this, insert the stopper holding the angled tube, and position the test tube in the experiment station as shown. The contents of the test tube will turn brown-black and you will observe the lively formation of gas. Allow the first gas bubbles to escape, then hold the water-filled test tube over the rising gas bubbles, which will displace the water from the tube. When the tube is full of gas, seal it under water with the stopper. Place the tube in the experiment station. Now hold the second water-filled test tube over the escaping gas bubbles. Once it is filled, seal it with a dense, moist wad of cotton. Place it in the experiment station as well.



Production of oxygen

Oxygen

Hydrogen peroxide  
+ copper sulfate  
+ sodium carbonate





Testing for oxygen (glowing splint test)



Iron wire burns in undiluted oxygen.

## EXPERIMENT 90

**Additional material:** Tealight candle, stick of wood (shish kebab skewer)

Place the tealight candle on a fireproof base, light it, and then light the stick of wood on the candle flame. After a few seconds, blow the flame out and, after removing the wad of cotton from the test tube you just filled, lower the still-glowing stick into it. It flares up, which indicates that oxygen is present. Test the contents of the test tube you filled first. Blow the flame out and place the stick on the fireproof base.



**Question 28.** Sometimes the oxygen detection doesn't work so well with the test tube you filled first. Can you think of why?

## What's happening here?



Hydrogen peroxide is also called dihydrogen dioxide. In Greek, the prefix *di* means "two." If you compare the formula of hydrogen peroxide,  $\text{H}_2\text{O}_2$ , with that of water,  $\text{H}_2\text{O}$ , you will see that peroxide has an additional oxygen atom — one oxygen for each hydrogen.

This additional oxygen is easily split off, for example by copper oxide, which was produced in Experiment 89 from copper sulfate and sodium carbonate. Other substances bring about decomposition, too, and — like the copper oxide produced — speed up the reaction merely by being present. Such substances are called **catalysts** (Greek *katalysis* = dissolution, a dissolving). The described effect was discovered in decomposition reactions like the one we just performed. The catalysts in motor vehicles contain reaction accelerants to reduce toxicity in the exhaust.

Experiment 90 shows that combustion processes are much livelier in pure oxygen than in air, which — as you know — contains only 21% oxygen. Other substances burn with a bright flame in oxygen, too. Thin iron wires glow and emit sparks. The **glowing splint** test is used to detect oxygen.

## The "featherweight" of gases

Maybe you know that boxers in the lowest weight class are called featherweights. Hydrogen is the featherweight of gases, and therefore the lightest substance of them all, of course. And you can trap this happy-go-lucky type in a test tube!



For **calcium hydroxide** and **sodium carbonate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

Be careful when twisting the angled tube into the stopper. Follow the instructions on p. 12/13! In case of injury: **First Aid 7** (inside front cover).

## EXPERIMENT 91

**Additional material:** Aluminum foil, wad of cotton

Have an adult help you with this experiment. The experimental setup is the same as in Experiment 89, and the hydrogen is trapped in the same way. This time, too, two test tubes lie in the basin filled with water.

Twist the angled tube into the stopper. Then add 5 spoonfuls each of calcium hydroxide and sodium carbonate as well as a scrap of aluminum foil measuring about  $4 \times 4$  cm. Seal the tube and shake vigorously.



Production of hydrogen according to the pressure electrolysis method  
(Photo: Lurgi, Frankfurt am Main)



Open the tube (do not breathe in the dust!) and pour 3 – 4 cm of very hot tap water into it. Place the stopper with the angled tube on the test tube and let the first gas bubbles pass through the water.

Now fill the first test tube with gas, seal it under water with the stopper and place it in the experiment station.

After filling the second test tube, seal it with the dense, moist wad of cotton. Place it in the experiment station as well but with the sealed opening facing down.

?

**Question 29.** Why face down?

## EXPERIMENT 92

**Additional material:**  
Tealight candle on  
fireproof base, stick

of wood (shish kebab skewer)

Light the candle and then light the stick of wood on the candle flame. Remove the stopper from the test tube you filled first and hold the burning stick of wood to the mouth of the test tube. Don't be alarmed! There's a whistling noise. If you can temporarily darken your work area a bit, you might be able to see a blue flame whooshing into the test tube. Extinguish the flame on the burning stick and let the candle keep burning for the next experiment.

## EXPERIMENT 93

**Additional material:** Like in the previous experiment plus large clothespin

Hold the second test tube you filled — with the mouth still pointed downward — in the clothespin, lift it out of the hole it is in, and remove the wad of cotton. Light the stick of wood on the candle and hold it to the mouth of the test tube after 5 to 6 seconds (count!). There's a blub sound again! or a soft puff. If it's not that bright in the room, you will see a blue flame in the mouth of the test tube. Don't forget to put out the flame on the burning stick.

## What's happening here?

In Experiment 91, you formed some soda lye from sodium carbonate, calcium hydroxide, and water that reacted with the aluminum and released hydrogen. In Experiments 92 and 93, you burned the hydrogen: It bonded to the oxygen in the air to form water.

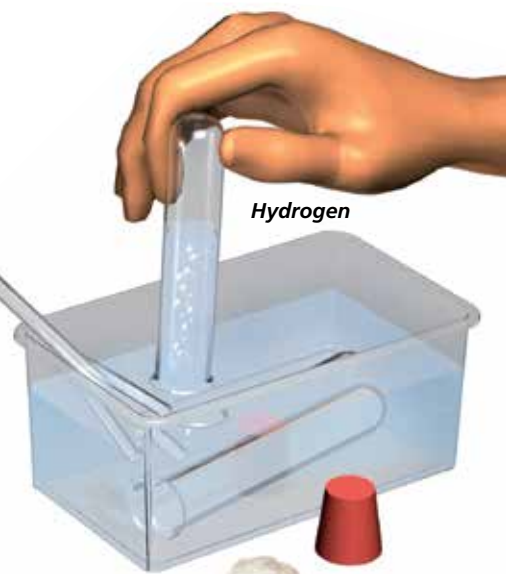
Experiment 93 also shows that the hydrogen in the upside-down test tube stays in the tube for several seconds. The featherweight gas floats to the top and hits the rounded bottom (which is now the top) of the test tube. After that, though, it begins to mix more and more with the air gases, which is to say the oxygen and nitrogen, which force the hydrogen out of the test tube.

Due to its lightness, hydrogen was used in earlier times to fill air balloons and airships, which led to catastrophic accidents due to the high flammability of the gas.



?

**Question 30.** Why is hydrogen a cleaner fuel than gasoline?



Calcium hydroxide  
+ sodium chloride  
+ aluminum scrap  
+ water

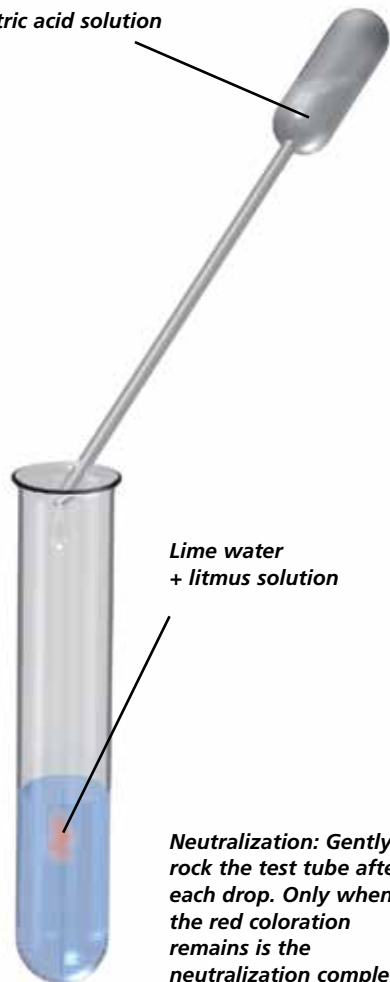


Even after 5 to 6 seconds, hydrogen is still in the upside-down test tube. Pull the burning stick away after lighting or you won't be able to see the little blue flame.



Hydrogen — light, but dangerous. The airship catastrophe in Lakehurst in 1937 (photo: dpa, Frankfurt am Main)

Citric acid solution



Lime water  
+ litmus solution

Neutralization: Gently  
rock the test tube after  
each drop. Only when  
the red coloration  
remains is the  
neutralization complete.

## Another way to produce water

In your experiments with magic blue, you came across the idea that an acidic- or alkaline-reacting solution has to be “neutralized” by the opponent before the color change can occur. Now let’s take a closer look at neutralization.



For **calcium hydroxide/lime water** and **citric acid**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

### EXPERIMENT 94

Add 3 drops of litmus solution to each of the two tubes.

In another test tube, dissolve 1 large level spoonful of citric acid in 4 cm of water (seal the tube and shake).

Now, while counting the drops, drip the citric acid solution into the test tube with 2 cm of lime water. Rock the test tube gently back and forth after each drop is added. After 13 – 15 drops, there will probably be a reddish cloud that disappears into the blue again after further shaking. After 1 – 2 more drops of citric acid, however, the color permanently changes to red. Did you count each drop? Make note of the number of drops needed.

### EXPERIMENT 95

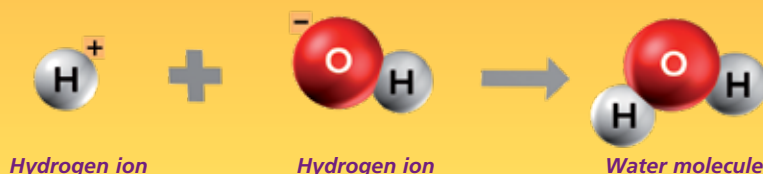
Now add the citric acid solution drop by drop into the test tube filled with 4 cm of lime water. This time, it takes longer until the red cloud appears and until the color changes permanently to red. You will have used twice as many drops.



## What’s happening here?

Here’s what happens during **neutralization**. Let’s take another look into the realm of atoms. Acids split off positively charged hydrogen atoms in aqueous solution. “Aha!” you say, “hydrogen ions.” Completely right. They cause the litmus solution to turn red. Alkaline-reacting solutions contain negatively charged hydroxide ions. These are particles that are composed of one hydrogen atom and one oxygen atom and cause red litmus solution to turn blue. In aqueous solution, calcium hydroxide splits hydroxide ions off. Hydrogen and hydroxide ions are aggressive, especially in larger quantities, and can cause burns. This is the reason why acids and bases are dangerous to the skin and eyes.

If you add an acid to a solution containing hydroxide ions (for example, lime water as in the last two experiments), then each hydrogen ion grabs a hydroxide ion and forms an uncharged, harmless water molecule with it.



The positively charged hydrogen ion and the negatively charged hydroxide ion unite to form an electrically neutral  $H_2O$  molecule.

The game is continued down to the last hydroxide ion. It should be clear that, for twice the number of hydroxide ions (Experiment 95), you need twice the amount of hydrogen ions, which is to say twice the amount of acid.

A common characteristic of all neutralization is the joining of hydrogen ions and hydroxide ions to form water. And what happens with the “leftovers”? That is, what happens with the citric acid without hydrogen ions and the calcium hydroxide without hydroxide ions (it has two)? These remainders (also ions) combine to form calcium citrate.

Here’s a general rule: **acid + hydroxide → water + salt**

# 9

## Old and new cleaning guys — the surfactants

In Chapter 5, you studied the phenomenon of surface tension through experimentation and became acquainted with the surfactants, which make mountains of water come crashing down and otherwise spell problems for surface tension.

The oldest surfactant — used for personal hygiene — is curd soap, which continues to be offered in the form of solid bars of soap. For some time now, though, liquid soaps and washing lotions have been threatening their preeminence. They have similar properties as the dishwashing detergents you've already used.

For the experiments in this chapter, you will need a bar of **bath soap** (light-colored, if possible), **liquid soap**, and a **dishwashing detergent**. You might also have a "**citrus all-purpose cleaner**" in the house. If your water in your house is hard (ask your parents, who will know from doing the laundry), you will also need **distilled water**.

Prepare 20 – 30 mL of lime water again according to the instructions given in Experiments 9 and 10.

### The litmus test provides clarity

When you use a piece of soap for washing, you know what you're dealing with. But if someone gave you a soap solution, it wouldn't be so easy to tell whether it is dissolved solid soap or liquid soap. The litmus test brings the mystery to light.



For **calcium hydroxide** and **lime water**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

#### EXPERIMENT 96

**Additional material:** White vinegar, knife, smooth paper

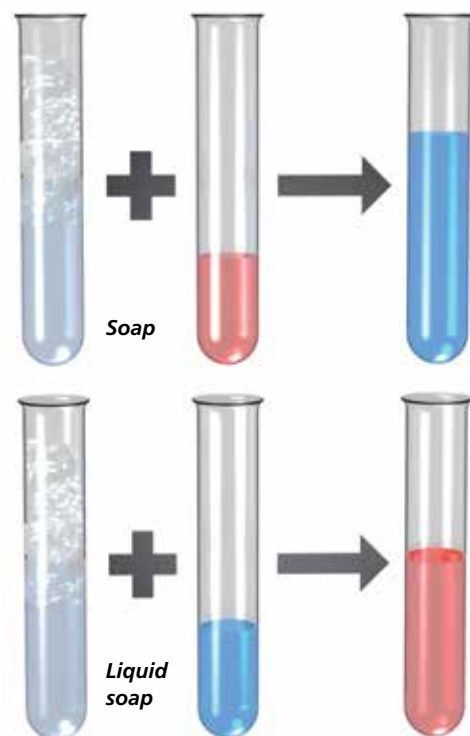
Using the knife, shave some "soap powder" from your piece of soap (place the paper underneath to catch it) and dissolve 2 spoonfuls of it in 4 cm of water. In a second test tube, prepare a red litmus solution out of 3 cm of water, 3 – 4 drops of litmus solution, and 1 drop of vinegar. Add the soap solution to the litmus solution: It turns blue.

#### EXPERIMENT 97

In a sealed graduated beaker, shake a few drops of liquid soap with 10 mL of water. Add the solution to diluted blue litmus solution (3 – 4 drops of litmus solution in 3 cm of water). Either you will see no change or a red color that changes to blue if you add 3 drops of lime water.

#### EXPERIMENT 98

If you have a "citrus all-purpose cleaner" available, repeat the previous experiment and replace the liquid soap with it. Look closely. The red coloration may be obscured by yellow dye in the cleaner. Add lime water to this drop by drop. At first, there is no change. Continue to add lime water until it turns blue.



The alkaline soap powder solution turns red litmus solution blue, and the weakly acidic liquid soap solution turns blue litmus solution red.



**Question 31.** Why do you need a lot more lime water to make the color turn blue with citrus all-purpose cleaner than with the soap solution in the previous experiment?





## What's happening here?

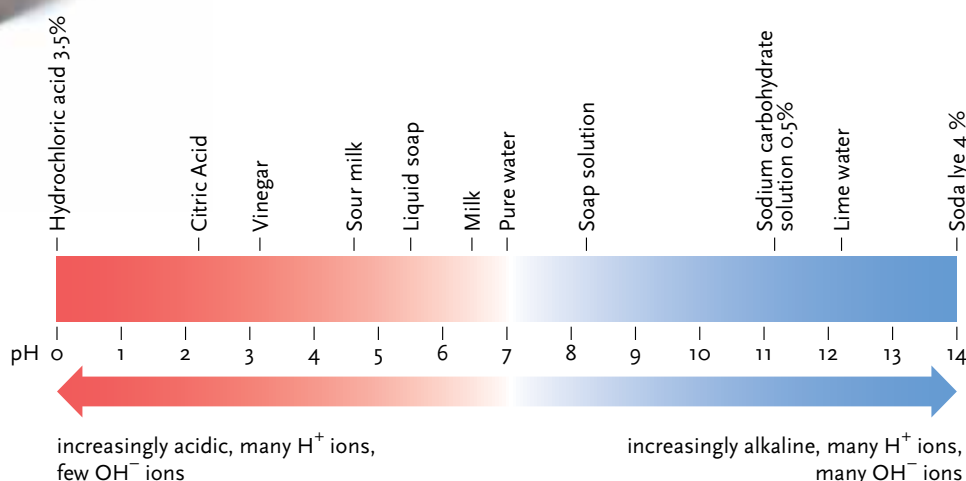
Solid curd soaps are obtained by converting the fatty acids contained in fats with alkalis, usually soda lye. Since the fatty acids are weak acids, the alkaline effect of the strong soda lye is predominant in the soap solution, which is why it turns the red litmus solution blue.

To make liquid soaps, acids are used that have similar structures to fatty acids and therefore act in similar ways. But they are stronger than the **fatty acids**, so liquid soap solution is neutral or acid-reacting. If the latter was the case, you neutralized the acid with a few drops of lime water (Experiment 97).

On bottles with liquid soaps, you will usually find the indication "pH-neutral" or "pH 5.5." The **pH value** indicates how strongly acidic or alkaline a solution is. Acidic solutions have pH values of 0 to 7, and alkaline solutions have pH values from 7 to 14. The pH value 7 is the golden mean: neutral reaction. For sensitive skin (especially in children), the neutral or weakly acidic soaps (pH 5.5) are preferable to solid soaps because they don't attack the skin. Alkaline soap solution has a slight irritant effect.



The pH value can be measured using a pH meter or an indicator that shows pH values through different colors after being added with the dropper.



The pH scale and pH values of some liquids and solutions; pH stands for *potentia Hydrogenii* = strength of the hydrogen, or more precisely, the concentration of the hydrogen ions that are responsible for the acidic effect.

## Hard water — unfortunately no fairy tale

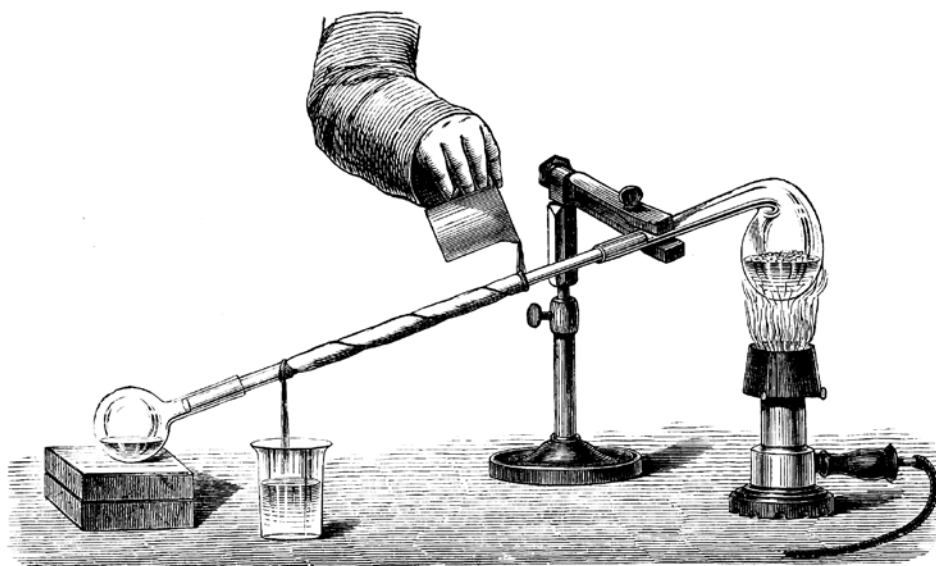
The following experiments are only truly compelling if you have distilled water available as mentioned above, unless very soft water flows out of the taps at your house. Distilled water is produced by vaporizing tap water or well water and cooling the vapor to liquefy it. This process is called **distillation** (from the Latin *destillare* = to drip off, referring to the re-liquefied water). Why all the fuss?

Depending on the region, tap water or well water contains different quantities of mineral salts in soluble, ionic form. They make the water "hard." Incidentally, the commercially available "distilled water" is usually not obtained through distillation, but through ion exchange. During full desalination (taking all of the salt out), all of the salt ions are exchanged for hydrogen ions and hydroxide ions, together producing — as you know from the previous chapter — water. It doesn't get any cleaner. Water obtained in this way is called "deionized" water.

To be able to study the effects of hard water, you will first need to prepare some hard water, especially if the water at your house is very soft, that is, has a low mineral salt content.



This ion exchanger removes all troublesome ions from drinking water by causing them to bond to special plastic resins (aqua purificata = purified water). The used-up, biodegradable cartridge is switched out for a new one (photo: LAB-IN, Kreuzwertheim).



*Historical distillation apparatus. To the right in the retort (the glass device with the bulb and the long conical neck), water is evaporated, liquefied again in the descending tube (additionally cooled by pouring cold water over it). The water liberated from mineral salts, which remain behind in the retort, is collected in the flask to the left. Adapted from Arendt-Doermer, *Techniques of Inorganic Experimental Chemistry* (4th edition 1910)*



For **calcium hydroxide**, **lime water**, and **citric acid**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

## EXPERIMENT 99

**Additional material:** Fresh mineral water

Add the mineral water in small portions to 5 cm of lime water. First the white precipitate of calcium carbonate forms that you know from Experiments 12 and 49. If you now continue to slowly add mineral water, the cloudiness becomes weaker and weaker until it finally disappears. The carbon dioxide in the mineral water has converted the poorly soluble calcium carbonate that was first formed into soluble and therefore invisible calcium hydrogen carbonate. Your hard water is ready.

## EXPERIMENT 100

Prepare a soap solution from toilet soap like in Experiment 96. In a test tube, shake 1 cm of the soap solution with 5 cm of highly distilled water. A lot of

foam is formed. Repeat the experiment with the hard water from the previous experiment. The mixture doesn't foam up. Instead, it precipitates a light, crumbly substance.

## EXPERIMENT 101

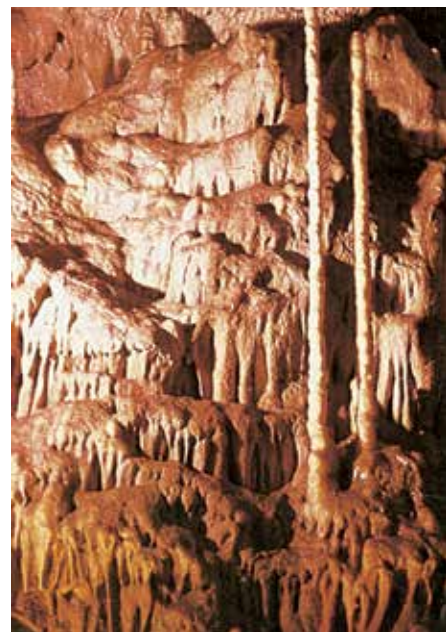
Shake a few drops of dishwashing detergent or liquid soap with the hard water. The mixture foams up and the crumbly precipitate doesn't form.

## EXPERIMENT 102

Now see how your tap water foams up with soap solution made from toilet soap.



*Soap solution with distilled water (left) and hard water*



*In stalactite caves, the hard water is disintegrated into "soft" water, carbon dioxide and calcium carbonate, which precipitates in the form of hanging stalactites or rising stalagmites.*

## What's happening here?



The hard water found in nature is formed in a manner similar to Experiment 99. When precipitates containing carbonic acid trickle through soil containing lime, they convert insoluble calcium carbonate (limestone) into soluble calcium hydrogen carbonate, which is the main cause of hard water. With conventional soap solution, insoluble lime soap is formed: the crumbly precipitate. No foam is formed. The precipitated **lime soap** is lost to the washing process. With dishwashing detergents and liquid soaps, calcium salts and other salts do not form any insoluble precipitates that reduce washing power. One way to tell is that the formation of foam is not impaired.





In a short time, the heat in the washing machine leads to the precipitation of limestone (calcium carbonate) on the heating element (photo: Benckiser Deutschland GmbH, Mannheim).



This "cage molecule" (zeolite A) contains sodium ions that are exchanged for the troublesome calcium ions contained in hard water. Sodium ions are not disruptive during the washing process (photo: Henkel KgaA, Duesseldorf).



The addition of acid leads to the precipitation of an insoluble substance (left), but not in liquid soap or dishwashing detergents

## TECHNOLOGY AND ENVIRONMENT



### Without softener, it's a no-go

Although today's liquid soaps and laundry and dishwashing detergents are immune to hard water, the softening of water is unavoidable. In hot steam boilers and washing machines, calcium hydrogen carbonate is broken down. The insoluble calcium carbonate regenerated in the process precipitates on boiler walls, in steam pipes, or on the heating elements of washing machines, which not only causes energy inefficiency but can lead to severe damage of the equipment. Through chemical research, softeners have been developed which

take calcium hydrogen carbonate and other hardening elements hostage and render them temporarily or permanently harmless. These softeners are contained in heavy-duty detergents but can also be added as needed (depending on how hard the water is). Bath soaps also contain many softeners in order to save soap. Then the reduction of foam formation is not so noticeable with moderately hard water (Experiment 102). However, softeners aren't up to the task of softening the hard water from Experiment 99 in the amounts added.

## Soap doesn't like it sour, either

### EXPERIMENT 103

**Additional material:** Same as Experiment 96

Prepare a soap solution from a bar of soap again and shake 1 – 2 cm of the solution together with five times the quantity of distilled water. When you add a few drops of vinegar, the foam collapses and a light-colored precipitate is formed.

### EXPERIMENT 104

Repeat the previous experiment with a spoon tip of citric acid (be careful not to let it get in your eyes).



**Question 32.** The precipitate is reminiscent of lime soap. Could it be lime soap?

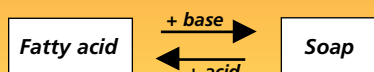
### EXPERIMENT 105

Repeat Experiment 103 with liquid soap or dishwashing detergent instead of bar soap. There is no impairment of foaming and no precipitate.

## What's happening here?



During the manufacture of solid soaps, the fatty acids combine with bases to form soap. If acid is added to soap solution, the reaction runs in the opposite direction: The weak fatty acids are released from the soap solution.



Since the fatty acids are practically insoluble in water, they settle as a fatty precipitate that has more of a soiling than a cleaning effect. The reduction of foam formation shows that washing power has been lost.

The modern surfactants contain stronger acids, are water-soluble, and are not precipitated when acid is added. This offers advantages to fabrics like wool and silk, which are attacked by alkaline soap solution. It is better to wash such textiles in weakly acid-reacting detergents.



## Reconciliation of water and fat

Water and fat (oil, for example) can't mix without help, and that causes problems when washing and cleaning.



**Question 33.** What observation from daily life shows that water and fat don't mix?

### EXPERIMENT 106

**Additional material:** Edible oil

To a test tube containing 5 cm of water, add about 1 cm of edible oil. The lighter oil floats on the water. Seal the test tube with the stopper and shake it thoroughly. It doesn't take long for the water and oil to form two layers again (Figure A). Water and fat repel each other. You will still need the test tube with its contents for Experiments 107 and 108.

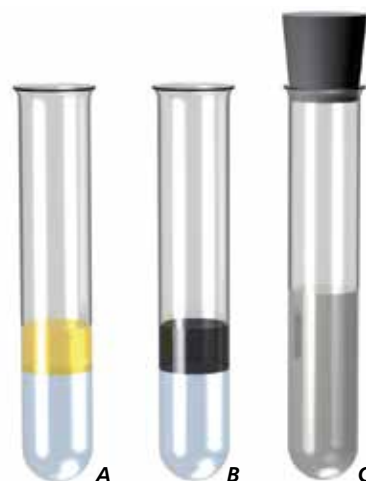
### EXPERIMENT 107

**Additional material:** Charcoal powder (from burnt matches, for example)

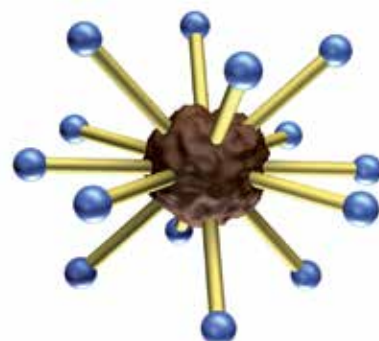
Add 1 – 2 small spoonfuls of the charcoal powder to the mixture from the previous experiment. Seal the test tube and shake. The charcoal powder collects in the oil layer and the water remains clear (Figure B).

### EXPERIMENT 108

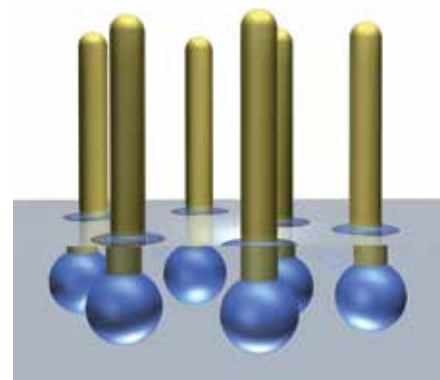
Add some soap solution (any kind) or dishwashing detergent to the mixture from the previous experiment and shake the sealed tube again. The water and oil don't separate so quickly, and the mixture now appears uniform. The grime is distributed in the mixture (Figure C). If it's made of matchstick charcoal, it will rise slowly to the top. If you pour the contents of the test tube into the drain, the glass will be practically clean. The grime has been washed away by the "washing lye."



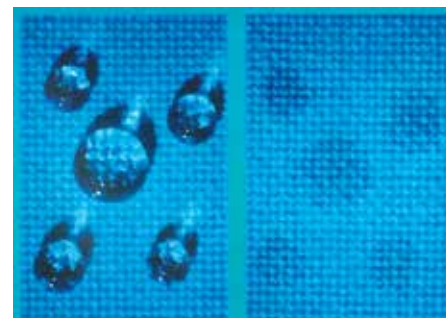
Soap solution mediates between water and fatty grime.



The surfactant particles point their fat-friendly end (yellow) toward the usually fatty grime and the water-friendly end (blue) toward the washing lye.



The surfactant particles collect on the water's surface. The weaker attractive forces between the fat-friendly rods reduce the surface tension.



Water beads away from the fabric. Surfactants ensure that there is sufficient wetting by reducing the surface tension.

## What's happening here?



### Little particles with dual functions

In Chapter 5, the surface tension of water was reduced with a dishwashing detergent, a surfactant. In the last experiment, you joined water and fat together using a surfactant. Is there a connection here?

Yes — and the connection comes from the structure of the surfactant particles. Imagine them as tiny rods that have a water-friendly and a fat-friendly end (the blue spheres and the yellow rods in the figure). The surfactant particles turn the fat-friendly end toward the usually fatty grime and the water-friendly one toward the water. They enclose the fatty particles of grime with a shell that is water-friendly on the outside, thus reducing the repelling effect between water and fatty grime. The grime distributed in the water is rinsed away with the washing lye.

The reduction of surface tension is also a result of the dual functions of the surfactant rods. On the water's surface, they arrange themselves as shown in the illustration. Since there are much weaker attractive forces between the fat-friendly rods protruding out of the water than between water molecules, the surface tension is reduced significantly. The water "relaxed" in this way is neither able to form a surface tension dome (Experiment 41) nor to carry the steely "ship" (Experiments 43 and 45). The tension required to form a drop shape isn't there, either. That's beneficial when washing, since relaxed water enables better wetting of the objects to be cleaned.

The soap particles and many other surfactant particles are negatively charged ions. The charge is at the water-friendly end (indicated by the blue spheres). But there are uncharged surfactant particles, too. These are called non-ionic surfactants, and you will find them in almost every package of laundry and dishwashing detergent.

# 10

## A first look at food chemistry



*Each and every day, foods and food products are checked in chemical investigational laboratories for quality, particularly for freshness and absence of harmful substances.*

What does food have to do with chemistry? Quite a lot, actually. Foods are substances or mixtures of substances and therefore the subject matter of chemistry, the science of the substances.

Food chemists deal with the composition of foods and their behavior when stored and processed. They check the quality of the various products and particularly that foods don't contain any additives that aren't allowed, such as adulterants or medications that are mixed into animal feed to increase the yield. The quality control also includes checking the state of freshness, the preservatives used, and a determination of the content of harmful substances (such as lead in lettuce or mercury in fish). Food chemists are indispensable.

### Black charcoal from white crystals

You now know that chemical compounds look quite different from the elements of which they are composed. This is the case with the sugar from the sugar bowl, too. In a few minutes, you are going to transform a white sugar cube into black charcoal. You will need a grown-up assistant for this experiment to make sure that the little fireworks don't cause any damage.

#### EXPERIMENT 109

**Additional material:** Aluminum foil as a base, matches, metal tweezers (or small tongs), tobacco ash or wood ashes, sugar cube

Grab hold of a piece of sugar cube with the tweezers (tongs) and hold it over a match flame. You won't be able to light the sugar very easily. Rub a few crumbs of the ash on a large surface of the piece of sugar and hold the prepared side over a match flame — with the tweezers (tongs) again, of course. The sugar melts, catches fire suddenly, and continues to burn even when moved away from the flame, frequently dripping a black, molten mass. Or a crispy, black substance remains stuck to the tweezers (tongs).

If no one at your house smokes, you can also use ash from wood or plants that you can obtain by burning dried lettuce or spinach leaves.



*Burning sugar cube*

### What's happening here?



Apparently, the tobacco or wood ash acts as a catalyst that promotes the burning of the sugar. The black substance making an appearance is carbon. The other elements of the sugar — hydrogen and oxygen — went up in smoke and flames or are still in the black mass that dripped off: sugar charcoal.

The fact that all foods contain carbon is a discovery that you can make through your daily experiences, even without chemical experimentation. For example, bread browns when you toast it in a toaster and meat browns when you fry it in a skillet.



## TECHNOLOGY AND ENVIRONMENT



### Sugar from cane and beets

When you think of "sugar," you naturally first think of sugar from the sugar bowl, **table sugar**, whether in the form of crystalline powder (refined, fine sugar) or sugar cubes. Depending on how it's obtained, a distinction is made between cane and beet sugar. Chemically speaking, it's the same substance: **sucrose**. The cane sugar cultivated above all in the Americas gets its name from the sugar cane, a type of sweet grass. The white core of the up to 7 meter- (23 foot-) high stalks contains 13 – 20% sucrose.

While sugar cane was cultivated in India as early as the 3rd century A.D., beet sugar is a relatively recent discovery. It was detected in beet juice in 1747 by the chemist Andreas Sigismund Marggraf (1709 – 1782). Beet sugar gained in importance when Napoleon blocked the importation of English goods from overseas in 1806, thus promoting the cultivation of beets. Production increased dramatically in the following years. Worldwide, though, cane sugar is still ahead with a market share of 65%.



*Sugar cane harvest: The up to 7 meter-high stalks contain a white core with 13 – 20% sucrose (photo: Rapunzel Naturkost AG, Legau).*



*These pictures show some of the stages of beet sugar production. In beet yards (1), the beets are washed, cut up, and imbibed. Then the raw juice is freed from harmful impurities (oxalic acid, among other things) through the addition of lime milk and carbon dioxide (2). The purified thin juice is thickened from 14% to 65% sugar content at a multistage evaporation station (3). The resulting thick juice goes to a special boiling apparatus where it is boiled under reduced pressure until crystallization begins. The separation of the crystallized sugar from the syrup is done in centrifuges (4), special spinners made for this purpose. The brown raw sugar obtained in this way is refined through washing and recrystallizing until the familiar white product is obtained (photos: Pfau/Süddeutsche Zucker AG).*





*Grape sugar (glucose) in natural and in concentrated form*

*Bee honey (right) is composed of grape sugar (glucose) and fruit sugar (fructose).*

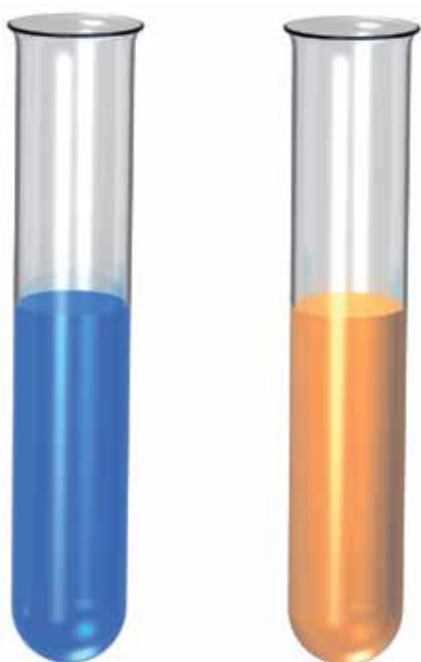


## Honey — a mixture of two sugars

Indeed, there are a whole bunch of substances that chemists call sugar besides the table sugar from which you made carbon. You may already be familiar with or have heard of two of these compounds: **grape sugar**, or **glucose**; and **fruit sugar**, or **fructose**. Grape sugar doesn't only come in the form of grapes; you can also get it as a powder or tablets that quickly provide power during physical exertion. Fruit sugar is used in diet foods. Both are responsible for the sweetness of honey.



For **copper sulfate** and **sodium carbonate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8 as well as the instructions for disposal for **copper sulfate** on p. 75.



*The dark-green or dark-blue solution slowly changes to an orange-colored precipitate.*

### EXPERIMENT 110

**Additional material:** Honey, small glass, teaspoon

Fill the small glass halfway with warm water and stir a teaspoon of honey in it (use the spoon handle to stir).

In a test tube, dissolve a small spoonful of copper sulfate in 2 cm of water and add 2 cm of the honey water you just prepared. The solution is light green (mixed color from the blue of the copper sulfate and the yellow of the honey). Add some sodium carbonate solution to this solution (2 spoonfuls of sodium carbonate in 3 cm of water). You will see the color change to dark green to dark blue, but no precipitate. The honey solution prevents the formation of precipitate.



**Question 34.** Have you ever observed a similar effect?

After a few minutes, the solution will become cloudy. A green precipitate forms that becomes lighter and lighter and turns a yellowish-red to brick-red color. Blue copper(II) sulfate has become orange-colored copper(I) oxide. Placing the test tube in hot tap water accelerates the reaction.

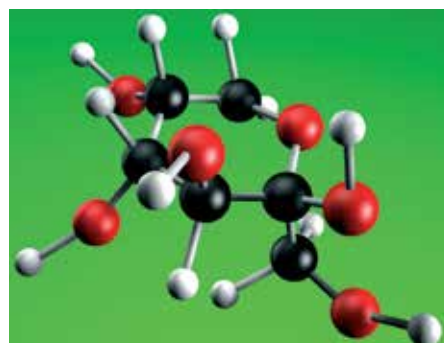
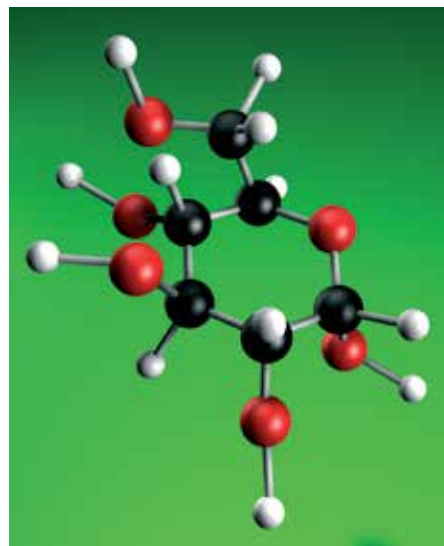
## What's happening here?






Table sugar (or sucrose in technical language), grape sugar (glucose), and fruit sugar (fructose), like a number of other sugars and starch, which we'll come to in the next section, belong to the carbohydrates. **Carbohydrates** are composed of carbon well as hydrogen and oxygen, the elements of water (Greek *hydor*).

The molecules of **glucose** and **fructose** contain the same number of "building blocks": 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. Both substances convert blue copper(II) sulfate into orange copper(I) oxide, but they differ in how the atoms in the molecule are joined together. This has far-reaching consequences. Like sucrose, glucose should be consumed with caution by diabetics or, better, not at all; in contrast, fructose is suitable as a sweetener for diabetics.

And how does the sucrose molecule look? In a sucrose molecule, a glucose and a fructose molecule are coupled together. Given that, Experiment 110 should work especially well. But it doesn't! As a result of the coupling, the "active places" on the two parts of the molecules are blocked. By heating it with acids, however, sucrose can be broken down into its active components. This is how artificial honey is made from sugar, in fact. Like bee honey, artificial honey is a mixture of glucose and fructose.



Here you can see the molecular models of grape sugar (glucose) and, below, fruit sugar (fructose).

-  Carbon atoms
-  Hydrogen atoms
-  Oxygen atoms

## Detecting sugar in apple juice

The unfermented juice squeezed out of pressed apples is called apple juice. Sugars that result from the apples' ripening process are responsible for the sweetness of the juice. You can prove this in the same way you proved there is sugar in honey.

### EXPERIMENT 111

#### **Additional material:** Apple juice

In a test tube, dissolve 1 small spoonful of copper sulfate in 2 cm of water, and then add 5 cm of apple juice.

The solution is likely to be more or less light green, depending on the color of the juice. Mix another solution using 2 spoonfuls of sodium carbonate in 3 cm of water. Adding this solution to the apple juice solution results in a dark green or dark blue solution. Any temporary precipitate that may have formed will dissolve. To speed up the reaction, put the test tube in a bath of hot tap water. After some time, the test tube contents will turn to a color between yellowish red and brick red.

In the last two experiments, you have performed a variant of the **Fehling's test**, which is used for the detection of glucose, fructose, and other substances. In Fehling's solution, the solution you made is mixed with a solution of sodium hydroxide and potassium sodium tartrate, which prevents the precipitation of copper through a process called complexing. In experiment 81, you achieved the same thing with citric acid.

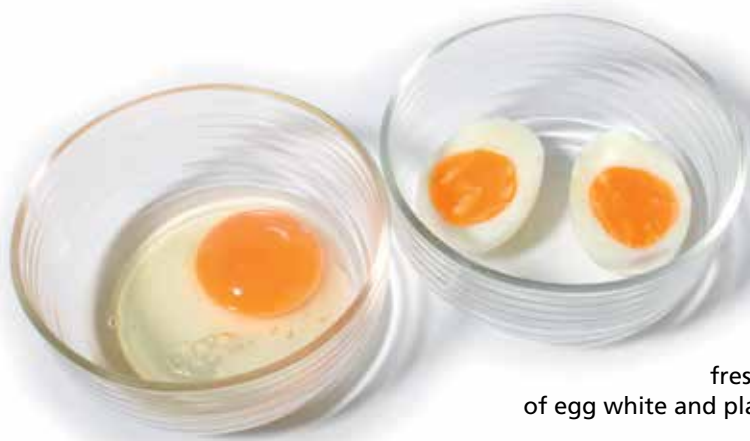
You can easily show that even the sweetest apple juice is acidic with a litmus test.

### EXPERIMENT 112

#### **Additional material:** Apple juice

Fill a measuring cup with 2 cm of water. Add 5 drops of litmus solution and 1 spoon tip of sodium carbonate, so that the solution appears translucent blue. Drop by drop, add some apple juice to the solution in the measuring cup, and immediately you will notice the blue turns to red, indicating the presence of an acid.

## Protein — the foundation of life



A raw egg with its gelatinous albumen (left) — and denatured eggs (right). Bon appetit!

When chemists talk about proteins, they're not just talking about meat and eggs, but a whole group of substances that practically form the foundation of life, since they participate in all of the processes in plant and animal organisms. This is why they're called **proteins**, actually (from the Greek *proton* = the first).

It makes sense that the egg in which a bird grows is especially rich in proteins. The viscous, transparent mass surrounding the yolk, the **egg white**, is made almost completely of protein, mostly **albumen** (Latin *albumen* = white of an egg) dissolved in water. Most of the fats in an egg are located in the yolk. Sometime when fresh eggs are being prepared in the kitchen, ask for 1 teaspoonful of egg white and place it in a graduated beaker.



For **copper sulfate**, **ammonium iron(III) sulfate**, and **potassium hexacyanoferrate(II)**, note the "Hazardous substances and mixtures" information on pp. 7 – 8, and for **copper sulfate** follow the instructions for disposal on p. 75.



Egg white congeals when copper sulfate or ammonium iron(III) sulfate solution is added, but not when potassium hexacyanoferrate(II) is added.

### EXPERIMENT 113

Dilute the egg white with 10 times the amount of water. Seal the beaker and shake vigorously. To a test tube with 3 – 4 cm of the diluted egg white solution, add a few drops of copper sulfate solution (dissolve 1 spoonful of copper sulfate in 3 cm of water). A bluish precipitate is formed. The egg white congeals. Save the diluted egg white solution and the copper sulfate solution for the following experiments.

### EXPERIMENT 114

Repeat the previous experiment with ammonium iron(III) sulfate instead of copper sulfate (dissolve a spoonful of ammonium iron(III) sulfate in 3 cm of water). Here, too, there is clearly precipitation as a result of the congealing of the egg white solution.

### EXPERIMENT 115

Dissolve 1 small spoonful of potassium hexacyanoferrate(II) in 3 cm of water and add the solution to 3 – 4 cm of the diluted egg white solution. The solution remains clear: no precipitate and no congealing of the egg white solution.



## What's happening here?

Proteins are very sensitive to chemical and physical influences. Not just copper sulfate, but other heavy metal salts such as ammonium iron(III) sulfate make proteins congeal. This is also the reason why these substances are toxic. In a congealed state, the proteins responsible for processes essential to life are no longer able to function. In potassium hexacyanoferrate(II), the iron is bound into a complex, that is, it is shielded, which is why congealing didn't occur.

You know that proteins are also sensitive to heat from your everyday experiences in the kitchen. When egg white is heated, it congeals in the same way as with heavy metal salts. What is happening is that the proteins are broken down and rearranged in a different order, changing the properties of the egg white in this case. Chemists call this process of breaking down and rearranging proteins **denaturation**. Up until now, you've just called it "cooking." The protein coagulates when cooking eggs!



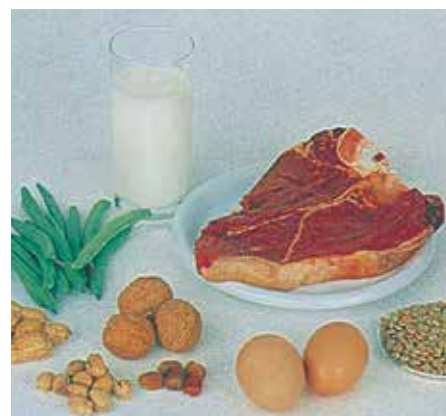
In addition to toxic chemicals like copper sulfate, there are other substances that can also cause proteins to coagulate, including stimulants like black tea.

## EXPERIMENT 116

### **Additional material: Black tea**

Have an adult brew a cup of black tea for this experiment. In a test tube, mix 3 – 4 cm of the diluted egg white solution with the same amount of the cooled black tea. The albumen coagulates.

Black tea contains substances that are used in tanning, which is the preservation of animal hides. These substances, called tannins, play an important role in preserving animal protein by curing it. You can read more about tannins on page 72.



*Protein-rich foods*

## Protein isn't only found in eggs

Proteins are not just contained in eggs, of course, but a variety of foods. You can detect them using special methods of detection like the following.



For **copper sulfate** and **sodium carbonate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8, and for **copper sulfate** follow the instructions for disposal on p. 75.

## EXPERIMENT 117

To a test tube, add 3 – 4 cm of the saved egg white solution and 1 spoonful of sodium carbonate. Seal the tube with the stopper (secure with your thumb!) and shake. If you now add a few drops of the copper sulfate solution from Experiment 113, the congealing you observed doesn't happen; rather, the solution gradually turns violet. This coloration is indicative of proteins.

## EXPERIMENT 118

### **Additional material: Raw meat or ground beef; cutting board, knife if needed**

Cut up some raw meat or simply take some ground beef and shake it together with a half test tube of water.

Separate the "meat water" by filtering off the remaining meat and test for proteins using the copper sulfate solution like in the previous experiment.

## EXPERIMENT 119

### **Additional material: Milk**

Shake together 2 – 3 cm of milk with 1 spoonful of sodium carbonate in a sealed test tube. After adding a few drops of copper sulfate solution, the violet coloration

appears. With whole milk (3.5 – 3.8% fat), the violet is somewhat fainter than with skim milk (low-fat milk).



*The Biuret reaction also indicates that protein is present in the milk.*

## What's happening here?



The detection reaction you just performed is called a **biuret reaction**, after the substance biuret in which this coloration was observed. Want to know why all proteins produce the same reaction?

Well, to understand you need to know a little something about the structure of proteins. Like starch, proteins form long chain molecules. But while the starch chain is made of glucose units, different amino acids alternate with each other in the protein chains. Besides carbon, hydrogen, and oxygen atoms, **amino acids** contain nitrogen atoms as a characteristic property (and some also contain sulfur atoms). The nitrogen atoms are located at the places in the chain at which the amino acids are linked together. This results in a characteristic grouping of atoms that matches with that in the biuret molecule.



Dried and fresh black tea leaves

## What makes black tea black

Black tea comes from the leaves of tea bushes that are mainly grown in parts of India, China, and other tropical and subtropical regions. People drink black tea because of its taste and because it contains antioxidants and, like coffee, the stimulant caffeine. In addition, tea contains a wealth of other substances. There are over 400 identified flavor substances in tea, including the previously mentioned tannins (Latin *tannare* = to tan, as in tanning leather). What could you add to black tea to make it even blacker? Iron!

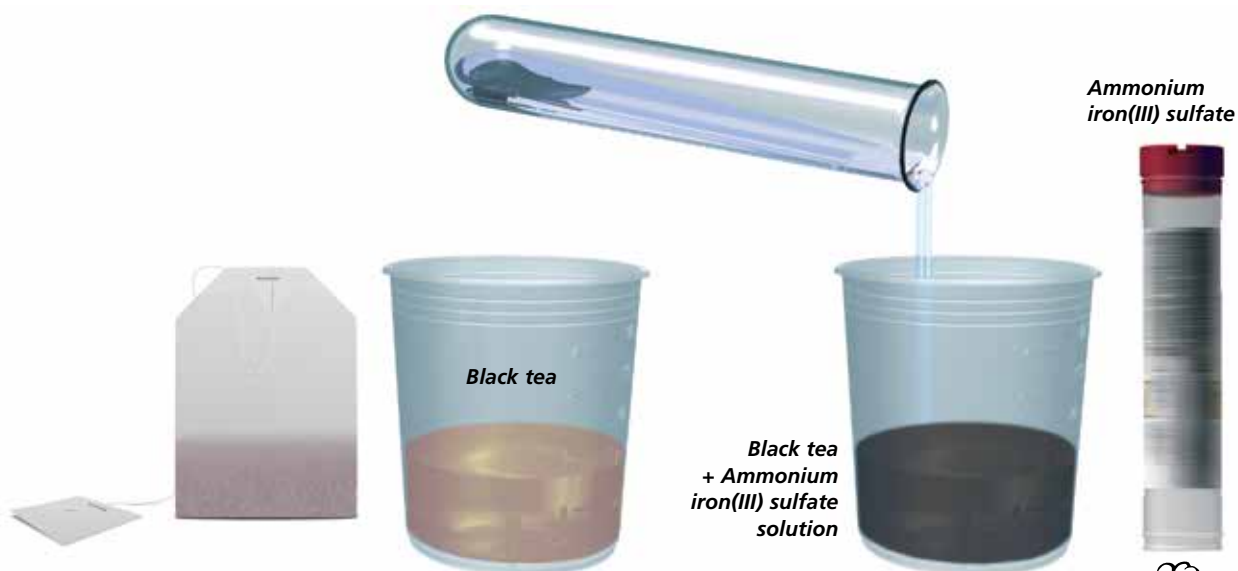


For **ammonium iron(III) sulfate** and **citric acid**, note the “Hazardous substances and mixtures” information on pp. 7 – 8.

### EXPERIMENT 120

#### Additional material: Black tea

Have an adult brew a mug of black tea for you. Let the tea cool and divide it into two measuring cups. The contents of one cup will be used as a control, so that you can make a comparison. In a test tube, mix a solution of 1 small spoonful of ammonium iron(III) sulfate and 2 cm of water. Pour the ammonium iron(III) sulfate solution into the second cup of tea. The tea will turn pitch black, sometimes with a bluish tinge.



Oak leaves with oak apple (oak galls)

## What's happening here?



**Tannins** are compounds with complex structures made of gallic acid and glucose. When combined with iron salts, tannins form a black, insoluble precipitate, which is called iron gall ink. This permanent ink is resistant to fading in the light and does not wash away with water. Because of this, iron gall ink has been used for centuries for important documents and signatures.

However, black tea is not used in the industrial production of iron gall ink. The main source of tannins for iron gall ink is oak apples or oak galls. You may have seen these spherical growths on oak trees in your neighborhood. Galls are abnormal tissue growths on plants and animals, usually caused by external factors such as parasites. Oak galls are caused by the larvae of wasps that lay their eggs on oak leaves.

## The darkness lightens

You have proved the presence of iron in a tannin solution (for example, black tea). Instead of Berlin blue, you get “gall black” — provided however, that the solution being examined is not too acidic. Try the following experiment to see for yourself.

## EXPERIMENT 121

### **Additional material:** white vinegar

Stir up the black solution from the previous experiment. Fill two test tubes with 3 cm each of the solution. To the first test tube, also add 1 cm of white vinegar. Seal the test tube with the stopper and shake it up. You will see almost no change. The vinegar does not affect the ink.

Put 1 small spoonful of citric acid into the second test tube, seal it, and shake it up. You will probably see some lightening of the solution and a precipitate that hovers finely distributed in the liquid.

## EXPERIMENT 122

### **Additional material:** Half of a lemon

Squeeze the juice from half a lemon and filter out the pulp. Add the filtered lemon juice to 3 cm of your black ink solution, seal the test tube, and shake. The contents of the test tube will be noticeably lighter and slightly yellow in color. After the solution sits for a while, a small amount of a light brownish precipitate will be visible. These are iron compounds.

In the same way that the citric acid reacted with the black ink, ascorbic acid (or vitamin C) in the lemon juice also reacts with the ink. Ascorbic acid slows the process that causes the permanence of the iron gall ink. If you have any vitamin C (tablets or powder) available to experiment with, you can repeat the experiment with a solution of lemon juice and vitamin C powder. In stronger acids (e.g., dilute hydrochloric acid) all freshly precipitated black precipitate immediately goes back into solution.



Tea and lemon

## Tannins in apple juice

Tannins are found in many foods. They often cause a tart, puckery feeling in the mouth. It's easy to taste them in a tart apple or bitter tea. But even in the sweetest apple juice, you can prove the presence of tannins.

## EXPERIMENT 123

### **Additional material:** Apple juice

Dissolve 1 small spoonful of ammonium iron(III) sulfate in a test tube containing 2 cm of water. To the solution, add 5 cm of apple juice. Seal and shake the test tube.



**Question 35.** Why would the black solution formed from the apple juice be lighter than the black solution formed from the tea?

## Calcium — a common food supplement

Among the food supplements offered at pharmacies, there are usually several calcium preparations, often together with magnesium and vitamin D, which have a favorable influence on the absorption and effectiveness of the calcium. Calcium plays a central role in the body's metabolic goings-on: It controls the heart's metabolism and muscular activity and is involved in bone growth. Many older people take calcium preparations in order to slow the bone loss associated with aging.

You will prepare a soluble calcium compound and use it to put a calcium detection method to the test. Then you can try out whether the detection method that the commercially available preparations use works.



For **calcium hydroxide** and **sodium carbonate**, note the "Hazardous substances and mixtures" information on pp. 7 – 8.

## EXPERIMENT 124

### **Additional material:** White vinegar

In a sealed test tube, shake together 1 spoonful of calcium hydroxide and 5 cm of vinegar (thumb on the stopper!). The solution will slowly become clear. Then filter it. The filtrate contains dissolved calcium acetate. You will detect it in the following experiment.



*In old age, calcium deficiency promotes osteoporosis, a progressive loss of bone mass. But calcium preparations alone don't help. Healthful nutrition and exercise are just as important.*





Calcium hydroxide  
+ vinegar

Preparation of calcium acetate (Experiment 124) and detection of calcium (Experiment 125)

## EXPERIMENT 125

In a test tube, dissolve 2 spoonfuls of sodium carbonate in 3 cm of water. Seal the tube and shake. Here, too, don't let any spray out! Slowly add the sodium carbonate solution to 2 cm of the filtrate from the previous experiment. A thick, white precipitate is formed that dissolves when vinegar is added, bubbling forcefully.

## What's happening here?



In Experiment 124, you produced a calcium acetate solution from calcium hydroxide and acetic acid. In the following experiment, the calcium acetate was converted with the sodium carbonate solution. This resulted in a switching of partners:



You know the precipitate from calcium carbonate from the CO<sub>2</sub> detection experiments. There, it detected the gas, and here it detected the calcium in calcium acetate. In calcium preparations, the calcium is often present in the form of calcium carbonate, that is, in the same poorly soluble form as in the precipitate. The calcium carbonate is then dissolved either by the stomach acid or by acids contained in effervescent tablets in order to be available to the organism.

Incidentally, calcium acetate like what you synthesized is used as a food additive in the food industry.

To detect calcium in preparations, it will make a difference whether you have an effervescent or a normal calcium tablet. There is probably an older person you know in your family or circle of acquaintances who can give you a calcium tablet. In any case, you will need a sodium carbonate solution (dissolve 2 spoonfuls in 3 cm of water) as a detection agent.

## EXPERIMENT 126

**Additional material:** Calcium preparation (effervescent tablet or normal tablet), white vinegar

Effervescent tablet: Dissolve the tablet in a glass of water and, when no more gas is generated, filter the solution if it is still cloudy. Normal tablet: Crush the tablet in a clean, folded sheet of paper using a hard object (hammer). Add 1 small spoonful of the powder and 4 – 5 cm of vinegar to a test tube. Shake and filter off the insoluble components.

Now slowly add the sodium carbonate solution to the filtrate. A precipitate is indicative of a soluble calcium compound. If no precipitate is formed, a calcium compound may be there that doesn't respond to your simple agents.

That concludes our little excursion through food chemistry. As the chapter heading suggested, it was no more than a first glance.

And, this also concludes the experiments in Chem C1000. Did you have fun with the chemistry kit? Would you like to learn more and perform more experiments? Then go to [www.thamesandkosmos.com](http://www.thamesandkosmos.com) to find more chemistry sets and other science kits.



Synthetic calcium acetate is used as a food additive.

# 11

## Where do I dispose of the waste?

You may wonder whether the chemicals contained in Chem C1000 and the resulting reaction products can harm the environment and what precautions should be taken when disposing of them.

Most of the substances with which you perform the experiments and use only in small quantities do not pose a noteworthy threat to the environment. You can simply rinse them down the drain with lots of water.

If you take a look at the list of “Hazardous substances and mixtures” on pages 7 and 8 again, you will notice the “Environmental” pictogram (the sick tree and dead fish) next to the information for **copper sulfate**. In experiments with this chemical, please follow the instructions for disposal below.

### Disposal of copper sulfate

In the experiments, there may be leftovers of the light-blue copper sulfate solution, or alkaline-reacting precipitates or dark-blue solutions may be formed. Add some household vinegar to the alkaline leftovers, thus forming a light-blue solution containing copper salts.

This solution and copper sulfate solution can be properly discarded by placing them in a screw-top jar half-filled with water. Add a shot of vinegar, a few **small iron nails** or some **steel wool** to the jar (the rust-free steel wool used in the kitchen is unsuitable). Close the jar and let it stand in a place out of the reach of small children for one day. Like in Experiment 68, copper precipitates on the iron or steel wool. The solution containing some iron sulfate can be poured into the drain, and the copperized iron can be discarded in the trash.

There is one more chemical in your set that is considered harmful to the environment (especially for aquatic life) but which doesn't bear an “Environment” pictogram: **potassium hexacyanoferrate(II)**. Since there is no simple way for you to safely dispose of it, for the experiments on page 19 we gave you the tip of using this chemical sparingly. You prepare one solution for several experiments and need only small quantities of it (usually just a few drops).

### A few more tips

**Filter papers, paper towels, or cotton balls** that are wetted or soaked with chemical solutions can be placed in the household trash (wear protective gloves).

Your parents or other adults can take **leftover chemicals** that you no longer need with them the next time they go to your local hazardous waste collection site.

These tips are based on the guidelines recommended by a conference of educational administrators focused on safety in science classes, particularly in relation to safe disposal in the drain or household trash.



Steel wool

Disposal of copper-containing waste

# 12

## Answers to the questions



Spring in Iceland containing sulfur



Berlin blue ring

1. Young children will likely find it difficult to open the vial by simultaneously pressing and turning without being taught how to do it.
2. First you prepare the red ink from 5 cm of water, 10 drops of litmus solution, and 1 – 2 drops of white vinegar. You add 1 spoon tip of sodium carbonate to the rounded bottom of the “empty” test tube.
3. The filtrate is obviously not pure water but an alkaline-reacting liquid. Pure water doesn’t turn red litmus solution blue. Perform a control experiment to check!
4. The “ferrate” in potassium hexacyanoferrate(II) means that it contains iron (from Latin *ferrum*).
5. The addressee places a paper towel soaked with potassium hexacyanoferrate(II) onto the secret message (let the solution drip onto the paper towel).
6. Berlin blue is not stable in sodium carbonate solution (or in any alkaline solution, for that matter). When treated with acid, the blue coloration fades.
7. The entire amount of water runs quickly through the filter made of filter paper, whereas the filter made of writing paper practically holds the water back. Individual drops fall in long intervals of time.
8. In the strip of filter paper, the water rises up, while in the case of the writing paper only the part dipped into the water softens. You can see the leading edge of the water more clearly with colored water.
9. The two salts took the same time, but the copper compound went farther. Since speed (or velocity) is the distance traveled in a certain period of time, the copper compound was faster.
10. Discoloration occurs when the sodium carbonate is transported from the drip point to the ring.
11. Place a beaker face-down and perpendicularly into a basin full of water. The water doesn’t go into the beaker. Only if you tip the beaker to the side does the air escape, allowing some water to enter.
12. When cooled off, the air contracts. Negative pressure (a vacuum) is created in the test tube and the outside air pressure is able to push the water in the angled tube further down.
13. The wire length needs to have enough margin so that the wires don’t shrink so much that they might break when it is cold. They are especially long when they’re hot and droop lower.
14. The formula  $\text{H}_2\text{O}$  reveals that a water molecule consists of two atoms of hydrogen (symbol H) and one atom of oxygen (symbol O).
15. As the name reveals, sodium hydrogen carbonate also contains hydrogen (Greek-Latin *hydrogenium*).
16. Acid getting into the tube of lime water with the foam might dissolve the precipitate.
17. The air heated by the candle flame rises, carrying the carbon dioxide with it, and away from the flame.



18. Like in the experiment relating to question 17, the heat of the flame cause the air and combustion gases to rise, this time into the jar.
19. Instead of lime water, you could pour diluted blue litmus solution into the jar which indicates carbon dioxide by turning red.
20. The air bubbles rising next to the jar tell us that the jar is full.
21. During the precipitation of copper out of the copper sulfate solution, some iron must have dissolved.
22. Blue coloration might have occurred due to the dissolving of copper. Many copper salts, not just copper sulfate, are blue or green.
23. No, noble gold, which is below zinc in the table, can't force zinc out of solution. But ignoble zinc is able to precipitate gold out of a solution (out of gold chloride, for example), thus helping it regain its radiant luster.
24. You may have used too many little crystals of citric acid for the red litmus solution.
25. "Ferrate" comes of course from the Latin word *ferrum* = iron.
26. The cubic shape of sodium chloride crystals reflects the structure of the ion lattice shown on page 50.
27. Sodium carbonate is a salt composed of ions. You already know the positively charged sodium ions; the negatively charged carbonate ions are composed of one carbon atom and three oxygen atoms. The ions are what make the solution conductive.
28. At the beginning of the experiment, there is still air in the reaction vessel which — as you know — contains only one-fifth oxygen. The air "dilutes" the pure oxygen.
29. The light, rising gas is best "caught" in this way so that its pathway upward is blocked.
30. When gasoline and other hydrocarbons are burned, the greenhouse gas  $\text{CO}_2$  is produced in addition to water.
31. "Citrus all-purpose cleaners" usually contain citric acid or other acids (take a look at the ingredients). Those have to be neutralized by the lime water before the color can change.
32. No. If you used distilled water, there were no soluble calcium compounds present anywhere that are required to form lime soap.
33. One example is blobs of fat floating in soup.
34. Yes, during the formation of the dark-blue copper complex in Experiment 81.
35. The apple juice either contains a lower concentration of tannins or more vitamin C than the tea.



**Malachite** — a mineral containing copper hydroxide and copper carbonate (cut and polished)



**The cube-shaped crystals of halite** — sodium chloride, chemically speaking

**Chemical elements** — how many can you identify?  
The illustration on p. 21 will help if you have trouble.



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# Notes





## Notes



## Kosmos Quality and Safety

More than one hundred years of expertise in publishing science experiment kits stand behind every product that bears the Kosmos name. Kosmos experiment kits are designed by an experienced team of specialists and tested with the utmost care during development and production. With regard to product safety, these experiment kits follow European and US safety standards, as well as our own refined proprietary safety guidelines. By working closely with our manufacturing partners and safety testing labs, we are able to control all stages of production. While the majority of our products are made in Germany, all of our products, regardless of origin, follow the same rigid quality standards.

### Instructions for use for the safety goggles (Art.-Nr. 717019)

**Manufacturer Frame:** Georg Schmerler GmbH & Co. KG / Reitweg 7 / 90587 Veitsbronn / Germany

**Manufacturer Lens:** IPB NV / Steenovenstraat 30 / 8790 Waregem / Belgium

**Certification Office:** ECS GmbH – European Certification Service / Hüttfeldstr. 50 / 73430 Aalen, DE / Germany / Notified Body Number: 1883

**GSF** – Safety goggles comply with the Regulation (EU) 2016/425 on personal protective equipment (PPE), are design tested, and bear the CE symbol. The Model 610 in this kit is suitable for protection against mechanical dangers.

#### Identification of the frame:

**GSF** = Code letter of manufacturer

**166** = Number of the standard

**S** = increased robustness

**CE** = EC Conformity symbol

#### Identification of the safety lenses:

**IPB** = Code letter of manufacturer

**1** = Optical class

**S** = increased robustness

If frame and lens have differing certification marks in regards to S, F, B, or A, the lower of the two applies. Wearers with extremely sensitive skin may experience allergic reactions upon contact with some materials. There are no spare parts or additional components available for this model. If damaged please discard the safety

goggles and discontinue use. Eye protectors used for protection against high-speed particles that are worn over standard ophthalmic spectacles may transmit impacts, thus creating a hazard to the wearer. If protection against high-speed particles at extreme temperatures is required then the selected eye protector should be marked with the letter T immediately after the impact letter, i.e. FT, BT, or AT. Otherwise the eye protector should only be worn and used at room temperatures.

**Accessories:** No accessories are available for this product.

**Storage:** Store in a dry and dust free place at room temperature.

**Cleaning:** Clean with clear water and household detergent. (Do not use solvents!) Avoid strong scrubbing as it can cause scratches.

**Disinfection:** Product can be disinfected with all regular commercial disinfectants.

**Disposal:** Pay attention to national regulations when disposing.

**Length of usage:** Do not use longer than 5 years after purchase date.

**Warnings:** Dispose of the goggles immediately if damaged. Only use eye protection lenses with optical class 1. Do not repair damaged safety goggles.

**Limitations of use:** These safety goggles in particular are not suitable for high-speed particles (e.g. cartridge tools), laser beam, temperatures above 55 °C, stray light arches, fusion metals, infectious substances, or organisms.

**Declaration of Conformity:** A Declaration of Conformity according to Regulation (EU) 2016/425 on PPE and the Directive 2001/95/EC on general product safety is available for this product on the following web address: <http://thamesandkosmos.com/declarations/declaration-717019.pdf>

