

MACHEREY-NAGEL

VISOCOLOR® School



- Handbuch
- Manual

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**MACHEREY-NAGEL**

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## 1. Introduction

The origin of water analysis dates back to the 16th to 18th century – among others to the physician Paracelsus, the English chemist Robert Boyle and the German physician Friedrich Hoffmann (noted for “Hoffmann’s anodyne”). They described the first precipitation and color reactions to detect dissolved substances in water. Around 1800 already, the first portable water analysis laboratories were introduced into the market – e.g. by the chemistry professor Johann F. A. Götting, who was actually supported by Johann W. Goethe, as “chemical trial closet”. Water analysis experienced an initial significant rise in the 19th century because of the development of titrimetric and colorimetric methods and due to the activities of still famous Carl Remigius Fresenius in his chemical laboratories in Germany, which were founded in 1848.

In the second half of the 20th century, environmental analysis and therefore also water analysis gained more and more importance due to increasing pollution. In spite of improvements in instrumental analytics, titrimetric and colorimetric analysis methods retained their significance in the laboratory and especially as mobile analysis methods at the point of interest. By now, these methods are standardized (according to DIN-, EN-, and ISO norms) and thus officially accepted. They are used for test strips as well as for colorimetric or photometric analysis and the necessary equipment is packed into portable reagent cases. Today, special reagents enable to specifically determine substances using colored compounds whose color intensity increases together with the concentration.

Water analysis test kits are based both on colorimetric and titrimetric methods. Titrimetrically the concentration is determined by using plastic pipets or by counting drops from drop bottles. Colorimetrically, the colors develop after adding reagents to the vessel containing the sample and are then compared to standardized colors, thus determining the respective concentration. For color comparison, either color scales or disks are used. Using the principle of color comparison (in so called comparators), pre-colored water samples can be analyzed as well. In this case, the comparator includes a water sample without reagent and one water sample with reagents and the respective (color) dye. The color scale includes two rows of colors, one representing the different concentrations and one being white or including a base color. To determine the concentration, the comparator is moved along the color scale, until both colors match. The length of the sample container (= optical path length in photometry) determines the sensitivity of the test. Thus, low concentrations of substances in water samples can be determined with sufficient accuracy.

The fundamental parameters for initially evaluating a water body are the pH value, total hardness and, with regards to possible contaminations, the nitrate, nitrite, ammonium and phosphate content.

## 2. General information

### 2.1 Refill pack and accessories

Description	REF
Refill pack for <i>VISOCOLOR</i> <sup>®</sup> School reagent case including all necessary reagents for ammonium, total hardness, nitrate, nitrite, phosphate and pH	933200
Color chart for <i>VISOCOLOR</i> <sup>®</sup> School reagent case	933300
Measuring glasses with screw caps, 10 pcs	931151
Sliding comparator, 2 pcs	931152
Titration test tube with 5 mL marking	915499
Sample beaker 25 mL	914498
Plastic measuring spoon, 10 pcs, 70 mm	914492
Syringe 5 mL	914661
Manual (german / english) for <i>VISOCOLOR</i> <sup>®</sup> School reagent case	933150

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### 2.2 Additional test papers







The case also has room for one pack of pH-Fix and one pack of QUANTOFIX<sup>®</sup> test strips.

pH-Fix test strips are high quality pH test strips. They are non-bleeding and thus prevent sample contamination. The strips are available in many different gradations.

QUANTOFIX<sup>®</sup> test strips are semi-quantitative and can be evaluated using a color scale. There are test strips for a wide variety of parameters available, e.g. ascorbic acid and iron, etc.

For additional information and how to order, please contact your local distributor, check our website ([www.mn-net.com](http://www.mn-net.com)) or contact MACHEREY-NAGEL directly.

### 2.3 Explanation of symbols

 Use by	 Lot number
 Catalog number	 Storage temperature
 Please read instructions	 Observe the safety precautions in instructions

### 3. Testing procedure

Determinations with *VISOCOLOR*<sup>®</sup> School test kits don't require any previous knowledge and are suited ideally for schools, students and teachers. Both colorimetric and titrimetric methods are used. Results can be read directly in mg/L or ppm (parts per million, mg/L  $\triangleq$  ppm). Generally accepted exemptions are various degrees of hardness (e.g. °e for english hardness or mmol/L).

All *VISOCOLOR*<sup>®</sup> School reagents are approved to be used in schools in Germany according to GUV-SR 2004.

#### 3.1 Colorimetric methods

Colorimetric analysis takes advantage of the properties of special reagents, which form colored compounds with the substance in question. The color intensity rises proportionally to the concentration of the sought-for substance.

All colors are then compared to a set of standard colors. Once colors have been assigned or matched, the result can be read directly.

Colorimetric *VISOCOLOR*<sup>®</sup> School test kits use two tubes, which are both filled with the sample. Reagents are only added to one of the tubes (tube B). Once all reagents have been added and the respective reaction time has elapsed, the comparator is moved along the color chart until the colors of tube A and B match (when looking from above through the samples). The result can be read in the gap on the upper side of the comparator.

#### 3.2 Titrimetric methods

There are some substances which are difficult or even impossible to convert into colored, analyzable compounds. In such cases, often titrimetric methods are employed. During such a volumetric analysis, a reagent, which reacts with the sought-for substance, is added dropwise to a define sample volume. As soon as the substance in the sample is entirely converted, adding more drops of titration solution would result in an excess of reagent. This point of completed reaction (equivalent point) is shown by a previously added indicator via a color change.

To use the titrimetric *VISOCOLOR*<sup>®</sup> School test kit, first an indicator is added to 5 mL of the sample. Then the titration reagent is added drop by drop until the sample changes color. The amount of drops necessary to achieve the color change of the indicator represents the amount of the sought-for substance in the water sample.

#### 3.3 Disposal

Information regarding disposal can be found in the safety data sheet. You can download the SDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

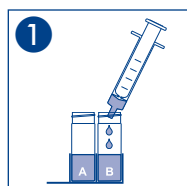
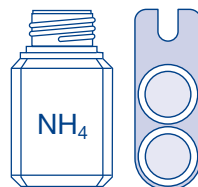
## 4. Ammonium

### Measuring range

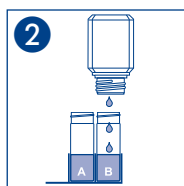
0.2–3 mg/L  $\text{NH}_4^+$

### Test instructions

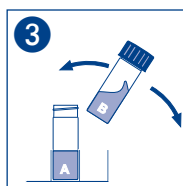
- 1 Fill a **5 mL water sample** into each of the measuring glasses using the plastic syringe. Place a measuring glass on position A in the comparator. Only add the **reagents to measuring glass B**.
- 2 Add **10 drops of  $\text{NH}_4$ -1**.
- 3 Seal the glass and mix.
- 4 Add **1 level measuring spoonful of  $\text{NH}_4$ -2**.
- 5 Seal the glass and shake the mixture until the powder has dissolved.
- 6 Wait for **5 min**.
- 7 Open the glass, add **4 drops of  $\text{NH}_4$ -3**.
- 8 Seal the glass and mix.
- 9 Wait for **7 min**.
- 10 Open the glass, place it in position B of the comparator and slide the comparator along the color scale until the colors match. Check the result in the gap on the upper side of the comparator. Mid-values can be estimated.



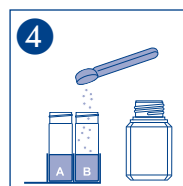
2 x 5 mL sample



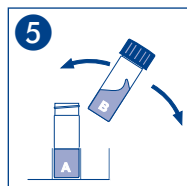
10  $\text{NH}_4$ -1



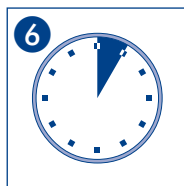
Mix



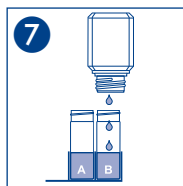
1  $\text{NH}_4$ -2



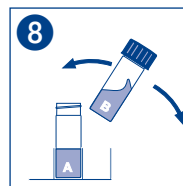
Mix



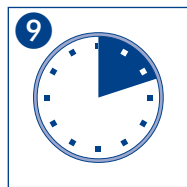
5'00 min



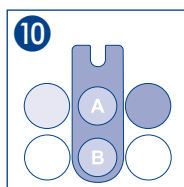
4  $\text{NH}_4$ -3



Mix



7'00 min



Measurement

After use, rinse both measuring glasses thoroughly and seal them.

This method can also be used for analyzing sea water after dilution (1 + 9).

## Background information

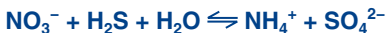
In nature, ammonia is formed due to volcanic activity and electrostatic discharge in the upper ranges of the atmosphere from nitrogen and water vapor. Additionally, ammonia is formed during decomposition processes, which means during the mineralization of nitrogen containing plant and animal proteins. Within the mineral kingdom, almost all magmatic rocks contain small amounts of ammonium salts. Large sources of ammonium chloride have been discovered more than 1000 years ago close to smoldering coal deposits. Close to the volcanoes Vesuvius and Mount Etna, ammonium salts can be detected as well. From these different sources, as well as from automobile and industry exhaust fumes, significant amounts of ammonium salts are fed into the soil as nitrogen fertilizer in the form of rainfall.

Both humans and animals excrete ammonia next to urea as nitrogen compounds in their feces. Ammonia is a cytotoxin for higher organisms, which has to be removed as quickly as possible. As metabolic intermediate, it is formed in brain, muscles, liver and kidneys. In the liver, it immediately reacts with carbon dioxide to urea, in the brain it is rendered harmless by conversion to glutamine. In contrast to ammonia, ammonium salts are not toxic.

Ammonium ions can also form as a result of microbial nitrate reduction. As fertilizer run-off or primary decomposition products of organic compounds within municipal waste water, these ions may end up in natural water bodies. Pure water bodies contain less than 0.1 mg/L ammonium ions, in polluted water however, more than 10 mg/L may be detected. Especially in light of hygienic aspects, the ammonium compound content needs to be evaluated critically, as ammonia may form in the decomposition of human or animal feces. Urea for example is decomposed to ammonia and carbon dioxide or to ammonium and hydrogen carbonate ions, respectively:



An exception are marsh water or special ground water with high iron and manganese content, in which iron sulfide and carbon dioxide form hydrogen sulfide under pressure in the earth, which then reduces nitrate to ammonia.



Ammonia is especially toxic to fish (from 0.5 mg/L) and is present at pH values above 7 due to the pH dependant balance between ammonium ions and ammonia.



At pH 6, the balance is almost exclusively on the side of ammonium, at pH 8, already 4 %, at pH 9 25 % and at pH 10 even 78 % ammonia are present (at a water temperature of 17 °C).

High ammonium contents also substantially strain the oxygen balance of water bodies. During the bacterial oxidation of ammonium to nitrate, oxygen is consumed.



This process is on the one hand an important step in the self-purification, however, on the other hand it may lead to fish-kill due to the oxygen consumption. The so-called nitrification process transforms ammonium salts via nitrite to nitrate. Nitrate again is used by higher plants in complex biosynthesis steps within their roots and leaves to create essential proteins. In biological waste water treatment facilities, monitoring the nitrogen compounds ammonium, nitrate and nitrite plays a critical part, the reason is that here the nitrification steps are the most important aspect of the cleaning process.



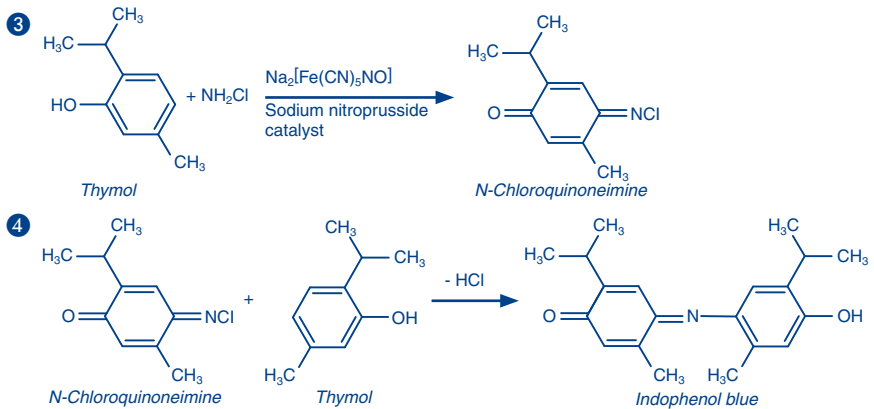
ture. The ammonium concentration in the waste water inflow determines the oxygen demand. The residual ammonium content in the waste water outflow provides information on the facility's effectiveness.

## Reaction basis

Monochloramine is derived from ammonium ions as a result of the effect of chlorine in the alkaline range. Combined with thymol, this forms a blue indophenol dye.

## Reaction equation

- 1  $R-Cl + NaOH \longrightarrow NaOCl + RH$   
(R-Cl = Dichloroisocyanuric acid)  
*Sodium hypochlorite*
- 2  $NH_3 + NaOCl \longrightarrow NaOH + NH_2Cl$   
*Monochloramine*



## Interferences

Primary amines react in the same way as ammonium ions and produce higher results. Depending on their concentration, substances which consume chlorine may reduce the measurement reading or suppress the reaction totally.

## Hazard warning

Information regarding safety can be found on the case's label and in the safety data sheet. You can download the MSDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

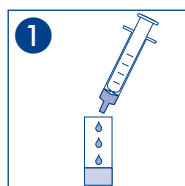
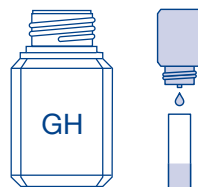
## 5. Total hardness

### Measuring range

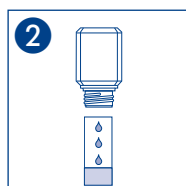
1 drop  $\triangleq$  1.3 °e  $\triangleq$  17.8 mg/L CaCO<sub>3</sub>

### Test instructions

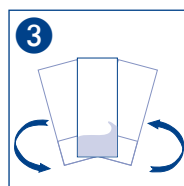
- 1 Fill a 5 mL water sample into the plastic tube using the plastic syringe.
- 2 Add 2 drops of GH-1.
- 3 Shake gently to mix the contents. The water sample turns **red**. If the water sample turns green, this means that there are no hardness-producing substances.
- 4 Hold the dropping bottle **GH-2** absolutely vertical and add the reagent drop by drop, shaking the specimen at the same time to mix until it turns **green**.
- 5 Count the number of drops. One drop corresponds to 1.3 degree of total water hardness (°e).



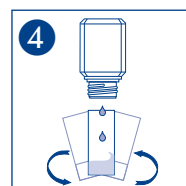
5 mL sample



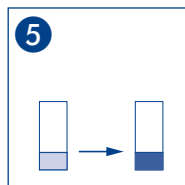
2 drops GH-1



Shake gently



1 drop GH-2 until coloration



1 drop  $\triangleq$  1.3 °e  
red  $\rightarrow$  green

After use, rinse out the plastic tube thoroughly.

Seal the dropping bottles immediately after use. Do not touch the dropping pipettes.

This method can be used also for the analysis of sea water after dilution (1 + 29).

### Background information

The name „hardness“ primarily leads back to the effect of calcium ions to lower the cleaning properties of soap by forming difficult or unsolvable lime soap (calcium salts of higher fatty acids such as palmitic acid). As such hardness components, calcium and magnesium are taken into account together. In natural water bodies, strontium and barium salts can be disregarded due to their very low concentrations (their carbonates and sulfates are hardly soluble).

The sum of calcium and magnesium salts is called *total hardness*. If these alkali earth ions are present as hydrogen carbonates the carbonates precipitate when the water is heated in the form of *limescale*:



Therefore, the water hardness caused by hydrogen carbonates is often referred to as temporary hardness. The salts of other acids (such as hydrochloric or sulfuric acid) remain in solution even when the water is heated. Such hardness is called *permanent hardness* or *sulfate-hardness*.

Depending on the amount of calcium and magnesium salts, water is classified as hard or soft in different gradations:

<b>Very soft:</b>	<b>0 – 3.8 °e</b>
<b>Soft:</b>	<b>5 – 8.8 °e</b>
<b>Medium hard:</b>	<b>10 – 13.8 °e</b>
<b>Fairly hard:</b>	<b>15 °e – 21.3 °e</b>
<b>Hard:</b>	<b>22.5 – 37.5 °e</b>
<b>Very hard:</b>	<b>more than 37.5 °e</b>

Calcium and magnesium ions can be found in almost all natural and unpolluted water bodies. In geochemical processes, carbon dioxide removes hydrogen carbonates from rocks such as dolomite, marble, chalkstone or plaster as hydrogen carbonates (or sulfate from plaster), dissolving it in the water. The water within the catchment area of such rock may have high amounts of hardness, sometimes above 540 mg/L CaCO<sub>3</sub> (more than 37.5 °e). Water from new sandstone areas in contrast contain only small amounts of hardness components. Because of carbon fixation (due to photosynthesis) so called biogenic decalcification may occur in plankton rich water, thus creating high levels of hardness. The following balance applies:



Taken extreme geological conditions into account, generally, water with hardness levels of more than 445 mg/L CaCO<sub>3</sub> (ca. 31.5 °e) are considered to be polluted. Such pollution may for example result due to run-off from landfill sites. Decomposition of plant matter forms carbon dioxide, which seeps away into the ground with rain water. Lime containing soil can thus elute calcium carbonate and calcium hydrogen carbonate (reversal of the limescale equation). Fertilizers may add calcium salts to ground water as well. A certain amount of hardness is desired in drinking water because of two reasons. First, calcium carbonate forms a protective coat inside the plumbing, so that the metal cannot be affected by free, so called aggressive carbon dioxide. Secondly, with respect to the human mineral requirements, an amount of 20 to 60 mg/L calcium is favorable. However, values above 280 mg/L calcium and 125 mg/L magnesium respectively should not be exceeded. Hard water has a fresher taste than soft or even “pure” distilled water. Drinking water of medium hardness levels may cover 10 % of the daily calcium requirement. Conversely, high levels of hardness lead to significant differences in the taste of coffee and tea.

*Carbonate hardness* (see pH value, page 46; interchangeably used with alkalinity) refers to the part of alkaline earth ions present in form of hydrogen carbonate or carbonates. Normally, a proportion of the calcium and magnesium salts are dissolved as sulfates, so that carbonate hardness is always smaller than total hardness. Water bodies are regarded as less fertile, if they contain carbonate hardness with an alkalinity value below 0.5 and as highly fertile with values above 1.5. Very high total hardness together with high alkalinity-values are found in the famous chalk streams in the UK,

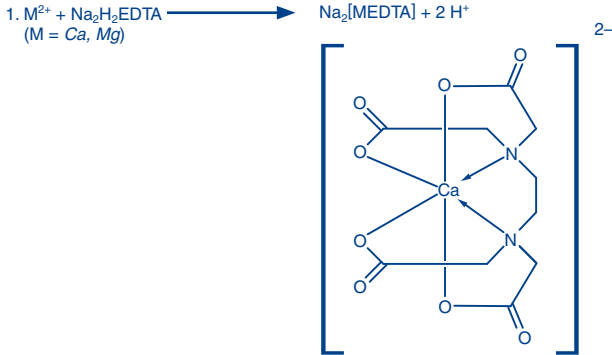
which are among the best salmonid fishing grounds in the world. In rather acidic water with little calcium carbonate, the fish stock is generally low.

## Reaction basis

Complexometric titration

Magnesium and calcium ions, which cause hardness, are combined by the complexing agent EDTA to form chelates. The test is carried out by titration using a metal indicator which changes color when all of the hardness-producing substances have combined (reaction base analog to DIN 38406-3 E3).

## Reaction equation



## Interferences

Copper(II) ions may delay the indicator change or even block this change if higher levels are present. Therefore, in the case of copper pipes, let the water run for a sufficient amount of time before taking the sample.

## Conversion table

°d	°e	°f	mg/L CaO	mg/L CaCO <sub>3</sub>	mmol/L CaCO <sub>3</sub>
1	1.3	1.8	10	18	0.18
2	2.5	3.6	20	36	0.36
3	3.8	5.4	30	54	0.54
4	5.0	7.1	40	71	0.71
5	6.3	8.9	50	89	0.89
6	7.5	10.7	60	107	1.07
7	8.8	12.5	70	125	1.25
8	10.0	14.3	80	143	1.43
9	11.3	16.1	90	161	1.61
10	12.5	17.8	100	178	1.78

## Hazard warnings

Information regarding safety can be found on the case's label and in the safety data sheet. You can download the MSDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

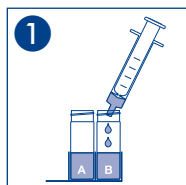
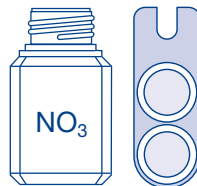
## 6. Nitrate

### Measuring range

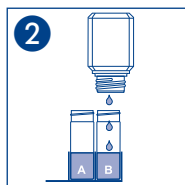
1 – 90 mg/L  $\text{NO}_3^-$

### Test instructions

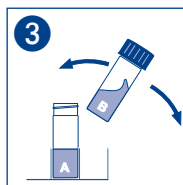
- 1 Fill a **5 mL water sample** into each of the measuring glasses using the plastic syringe. Place a measuring glass on position A in the comparator. Only add the **reagents to measuring glass B**.
- 2 Add **5 drops** of  $\text{NO}_3\text{-1}$ .
- 3 Seal the glass and mix.
- 4 Add **1 level measuring spoonful** of  $\text{NO}_3\text{-2}$ .
- 5 Seal the glass and **immediately shake** the mixture **well for 1 min**.
- 6 Wait for **5 min**.
- 7 Open the glass, place it in position B of the comparator and slide the comparator along the color scale until the colors match. Check the result in the gap on the upper side of the comparator.



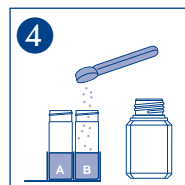
2 x 5 mL sample



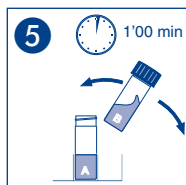
5 drops  $\text{NO}_3\text{-1}$



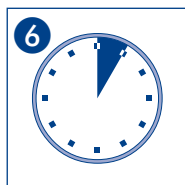
Mix



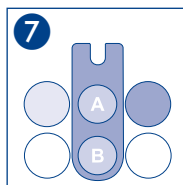
1 spoonful  $\text{NO}_3\text{-2}$



Shake well



5'00 min



Measurement

After use, rinse both measuring glasses thoroughly and seal them.

This method can be used also for analyzing sea water (see conversion table).

### Background information

Sodium nitrate is the most important, naturally present nitrate. The main deposit is in the dry areas of northern Chile. Therefore, it is sometimes referred to as *Chile salt-peter*. Additional small deposits are known in Egypt, Colombia and California. Due to weathering of potassium rich soils, potassium nitrate mainly occurs in India, China and Egypt. Formation of nitrates can be explained mainly due to the conversion of proteins (e.g. from dead sea algae in the case of sodium nitrate) as a result from microbial nitrification. A special formation is calcium nitrate (e.g. saltpeter from masonry). From decomposing proteins, bacteria form in oxidizing processes first ammonia and then nitric acid. Nitric acid then forms together with the calcium carbonate within the

masonry calcium nitrate. A group of plants that indicate nitrate in soil and, to some extent, can store nitrate very well are called nitrate plants (nitrophilous). Examples for nitrophilous plants are willowherbs, nettles, hogweeds and cow parsley, which occur in large populations in wet lands or tilled soil with high nitrate content.

Some plants for human consumption also have a high nitrate content, e.g. spinach, soybeans, chard, beetroot and radish. Nitrate itself is rather non-toxic for humans and animals alike. The danger posed by nitrates develops in the human body, specifically in the digestive system (in saliva already); where bacteria can form nitrites from nitrate (see Nitrite, page 40).

In natural and uncontaminated water, the nitrate ion content is between 0.4 to 8 mg/L, the nitrite ion content at most 0.01 mg/L. From fertilizers (saltpeter) as well as from the aerobic decomposition of nitrogen compounds (e.g. proteins), large amounts of nitrate can get into the water.

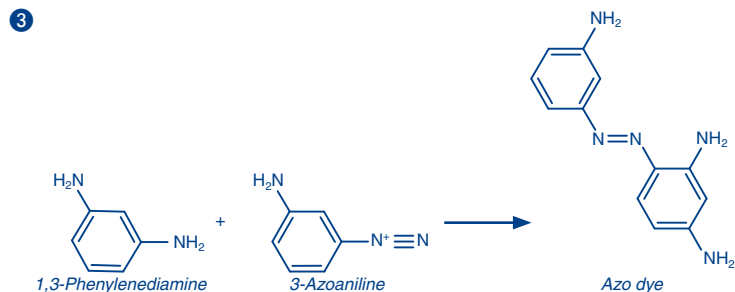
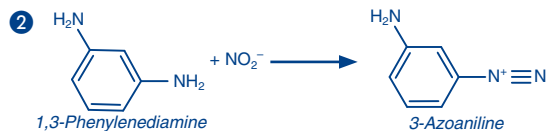
In polluted water bodies, nitrate contents can range from 50 to 150 mg/L, even higher values are possible. If such high nitrate contents cannot be traced back to geological saltpeter deposits (especially when concentrations in ground water are high), a contamination is always at hand.

To evaluate the self-purification capacity of a water body, it is important to test if there are high amounts of ammonium and nitrite present as well. If these other parameters are not high, the self-purification capacity to mineralize organic compounds is sufficient. This relation also holds true for the outflow of biological waste water treatment facilities. In drinking water, 50 mg/L nitrate are set as the limit by WHO, European Union and the German water authorities alike.

## Reaction basis

Nitrate ions are reduced to nitrite ions in an acidic medium. Combined with a suitable aromatic amine, an orange-yellow azo dye is formed.

## Reaction equation



## Interferences

Depending on their concentration, oxidizing substances may reduce the measurement reading or suppress the reaction completely. Chlorine  $\leq 10$  mg/L does not interfere. Nitrite interferes (same reaction). This can be circumvented by addition of amido sulphonic acid (REF 918973) prior to analysis.

The water sample should be between 18 and 30 °C. At lower temperatures the reaction takes place at a significantly slower rate, and the results are limited.

## Conversion table

mg/L NO <sub>3</sub> <sup>-</sup>	mg/L NO <sub>3</sub> -N (Nitrate-nitrogen)	mmol/m <sup>3</sup> NO <sub>3</sub> <sup>-</sup>	mg/L NO <sub>3</sub> <sup>-</sup> in sea water
1	0.2	16	1
5	1.1	81	5
10	2.3	160	12
20	4.5	320	25
50	11	810	65
90	20	1450	120

## Hazard warnings

Information regarding safety can be found on the case's label and in the safety data sheet. You can download the MSDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

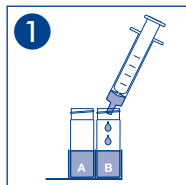
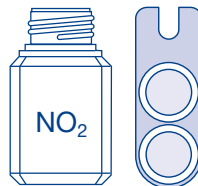
## 7. Nitrite

### Measuring range

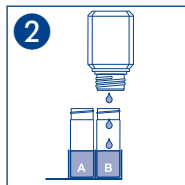
0.02–0.5 mg/L  $\text{NO}_2^-$

### Test instructions

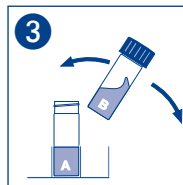
- 1 Fill a **5 mL water sample** into each of the measuring glasses using the plastic syringe. Place a measuring glass on position A in the comparator. Only add the **reagents to measuring glass B**.
- 2 Add **4 drops** of  **$\text{NO}_2-1$** .
- 3 Seal the glass and mix.
- 4 Add **1 level measuring spoonful of  $\text{NO}_2-2$** .
- 5 Seal the glass and shake the mixture until the powder has dissolved.
- 6 Wait for **10 min**.
- 7 Open the glass, place it in position B of the comparator and slide the comparator along the glass, place it in position B of the comparator and slide the comparator along the glass, place it in position B of the comparator and slide the comparator along the glass, place it in position B of the comparator and slide the comparator along the glass until the colors match. Check the result in the gap on the upper side of the comparator. Mid-values can be estimated.



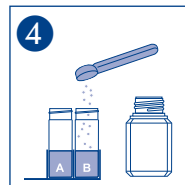
2 x 5 mL sample



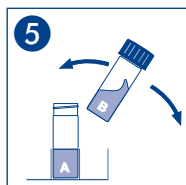
4  $\text{NO}_2-1$



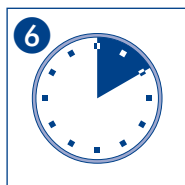
Mix



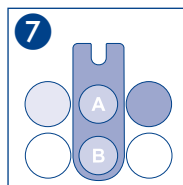
1  $\text{NO}_2-2$



Mix



10'00 min



Measurement

After use, rinse both measuring glasses thoroughly and seal them.  
This method can be used also for analyzing sea water.

### Background information

Adults convert nitrate to nitrite in the small intestines, babies however do so in the stomach already, as their stomach acid production is not yet fully developed. Stomach acid prevents, among others, nitrate reducing bacteria to colonize the small intestines. In the blood, nitrite oxidizes hemoglobin, which is responsible for transporting essential oxygen. Especially for children and babies life threatening conditions with acute danger of suffocation may occur.

Nitrite forms as an intermediate during ammonium oxidation (see Ammonium, page 31) or nitrate reduction. It is highly toxic to fish and presents a pre-stage of carcinogenic *N*-nitroso compounds (e.g. reaction with amines to nitrosamines).



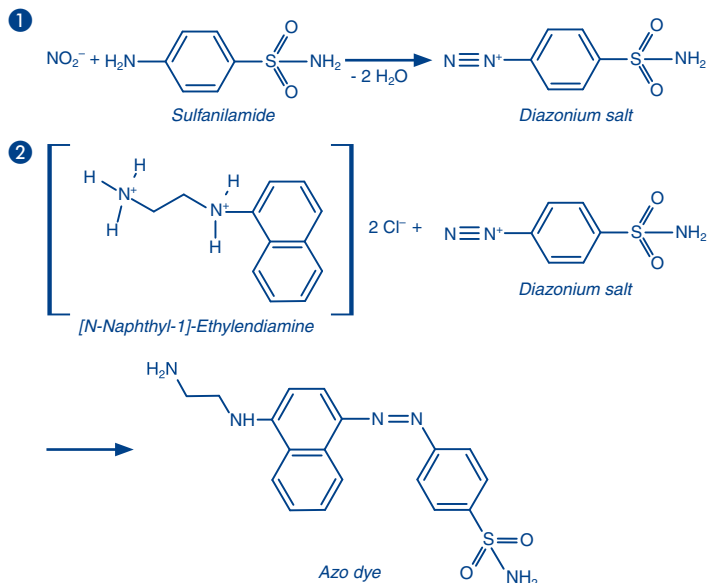
The nitrite ion concentration in ground and surface water is low. Waste water can contain higher amounts of nitrite, which may be caused either by industrial waste water from the metal or chemical industry, or due to the contamination with feces. Another source of nitrite from nitrate may be zinc-plated iron plumbing from domestic pipe systems. Concentrations of up to 1 mg/L are regarded as non-hazardous.

The German drinking water regulations demand a limit of 0.1 mg/L at the water works outflow and 0.5 mg/L at the tap.

## Reaction basis

Sulphanilamide is diazotized by nitrite in acidic solution. The diazonium salt is coupled with a naphthylamine to form a reddish-violet azo dye.

## Reaction equation



## Interferences

Chromium(VI) and iron(III) ions present in concentrations of > 3 mg/L simulate nitrite values which are too high. Chlorine interferes even in minor concentrations.

## Conversion table

mg/L NO <sub>2</sub> <sup>-</sup>	mg/L NO <sub>2</sub> -N (Nitrite-nitrogen)
0.02	0.006
0.05	0.015
0.1	0.03
0.2	0.06
0.5	0.15

## Hazard warnings

Information regarding safety can be found on the case's label and in the safety data sheet. You can download the MSDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

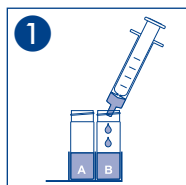
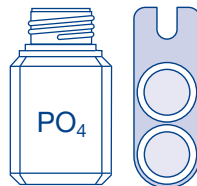
## 8. Phosphate

### Measuring range

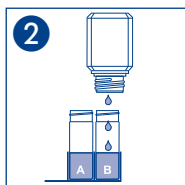
0.6 – 15 mg/L  $\text{PO}_4^{3-}$

### Test instructions

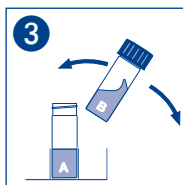
- 1 Fill a **5 mL water sample** into each of the measuring glasses using the plastic syringe. Place a measuring glass on position A in the comparator. Only add the **reagents to measuring glass B**.
- 2 Add **6 drops of  $\text{PO}_4$ -1**.
- 3 Seal the glass and mix.
- 4 Add **6 drops of  $\text{PO}_4$ -2**.
- 5 Seal the glass and mix.
- 6 Wait for **10 min**.
- 7 Open the glass, place it in position B of the comparator and slide the comparator along the color scale until the colors match. Check the result in the gap on the upper side of the comparator.



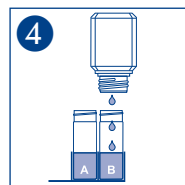
2 x 5 mL sample



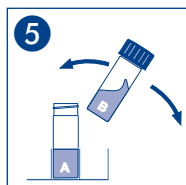
6  $\text{PO}_4$ -1



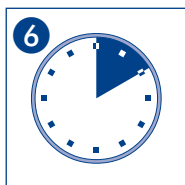
Mix



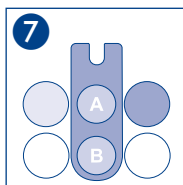
6  $\text{PO}_4$ -2



Mix



10'00 min



Measurement

After use, rinse both measuring glasses thoroughly and seal them. This technique can also be used for analyzing sea water.

### Background information

Clean water bodies, especially in mountains, have phosphate contents of less than 0.1 mg/L, often even below 0.03 mg/L. If concentrations are higher than 0.1 mg/L  $\text{PO}_4^{3-}$ , contamination is only present if other contamination indicators are positive. Phosphate concentrations exceeding 0.3 mg/L provide strong suspicion for contamination. An exception is marsh water, which may contain up to 1 mg/L.

Because of contamination from domestic waste water (kitchen waste or especially from feces), large amounts of phosphate get into the water. In the case of fecal contamination, high phosphate contents are a reliable chemical indicator.

Chemical fertilizers raise the phosphate content in ground water. Some soils are capable of absorbing phosphate, so contaminated and deep running ground water may exhibit normal amounts as well. Moreover, high phosphate concentrations may have a geological cause, but would then be on a constant level (up to 7 mg/L). In case of a contamination, the phosphate concentration fluctuates strongly over time.

For plumbing and pipe systems used for aggressive water, phosphate is important in creating a protective layer, however, concentrations of at least 0.1 mg/L (up to about 1 mg/L) are sufficient. In cooling and boiler feed water, exactly dosed phosphate additions prevent boiler scale, as they link to calcium ions.

The German drinking water regulation from 2001 does not define a limit for phosphate anymore. Increased phosphate levels in water bodies lead to eutrophication (over-fertilization), which means an intensified growth of aquatic plants, especially algae, as a result from the large nutrient supply. Due to excessive bacterial decomposition of dead algae, this process, also known as water-bloom, leads to an increase in oxygen consumption. If as a result anaerobic conditions occur, hydrogen sulfide may be released. With regards to elemental phosphorous, there is a distinction between inorganic phosphorous (phosphate and hydrogenphosphate ions and polyphosphates) and organically bound phosphorous, which is only released after digestion (decomposition) of the organic material.

For humans, phosphate plays an important role in bone development and energy metabolism. In drinking water high phosphate concentrations are undesirable though, as they can lead to digestive disorder.

## Reaction basis

Ammonium molybdate and phosphate ions form phosphomolybdic acid, which is reduced to phosphomolybdenum blue.

(Reaction basis analog to DIN EN ISO 6878-D11).

## Reaction equation



## Interferences

Larger amounts of oxidizing reagents inhibit formation of the blue color complex and have to be destroyed. H<sub>2</sub>S interferes in concentrations above 2 mg/L, but can be expelled after acidification of the water sample. Heavy metals in excess of 10 mg/L can slightly decrease the intensity of the color (vanadium causes an increase in color). Silica interferes in excess of 10 mg/L Si.

### Conversion table

mg/L PO <sub>4</sub> <sup>3-</sup>	mg/L PO <sub>4</sub> -P (Phosphate-phosphorous)	mg/L P <sub>2</sub> O <sub>5</sub>
0.6	0.2	0.5
0.9	0.3	0.7
1.5	0.5	1.1
2.1	0.7	1.6
3	1	2
6	2	5
9	3	7
15	5	12

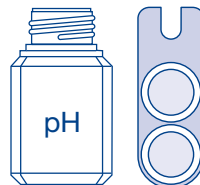
### Hazard warnings

Information regarding safety can be found on the case's label and in the safety data sheet. You can download the MSDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

## 9. pH value

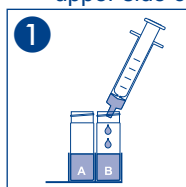
### Measuring range

pH 4.0–9.0

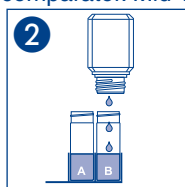


### Test instructions

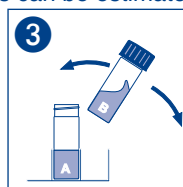
- 1 Fill a **5 mL water sample** into each of the measuring glasses using the plastic syringe. Place a measuring glass on position A in the comparator. Only add the **reagent to measuring glass B**.
- 2 Add **4 drops** of pH-1.
- 3 Seal the glass and mix.
- 4 Open the glass, place it in position B of the comparator and slide the comparator along the color scale until the colors match. Check the result in the gap on the upper side of the comparator. Mid-values can be estimated.



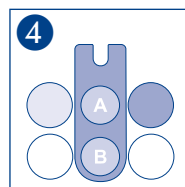
2 × 5 mL sample



4 drops pH-1



Mix



Measurement

After use, rinse both measuring glasses thoroughly and seal them. This method can be used also for analyzing sea water.

### Background information

At room temperature, pure water with the formula  $\text{H}_2\text{O}$ , which contains solved traces of other substances, is to a small degree divided (dissociate) into hydrated hydrogen ions – known as hydronium ions –  $\text{H}_3\text{O}^+$  and the counter ion  $\text{OH}^-$  – the hydroxide ion:



1 liter of water contains 0.000001 mole ( $10^{-7}$  mole; given the hydrogen atom weight of 1 g/mol also the same amount in gram) hydrogen ions  $\text{H}^+$  and at the same time 0.0000001 mole (multiplied by 17 g/mol as mole weight for  $\text{OH}^-$  in gram) of the hydroxide ion. To simplify notation of these low concentrations, the pH value only indicates the exponent, in this case the number 7 (with positive algebraic sign). Mathematically, the pH value is defined as the negative common logarithm:  $\text{pH} = -\lg [\text{H}_3\text{O}^+]$ . Thus, a pH value of 7 at room temperature means that the concentration of hydrogen ions corresponds to the concentration from the dissociation of pure water (in balance with hydroxide ions). If common salt (sodium chloride  $\text{NaCl}$ ) is dissolved in water, the pH value does not (or hardly) change, as no new hydrogen ions are generated or used up (bound). In water, common salt breaks down into sodium and chloride ions. Therefore, the water is still neutral, even after dissolving common salt. If a strong acid, such as hydrochloric acid ( $\text{HCl}$ ), is dissolved in distilled water in very low quantities, the amount of hydrogen ions increases (coming from the hydrochloric acid). As a strong acid, hydrochloric acid is almost completely broken down into hydrogen and chloride ions. The resulting increase in hydrogen or hydronium ions respectively is shown in the decreasing pH value. To keep the above mentioned water equation in balance, the

concentration of hydroxide ions has to decrease. Therefore, the sum of the exponent of hydrogen and hydroxide ion concentration is always 14.

A one unit decrease in pH value signifies an increase in the hydrogen ion concentration by the power of ten in comparison to the concentration before. Simultaneously, the hydroxide ion concentration decreases by the power of ten.

Weak acids, such as acetic acid are not entirely broken down into their ions. At the same mole concentration (with regards to the hydrogen within the molecule), weak acids will not decrease the pH value as much as strong acids, such as hydrochloric or sulfuric acid. Bases feature the same processes, but in the other direction. The hydroxide ion is released, therefore the hydrogen ion concentration decreases and the pH value is raised above 7.

Compounds reacting with water may cause a change in pH value as well, even if they cannot provide hydrogen or hydroxide ions in the water themselves. If carbon dioxide is dissolved in water, small, but measurable amounts of hydrogen carbonate and hydrogen ions form according to the following equation:



Water and carbon dioxide create one hydronium and one hydrogen carbonate ion each. The pH value of water, which always contains some dissolved carbon dioxide due to contact with air, is therefore always below 7. The solution becomes acidic.

Salts dissolved in water can also cause a change in pH value. If sodium carbonate (soda) is dissolved in water, the following reaction occurs between water and carbonate ions:



As hydroxide ions are formed, the pH value rises above 7. The solution becomes alkaline.

In this context, the concept of acid binding capacity or alkalinity plays an important role in hydro chemistry. Carbonate alkalinity is based on the fact that carbonate ions can bind hydronium ions, thus having a buffering effect. The so called buffer capacity of natural water is therefore mainly determined by the amount of dissolved carbon dioxide, hydrogen carbonate and carbonates (see Total hardness, page 34).



If ammonium chloride is dissolved in water, the following reaction between ammonium ions and water molecules sets in:



The ammonium ion releases one hydrogen ion to the water, a hydronium ion is created. Therefore, the pH value decreases and the solution becomes acidic.

Carbonic acids (from carbon dioxide in water) as well as ammonia are both a weak acid and a weak base, respectively. Therefore, salts react with strong bases or strong acids not neutrally, but either alkaline (weak acid within the salt) or acidic (weak base within the salt). Such reactions with water are called hydrolysis.

The pH scale ranges from 0 to 14. At the lower end are the strong mineral acids, such as hydrochloric, sulfuric and nitric acid. On the upper end of the scale are caustic soda and caustic potash. Starting with diluted hydrochloric acid (concentration 1.0 mol/L = 36.5 g/L = 3.65 %) and a pH value of 0, for acids and bases as well as for everyday products the following pH scale can be devised:

<b>pH 0:</b>	hydrochloric acid – 3.65 %
<b>pH 0.9 – 1.5:</b>	stomach acid (also diluted hydrochloric acid)
<b>pH 2.3:</b>	lemon juice
<b>pH 3.1:</b>	vinegar
<b>pH 3.2 – 4.6:</b>	acidic vegetables
<b>pH 4.5:</b>	beer
<b>pH 7:</b>	pure water
<b>pH 8.3:</b>	sea water
<b>pH 8 – 10:</b>	soap water
<b>pH 12.3:</b>	saturated lime water
<b>pH 14:</b>	caustic soda – 4.0 %

Determining the pH value of a water body gives initial and basic indications as to water quality in general. The pH value allows evaluation of the degree of aggressiveness on construction material. Moreover and especially, the pH value gives an indication of the water's effect on plant and animal life in water bodies or even in waste water treatment plants.

Fish can only tolerate a certain pH range. At the limits, both at the low and high end, skin and gills are negatively affected. The limits for carp for example are at pH 4.5 in the acidic and 10.8 in the alkaline range. Living for a prolonged time at these limits will lead to the fishes' death. The brown trout for example can only accept a more narrow range between pH 5.5 and 9.4. Ideal fish water has a pH value between 6.5 and 8.

Additionally, sewage networks as well as iron or concrete pipes are negatively affected at pH values below 5 and above 10.

Acids in the form of rain water (e.g. from nitric, sulfuric or hydrochloric acid precipitation) or in the soil and bases (e.g. from the decomposition of organic material or caustic soda from industrial effluents) may change the pH value of water. Additionally, salts formed from both weak acids and strong bases (sodium carbonate) or from weak bases and strong acids (calcium chloride) may alter the pH value due to hydrolysis (molecule reaction with water s.a.). So, calcium and magnesium chloride lower the pH value as they create an acidic reaction. Similarly reactions may occur due to iron and aluminum salts from sulfuric, hydrochloric or nitric acid, which may originate from metal processing industry effluents. Carbonates of sodium, potassium, magnesium as well as calcium provide alkaline reactions. They can stem from rock formations or domestic waste water.

Normally, water bodies show a pH value between 6.7 and 7.5. Deviations from pH 7 are mainly caused by carbon dioxide (see above). In form of carbonate, carbon dioxide increases, in its free form as carbonic acid, it decreases the pH value. Marsh water shows pH values in the acidic range due to organic acids. Domestic waste water is normally neutral to alkaline, whereas industrial waste water is most often acidic, e.g. etching wastes from iron processing. If the pH value is below 5, concrete is damaged. Already at a pH value of 5.5, the biological cleaning stage in waste water treatment plants is disrupted, at pH 6 or 8 the cleaning stage is at least affected. Drinking water



should have a pH value between 6.5 and 8.5 (the German drinking regulations set 9.5 as the maximum pH value).

### **Reaction basis**

A special mixture of indicator dyes produces a specific and characteristic color for every pH value covered.

### **Interferences**

The favorable ratio between indicator and sample minimizes the indicator error. This means that perfect measuring results are ensured even for weakly buffered samples. High concentrations of neutral salts and colloids as well as organic solvent contents above 10 % can cause wrong results.

Notice: If a wider measuring range, or tighter gradations are necessary, pH-Fix test strips are the ideal solution. The strips come in a wide variety of ranges and gradations and the case has room for one package. For a detailed list of options check our catalog or go online to [www.mn-net.com](http://www.mn-net.com).

### **Hazard warnings**

Information regarding safety can be found on the case's label and in the safety data sheet. You can download the MSDS from [www.mn-net.com/SDS](http://www.mn-net.com/SDS).

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