TU Dresden **Institute of Hydrochemistry**	**Instruction I** **Determination of the Acid- und Base capacity**	**I 1** **2005**

Practical Course: Basics of Hydrochemistry

Instruments:
Erlenmeyer flasks (250 mL, 300 mL)
transfer pipettes (100 mL)
burette (50 mL)

Chemicals:
hydrochloric acid-solution c = 0,1 mol/L
sodium hydroxide-solution c = 0,1 mol/L
methyl orange-solution (indicator)
phenolphthalein-solution (indicator)

1. Determination of the acid capacity $K_{S\ 4,3}$ (+m-Wert)

Completion:
- 3 drops of methyl orange solution are added to 100 mL of the sample,
- this solution is titrated with a 0,1 molar hydrochloric acid solution till you observe a change in color (yellow to light orange)

Calculation:
The acid capacity of the sample should be calculated of the following equation:

$$K_{S\ 4,3} = \frac{V_1 * c_1 * f}{V_P}$$

here mean:

$K_{S\ 4,3}$	acid capacity till a pH-value of 4,3 in mmol/L This value corresponds to the hydrogen carbonate concentration in natural waters
V_1	volume of the spent hydrochloric acid-solution in mL
V_P	volume of the undiluted water sample in mL
c_1	concentration of the hydrochloric acid-solution in mol/L
f	conversion factor (f = 1000 mmol/mol)

2. Determination of the base capacity $K_{B\ 8,2}$ (-p-Wert)

Completion:

- 0,2 mL of phenolphtalein solution are added to 100 mL of the sample (pH-value has to be under 8,2!!),
- this solution is titrated with a 0,1 molar sodium hydroxide solution till you observe a change in color (colourless to violet)

Calculation:
The base capacity of the sample should be calculated by the follwing equation:

$$K_{B\ 8,2} = \frac{V_2 * c_2 * f}{V_P}$$

here mean:

$K_{B\ 8,2}$	base capacity till a pH-value of 8,2 in mmol/L
V_2	volume of the spent sodium hydroxide-solution in mL
V_P	volume of the undiluted water sample in mL
c_2	concentration of the sodium hydroxide-solution in mol/L
f	conversion factor (f = 1000 mmol/mol)

TU Dresden **Institute of Hydrochemistry**	**Instruction II** **Determination of Calcium- und Magnesium-Concentration by DIN 38 406 part 3**	**I 2** **2005**
Practical Course: Basics of Hydrochemistry		

1. Complexometric determination of Calcium

Instruments:
Erlenmeyer flasks (250 mL)
transfer pipettes (50 mL, 100 mL)
burette (50 mL)

Chemicals:
sodium hydroxide-solution c = 2 mol/L
EDTA-solution   c = 0,01 mol/L
calconcarbonic acid-indicator

Preparation of the sample:
The sample should be contained between 2 and 100 mg/L calcium corresponding 0,05 und 2,5 mmol/. If a concentration above 100 mg/L is expected, the sample should be diluted with 50 mL of pure water. If concentrations are expected under 2 mg/L higher volumes of the pure sample should be used.

Completion:
2 mL of the sodium hydroxide-solution and about 0,2 g calconcarbonic acid-indicator (a tip of a spatula) should be added to 50 mL of the water sample.

This solution is titrated immediately with a 0,01 molar EDTA-solution till you observe a change in color (red to blue)

Evaluation:

The concentration of calcium should be calculated by the following equation:

$$c_{Ca} = \frac{V_E * c_E * f}{V_P}$$

here mean:

c_{Ca}	concentration of calcium-Ions in mmol/L
V_E	volume of the spent EDTA-solution in mL
V_P	volume of the undiluted water sample in mL
c_E	concentration of the EDTA-solution in mol/L
f	conversion factor (f = 1000 mmol/mol)
	(1 mmol/L Ca^{2+} = 40,08 mg/L Ca^{2+})

2. Complexometric determination of the total concentration of Calcium and Magnesium

Instruments:
Erlenmeyer flasks (250 mL)
transfer pipettes (50 mL, 100 mL)
burette (50 mL)

Chemicals
- EDTA-solution c = 0,01 mol/L
- eriochrome black-solution (indicator)
- buffer-solution, ammonium chloride / ammonia-solution pH = 10

Preparation of the sample:
If a total concentration above 4 mmol/L is expected, the sample should be diluted with of pure water.

Completion:
The titration should be performed very fast to avoid faults by origin of carbonates. Therefor the approximate consumption should be determined by preliminary titration .

Preliminary titration:
4 mL of the buffer-solution (pH- value 10) and 3 drops of eriochrome black-indikator should be added to 50 mL of the water sample.
This solution is titrated (in the beginning faster, at the end slower) with a 0,01 molar EDTA-solution till you observe a change in color (violet to blue). The end point is reached when the last red glimmer is disappearing.
If the consumption of the EDTA-solution is bigger then 20 mL, the titration should be repeated by a smaller volume. Therefor the sample is diluted with 50 mL of pure water.
If the consumption of the EDTA-solution is smaller then 4,5 mL, the titration should be repeated with a bigger volume (100 mL) of the sample.

Main titration:
0,5 mL below the consumed EDTA-solvent of the preliminary titration should be added at first to 50 mL of the water sample. After that 4 mL buffer-solution and 3 drops eriochrome black-solution should be added to the sample. The equivalence point is reached by an observed colour change (violet to blue).
If the difference of consumption between the two titrations is more then 0,1 mL, a third tiration should be performed. is more then

Evaluation:
The total concentration of magnesium- and calcium ions in the sample should be calculated by the following equation:

$$c = \frac{V_E * c_E * f}{V_P}$$

Here mean:
c total concentration of Ca- and Mg-ions in mmol/L
V_E volume of the spent EDTA-solution in mL
V_P volume of the undiluted water sample in mL
c_E concentration of the EDTA-solution in mol/L

f conversion factor (f = 1000 mmol/mol)
(1 mmol/l Ca^{2+} = 40,08 mg/L Ca^{2+}, 1 mmol/l Mg^{2+} = 24,31 mg/L Mg^{2+}

3. Calculation of the magnesium concentration

The magnesium concentration has to be calculated by the difference of the measured total concentration of magnesium and calcium (2) and the concentration of the measured concentration of calcium (1).