Ion-selective polymeric membrane electrodes for potentiometric trace level measurements

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Ion-Selective Polymeric Membrane Electrodes for Potentiometric Trace Level Measurements

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Presented by

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1. SUMMARY

The present thesis describes methods to improve the performance of polymer membrane ion-selective electrodes (ISEs). By avoiding biases through uncontrolled transmembrane ion fluxes, both the lower detection limit and the selectivity values are massively improved so that ISEs can now be used for environmental trace metal analysis. The fundamental membrane processes have been investigated with Pb\(^{2+}\) - and Ca\(^{2+}\)-ISEs, which were based on plasticized poly(vinyl chloride) membranes and commercially available ionophores and lipophilic ionic sites.

In the first part of the work, the complex formation of a series of available Pb\(^{2+}\) ionophores with many cations has been studied in the polymeric membrane phase. It was found that most of the determined selectivities were far better than the ones listed in earlier publications. Surprisingly, amide ionophores form stronger Pb\(^{2+}\) complexes than the respective thioamides, but membranes based on the latter exhibit better selectivities toward alkali and alkaline earth metals. On the other hand, thioamides suffer from severe interferences by Hg\(^{2+}\), Ag\(^{+}\), Cu\(^{2+}\) and Cd\(^{2+}\). Based on these studies, a thioamide functionalized calix[4]arene was selected as the ionophore possessing the best selectivity for Pb\(^{2+}\), especially over alkali and alkaline earth metals.

Subsequently, the influence of the inner filling solution and various membrane parameters, including thickness, viscosity, and concentration of ionic sites on the ISE response has been investigated. Theoretical models predict that zero-current transmembrane ion-fluxes can considerably influence the composition of the sample in the sensed Nernstian layer and, as a consequence, the lower detection limit of the ISEs. The model predictions have been verified and, at least qualitatively, confirmed for Ca\(^{2+}\) - and Pb\(^{2+}\)-ISEs. It has been demonstrated that with appropriate compositions of
membrane and inner filling solution, the extent of transmembrane ion fluxes
can be reduced to obtain lower detection limits that are, with \( \sim 10^{-9} \) M, by at
least three orders of magnitude better than the earlier reported values.

In the last part of this work, the Pb\(^{2+}\)-ISE and the measuring procedure were
optimized in order to achieve rugged sensors with lower detection limit
values that are at least 10 times smaller than the action level of the U.S.
Environmental Protection Agency (15 \( \mu \)g/L in drinking water). A large
number of native and spiked natural water samples, containing total Pb\(^{2+}\)
concentrations varying from \( 10^{-8} \) to \( 10^{-10} \) M, were acidified to pH 4 and
potentiometrically assessed for Pb\(^{2+}\). For lead concentrations above \( 10^{-9} \) M,
the values compare well with ICPMS measurements. Below \( 10^{-9} \) M Pb\(^{2+}\), the
ISE response is dictated by H\(^{+}\) or Cu\(^{2+}\) interference. The sensors are also
shown to be applicable for the speciation analysis of lead. The obtained
results clearly demonstrate that polymer membrane electrodes are useful
analytical tools for potentiometric heavy metal ion determination in
environmental samples at nanomolar levels.
2. ZUSAMMENFASSUNG

Die vorliegende Arbeit beschreibt verschiedene Methoden zur Verbesserung des Ansprechverhaltens ionenselektiver Flüssigmembranelektroden (ISE). Es wird gezeigt, dass durch die Reduktion unkontrollierter Ionenflüsse innerhalb der Membran sowohl die untere Nachweisgrenze als auch das Selektivitätsverhalten der ISE verbessert werden können. Die beschriebenen Methoden ermöglichten damit den Einsatz dieser Sensoren zur Bestimmung von Schwermetallionen in der Umweltanalytik. Die Membranprozesse wurden dabei mit Hilfe von Pb\textsuperscript{2+} und Ca\textsuperscript{2+}-selektiven ISE untersucht, deren Membranen auf Poly(vinylchlorid) mit Weichmachern und käuflichen Ionophoren und lipophilen Salzen basierten.

Im ersten Teil der Arbeit wurde die Komplexbildung zahlreicher Pb\textsuperscript{2+}-selektiver Liganden mit verschiedenen Kationen in der Membranphase untersucht. Diese Experimente zeigten, dass die meisten Selektivitäten bedeutend besser sind, als die Werte, die bisher in der Literatur beschrieben wurden. Überraschenderweise wurde festgestellt, dass Ionophore mit Amidfunktionen bedeutend stärkere Pb\textsuperscript{2+}-Komplexe bilden als die entsprechenden Thioamidverbindungen. Einerseits zeigten Membranen basierend auf Ionophoren mit Thioamidgruppen zwar bessere Selektivitäten gegenüber Alkali- und Erdalkalimetallionen. Andererseits wurden aber für die thioamidhaltigen Liganden starke Interferenzen durch Hg\textsuperscript{2+}, Ag\textsuperscript{+}, Cu\textsuperscript{2+} und Cd\textsuperscript{2+} beobachtet. Ausgehend von diesen Untersuchungen, wählte man dasjenige Thioamid-Calix[4]aren als Pb\textsuperscript{2+}-Ionophor für die weiteren Experimente aus, welches die besten Selektivitäten gegenüber Alkali- und Erdalkalimetallionen zeigte.

Man untersuchte zudem den Einfluss der Innenlösung und zahlreicher Membranparameter wie Dicke, Viskosität und Konzentration der anionischen Additive auf das Ansprechverhalten der ISE. Theoretische Modelle sagten

Im letzten Teil der Arbeit wurden die entwickelten Pb\(^{2+}\)-selektiven ISE und das Messverfahren optimiert. Es wurden ISE hergestellt, die bei Messungen von Umweltproben untere Nachweisgrenzen zeigten, die mindestens zehnmal kleiner waren, als der von der Amerikanischen Umweltschutzbehörde vorgeschlagene Grenzwert (15 µg/L in Trinkwasser). Zahlreiche Trinkwasser- und Umweltproben mit Bleikonzentrationen im Bereich von 10\(^{-8}\)–10\(^{-10}\) M wurden angesäuert (pH 4) und anschließend mit ISE bezüglich der Pb\(^{2+}\)-Konzentration analysiert. Dabei zeigten die mit ISE gemessenen Pb\(^{2+}\)-Konzentrationen oberhalb von 10\(^{-9}\) M eine gute Übereinstimmung mit den Messresultaten, die mittels ICPMS ermittelt wurden. Lag die Pb\(^{2+}\)-Konzentration unterhalb von 10\(^{-9}\) M, zeigten die ISE dagegen H\(^{+}\)- bzw. Cu\(^{2+}\)-Interferenz. Die erzielten Resultate zeigen eindeutig, dass die in dieser Arbeit entwickelten ISE für die Schwermetallbestimmung in der Umweltanalytik eingesetzt werden können.
3. ION-SELECTIVE ELECTRODES

Ion-selective electrodes (ISEs) are electrochemical sensors that allow the potentiometric determination of the activity of an ionic species in the presence of other kinds of ions [1].

3.1 Introduction

A chemical sensor can be defined as a small device that allows the transformation of chemical information into an optical or electrical signal that can be processed by an instrument. ISEs are potentiometric ion sensors based on ion-selective membranes. They respond to the activity of the analyte ion, whose logarithmic value is proportional to the membrane electrical potential measured relative to a reference electrode. Several kinds of ion-selective membrane are known including glass [2], plasticized polymers [3], or various crystalline materials. The best known example is probably the pH glass electrode which is still one of the most important standard laboratory devices. However, at present, the most versatile ion-selective membrane consists of an organic polymeric matrix containing a lipophilic ligand and a lipophilic ionic species. The key components of the membrane are the lipophilic ion and the ionophore. The former guarantees the operation of the ISE by keeping the total amount of measuring ions inside the membrane constant, while the latter assures a selective response of the ISE to the target ion.

Ion-selective electrodes are cheap and simple devices that can be miniaturized, allow on-line and in-situ measurements and may provide speciation information. Ideally they consume no analyte during the measurement and usually do not need sample preparation. They are widely applied, especially in clinical analysis.
3.2 Components of the polymeric ion-selective membrane

The ion-sensitive solvent polymeric membrane is ideally a water-immiscible liquid of high viscosity containing the components listed below.

*Polymeric matrix*

The polymeric matrix provides mechanical stability to the membrane. In ideal cases it is inert and has no chemical interaction with the sensed ions. Polymeric ISE-membranes are commonly prepared with poly(vinyl chloride) (PVC).

*Membrane solvent or plasticizer*

The membrane solvent reduces the viscosity and ensures a relatively high mobility of the membrane constituents. In order to provide a homogeneity of the organic phase it must be compatible with the other membrane components and acts as a plasticizer. Depending on its polarity and dielectric constant, it can influence the ion-exchanger characteristics of the membrane and, as a consequence, the selectivity. The plasticizers that are usually used in ISE-membranes are bis(2-ethylhexyl) sebacate (DOS, apolar) or ortho-nitrophenyl octyl ether (o-NPOE, polar) [4].

*Lipophilic ion*

The prerequisite for a theoretical Nernstian response of the ISE is that no significant amount of primary ions may be coextracted together with counter ions (ions with opposite charge sign of the measuring ion) from the sample into the membrane phase. This means that the membrane is permeable only for ions with the same charge sign of the measuring ion. This membrane characteristic is called permselectivity or Donnan exclusion. The presence in
the membrane phase of non-exchangeable lipophilic ions guarantees the operation of the ISE by keeping the total concentration of measuring ions in the membrane much higher than the coextracted amount, i.e. constant. Cation-selective ISE-membranes contain normally as lipophilic site a tetraphenylborate derivate, anion-selective membranes a tetraalkylammonium salt.

*Ionophore*

The ionophore, or ion carrier, or ligand, has the main influence on the selectivity of the ISE membrane. It ideally forms relatively strong, selective and reversible complexes only with the target ion, so that in the membrane no ion-exchange between measuring and interfering ions, i.e., no interference, occurs. Additionally, the ionophore structure must contain numerous lipophilic groups in order to keep its leaching rate from the membrane to the sample phase as low as possible.

*Lipophilic salt*

The addition of a lipophilic salt without ion-exchanger properties reduces the electrical resistance of the membrane. Additionally, it renders the ion-selective membrane more selective for divalent over monovalent ions, by increasing the ionic-strength in the membrane [5].

### 3.3 Response Mechanism

The electrochemical measuring cell (Figure 3.1) consists of two galvanic half-cells: the ion selective electrode and the reference electrode [1].
The total potential difference (electromotive force, $EMF$) measured under zero-current conditions between the two electrodes is the sum of local potential differences, arising at each electrochemical interface. Considering that only the membrane potential, $E_M$, and the liquid junction potential of the reference electrode, $E_{D,\text{ref}}$, are sample dependent, the other terms can be expressed as a constant contribution, $E_{\text{const}}$ [6]:

$$EMF = E_{\text{const}} + E_M + E_{D,\text{ref}}$$ (3.1)

The potential $E_{D,\text{ref}}$ originates from the different mobilities of ionic species at the boundary between sample and bridge electrolyte of the reference electrode. It can be kept small and constant by using an equitransferent bridge electrolyte of high concentration (e.g., 1 M KCl, NH$_4$NO$_3$ or LiOAc), or by
using samples with constant ionic strength. If necessary, the magnitude of
$E_{D,\text{ref}}$ can be estimated according to the Henderson formalism [7, 8]:

$$E_{D,\text{ref}} = \frac{RT}{F} \sum_i \frac{z_i u_i (a_{i,\text{ref}} - a_{i,S}) \ln \frac{\sum z_i^2 a_{i,\text{ref}}}{\sum z_i^2 a_{i,S}}}{\sum z_i^2 a_{i,S}}$$  \hspace{1cm} (3.2)

with:

- $R$ molar gas constant (8.314 J K$^{-1}$ mol$^{-1}$)
- $T$ absolute temperature [K]
- $F$ Faraday constant (96487 C mol$^{-1}$)
- $z_i$ charge of the ion $i$
- $u_i$ absolute mobility of the ion $i$ [cm$^2$ mol$^{-1}$ s$^{-1}$ J$^{-1}$]
- $a_i$ activity [mol l$^{-1}$] of the ion $i$ in the sample (S) and in the bridge
electrolyte solution of the reference electrode (ref). The ionic activities
can be calculated by the Debye-Hückel approximation [9, 10].

The membrane potential, $E_M$, can be divided into three separate
contributions, i.e., the phase boundary potential at the membrane/inner filling
solution interface, $E_{\text{PB}^+}$, the diffusion potential within the membrane, $E_{D,M}$,
and the phase boundary potential at the membrane/sample interface, $E_{\text{PB}^-}$ [6].
The inner phase boundary potential is usually independent of the sample. The
diffusion potential arises from charge separations within the membrane due to
the different mobilities of ionic species. In most cases of practical relevance,
the contribution of $E_{D,M}$ to $E_M$ can be neglected [11-14] since membranes
showing a Nernstian response contain only homogeneously distributed
primary and lipophilic ions of opposite charge sign. Moreover, even if more
than one kind of exchangeable ion are simultaneously present in the
membrane, they are usually quantitatively complexed by the ionophore, and,
therefore, have the same mobility. It is consequently possible to describe the membrane potential as [6]:

$$E_M = E_{M,\text{const}} + E_{PB'}$$  \hspace{1cm} (3.3)

The phase boundary potential difference arises from a charge separation, due to a non-uniform distribution of ionic species between the organic membrane and the aqueous sample phase. Taking into account chemical and electrical potential contributions, the electrochemical potential $\tilde{\mu}_1$ for the primary ion $I$ in a solution can be formulated as:

$$\tilde{\mu}_1 = \mu_1 + z_1 F \phi = \mu_1^0 + RT \ln a_1 + z_1 F \phi$$  \hspace{1cm} (3.4)

where $z_1$ and $a_1$ are the charge and activity of the ion $I$; $\mu_1^0$ is its standard free energy, $\mu_1$ the chemical potential and $\phi$ the electric potential. At thermodynamic equilibrium the electrochemical potential of the ion $I$ in the membrane, $\tilde{\mu}_{1,M}$, and in the sample, $\tilde{\mu}_{1,S}$, are equal, so that the phase boundary potential is obtained as:

$$E_{PB'} = \phi_M - \phi_S = \frac{\mu_{1,S}^0 - \mu_{1,M}^0}{z_1 F} + \frac{RT}{z_1 F} \ln \frac{a_{1,S}}{a_{1,M}}$$  \hspace{1cm} (3.5)

where the subscripts S and M indicate the aqueous sample and the organic membrane phase, respectively. The free energy difference, $\mu_{1,S}^0 - \mu_{1,M}^0$ (free energy of transfer), can be replaced by the so called single ion distribution coefficient, $k_1$:

$$k_1 = \frac{1}{RT} \exp \left\{ \mu_{1,S}^0 - \mu_{1,M}^0 \right\}$$  \hspace{1cm} (3.6)
The parameter $k_i$ is not a thermodynamic equilibrium constant, but it is a measure of the ion lipophilicity. It can however be combined with the corresponding values, $k_j$, for another ion, $J$, in an equilibrium constant $K_{ij}$ that describes the ion-exchange of the two uncomplexed ions between sample and membrane (cf. Figure 3.2):

$$K_{ij} = \frac{(k_j)^{z_i/z_j}}{k_i}$$  \hspace{1cm} (3.7)

Figure 3.2. Schematic view of the equilibria involved when a cation-selective membrane is in contact with a sample solution. The symbols $L$ and $R^-$ indicate the ionophore and the lipophilic anion; $\beta_{IL}$ and $\beta_{IL}$ are the formation constants for the complexes $IL^+$ and $JL^+$ between the ionophore and the primary, $I^+$, and interfering, $J^+$, respectively; $K_{ij}$ is the equilibrium constant for the ion-exchange of two uncomplexed ions between sample and membrane phase.

The Nernst equation (3.8) for the membrane-depending potential difference of the measuring cell is obtained from equations 3.1 and 3.5 if the activities of the ions inside the membrane are sample independent and, therefore, constant:
\[ E_1 = EMF - E_{D,\text{ref}} = E_1^0 + s_1 \log a_{1,S} \]  

(3.8)

where \( E_1 \) is the membrane-depending potential difference of an ideal measuring cell consisting of a reference electrodes and an ISE that only responds to changes in \( a_{1,S} \); \( E_1^0 \) represents the sum of all constant potential differences in the measuring cell and corresponds to the intercept of the linear response function of the ISE; \( s_1 = \ln 10/RT/(z_jF) \) is the Nernstian slope of the ISE response function (59.16/\( z_j \) mV at 25°C).

### 3.4 Selectivity

Selectivity is one of the most important characteristics of an ISE, as it determines its sensitivity when the sample contains interfering ions, \( J \), together with the primary ions, \( I \). The response of an ISE membrane in the presence of several kinds of ions is related to the equilibrium constants of the exchange reactions of primary and interfering ions between the organic and aqueous phase (cf. Figure 3.2 and Section 5.2). However, the resulting membrane potential can be exactly described with the help of potentiometric selectivity coefficients [15], without knowing any equilibrium constants.

The selectivity coefficient can be defined from the linear part of the response functions of the primary and interfering ions. The response function of a membrane to an interfering ion \( J \) is defined in complete analogy to the Nernst equation for a primary ion:

\[ E_J = E_J^0 + \frac{RT}{z_jF} \ln a_{J,S} \]  

(3.9)

where \( E_J \) and \( E_J^0 \) have analogous meaning of \( E_1 \) and \( E_1^0 \). If the response to the interfering ion \( J \) is described by using the standard potential \( E_J^0 \) of
equation 3.8 the following equation is obtained:

\[
E_j = E_i^o + (E_j^o - E_i^o) + \frac{RT}{z_j F} \ln a_{jS} \tag{3.10}
\]

The selectivity coefficient, also known as Nikolsky coefficient, is related to the \( EMF \) difference, \( (E_j^o - E_i^o) \), at the intercept of the linear response functions of the primary and interfering ions:

\[
K_{i,j}^{\text{pot}} = \exp \left\{ \frac{z_j F}{RT} \left( E_j^o - E_i^o \right) \right\} \tag{3.11}
\]

The magnitude of \( K_{i,j}^{\text{pot}} \) depends on the membrane composition, on the formation constants between ligand and investigated ions, and on the lipophilicity of the ions under study. The three factors and the experimental determination of \( K_{i,j}^{\text{pot}} \) are discussed in detail in Chapter 4. In order to calculate the selectivity coefficients from the measured \( EMF \) values the following equations can be used:

\[
\log K_{i,j}^{\text{pot}} = \frac{(E_j^o - E_i^o)}{s_i} \tag{3.12}
\]

or

\[
\log K_{i,j}^{\text{pot}} = \frac{E_j - E_i}{s_i} + \log a_i(I) - \frac{z_i}{z_j} \log a_j(J) \tag{3.13}
\]

where \( a_i(I) \) and \( a_j(J) \) are the activities of two different solutions, containing only a salt of the primary and interfering ion, respectively, and inducing the potentials \( E_i \) and \( E_j \). A graphical representation of the selectivity coefficient is given in Figure 3.3.
Figure 3.3. Determination of the potentiometric selectivity coefficients according to the separate solution method (SSM). The selectivity coefficient $K_{i,j}^{\text{pot}}$ corresponds to the vertical distance between the separately measured linear response functions for primary and interfering ions at $\log a = 0$, divided by the theoretical slope of the response function of the primary ion.

When more than two ionic species are simultaneously present in the sample, their contributions to the mixed response of the ISE, $E_{\text{ii}}$, can be estimated according to the semi-empirical Nikolsky-Eisenman formalism [16, 17]:

$$E_{\text{ii}} = E_i^0 + s_i \log \left\{ a_{i,S} + \sum_j K_{i,j}^{\text{pot}} a_{j,S}^{z_i/z_j} \right\}$$

(3.14)

Unfortunately, this simple equation provides correct results only if all involved ions have the same charge. When primary and interfering ions have different charges, the Nikolsky-Eisenman equation is inconsistent and leads to an incorrect description of the observed membrane potential in the mixed-ion response range, i.e., in the region where a deviation from the linear response
of the ISE is observed. This has been previously observed by Morf et al. [1], but only recently self-consistent models for two ions of any charge [18] and for any number of ions of any charge [15] have been published. Equation 3.15 describes the potentiometric response of an ISE in a sample solution containing mono- and divalent ions:

$$E_{iL} = E^{o}_{iL} + \frac{RT}{F} \ln \left[ \frac{1}{2} \sum_{i(1)} K_{i,i(1)}^{pot} \frac{1}{z_i} a_{i(1),S} + \frac{1}{2} \sum_{i(1)} K_{i,i(1)}^{pot} \frac{1}{z_i} a_{i(1),S} \right]$$

(3.15)

where the indexes $i(1)$ and $i(2)$ indicate monovalent and divalent primary or interfering ions, respectively ($K_{i,i}^{pot} = 1$).

### 3.5 Detection Limits

Deviations of the electrode function from the linear response are normally observed at high and low activities of the measuring ion. According to the IUPAC recommendations [25], the corresponding upper and lower detection limits are defined by the cross section of the two extrapolated linear segments of the calibration curve (Figure 3.4).

The upper detection limit is caused by the loss of membrane permselectivity, owing to the coextraction of the primary ion together with a counterion from the sample into the membrane. As a consequence, $a_{i,M}$ increases with increasing $a_{i,S}$ and the $EMF$ difference is smaller than expected by the Nernst equation (cf. equations 3.5 and 3.8 and Section 5.2) [19-23].
The lower detection limit can be dictated by two processes:
- the interference of a competing ion present in the sample. In this case the detection limit $a_t(DL)$ is related to the Nikolsky equation 3.14 and is given by:

$$a_t(DL) = K_{i,j}^{\text{pot}} a_j^{z_j/z_i}$$  \hspace{1cm} (3.16)

- the leaching of primary ions from the membrane phase into the aqueous surface layer. This process is treated in detail in Chapters 5 and 6.

Figure 3.4. The detection limits of an ion-selective electrode are defined according to the IUPAC recommendations [25] by the cross section of the two extrapolated linear segments of the calibration curve.
4. DETERMINATION OF FORMAL COMPLEX FORMATION CONSTANTS AND STOICHIOMETRIES OF VARIOUS Pb\(^{2+}\) IONOPHORES IN THE PLASTICIZED MEMBRANE PHASE

The formal complex formation constant and the stoichiometry of the ion-ionophore complexes in the membrane phase, together with the lipophilicity of the respective ions and the composition of the membrane, determines the selectivity of carrier-based potentiometric sensors. These parameters are the fundamental thermodynamic properties required for structure-selectivity correlations. Moreover, their knowledge allows one to optimize the membrane composition for a specific application.

4.1 Introduction

The most important characteristic of a chemical sensor is its selectivity. It is, therefore, not astonishing that much of the research work in the field of potentiometric ion sensors has been invested in developing selective lipophilic complexing agents [3, 24]. Such ionophores have traditionally been characterized in ISE membranes by measuring the potentiometric selectivity coefficients as the relevant parameters. If correctly determined, i.e., under conditions where the response to each of the ions in the investigated activity range is Nernstian, the \(K_{i,j}^{\text{pot}}\) values are constant for a given membrane composition [6]. Their magnitude is above all dictated by the complex formation constants (equation 4.4) and the stoichiometries of the ion-ionophore complexes within the membrane phase, but also, in a minor extent, by the ionophore and sites concentrations and by the lipophilicity of the ions under study (cf. equation 3.7).
For strongly discriminated ions, the conventional methods (separate solution Method, SSM or fixed interference method, FIM) proposed by IUPAC [25] for determining potentiometric selectivity coefficients yield biased values because of experimental artifacts. This was realized many years ago [26-28], but only recently an experimental procedure proposed by prof. E. Bakker allowed one to obtain unbiased selectivity coefficients for any discriminated ion [29]. According to this method, calibration curves for discriminated ions are taken before any contact of the membrane with the primary ion, so that each ion gives a Nernstian electrode slope. The selectivity coefficients are calculated from the measured potentials according to the separate solution method [25]. Later, it was shown that unbiased selectivity coefficients are also obtained with ISEs having inner solutions that prevent leaching of primary ions from the membrane into the sample solution [30].

The knowledge of true $K_{ij}^{pot}$ values of a given ISE membrane lets one to predict its exact response function in mixed solutions containing any kind of interfering ions of any charge [15, 18]. However, without knowing the stoichiometry of the ion-ionophore complex, it is not possible to predict the selectivity behavior of another ISE membrane with the same components but in a different proportion. Moreover, $K_{ij}^{pot}$ values are only relative quantities and cannot be directly interpreted in terms of complex formation constants. To establish structure-selectivity correlations with a view to ionophore design and to optimize membrane compositions it is, therefore, desirable to know the more fundamental thermodynamic parameters, i.e., the stoichiometries and formation constants of the complexes in the membrane.

Formal complex formation constants of ionophores in the organic membrane phase were first determined with an optical method [31], by comparing the responses of two bulk optode membranes (see Section 4.6), one containing the usual components, i.e., the ionophore ($L$), the $\text{H}^+$-selective chromoionophore ($C$), and the lipophilic ionic site ($R_T$), and the other without $L$, but otherwise of the same composition [32]. The ratio of the
corresponding constants of equilibrium for the exchange of \( \text{H}^+ \) and \( I^i \) between sample and membrane phase correspond to the ion-ionophore complex formation constant (Equation 4.9). This method relies on the basic assumption that the chromoionophore does not bind other ions than \( \text{H}^+ \).

In ISEs, theoretically, the difference between the phase boundary potentials of two permselective membranes, one with and one without an ionophore, measured in the same sample solution, directly reflects the degree of ion complexation in the membrane phase. In fact, according to equation 3.5, ion-selective membranes respond to changes in the activity of the uncomplexed ions in the membrane phase, if the activity in the sample is kept constant. Since however, the same potential change simultaneously occurs at both sides of the membrane, no net potential change is be observed. For determining complex formation constants, the inner phase boundary potential must be kept independent of the presence of the ionophore in the membrane.

The easiest way to eliminate the influence of the inner membrane potential would be the use of solid-contact electrodes, since the potential between membrane and inner electrode should not depend on the incorporated ionophore. Unfortunately, those systems are still not perfectly well understood, and there are not available simple techniques to manufacture them [33].

Another possibility is to choose a reference ion that in both membranes (with and without ionophore) induces the same potential. It has been shown that formal complex formation constants can be determined by comparing the potentiometric responses of a pair of ISE membranes charged with the ligand under investigation and a \( \text{H}^+ \) selective ionophore [34, 35]. Until the basic assumption that the chromoionophore does not bind other ions than \( \text{H}^+ \) is valid, the potential induced by the hydrogen ion is the same in both membranes (with and without ionophore) and it can be used as reference point. Consequently, the ratio of the selectivity values \( K_{i,\text{H}}^{\text{pot}} \) of the membranes
with and without ionophore is a function of the formal complex formation constant of $IL^{2+}$.

More recently, a new method based on segmented sandwich membranes has been introduced [36]. The problem of the inner membrane potential was in this case solved by using a two layer membrane, that made possible to obtain a membrane with the inner and outer surfaces of different composition. It has been shown that the initial membrane-potential difference across such a two-layer sandwich membrane, one with and the other without ionophore, contacted with the same solution on both sides, directly reflects the degree of ion complexation in the membrane. After reaching the equilibrium, i.e. after the ionophore is homogeneously distributed in both membrane layers, the potential across the membrane is zero, so that the measured potential only depends on the other contributions of the cell. It therefore has the same value as the ones obtained with the two individual membranes. For practical purposes of saving time, the complex formation constant are calculated from the $EMFs$ of the sandwich and single membranes measured under the same experimental conditions.

The topic of this chapter is the determination of formal complex formation constants of various $Pb^{2+}$-selective ionophores in the membrane phase. Over the last 20 years more than 20 of such ionophores have been described (cf. [3]) and the possible use of some of them in view of monitoring the quality drinking water has been demonstrated with optodes [37]. However, so far, no system has been achieved with optimal response times, signal stability and selectivities [38]. The recent relevant improvement of lower detection limits of ISEs [39] (cf. Chapters 5 and 6) make potentiometric sensors a promising alternative for this task. The knowledge of formal complex formation constants and stoichiometries is required for the rational optimization of the membrane and inner solution composition for a given application. Otherwise, the sensor design remains at the level of a trial and error process.
Unfortunately, the potentiometric method used so far [34] is based on the concurrence of \( \text{H}^+ \) an \( \text{I}^{2+} \) in the membrane, and cannot be directly applied to \( \text{Pb}^{2+} \)-selective ionophores because the measuring procedure requires solutions of high pH, whereas lead is only soluble at low pHs. On the other hand, the method based on sandwich membranes has been published after this work was done, and, compared with the procedure presented below, it has the disadvantage that a new membrane is required for each investigated ion.

In this work, tetramethylammonium ion (TMA\(^+\)) is chosen as the reference ion and the selectivities of a ligand-free ion-exchanger-based membrane are compared with those of the ionophore-based membrane. It is shown that this method is generally applicable in order to characterize ionophores, as long as no complexation between TMA\(^+\) und ionophores occurs. The changes in \( K_{\text{I,TMA}}^{\text{pot}} \) values upon varying the ionophore concentration also provide information on the stoichiometry of the complexes formed in the membrane. Finally the results are validated by reference measurements with bulk optodes [32, 38].

### 4.2 Theory

The potentiometric determination of the complex formation constant of an ionophore, \( L \), with an ion, \( \text{I}^{2+} \), is based on the comparison of \( K_{\text{I,L}}^{\text{pot}} \) values obtained with two ISE membranes, one with the ionophore to be characterized and the other without. The reference ion, \( \text{J}^{2+} \), must not interact with \( L \).

The model is based on the following assumption, that have already been listed in [32] and [18]:

- The solvent polymeric membrane phase behaves like a homogeneous organic phase and its surface is in equilibrium with the contacting aqueous solution.
- The membrane response is governed by the phase boundary potential at the sample/membrane interface, and the diffusion potential within the membrane is neglected.
- The ionophores and the anionic additives are confined to the membrane phase.
- Ion pairs between the lipophilic ionic additives and their counterions are either equal for all ions or negligible.
- Ionophores form stable complexes of not more than one stoichiometry at a time.
- The equilibrium constants determined are related to concentrations of the species in the membrane phase and to activities in the aqueous phase.

Since the model assumes that concentrations can be used in the membrane and neglects the formation of ion pairs, the obtained values are formal complex formation constants only.

According to the phase boundary potential model (equation 3.5), the potentiometric selectivity coefficient, $K_{i,j}^{\text{pot}}$, of a cation-selective ISE, is related to the concentration of the uncomplexed cations, $I^{z_i}$ and $J^{z_j}$, in the membrane surface layer as follows [6]:

$$K_{i,j}^{\text{pot}} = K_{i,j} \frac{[I^{z_i}]}{[J^{z_j}]}^{z_i/z_j}$$  \hspace{1cm} (4.1)

This simple equality can be derived from equations (3.5-3.9) and (3.11) by assuming that:

$$a_{i,M} \approx [I^{z_i}] \text{ and } a_{j,M} \approx [J^{z_j}]$$  \hspace{1cm} (4.2a, 4.2b)

The concentrations $I^{z_i}$ and $J^{z_j}$ in equation 4.1 refer to two different measurements under conditions ensuring that in each case only one kind of
cation is present in the membrane. This is easy to verify by experiment because it is the prerequisite of Nernstian response of the ISE to an ion. If the ISE membrane contains an ionophore that forms a complex with \( I^{z_1} \) with the stoichiometric factor \( n_1 \), its concentration in the membrane phase can be expressed as:

\[
[I^{z_1}] = \frac{[IL^{z_1}_{n_1}]}{\beta_{II_{n_1}}[L]^{n_1}}
\]  

(4.3)

with \( \beta_{II_{n_1}} \) as the formation constant of the complex \( IL^{z_1}_{n_1} \). An analogous equation is valid for \( J^{z_j} \) if it also forms a complex. If, however, TMA\(^+\) is used as \( J^{z_j} \), its concentration in the membrane is defined by the one of the lipophilic anionic sites, \( R_T \):

\[
[TMA^+] = R_T
\]  

(4.4)

From the selectivity coefficients \( K^\text{pol}_{l,TMA}^{\text{IE}} \) and \( K^\text{pol}_{l,TMA}^{\text{IE}} \), obtained with two membranes, one based on an ion exchanger only (IE) and the other containing \( L \) also, the complex formation constant for \( IL^{z_1}_{n_1} \) can be calculated accordingly:

\[
\beta_{II_{n_1}} = \frac{K^\text{pol}_{l,TMA}^{\text{IE}}}{K^\text{pol}_{l,TMA}^{\text{IE}}[L_T - n_1 R_T / z_1]^{n_1}}
\]  

(4.5)

Equation 4.5 is closely related to those described in [34, 40] and is valid only if strong complexes are formed, i.e. for \( \beta_{II_{n_1}} [L]^{n_1} \gg 1 \), so that:

\[
[IL^{z_1}_{n_1}] = R_T / z_1 \quad \text{and} \quad [L] = (L_T - n_1 R_T / z_1)
\]  

(4.6a, 4.6b)
where \( L_T \) is the total ionophore concentration. When selectivity coefficients have been calculated in relation to one selected primary ion, such as \( \text{Pb}^{2+} \) in this work, they must be recalculated since equation 4.5 holds for \( \text{TMA}^+ \) as \( J^{z_1} \):

\[
\log K_{1,TMA}^{\text{pot}} = \frac{z_1}{2} \left( \log K_{\text{Pb},TMA}^{\text{pot}} - \log K_{\text{Pb},1}^{\text{pot}} \right) \quad (4.7)
\]

Equation 4.5 is based on the assumption that only one kind of complex is formed under the experimental conditions. Its stoichiometry can be supposed for ionophores that build a well-defined cavity (e.g. valinomycin) or if independent evidence is available. In some cases, it has been gathered from the shape of the response function of bulk optodes \([38, 41, 42]\). In this work, \( n_1 \) is determined from the \( K_{1,j}^{\text{pot}} \) values of two membranes differing only in their ionophore concentration. Again, assuming that \( L \) does not complex \( \text{TMA}^+ \), according to the equations 4.1 and 4.6 the ratio of the selectivity coefficients obtained with ISE membranes 1 and 2 corresponds to:

\[
\frac{K_{1,TMA}^{\text{pot}}(1)}{K_{1,TMA}^{\text{pot}}(2)} = \left[ \frac{L_T(2) - n_1 R_T / z_1}{L_T(1) - n_1 R_T / z_1} \right]^{n_1} \quad (4.8)
\]

This equation can be used with different values of \( n_1 \) to predict \( K_{1,j}^{\text{pot}}(2) \) for a membrane with \( L_T(2) \) from \( K_{1,j}^{\text{pot}}(1) \) of another membrane with \( L_T(1) \) (cf. Figure 4.4). The stoichiometric factor \( n_1 \) can be estimated through the comparison of calculated and experimental values. Alternatively, using the selectivity coefficients obtained with the two membranes, equation 4.8 can be solved iteratively for \( n_1 \) (cf. Table 4.2).
4.3 Determination of the selectivity coefficients

Seven out of more than 20 Pb\(^{2+}\)-selective ionophores described in the literature [3, 43-47] have been selected for the investigation in this work (Figure 4.1). Five of them (I, IV-VII) are commercially available as lead ionophores. Compound I (ETH 322) was used in the first reported Pb\(^{2+}\)/PbA\(^+\) (A\(^-\): a sample anion) polymeric membrane ISE [44]. Ionophore II (ETH 5428) [37] has the same central unit and a similar lipophilicity as I, but different aliphatic chain lengths. The replacement of the two amide oxygens of II by sulfur atoms leads to the dithioamide IV (ETH 5435), which was shown to have an improved Pb\(^{2+}\) selectivity in optode membranes with respect to alkali and alkaline earth metal ions, but suffered severe interference from Hg\(^{2+}\), Ag\(^+\), Cu\(^{2+}\) and Cd\(^{2+}\) [37]. More recently, the corresponding monothioamide III (ETH 5439) [38] showed some improvement in this respect, while maintaining sufficient Pb\(^{2+}\) selectivities in relation to alkali and alkaline earth metal ions. The series II-IV allows one to check which is the best coordinating atom for Pb\(^{2+}\) by evaluating the influence of a stepwise replacement of amide O by S on the complex formation constants. Compound VI [47] currently seems to be the best ionophore for Pb\(^{2+}\) in spite of the fact that interference by Ag\(^+\) was severely underestimated in the original work. Finally, although VII is still offered as a lead selective ionophore, in agreement with other independent recent studies [29] it is shown to be only adequate for Ag\(^+\) measurements.
Figure 4.1. Structural formulas of the Pb$^{2+}$-selective ionophores investigated.
To determine the selectivity coefficients of the ISE membrane containing the cation exchanger, NaTFPB, but no ionophore, the originally proposed FIM [25] has been extended as follows. First, TMA\(^+\) is diluted in a constant background of \(I^z\) until the ISE no longer shows any potential change (Figure 4.2, right side). The assumption that at this point \(I^z\) is the potential determining ion, which is essential for obtaining unbiased \(K_{\text{IE}}^{\text{pot}}\) values, is then checked by further diluting the solution with pure water. As the plots in Figure 4.2 (left side) show, a nearly Nernstian slope is obtained for all investigated cations. The corresponding \(K_{\text{IE}}^{\text{pot}}\) values, recalculated for Pb\(^{2+}\) as primary ion, are given in Table 4.1.

![Figure 4.2](image_url)

**Figure 4.2.** EMF response curves obtained by the extended FIM on an ISE with an ion-exchanger membrane (NaTFPB, DOS, PVC) as basis for determining \(K_{\text{IE}}^{\text{pot}}(IE)\). First, the calibration curve was taken by diluting TMA\(^+\) with a constant background of the ion, \(I^z\), under study (right side). Then, the solution was further diluted with pure water to confirm that \(I^z\) was the potential determining ion (left side).
Table 4.1. Potentiometric selectivity coefficients, log $K_i^{ref}$, obtained with DOS/PVC membranes based on the ion exchanger NaTFPB or ionophore I-VII.

<table>
<thead>
<tr>
<th>Ion</th>
<th>NaTFPBa</th>
<th>I</th>
<th>II</th>
<th>III</th>
<th>IV</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
</tr>
</thead>
<tbody>
<tr>
<td>N(CH₃)₄⁺</td>
<td>7.0</td>
<td>−6.3</td>
<td>−6.0</td>
<td>−4.9</td>
<td>−3.9</td>
<td>−1.0</td>
<td>−6.5</td>
<td>6.7</td>
</tr>
<tr>
<td>H⁺</td>
<td>5.4</td>
<td>−3.1</td>
<td>−3.0</td>
<td>−4.8</td>
<td>−5.8</td>
<td>2.5</td>
<td>−7.5</td>
<td>4.4</td>
</tr>
<tr>
<td>Na⁺</td>
<td>3.8</td>
<td>−2.4</td>
<td>−2.1</td>
<td>−4.5</td>
<td>−5.6</td>
<td>−1.0</td>
<td>−7.5</td>
<td>3.5</td>
</tr>
<tr>
<td>K⁺</td>
<td>4.7</td>
<td>−3.0</td>
<td>−2.8</td>
<td>−4.4</td>
<td>−6.5</td>
<td>−1.1</td>
<td>−6.9</td>
<td>4.5</td>
</tr>
<tr>
<td>Ag⁺</td>
<td>3.8</td>
<td>−0.3</td>
<td>−0.1</td>
<td>15.0e</td>
<td>22.0e</td>
<td>0.9</td>
<td>9.5d</td>
<td>21.8e</td>
</tr>
<tr>
<td>Mg²⁺</td>
<td>−0.6</td>
<td>−6.0</td>
<td>−6.2</td>
<td>−9.1</td>
<td>−11.6</td>
<td>−3.2</td>
<td>−13.9</td>
<td>−1.7</td>
</tr>
<tr>
<td>Ca²⁺</td>
<td>−0.5</td>
<td>−0.7</td>
<td>−0.9</td>
<td>−5.8</td>
<td>−10.9</td>
<td>−1.2</td>
<td>−13.1</td>
<td>−1.1</td>
</tr>
<tr>
<td>Cu²⁺</td>
<td>−0.5</td>
<td>−3.5</td>
<td>−3.6</td>
<td>−2.7</td>
<td>−0.4</td>
<td>−2.5</td>
<td>−4.0</td>
<td>0.3</td>
</tr>
<tr>
<td>Cd²⁺</td>
<td>−0.6</td>
<td>−0.7</td>
<td>−1.0</td>
<td>0.7</td>
<td>−1.0</td>
<td>−2.5</td>
<td>−6.3</td>
<td>−0.3</td>
</tr>
</tbody>
</table>

aExtended FIM; mean values from measurements with six ISE; standard deviations: 0.1-0.2.
bSSM; mean values from a total of nine measurements on three membranes, each at three different concentrations between $5 \times 10^{-2}$ and $5 \times 10^{-4}$ M; standard deviations: 0.1-0.2.
cMean values from measurements on three ISE membranes at $a_{Ag} = 10^{-5}$ M (III) or $a_{Ag} = 10^{-6}$ M (IV) and $a_{Pb} = 10^{-2}$ M.
dObtained with o-NPOE as plasticizer [30] (it was not possible to obtain a cationic response with DOS membranes).

The selectivity coefficients for ISE membranes with ionophores I-VII have been determined by the SSM. For each ion, the Nernstian response was tested (see Figures 4.3a-d for four ionophores). Of the cations studied, only Ag⁺ did not fulfil this condition in that at higher activities, anionic responses were obtained with the S-containing ionophores. This is a consequence of the formation of too strong complexes inducing a coextraction of Ag⁺ and its counterion (Donnan failure). Except for VI, which exhibited a nearly horizontal (activity independent) response, in EDTA-buffered solutions cationic responses have been obtained in the range of $10^{-4}$-$10^{-8}$ M free Ag⁺ (not shown in Figure 4.3b-d). The corresponding EMF values were used for
calculating $K_{Pb,Ag}^{pot}$. All other selectivity coefficients in Table 4.1 represent mean values from three ISE at three concentrations between $5 \times 10^{-2}$ and $5 \times 10^{-4}$ M of primary or interfering cation.

In contrast to the original paper describing ETH 322 (I) as a PbA$^+$ selective ionophore [44], a Nernstian slope is observed here for Pb$^{2+}$. The reason is probably the use of lipophilic anionic sites in this work. It is, in fact, well known that an increased ionic strength in the membrane has a beneficial effect on the selectivity of divalent over monovalent ions [5]. The obtained response to Pb$^{2+}$ is in agreement with earlier optodes measurements performed by M. Lerchi [37].

**Figures 4.3a.** Response of ISE membranes (DOS/PVC/NaTFPB) based on the ionophore II (ETH 5428) to the cations under study.
The ligand ETH 5428 (II) yields very similar selectivities to ETH 322 so that its binding properties are apparently not influenced by the length of the aliphatic chains. As expected, alkali and alkaline earth metal ions are strongly discriminated by all S-containing ionophores. Interestingly, however, the discrimination of TMA⁺ worsens when the amide O in the ligand is replaced by S. If the assumption that the ionophores used do not complex TMA⁺ holds, this indicates that the Pb²⁺ complexes become weaker (see Section 4.5).

**Figures 4.3b.** Response of ISE membranes (DOS/PVC/NaTFPB) based on the ionophore III (ETH 5493) to the cations under study.
The selectivity coefficients obtained with the ISEs based on ligand VI are up to 8 orders of magnitude better than the original literature values [47], and it turns out that, so far, the calix[4]arene derivate is the best available Pb\(^{2+}\)-selective ionophore. It is important to realize that the improvement is only due to the new measuring procedure for the determination of unbiased selectivity coefficients described in [48].

**Figure 4.3c.** Response of ISE membranes (DOS/PVC/NaTFPB) based on the ionophores IV (ETH 5435) to the cations under study.

The application of VI in a potentiometric sensor for the measurements of Pb\(^{2+}\) in environmental water samples is discussed in Chapter 6. Compound VII was found to be only Ag\(^{+}\) selective, in agreement with other recently published data [29, 42].
4.4 Determination of the stoichiometries

With the experiments described above, the influence that have the presence of an ionophore on the selectivity of the membrane, i.e. on free cation activity within the membrane, could be evaluated. However, in order to calculate complex formation constants, the stoichiometry, $n_i$, of the ion-ionophore complex must also be known (cf. equation 4.5). In previous publications discussing the determination of formation constant in polymeric membranes, the stoichiometric coefficients have always been assumed, or, in a few cases, determined by optode measurements.
In this work the influence of the total ligand concentration on the selectivity coefficient of ISE membranes differing only in their ionophore concentration have been investigated, in order to collect information about the complex stoichiometry. As shown by equation 4.8, the ratios of the selectivities of such two membranes are related to the stoichiometric coefficient. In the case of ionophore III, the calibration curves obtained for Ca$^{2+}$ and Pb$^{2+}$ (normalized for TMA$^+$) are shown in Figure 4.4 with ISE membranes having $L_T(1) = 10.3$ and $L_T(2) = 30.9$ mmol/kg ($R_T$, and thus the ionic strength [5], is kept at a constant value of 3 mmol/kg). The dotted lines give the
theoretical response of the membranes with $L_T(2)$ calculated from the response of those with $L_T(1)$ for three different $n_i$ values. Evidently, for both cations, $n_i = 2$ gives a reasonable agreement between the calculated and experimental data. For a number of selected cases, selectivity coefficients have been determined with two membranes of different ionophore concentrations. The stoichiometric factors calculated by iteratively solving equation 4.8 are given in Table 4.2. Except for weak complexes, where the approximation involved ($\beta_{[L]}^n \gg 1$) does not hold, the obtained values are mostly a little above the nearest integer number. This can be explained either by the simultaneous occurrence of two complexes of different stoichiometries or by one of the underlying assumption (e.g., no ion pairs considered) being violated. The value determined for the Pb$^{2+}$ complex of III agrees with the one obtained earlier from the shape of the optode response function [38]. In the cases not figuring in Table 4.2, the $n_i$ values of table 4.3 were estimated as learned guess, e.g., by assuming that $n_{Mg} \leq n_{Ca}$, $n_{Cd} = n_{Cu} = n_{Pb}$, and that all $n_i$ values are the same for I and II (for VII, cf. [41]).

Table 4.2. Stoichiometric factors calculated with equation 4.8 for various ion-ligand complexes.

<table>
<thead>
<tr>
<th>Ion</th>
<th>II (DOS)</th>
<th>II (o-NPOE)</th>
<th>III (DOS)</th>
<th>IV (DOS)</th>
<th>VI (DOS)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H$^+$</td>
<td>1.08</td>
<td></td>
<td>(1.05)</td>
<td>(0.38)</td>
<td></td>
</tr>
<tr>
<td>Na$^+$</td>
<td>1.26</td>
<td>1.31</td>
<td>1.17</td>
<td>(0.46)</td>
<td>0.89</td>
</tr>
<tr>
<td>K$^+$</td>
<td>1.04</td>
<td></td>
<td>0.99</td>
<td>(0.46)</td>
<td>(0.76)</td>
</tr>
<tr>
<td>Ag$^+$</td>
<td>1.13</td>
<td></td>
<td></td>
<td>2.21</td>
<td></td>
</tr>
<tr>
<td>Ca$^{2+}$</td>
<td>1.39</td>
<td>1.43</td>
<td>2.15</td>
<td>(1.38)</td>
<td>(0.96)</td>
</tr>
<tr>
<td>Pb$^{2+}$</td>
<td>2.33</td>
<td>2.34</td>
<td>2.27</td>
<td>2.14</td>
<td>0.94</td>
</tr>
</tbody>
</table>

Values in parentheses refer to weak complexes (cf. Table 4.3) for which the approximations in equation 4.8 does not hold.
4.5 Determination of the formal formation constants

The formal complex formation constants calculated with the stoichiometries given above are listed in Table 4.3. To validate the results, bulk optode measurements were made for Pb\textsuperscript{2+} (see Figure 4.5). In spite of the fact that in optodes H\textsuperscript{+} instead of TMA\textsuperscript{+} is used as a reference ion, the values determined with the two methods are in excellent agreement. Obviously, the assumption that TMA\textsuperscript{+} does not interact with the ionophores holds. Moreover, the agreement of the data obtained with both sensor types shows that the response of the ionophores is not influenced by kinetic limitations, since contrary to steady-state conditions in ISE membranes, optode measurements are carried out in a situation of equilibrium.

\textit{Table 4.3a.} Formal complex formation constants, log\(\beta_n\), for \(\text{IL}_n\), and stoichiometries, \(n\), obtained with ionophores I – IV in DOS/PVC membranes.

<table>
<thead>
<tr>
<th>Ion</th>
<th>I</th>
<th>II</th>
<th>III</th>
<th>IV</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>(n)</td>
<td>(\log\beta_n)</td>
<td>(n)</td>
<td>(\log\beta_n)</td>
</tr>
<tr>
<td>H\textsuperscript{+}</td>
<td>1 (^b)</td>
<td>4.3</td>
<td>1</td>
<td>4.2</td>
</tr>
<tr>
<td>Na\textsuperscript{+}</td>
<td>1 (^b)</td>
<td>5.5</td>
<td>1</td>
<td>5.4</td>
</tr>
<tr>
<td>K\textsuperscript{+}</td>
<td>1 (^b)</td>
<td>4.7</td>
<td>1 (^a)</td>
<td>4.6</td>
</tr>
<tr>
<td>Ag\textsuperscript{+}</td>
<td>1 (^b)</td>
<td>6.5</td>
<td>1</td>
<td>6.4</td>
</tr>
<tr>
<td>Mg\textsuperscript{2+}</td>
<td>1 (^b)</td>
<td>9.7</td>
<td>1 (^a)</td>
<td>9.1</td>
</tr>
<tr>
<td>Ca\textsuperscript{2+}</td>
<td>1 (^b)</td>
<td>14.9</td>
<td>1</td>
<td>14.3</td>
</tr>
<tr>
<td>Pb\textsuperscript{2+}</td>
<td>2 (^b)</td>
<td>17.2</td>
<td>2</td>
<td>16.8</td>
</tr>
<tr>
<td></td>
<td>(17.4) (^c)</td>
<td>(17.2) (^c)</td>
<td>(15.7) (^c)</td>
<td>(14.6) (^c)</td>
</tr>
<tr>
<td>Cu\textsuperscript{2+}</td>
<td>2 (^b)</td>
<td>14.2</td>
<td>2 (^a)</td>
<td>13.7</td>
</tr>
<tr>
<td>Cd\textsuperscript{2+}</td>
<td>2 (^b)</td>
<td>17.1</td>
<td>2 (^a)</td>
<td>16.5</td>
</tr>
</tbody>
</table>

\(^a\) Standard deviations < 0.2 (for II-Pb\textsuperscript{2+}: 0.27); \(N = 9\).

\(^b\) Assumed values

\(^c\) From optode measurements
Table 4.3b. Formal complex formation constants, \( \log \beta_n \), for \( II_{L_n}^{+} \) and stoichiometries, \( n \), obtained with ionophores V – VII in DOS/PVC membranes.

<table>
<thead>
<tr>
<th>Ion</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>( n )</td>
<td>( \log \beta_n )</td>
<td>( n )</td>
</tr>
<tr>
<td>H(^+)</td>
<td>1 (^a)</td>
<td>&lt;3</td>
<td>1</td>
</tr>
<tr>
<td>Na(^+)</td>
<td>1 (^a)</td>
<td>3.2</td>
<td>1</td>
</tr>
<tr>
<td>K(^+)</td>
<td>1 (^a)</td>
<td>4.2</td>
<td>1 (^a)</td>
</tr>
<tr>
<td>Ag(^+)</td>
<td>1 (^a)</td>
<td>6.0</td>
<td>1</td>
</tr>
<tr>
<td>Mg(^{2+})</td>
<td>2 (^a)</td>
<td>8.6</td>
<td>1 (^a)</td>
</tr>
<tr>
<td>Ca(^{2+})</td>
<td>2 (^a)</td>
<td>10.5</td>
<td>1</td>
</tr>
<tr>
<td>Pb(^{2+})</td>
<td>2 (^a)</td>
<td>11.4</td>
<td>1</td>
</tr>
<tr>
<td></td>
<td></td>
<td>11.5 (^c)</td>
<td></td>
</tr>
<tr>
<td>Cu(^{2+})</td>
<td>2 (^a)</td>
<td>9.1</td>
<td>1 (^a)</td>
</tr>
<tr>
<td>Cd(^{2+})</td>
<td>2 (^a)</td>
<td>9.3</td>
<td>1 (^a)</td>
</tr>
</tbody>
</table>

\(^a\) Standard deviations < 0.2 (for II-Pb\(^{2+}\): 0.27); N = 9.

\(^b\) Assumed values

\(^c\) From optode measurements

The comparison of the formal complex formation constants of ionophores II-IV surprisingly shows that each successive replacement of an amide O by S induces a reduction in the stability of the corresponding Pb\(^{2+}\) complex by about one order of magnitude. The selectivity improvement, thus, is not the consequence of a stronger complexation of Pb\(^{2+}\) with the ligand, but of a weaker one of the interfering alkali and alkaline-earth metal ions. Another surprising result is that the stoichiometry of the Ca\(^{2+}\)–II (and, for structure-analogy reasons, of Ca\(^{2+}\)–I) complex in DOS/PVC and o-NPOE/PVC membranes is 1:1, whereas that of the Pb\(^{2+}\) complex is 1:2 (Table 4.2). This is in contrast with previous results obtained with other Ca\(^{2+}\) ionophores [28], showing that Ca\(^{2+}\) tends to form complexes of stoichiometries higher than 1. Further investigations are necessary to explain the contradiction.
4.6 Ion-selective bulk optodes

Since ion-selective bulk optodes are not a topic of this work, only a brief introduction is given here. The measurements with optodes presented here validate the assumption that TMA$^+$ does not interact with the investigated ionophores. More detailed informations about optodes can be found in the references [31, 49-51].

In ion-selective bulk optode membranes the degree of protonation of an incorporated H$^+$-selective chromoionophore, i.e., an ionophore that changes its spectral properties upon protonation, is measured spectrophotometrically. The thin membrane (2-4 μm), cast onto a lipophilized glass plate, contains in addition to the chromoionophore, an ionophore and a lipophilic anionic additive. The optode signal depends on the cation exchange equilibrium of the primary ion and H$^+$ between the aqueous and membrane phases. Activities of the primary ion in the sample are accessible, if the sample pH is measured or kept constant by buffering. The response curve of an optode membrane is described by the following equation:

$$a_1 = \left( \frac{a_{\text{H}^+}^{z_i}}{K_{\text{exch}}^1} \right)^{z_i} \left( \frac{\alpha}{1 - \alpha} \right)^{z_i} \frac{R_T - (1 - \alpha)C_T}{z_1 \left( L_T - \frac{n_i}{z_i} \left( R_T - (1 - \alpha)C_T \right) \right)^{n_i}}$$

(4.9)

where \(\alpha\) is the degree of deprotonation of the chromoionophore, \(K_{\text{exch}}^1\) is the constant of equilibrium for the exchange of H$^+$ and I$^{z_i}$ between sample and membrane phase and \(C_T\) is the total concentration of the chromoionophore. The other symbols have the same meaning as in the equations for ISE (see above).

The complex formation constant was obtained from the ratio of the two exchange constants determined for membranes with (L) and without (IE, ion-exchanger) cation-selective ionophore:
\[ \beta_{IL_{a1}} = \frac{K_{\text{exch}}(L)}{K_{\text{exch}}(IE)} \]  

(4.10)

**Figure 4.5.** Responses to Pb^{2+} (normalized to chromoionophore ETH 5418; pK_a 9.0 [51] and pH 4.7) of six DOS/PVC optode membranes based on the Pb^{2+}-selective ionophores as indicated (cf. Figure 4.1) and the chromoionophore ETH 5418 or ETH 5315. The data points were fitted with equation (4.9) using the stoichiometries from Table 4.3.
4.7 Experimental

**Reagents.** Poly(vinyl chloride) (PVC), bis(2-ethylhexyl) sebacate (DOS), sodium tetrakis-[3,5-bis(trifluoromethyl)phenyl]borate (NaTFPB), Methylene-bis-N,N-diisobutylthiocarbamate (MBDiBDTC, VII), 4-tert.-butyl-calix[4]arene-tetrakis(N,N-dimethylthioacetamide) (VI), and tetrahydrofuran (THF) were from Fluka AG (CH-8071 Buchs). The ionophores (cf. Figure 1), N,N'-diheptyl-N,N',6,6-tetramethyl-4,8-dioxoundecanediamide) (ETH 295, V), N,N'-dioctadecyl-N',N'-dipropyl-3,6-dioxaoctanediamide) (ETH 322, I), N,N,N',N'-tetradodecyl-3,6-dioxaoctanediamide) (ETH 5428, II), N,N,N',N'-tetradecyl-3,6-dioxaoctane-1-thio-8-oxodiamide) (ETH 5493, III), and N,N,N',N'-tetradodecyl-3,6-dioxaoctanedithioamide) (ETH 5435, IV) as well as the chromoionophores, 4-(octadecylamino)azobenzene (ETH 5315) and 9-dimethylamino-5-[4-(15-butyl-1,13-dioxo-2,14-dioanonadecyl)phenylimino]-5H-benzo[α]phenoxazine (ETH 5418) were synthesized in our laboratory as described in [37, 38, 52].

**ISE membranes.** Membranes of ca. 200 µm thickness were obtained by pouring a solution of ca. 250 mg of the membrane components, dissolved in ca. 2.5 mL of THF, into a glass ring (28 mm i.d.) fixed on a glass plate. The ion-exchanger membrane contained NaTFPB (0.45 wt%, 5.00 mmol/kg), DOS (66.15 wt%), and PVC (33.40 wt%). The membranes with an ionophore had the same amount of NaTFPB (5.00 mmol/kg), 65-66 wt% plasticizer, and 32-33 wt% of PVC. Ionophore concentrations were: ETH 322, 1.19 wt% (15.51 mmol/kg); ETH 5428, 1.30 wt% (15.29 mmol/kg); ETH 5493, 1.41 wt% (16.26 mmol/kg); ETH 5435, 1.43 wt% (16.24 mmol/kg); ETH 295, 1.35 wt% (28.64 mmol/kg); ionophore VI, 1.57 wt% (14.91 mmol/kg); MBDiBDTC, 1.24 wt% (29.31 mmol/kg). To determine the stoichiometries of ion-ligand complexes, pairs of membranes with constant amounts of NaTFPB (0.28 wt%, 3.20 mmol/kg), DOS (65-66 wt%),
and PVC (32-33 wt%) but two different concentrations of the ionophore under study were used. These were 0.9 and 2.6 wt% (ca. 10 and 30 mmol/kg, respectively) for ETH 5428, ETH 5493, and ETH 5435. For ionophore VI, owing to its limited solubility in DOS, they were 0.17 and 0.46 wt% (1.58 and 4.32 mmol/kg, respectively), with a constant NaTFPB concentration of 0.04 wt% (0.47 mmol/kg).

**EMF measurements.** Disks of 5 mm diameter were punched from the master membranes und glued with a THF/PVC slurry to a plasticized PVC tubing of 4 mm i.d. Three ISEs were prepared for each membrane composition. The internal filling solutions were in contact with the reference half-cell (Ag|AgCl, 3 M KCl) through a 1 M KCl bridge electrolyte. They consisted of $10^{-3}$ M N(CH$_3$)$_4$Cl (TMACl) for ISEs without ionophore, and of $10^{-1}$ M ethylenediamine tetraacetic acid disodium salt (Na$_2$EDTA) with $10^{-3}$ M Pb(NO$_3$)$_2$ for those with ionophore. The ion-exchanger-based ISEs were conditioned for about 12 h in $10^{-3}$ M TMACl, the others in a $10^{-3}$ M chloride solution of the most discriminated cation (Na$^+$ or TMA$^+$). All measurements were carried out at room temperature (ca. 21°C) with cells of the type Hg|Hg$_2$Cl$_2$|3 M KCl|1 M LiOAc|sample||membrane||inner filling solution|AgCl|Ag, with a free-flowing saturated calomel reference electrode [53]. Measuring solutions were prepared by successive automatic 5-fold dilution of stock solutions with a Liquino 711 and two Dosinos 700 (Metrohm AG, CH-9101 Herisau) equipped with 50-mL burettes.

For the ion-exchanger membranes, $K_{\text{Pct}}$ values were determined by the fixed interference method (FIM) as follows: The ISEs were equilibrated for 10 min in the initial sample containing $10^{-2}$ M TMACl and 2 x $10^{-3}$ M chloride of the interfering ion under study. Then, keeping the concentration of the interfering ion constant, the sample was successively diluted 5 times each, until further dilution did not give any potential change. At this point, the sample was diluted twice 5-fold with pure water to ascertain that the ISEs
responded to the interfering ion. After each dilution step, the potential was measured for 3 min.

For ionophore-based membranes, $K_{\text{TMA} j}^{\text{px}}$ values were obtained with the method described in [29], measuring the response curve in the range of $5 \times 10^{-2}$-$10^{-4}$ M for each ion, in sequence of increasing selectivity coefficients, starting with the most discriminated ion. In this work, the internal phase boundary potential was kept constant by using an EDTA-buffered Pb$^{2+}$ solution ($10^{-3}$ M Na$_2$EDTA, $10^{-5}$ M Pb(NO$_3$)$_2$). Initial concentrations of HCl, NaCl, KCl, Ca(NO$_3$)$_2$, Mg(NO$_3$)$_2$, and TMACI in the samples were $5 \times 10^{-2}$ M, those of Pb(NO$_3$)$_2$, Cu(NO$_3$)$_2$, Cd(NO$_3$)$_2$, and AgNO$_3$ were $10^{-2}$ M with $10^{-3}$ M NaOAc and $10^{-3}$ M HOAc (pH 4.70). After each dilution step, the potential was measured for 5 min. During this period of time, drifts were around 0.1 mV in most cases and never above 0.5 mV.

**Optode measurements.** The composition of the optode films was 10.8 mmol/kg (1.00 wt%) of NaTFPB, 26.3 mmol/kg (1.1-2.3 wt%) of ionophore (I-VI), 9.2 mmol/kg (0.65 wt%) of the chromoionophore ETH 5418, 65-66 wt% of DOS, and 32-33 wt% of PVC. The corresponding ionophore-free film had the same amounts of the remaining constituents. The membrane components totalling 150 mg were dissolved in 1.5 mL of THF and optode films of 2-3 μm thickness were prepared with the help of a spin-on device [54]. UV-vis spectra were measured with a UVIKON Model 810 double-beam spectrophotometer (Kontron AG, CH-8010 Zürich) controlled by a Macintosh SE and using a flow-through cell [55]. The pH values were determined with a pH meter Orion Model 290 A using a pH triode electrode Orion 91-57 BN (Hügli-Labortec AG, CH-9030 Abtwil). A peristaltic pump (Guldener Vario Perpex II, Werner Meyer AG, CH-6000 Luzern) was used with the flow-through system. The measuring cells were kept at 25°C by means of a thermostat (Gebrüder Haake KG, Berlin, Germany). The Pb(NO$_3$)$_2$ solutions ($10^{-2}$-$10^{-6}$ M) were buffered with $2.5 \times 10^{-3}$ M Mg(OAc)$_2$
and $5 \times 10^{-3}$ M HOAc to pH 4.70. For measurements with the ionophore-free membrane, unbuffered $10^{-1}$ M Pb(NO$_3$)$_2$ was used (pH 3.90). Sample activities were calculated according to the Debye-Hückel approximation and taking into account the complex formation of Pb$^{2+}$ with AcO$^{-}$, NO$_3$$^{-}$ and OH$^{-}$ [56, 57].
5. INFLUENCE OF KEY PARAMETERS ON THE LOWER DETECTION LIMIT AND RESPONSE FUNCTION OF SOLVENT POLYMERIC MEMBRANE ION-SELECTIVE ELECTRODES

The lower detection limit of conventional polymeric membrane ion-selective electrodes in diluted samples is given by the release of measuring ions from the membrane into the aqueous phase. It can be greatly improved by using inner solutions of appropriate composition [30, 39, 58].

5.1 Introduction

In most cases described so far, the lower detection limits of solvent polymeric ion-selective membrane electrodes were reported to lie in the micromolar range [3, 24]. Significantly lower values were found when the analyte ion activities were kept under control with the help of ion buffers [26, 28]. Even if it has been known for decades that the analyte concentrations in the aqueous surface layer, to which the ISE responds, and the sample bulk may differ significantly [1, 59-61], the origin of the improvements by using ion buffers was not explicitly traced at that time. A more intensive research in this field has only started more recently, when it was recognized that optical sensors, whose response mechanism is closely related to that of ISEs [6], had much better lower detection limits [42, 62]. Only in recent years, it became evident that the achievable lower detection limits of ISEs had been heavily underrated. Their values in conventional systems are biased by a continuous release of primary ions from the sensing membrane into the sample solution. This causes their activity in the sensed aqueous Nernst layer to be of the order of $10^{-6}$ M even in highly diluted samples. The reported selectivity coefficients for strongly discriminated ions obtained with conventional ISEs are in most cases biased too. In fact, they are based on EMF values measured in solutions
containing only the interfering ions, but with the ISE still responding to the activity of the primary ions released from the inner solution through the membrane. The values of the selectivity coefficients thus obtained are too pessimistic and may be worse by many orders of magnitude than the true ones. At present, there are various means to reduce or avoid such effects and to obtain unbiased selectivity coefficients. For example, an ISE membrane that was never in contact with primary ions before recording the calibration curve for a discriminated interfering ion, exhibits Nernstian slope to the latter and, therefore, provides unbiased thermodynamic selectivity coefficients [29, 48, 63]. The same results can be obtained with ISEs in which the release of primary ions from the inner filling solution to the sample is prevented [30, 39].

The primary ion release from the membrane is a consequence of zero-current ion fluxes that have their origin in the coextraction [64] of primary ions and their counterions from the inner solution into the membrane and/or ion exchanges at the sample- and inner solution-membrane interface [30, 58]. Both coextraction and ion exchange effects can be efficiently compensated by an appropriate choice of the inner filling solution. It has been shown that by lowering the analyte activity in the inner compartment through an appropriate ion buffer, while keeping the interfering-ion concentration high, the lower detection limits can be improved by many orders of magnitude [39, 65]. However, the composition of the inner solution must be carefully adjusted. While a too high concentration of primary ions on the inner membrane side has a negative influence on the lower detection limits, a too low value induces a primary ion uptake from the sample into the membrane. The corresponding response functions show a large super-Nernstian step at a critical sample activity, below which the ISEs are not sensitive to the primary ion [30]. Both processes, i.e., ion release and ion uptake, have been successfully described by a theoretical model worked out for the steady-state situation [58, 66, 67]. The influence of the composition of the inner solution and of various
membrane parameters on the response function of ISEs has been predicted with this model and verified by experiments shown in this chapter [30, 68].

5.2 Theory of ion-exchange, coextraction, and diffusion in solvent-polymeric ion-selective membranes

In recent years, several models have been developed for the prediction of the response function of ISEs in diluted samples [58, 66, 67, 69]. They consider the influence of ion-exchange between sample and membrane, of coextraction from the sample or inner solution into the membrane, and of zero-current ion-fluxes through the membrane and the aqueous phase boundary layer. In this Section only the basic equilibria and equations describing those processes are explained.

Ion-exchange

Uncomplexed interfering ions exchange with the primary ones at the sample-membrane interface according to the following ion-exchange equilibrium:

\[
\frac{Z_i}{Z_j} J_i^j (S) + I_i^j (M) \rightleftharpoons \frac{Z_i}{Z_j} J_i^j (M) + I_i^j (S) \quad (5.1)
\]

where \((S)\) and \((M)\) refer to species in the sample and membrane phase, respectively. The equilibrium constant, \(K_{ij}\) (cf. equation 3.7) is given by:

\[
K_{ij} = \left( \frac{a_{i,S}}{a_{i,M}} \right)^{Z_i/Z_j} \left( \frac{a_{j,M}}{a_{j,S}} \right) \quad (5.2)
\]
By considering the complexation between the ions and the ionophore in the membrane (cf. equations 4.1/4.2 and Figure 3.2), the overall ion-exchange equilibrium constant is obtained as follows (cf. Figure 3.2):

\[
K_{\text{exch}} = K_{ij} \frac{(\beta_{jL_{n_{ij}}})^{z_i/z_j}}{\beta_{iL_{n_{ij}}}} = a_{i,S} \left[ \frac{JL_{n_{ij}}^{z_j}}{[L]^{n_i}} \right] \left( \frac{[L]^{n_i/z_j}}{[L]} \right) \quad (5.3)
\]

The exchange constant \( K_{\text{exch}} \) is related to the selectivity coefficient and can be calculated without knowing the magnitude of \( K_{ij} \) and of the stability constants, if all the membrane parameters are known:

\[
\frac{1}{K_{\text{exch}}} = \frac{1}{K_{i,j}^{\text{pot}}} \frac{R_T}{z_i} \left( \frac{z_j}{R_T} \right)^{z_i/z_j} \left( \frac{L_T - n_j R_T / z_j}{L_T - n_i R_T / z_i} \right)^{n_j/z_j} \quad (5.4)
\]

The equation above can be obtained from equations 4.1 and 4.3 by inserting the appropriate mass and charge balances (equations 4.6a and 4.6b).

At equilibrium, the concentrations \([IL_{n_{ij}}^{z_i}]\) and \([JL_{n_{ij}}^{z_j}]\) in the membrane can be calculated if the activities \(a_{i,S}\) and \(a_{j,S}\) in the samples are known, by combining equation 5.3 with the following mass (equation 5.5) and charge (equation 5.6) balances:

\[
L = L_T - n_i \left[ \frac{IL_{n_{ij}}^{z_i}}{[L]} \right] - n_j \left[ JL_{n_{ij}}^{z_j} \right] \quad (5.5)
\]

\[
R_T = z_i \left[ \frac{IL_{n_{ij}}^{z_i}}{[L]} \right] + z_j \left[ JL_{n_{ij}}^{z_j} \right] \quad (5.6)
\]

The equation system has an analytical solution only for the simple case where \(z_i = z_j = 1\) and \(n_i = n_j = 1\). This has been used for the description of the
response function of ionophore-based ISEs by considering steady-state ion-fluxes under zero-current conditions [58]. If the charges or the stoichiometries have other values than 1, the equation system has to be solved numerically (cf. Section 5.4).

**Coextraction**

The word coextraction is used to describe the partitioning of the uncomplexed analyte cation together with a counter ion between the sample and the membrane phase.

\[
\frac{z_1}{-z_X} X^{z_X} (S) + I^{z_I} (S) \rightleftharpoons \frac{z_1}{-z_X} X^{z_X} (M) + I^{z_I} (M) \tag{5.7}
\]

The thermodynamic equilibrium of this process is described with the coextraction constant, \( K_{1X} \) [19]:

\[
K_{1X} = \frac{k_i}{(k_X)^{z_1/z_X}} \frac{\left[ I^{z_I} \right]}{a_{1,S}} \left[ \frac{X^{z_X}}{a_{X,S}} \right]^{-z_1/z_X} \tag{5.8}
\]

where \( X \) is a counter-ion of the measuring ion \( I \) and \( k_X \) is defined in complete analogy to \( k_i \). The overall coextraction constant \( K_{\text{coex}} \) is used to quantify the same process, when the complexation between ionophore and primary ion is considered:

\[
K_{\text{coex}} = \beta_{1L_{a_1}} \frac{k_i}{(k_X)^{z_1/z_X}} \frac{\left[ I^{z_I} \right]}{a_{1,S}} \left[ \frac{L}{a_{X,S}} \right]^{-z_1/z_X} \tag{5.9}
\]

The influence of the coextraction on the theoretical response function of ISE is treated in detail in references [19, 58, 64, 66].
Diffusion

Mass transfer in solution can occur because of a gradient in the electrochemical potential $\bar{\mu}$ (i.e., by diffusion and migration) or because of hydrodynamic reasons (i.e., by convection). A general description of the ion transport in an electrochemical cell at constant pressure and temperature, is given by the Nernst-Planck equation [70]:

$$J_i(x) = -D_i \frac{\partial c_i(x)}{\partial x} - \frac{z_i F}{RT} D_i c_i \frac{\partial \phi(x)}{\partial (x)} + c_i \dot{v}(x)$$

with:

- $J_i(x)$ rate of transfer of the species $i$ per unit area of the section in the x-coordinate, flux [mol cm$^{-2}$ s$^{-1}$]
- $c_i$ concentration of the diffusing species [mol cm$^{-3}$]
- $x$ space coordinate measured perpendicularly to the section [cm]
- $D_i$ diffusion coefficient for the diffusing species $i$ [cm$^2$ s$^{-1}$]
- $\phi(x)$ electrical potential in x-coordinate
- $\dot{v}(x)$ hydrodynamic velocity in x-coordinate [cm s$^{-1}$]

The general Nernst-Planck flux equation reduces to Fick’s first law of diffusion if the migration and convection terms can be neglected:

$$-J_i(x) = D_i \frac{\partial c_i(x)}{\partial x}$$

Diffusion is defined as the natural movement in solution, without effect of the electric field. The flux of a substance $i$, $J_i(x)$, represents the number of
moles of $i$ that pass at a given location per second and cm$^2$ of area normal to the direction of diffusion [70].

For a one-dimensional steady-state system, the differential equation 5.11 has the simple solution:

$$J_i = D_i \frac{\Delta c_i}{\Delta x} \tag{5.12}$$

where $\Delta c_i / \Delta x$ is the concentration gradient of the species $i$. Steady-state means that the concentration of the ion $i$ at the coordinate $x$ is not time dependent.

The equation above can be used for modeling the influence of ion-fluxes on the ISE response, if the laws of mass and charge conservation are considered, and if the ionic concentrations in the sample bulk and in the inner membrane boundary layer are assumed to be constant [58, 66, 67]:

- the law of mass conservation requires equal fluxes of the species $i$ in both the Nernstian layer and the membrane phase:

$$\Delta c_{i,S} \frac{D_{i,S}}{\delta_S} = \Delta c_{i,M} \frac{D_{i,M}}{\delta_M} \tag{5.13}$$

with $\delta_S$ and $\delta_M$ denoting the thickness of the aqueous Nernstian boundary layer and of the membrane, respectively;

- the law of charge conservation for zero-current measurements forbids a net charge-transport:

$$\sum_i z_i J_i(x) = 0 \tag{5.14}$$

Such system of equations can usually be solved only numerically. Simple explicit solutions exist only for special systems considering only two
monovalent ions. Simulations of the ISE response that takes into account migration are also reported in the literature, but they need a much more complicated approach [71, 72].

### 5.3 Leaching and depletion at the sample/membrane interface

Depending on the compositions of sample and inner solution, concentration gradients of primary and/or interfering ions are induced toward either one side of the ISE membrane and within the adjacent unstirred Nernstian layer of the aqueous sample.

Under zero-current conditions, the flux of primary ions across the membrane is accompanied by that of coextracted counterions in the same direction or by a counterflux of interfering ions entering the membrane through an ion-exchange process. The concomitant ion fluxes cause a difference, \( |\Delta c_{1,S}| = c_{1,\text{PB}} - c_{1,\text{bulk}} \), between the concentration of the primary ion in the bulk of the sample, \( c_{1,\text{bulk}} \), and in the sensed surface layer near the membrane, \( c_{1,\text{PB}} \). Their influence on the lower detection limit of ISEs is represented in Figure 5.1.

Usually, in samples with a primary ion concentration higher than about \( 10^{-6} \) M, the magnitude of \( \Delta c_{1,S} \) is much smaller than \( c_{1,\text{bulk}} \) and the response function of the ISE is not influenced by ion-fluxes. Only at lower bulk activities the analyte concentration at the phase boundary can be substantially higher or lower than that of the bulk, leading biased, non-linear response function.
Figure 5.1. Schematic representation of the processes influencing the lower detection limit of an ISE, which is based on an ionophore (L) that forms complexes with the primary (I⁺) and interfering (J⁺) ions. Concentration gradients are generated in the aqueous Nernstian layer (δₕ) either by a coextraction of I⁺ and the counterion X⁻ from the inner solution (A), or by partial ion-exchange at the sample and inner membrane sides (B and C).
When the flux of primary ions is directed from the membrane to the sample (Figure 5.1, schemes A and B), the lower detection limit is determined by \( |\Delta c_{1,s}| \), since at low bulk concentrations \( |c_{1,\text{PB}} - c_{1,\text{bulk}}| \approx |c_{1,\text{PB}}| \). Scheme A describes the situation where cations and anions coextracted from the relatively concentrated inner solution are transported through the membrane and reach the dilute sample. Scheme B shows that ion fluxes in direction of the sample may occur even when coextraction from the inner solution is negligible. In this case, the concentration gradients are induced by the exchange between primary and interfering ions at both membrane surfaces. In the reverse case, the Nernstian layer of the sample is depleted of the primary ions since their concentration in the membrane decreases toward the inner solution (Figure 5.1, scheme C). In an experiment with successive sample dilutions, the concentration in the aqueous phase boundary decreases abruptly at a given bulk concentration of the primary ion, \( c_{1,\text{bulk}} \). At this bulk activity the concentration gradient \( \Delta c_{1,s} \) approaches \( c_{1,\text{bulk}} \), since \( c_{1,\text{PB}} \ll c_{1,\text{bulk}} \), and the observed potential is much lower than expected for the concentration \( c_{1,\text{bulk}} \). As a consequence, the response function of the ISE has an apparently super-Nernstian slope.

At steady-state, the concentration difference between both sides of the membrane, \( \Delta c_{1,M} \), and of the Nernstian layer, \( \Delta c_{1,S} \), is related to the diffusion coefficients, \( D_{1,M} \) and \( D_{1,S} \), of the primary ion in the two phases and to the thickness of the respective potential determining diffusion layers \( \delta_M \) and \( \delta_S \) (cf. equation 5.13):

\[
\frac{\Delta c_{1,S}}{\Delta c_{1,M}} = \frac{D_{1,M}}{D_{1,S}} \frac{\delta_S}{\delta_M} = q \tag{5.15}
\]

The diffusion coefficients of cations in water are about \( 10^{-5} \text{ cm}^2 \text{ s}^{-1} \) (for example, for \( \text{Ca}^{2+} \) it is \( 8 \times 10^{-6} \text{ cm}^2 \text{ s}^{-1} \)) [73], while those of the ionophore complexes in the membrane are in the order of \( 10^{-8} \text{ cm}^2 \text{ s}^{-1} \) [74]. Since
lipophilic ionic sites and, as consequence, the measuring or interfering ion concentrations in the membrane are rather high (typically ca. 5 mmol kg\(^{-1}\), see Section 5.4) the overall ion-transport in the membrane, \( \Delta c_{1,M} D_{1,M} \), can be much higher than the one in a diluted sample, \( \Delta c_{1,S} D_{1,S} \). As a consequence, the smaller diffusion coefficients in the membrane are compensated by higher ionic concentrations, so that the slow ion release or uptake by the membrane can influence the composition of the Nernstian layer, even though the diffusion coefficient in water is much higher.

### 5.4 Influence of membrane parameters, inner filling solution and sample composition on the response function of Ca\(^{2+}\)-ISEs

The steady-state model outlined above, which considers ion exchange and coextraction on both sides of the membrane as gradient generating effects [58, 66, 67], was applied to predict how various parameters of an ISE influence its response function. In this Section the effects on ISE-response of the composition of the inner solution, the presence of interfering ions in the sample, the concentration of lipophilic anions in the membrane, the use of a flow-through cell, and the thickness and polymer content (viscosity) of the membrane are studied experimentally. Their influence on the response function of Ca\(^{2+}\)-selective electrodes with four different inner filling solutions (A to D) is reported, whereby the same ISE C with a conventional membrane (membrane thickness: 150 \(\mu\)m, \( L_T \): 13.1 mmol kg\(^{-1}\), \( R_T \): 4.5 mmol kg\(^{-1}\); represented with full circles in Figures) is used for direct comparison. Its inner solution generates a significant concentration gradient in the membrane as long as no significant ion exchange occurs on the sample side (80% of Ca\(^{2+}\) is replaced by Na\(^{+}\) at the inner side of the membrane). For convenience, the response curves are shifted so that potentials in the first sample solution (10\(^{-3}\) M Ca\(^{2+}\)) coincide. The potentiometric selectivity coefficients of the different ISEs are summarized in Table 5.1. Since quasi-
Nernstian responses were obtained for all investigated ions, the selectivity coefficients correspond to thermodynamic values [67], so that it is safe to use them for calculating equilibrium distributions between the inner solution and the adjacent membrane surface layer (cf. Section 5.2, Ion-exchange).

Table 5.1. Selectivity coefficients of various membranes containing different amounts of PVC and NaTFPB ($R_T$) and calculated percentage of Ca$^{2+}$ replaced by Na$^+$ in the inner membrane surface layer for the inner solutions A, B, C and D. The ratio $L_T : R_T$ is 3 : 1 for all the membranes.

<table>
<thead>
<tr>
<th>PVC wt%</th>
<th>$K_{Ca,Na}^{pot}$</th>
<th>$K_{Ca,H}^{pot}$</th>
<th>% of Ca$^{2+}$ replaced by Na$^+$ in inner membrane surface layer</th>
</tr>
</thead>
<tbody>
<tr>
<td>32.2</td>
<td>4.6</td>
<td>-4.7</td>
<td>&lt;0.1 1.7 79.8 &gt;99.9</td>
</tr>
<tr>
<td>20.8</td>
<td>4.6</td>
<td>-5.1</td>
<td>65.0</td>
</tr>
<tr>
<td>49.5</td>
<td>4.6</td>
<td>-5.6</td>
<td>46.0</td>
</tr>
<tr>
<td>32.2</td>
<td>0.45</td>
<td>-3.7</td>
<td>97.2</td>
</tr>
<tr>
<td>32.2</td>
<td>0.04</td>
<td>-1.5</td>
<td>99.8</td>
</tr>
</tbody>
</table>

The flow-through cell

The influence of the various membrane and sample parameters on the response function of the ISEs has been investigated using a flow-through system (Figure 5.2) instead of the conventional stirred solutions in a beaker. There are two motivations for this choice:
- in the flow-through system, the geometry of the measuring cell and the thickness of the Nernstian layer are exactly defined so that the experimental conditions are perfectly reproducible;
- in the flow-through system, the sample solution contacting with the membrane is continuously renewed, so that even at low Ca$^{2+}$ activities the
bulk concentration is not affected by ion-exchange processes between membrane or beaker wall and sample. The importance of this precaution is illustrated by the following. A typical membrane has a mass of 3 mg and contains 5 mmol kg\(^{-1}\) lipophilic anionic sites and thus \(1.5 \times 10^{-8}\) mol Ca\(^{2+}\). If only 10% of its Ca\(^{2+}\) content is exchanged by Na\(^+\), a 50 ml sample solution will be contaminated with \(3 \times 10^{-8}\) M Ca\(^{2+}\).

**Figure 5.2.** Schematic representation of the flow-through cell.

To reduce the noise caused by the pulsation of the peristaltic pump and by diffusion potentials that can arise in the small flow-channel, the reference electrode and the ISE were placed in front of each other in the flow-through
cell and the potentials were measured with a three electrode assembly (working ISE, reference and counter Pt electrode, cf. Figure 5.2).

**Influence of the inner filling solution**

The influence of the composition of the inner filling solution on the response function of the ISEs is shown in Figure 5.3. With the inner solution A (1 M Ca(NO₃)₂, 0.1 M NaCl) some coextraction of Ca(NO₃)₂ into the conventional membrane was expected [64]. This generates a concentration gradient and a steady release of Ca²⁺ on the sample side for all investigated samples. Nevertheless, a rather good lower detection limit of 10⁻⁷ M Ca²⁺ is obtained, due to the use of the flow-through cell (cf. [75]). In the inner filling solutions B, C, and D the Ca²⁺ activity in the inner solution gradually diminishes. Therefore, increasing amounts of Ca²⁺ are exchanged by Na⁺ in the inner phase boundary layer of the membrane. From the composition of the inner solution and the potentiometric selectivity coefficients, the amount of Ca²⁺ replaced by Na⁺ in the membrane containing conventional amounts of PVC, ionophore and sites was calculated as 1.7, 79.8, and >99.9% (ISE B, C and D, respectively; cf. Table 5.1 and Figure 5.3). A close to theoretical response with a lower detection limit of 10⁻⁸ M is obtained with the inner solution B. The fluxes generated by replacing 79.8% of Ca²⁺ in the inner membrane surface layer of ISE C cause a depletion of the sample at low Ca²⁺ activity so that a strong (super-Nernstian) EMF change is observed in the range of 10⁻⁷-10⁻⁸ M Ca²⁺. Below this concentration, the potential does not depend on aₙa,bulk, and Ca²⁺ at the sample side of the membrane is slowly exchanged by Na⁺ (cf. Response times). The steep EMF change at the critical Ca²⁺ activity was predicted by theory, but earlier experimental results [30] showed less abrupt changes than those presented here, and could have wrongly indicated the possibility of activity measurements with increased sensitivity (slope) in this region.
Figure 5.3. Response of four Ca$^{2+}$ ISE based on the ionophore ETH 5234 with identical conventional membrane composition but different inner solutions: (A) 1 M Ca(NO$_3$)$_2$, 0.1 M NaCl; (B) 10$^{-3}$ M CaCl$_2$, 0.1 M NaCl, pH 5.7; (C) 10$^{-3}$ M CaCl$_2$, 5 x 10$^{-2}$ M Na$_2$EDTA, pH 5.4; (D) 10$^{-3}$ M CaCl$_2$, 5 x 10$^{-2}$ M Na$_2$EDTA, pH 9.

In the present work, steeper response curves are obtained because at the corresponding concentrations the potential was recorded for a longer time (2 h), until the drift was < 10 mV h$^{-1}$ (cf. Response times). In the inner surface layer of the membrane of ISE with the inner solution D, Ca$^{2+}$ is quantitatively replaced by Na$^+$, so that larger concentrations gradients are generated than in the membrane of ISE with C. Therefore, the super-Nernstian response occurs already at a higher Ca$^{2+}$ activity and the EMF step exceeds that of ISE C by about 50 mV. Recently, it was proposed to monitor the potential difference of a pair of sensors such as ISEs C and D to obtain a strong signal in the critical activity range between the two super-Nernstian EMF-steps [76].
Influence of the concentration of interfering ions in the sample

Figure 5.4 demonstrates that the response function of ISE C is strongly affected by the ionic background of the sample. With 0.1 M NaCl as background, a lower detection limit of $1.6 \times 10^{-7}$ M Ca$^{2+}$ is obtained. This is close to the value expected from the Na$^+$ interference according to the traditional (static) model ($K_{Ca,Na}^{\text{pot}} a_{Na}^2 = 2 \times 10^{-7}$, cf. Table 5.1) [25]. On the contrary, with $10^{-3}$ and $10^{-5}$ M NaCl and without Na$^+$ background, the lower detection limits are higher than the values predicted in the absence of ion fluxes by the Nikolsky-Eisenmann formalism. The observed differences in the Ca$^{2+}$ responses at activities $< 10^{-7}$ M can be explained by a recently introduced dynamic selectivity model [77, 78] (cf. Chapter 6).

![Graph showing EMF response to Na$^+$ activity](image)

**Figure 5.4.** Influence of changes in the activity of the interfering Na$^+$ in the sample on the response function of the ISE with inner filling C. Much smaller influences would be expected from the Nicolosky-Eisenman equation, except in the case with 0.1 M Na$^+$. 
The exchange of only minor amounts of Ca$^{2+}$ by Na$^+$ or H$^+$ (ca. 6% and 1% at 10$^{-8}$ M Ca$^{2+}$ in solutions of pH 5 containing 10$^{-3}$ and 10$^{-5}$ M NaCl, respectively) on the sample side of the membrane modifies the gradients and reduces the uptake of primary ions from the Nernstian layer. As shown in Figure 5.4 the super-Nernstian slope, i.e., primary ion uptake, is successively reduced with increasing amounts of interfering ions.

Influence of the concentration of lipophilic ionic sites in the membrane

Not only the degree of ion-exchange in the inner (Figure 5.3) and outer (Figure 5.4) surface layers of the membrane, but also the concentration of lipophilic ionic sites define the extent of concentration gradients within the membrane. It is therefore expected that a decrease in the concentration of the ion exchanger, $R_T$, will reduce ion fluxes in either direction [66].

![Figure 5.5. Influence of 10fold changes in the concentration of ionic sites on the response of the ISE with inner solution C. The concentration ratio of ionophore to ion exchanger, $L_T/R_T$, is the same for all membranes.](image-url)
In the experiments shown in Figure 5.5 the concentration ratio of ionophore to ion-exchanger, $R_T/L_T$, was kept constant. A 10fold reduction in their concentrations improves the response of ISE C (full circles, same as in Figures 5.3 and 5.4), and a nearly Nernstian electrode function is obtained till $10^{-9}$ M Ca$^{2+}$. However, with a further 10fold reduction of $R_T$ and $L_T$, the lower detection limit increases. This can be explained by the strongly reduced H$^+$ selectivity of this membrane (see Table 5.1). As a general rule, the divalent to monovalent ion selectivity declines with decreasing ionic strength of the membrane [1]. In this case, the loss of selectivity due to a reduced amount of ion-exchanger was partly counterbalanced by the addition of a non-exchanging lipophilic salt [5]. Nevertheless, with $R_T = 0.05$ mmol kg$^{-1}$, $\log K_{Ca,H}^{\text{rel}}$ is +1.0 and the lower detection limit is mainly defined by H$^+$ interference ($K_{Ca,H}^{\text{rel}}a_H^+ = 10^{-9}$). For practical applications, it is probably expedient to reduce only the amount of lipophilic anionic sites, while keeping constant the concentration of the ionophore. This can on the one hand increase the coextraction of primary ions at high concentrations, but on the other hand will positively influence the selectivity of the membrane.

*Influence of the thickness of the Nernstian-layer*

The essential problem of submicromolar measurements with ISEs are the limited ion fluxes through the unstirred Nernstian layer of the sample. A reduction of the thickness of this layer is beneficial to the detection limit of membranes, both with release or uptake of primary ions, since a decrease of $\delta_S$ implies an increase of the net ion-transport (cf. equation 5.12) in the aqueous layer and thus reduces the influence of the ion-transport through the membrane. These theoretical predictions are fully confirmed by the experiments shown in Figure 5.6. The response curves obtained with ISE A, for which primary ions leach from the membrane into the sample, and with C, which shows a significant ion uptake, are Nernstian in a wider activity range when a flow-through cell is applied (solid lines) than in unstirred
samples (dotted lines). Similar improvements of the lower detection limits were observed earlier with a wall-jet system by Lindner et al. [75].

![Graph](image)

**Figure 5.6.** Response of Ca²⁺ ISEs with inner filling solution A (membrane with strong primary ion release) and C (membrane with moderate primary ion uptake) in unstirred samples and in the flow-through cell.

**Influence of the thickness and the viscosity of the membrane**

A decrease of the ion fluxes through the membrane has the same beneficial effect as their increase in the aqueous Nernstian layer. Experimentally it is possible to control the magnitude of ion fluxes through the ISE membrane by varying its thickness and/or the diffusion coefficient of the transported ions. A reduction of the concentration gradients by a 4fold increase of the thickness of the membrane leads to a nearly ideal response of ISE C down to 10⁻⁹ M Ca²⁺. On the other hand, its reduction by a factor of two shifts the super-Nernstian response step to higher activities on one logarithmic unit (Figure 5.7).
Figure 5.7. Influence of the thickness of the membrane on the response of the ISE with inner filling solution C.

Diffusion coefficients in the membrane can be widely varied by changing the concentration ratio of polymer to plasticizer [79, 80]. An increase of the polymer content of the membrane of ISE C from the usual value of 33% to 50% also leads to an almost ideal response down to the nanomolar range (Figure 5.8). On the other hand, its decrease shifts the super-Nernstian response to higher Ca$^{2+}$ activities in spite of the fact that the concentration gradients are smaller in this membrane than in the others because of its improved Na$^+$ selectivity (see Table 5.1).
Figure 5.8. Influence of the viscosity of the membrane on the response of the ISE with inner filling solution C.

Response times

All calibration curves were measured from the higher ($10^{-3}$ M Ca$^{2+}$) to the lower concentrated ($10^{-9}$ M Ca$^{2+}$) solutions. The time traces of the calibration curves of two ISEs (cf. Figure 5.3), one with a super-Nernstian step (C) and one without (B) are represented in Figures 5.9a and 5.9b. The response times in samples with concentrations higher than $10^{-6}$ M Ca$^{2+}$ were of the order of seconds for both kinds of the investigated ISEs, whereas in more diluted samples the times needed to reach a stable potential were considerably longer. For the ISE without super nernstian response the long equilibration time can originate from a small exchange ratio of Ca$^{2+}$ by H$^+$ at the sample side of the membrane, that leads to a small flux of Ca$^{2+}$ ions from the membrane to the sensed layer (in $10^{-8}$ M Ca$^{2+}$ solution at pH 5 0.6 % of Ca$^{2+}$ is exchanged by H$^+$).
Figure 5.9a. Time traces of the response of ISE B.

Figure 5.9b. Time traces of the response of ISE C.

The ISE with a strong gradient had a very long response time at the concentration where the super-Nernstian step occurred (after 2 h the potential drifts were still of about -10 mV h\(^{-1}\)). At this bulk concentration a complete ion exchange between Ca\(^{2+}\) and Na\(^{+}\) or H\(^{+}\) takes place, since the Ca\(^{2+}\)
concentration in the sensed layer is lower than $K_{1,j}^{\text{pot}} (a_{1,s})^{2/z_{1,j}}$. Consequently, the concentration gradient within the membrane changes, leading to potential drifts that reduce to the conventional values only when the membrane has reached the steady state (after ca. 10 h for a PVC content of ca. 33% [74]).

**Conclusions**

While the variation of membrane thickness and polymer content will be valuable means when optimizing ISEs for practical trace level applications, it must be kept in mind that a steady state situation is reached much more slowly with high polymer content and large thickness [80]. For example, the diffusion coefficient is about 10 times smaller when the PVC:DOS ratio is increased from 1:2 to 1:1 [33] and the time needed to reach steady state is 16 times longer when the membrane has a 4fold thickness. Therefore, the influence of other measures, i.e. the reduction of site concentration and efficient stirring or the use of a flow-through system should be fully exploited first.

5.5 Pb$^{2+}$-ISEs with strong leaching, strong depletion and with a linear response function down to 5 x 10$^{-9}$ M

In the previous Section, the influence of various key parameters on the response function of Ca$^{2+}$-ISEs has been systematically investigated. It has been shown that zero-current ion fluxes through ion-selective membranes in contact with dilute samples, may induce significant concentration polarizations in the aqueous Nernst diffusion layer, leading to a non-ideal response of the ISEs. On the basis of the theoretical models and the obtained results, it is now possible to carefully adjust the various parameters in order to obtain higher or lower detection limits, or a super-Nernstian response with any kind of polymer membrane ISE. Three kind of ISEs (E, F, and G) based
on the Pb$^{2+}$-selective ionophore VI described in Chapter 4 are used here as examples. Their responses in water solution containing a background of $10^{-3}$ M CaCl$_2$ and $10^{-4}$ M HNO$_3$ are shown in Figure 5.10. All ISEs were conditioned in the same solution and the calibration curves were measured all together in the same samples.

_Leaching of Pb$^{2+}$ ions from the membrane into the sample: ISE E_

The ion-selective membrane of ISE E contained the conventional amounts of lipophilic sites and ionophore, and had a thickness of about 150 μm. It had $10^{-1}$ M Pb(NO$_3$)$_2$ as inner filling solution and o-NPOE as plasticizer. The lower detection limit of ca. $2 \times 10^{-6}$ M is dictated by the strong coextraction of Pb(NO$_3$)$_2$ at the inner membrane side, due to the lipophilicity of NO$_3^-$, and by the high diffusion coefficient of ionic species in membranes plasticized with o-NPOE [74].

_Pb$^{2+}$-ISE with a Nernstian response down to $5 \times 10^{-9}$ M: ISE F_

In order to obtain a reproducible calibration curve down to $10^{-9}$ M the Pb$^{2+}$-selective membrane of ISE F was modified as follows. The plasticizer was DOS instead of o-NPOE and the content of lipophilic ionic sites as well as ionophore was about 20 times smaller compared with the previous one. Additionally, the Pb$^{2+}$ activity of the inner filling solution was kept at a constant value of ca. $10^{-9}$ M with NTA as ion buffer, so that the ion-exchange at the inner membrane surface and, as a consequence, the ion fluxes in the membrane were not strong enough to bias the response function of the ISE. This membrane is the same used in Chapter 6 for the measurements of Pb$^{2+}$ in drinking and natural water samples.
Depletion at the outer membrane surface: ISE G

The membrane composition of ISE G was the same of ISE E, but the inner filling solution contained a sufficiently lower Pb$^{2+}$ and higher Na$^+$ concentration to induce a complete exchange of Pb$^{2+}$ by Na$^+$ in the inner membrane surface layer. This caused a strong Pb$^{2+}$ gradient in the membrane directed toward the inner compartment and, as a consequence, a super-Nernstian response of the ISE between the Pb$^{2+}$ concentrations $5 \times 10^{-8}$ and $5 \times 10^{-7}$ M. At sample activities below $5 \times 10^{-8}$ M, virtually no Pb$^{2+}$ is present in the membrane-contacting aqueous layer as shown by the independence of the EMF on $a_{\text{Pb}^{2+}}$.

Figure 5.10. Response functions of ISEs E, F, and G in Pb$^{2+}$ solutions with a background of $10^{-3}$ M CaCl$_2$ and pH 4.
**Response times**

The Pb$^{2+}$ measurements were done in a 500 ml polypropylene beaker and the calibration curves was measured from lower to higher concentrations. Figure 5.11 shows that measurements at trace levels are possible even in the presence of a high concentration of ionic background (10$^{-3}$ M CaCl$_2$, pH 4), and that the response of the ISE is reasonably fast. Only ISE G has a very long response time when the concentration of the solution changes from 5 x 10$^{-7}$ to 5 x 10$^{-8}$ M. This is due to an ion exchange at the membrane surface contacting the sample, with a concomitant change of the ion gradients in the membrane.

**Figure 5.11.** Time traces of the responses of ISEs E, F, and G.
5.6 Experimental

**Reagents.** Poly(vinyl chloride) (PVC), bis(2-ethylhexyl) sebacate (DOS), 2-nitrophenyl octyl ether (o-NPOE), sodium tetrakis[3,5-bis(trifluoromethyl)phenyl]borate (NaTFPB), tetradecylammonium tetrakis(4-chlorophenyl)borate (ETH 500), the Ca$^{2+}$-ionophore $N,N$-dicyclohexyl-$N,N$-dioctadecyl-3-oxapentanediomide (ETH 5234), the Pb$^{2+}$ ionophore 4-tert-butylicalix[4]arene-tetrakis(thioacetic acid dimethylamide) (lead ionophore IV), and tetrahydrofuran (THF) were Selectophore® from Fluka AG (CH-9470 Buchs, Switzerland). Aqueous solutions were prepared with freshly deionized water (specific resistance, 18.1 MΩ cm) obtained with a NANOpure reagent-grade water system (Barnstead, CH-4009 Basel, Switzerland). Nitrilotriacetic acid (NTA) (Microselect), disodium ethylenediamine tetraacetic acid (Na$_2$EDTA) (puriss p.a.), nitric acid (Traceselect®) and CaCl$_2$ (puriss. p.a.) were from Fluka, NaCl and CaCl$_2$ (Suprapur®) from Merck (Darmstadt, Germany).

**Ca$^{2+}$ ISE Membranes.** Conventional Ca$^{2+}$-selective membranes contained ETH 5234 (1.0 wt.%, 13.1 mmol kg$^{-1}$), NaTFPB (0.4 wt.%, 4.5 mmol kg$^{-1}$), DOS (66.4 wt.%), and PVC (32.2 wt.%) and had a thickness of about 150 μm. Until not specified experiments were performed with this membrane. Membranes of ca. 75, 150, 600 μm thickness were obtained by casting respective solutions of 200, 400, 1600 mg of the membrane components (dissolved in ca. 3, 5, 16 mL of THF) into glass rings of 45 mm i.d. fixed on a glass plate. Membranes differing in their viscosity from the conventional ones contained the same amounts of ETH 5234 and NaTFPB as the others, but a PVC/DOS ratio of 1:3.7 for the softer and 1:1 for the harder ones; solutions of 180 mg of membrane components poured into glass rings of 28 mm i.d. gave membranes of ca. 160 μm thickness. The two membranes with lower concentrations of ETH 5234 ($L_T$) and NaTFPB ($R_T$) were prepared by the same procedure as the conventional ones but additionally contained ETH
500 (0.6 wt.%, 5.6 mmol kg\(^{-1}\)), with \(L_T\) and \(R_T\), respectively, corresponding either to 0.11 and 0.04 wt.% (1.38 and 0.45 mmol kg\(^{-1}\)) or to 0.01 and 0.004 wt.% (0.14 and 0.05 mmol kg\(^{-1}\)).

**Ca\(^{2+}\) Measurements.** Disks of 5 mm diameter were punched from the master membranes and glued with a THF/PVC slurry to a plasticized PVC tubing of 4 mm i.d. Four different inner filling solutions were used. They consisted of 1 M Ca(NO\(_3\))\(_2\), 10\(^{-1}\) M NaCl (inner solution A), 10\(^{-3}\) M CaCl\(_2\), 0.1 M NaCl, pH 5.7 (inner solution B), 10\(^{-3}\) M CaCl\(_2\), 5 \times 10\(^{-2}\) M Na\(_2\)EDTA, adjusted to pH 5.4 with 10\(^{-2}\) M NaOH (inner solution C), and 10\(^{-3}\) M CaCl\(_2\), 5 \times 10\(^{-2}\) M Na\(_2\)EDTA, adjusted to pH 9.0 with 6 \times 10\(^{-2}\) M NaOH (inner solution D). ISE A, B, C, D were based on the same conventional membrane and the respective inner solutions. The calculated activities of free Ca\(^{2+}\), \(a_{Ca^{2+}}\), were 4 \times 10\(^{-1}\) M, 4 \times 10\(^{-4}\) M, 5 \times 10\(^{-8}\) M, and 2 \times 10\(^{-12}\) M for ISEs A, B, C, and D, respectively. Between internal solution and inner half-cell (Ag|AgCl in 10\(^{-1}\) M NaCl) was 10\(^{-1}\) M NaCl as bridge electrolyte. The ISEs were conditioned for at least 4 d in a solution with 10\(^{-3}\) M CaCl\(_2\), 10\(^{-5}\) M NaCl and 10\(^{-5}\) M HNO\(_3\). Those with membranes of 600 \(\mu\)m thickness or 50 wt.% PVC were conditioned for 15 d. Sample solutions consisted of CaCl\(_2\) (concentration range 10\(^{-3}\)-10\(^{-9}\) M) with a fixed ionic background of 10\(^{-5}\) M NaCl and 10\(^{-5}\) M HNO\(_3\). The measuring time for sample concentrations above 10\(^{-6}\) M was 10 min. At concentrations at which super-Nernstian response was observed (10\(^{-6}\) or 10\(^{-7}\) M), the \(EMF\) was recorded for 2 h, at lower concentrations for 30 min.

Potentials were measured with a custom-made 16-channel electrode monitor at room temperature (20-21 °C) in a flow-through cell (see Figure 5.2) or in unstirred solutions. The reference electrode was in both cases a double-junction free-flowing calomel electrode with 1 M KCl as bridge electrolyte. All \(EMF\) values were corrected for the liquid-junction potential using the Henderson equation. The selectivity coefficients given in Table 5.1 were
determined by the method described in [29]. Activity coefficients were calculated according to [9].

**Pb$^{2+}$ ISE Membranes.** The membranes for LDL ISE contained 0.07 wt % (0.70 mmol kg$^{-1}$) lead ionophore IV, 0.03 wt % (0.35 mmol kg$^{-1}$) NaTFPB, 1.23 wt % (10.74 mmol kg$^{-1}$) ETH 500, 61.66 wt % DOS, 37.00 wt % PVC, and 0.17 mmol kg$^{-1}$ Pb(NO$_3$)$_2$. The membranes for conventional and super nernstian ISE contained 1.09 wt % (10.37 mmol kg$^{-1}$) lead ionophore IV, 0.44 wt % (4.94 mmol kg$^{-1}$) NaTFPB, no ETH 500, 65.65 wt % o-NPOE, 32.82 wt % PVC, and 0.17 mmol kg$^{-1}$ Pb(NO$_3$)$_2$. The parent membranes were prepared by dissolving ~450 mg of membrane components (except Pb(NO$_3$)$_2$) in ~5 mL THF. The lead salt was successively added to the membrane cocktail as 50 μL of aqueous Pb(NO$_3$)$_2$ solutions (the Pb$^{2+}$ amount in mol in the cocktail has to be one half of the borate amount). This membrane solution was stirred for about 1 h. A membrane of ca. 150 μm was obtained by casting the solution into a glass ring (7.0 cm i.d.) fixed on a glass plate.

**Pb$^{2+}$ Electrodes.** A disk of 4 mm diameter was punched from the above ion-selective membrane and glued to a plasticized PVC tubing with a THF/PVC slurry. The internal filling solutions were 10$^{-1}$ M Pb(NO$_3$)$_2$ (ISE E, conventional), 10$^{-4}$ M Pb(NO$_3$)$_2$, 10$^{-3}$ M NTA, 2.1 x 10$^{-3}$ M NaOH, pH 7, (ISE F, optimal low detection limit), and 10$^{-5}$ M Pb(NO$_3$)$_2$, 10$^{-1}$ M Na$_2$EDTA, 10$^{-1}$ M NaOH, pH 7.5, (ISE G, with super-Nernstian response). A bridge electrolyte of 10$^{-1}$ M NaCl (ISE E and ISE G) or 2 x 10$^{-3}$ M NaCl separated the inner filling solution from the inner reference half cell (Ag/AgCl, 3M KCl). The electrodes were conditioned for ca. 20 h in 5 x 10$^{-9}$ M Pb(NO$_3$)$_2$, 10$^{-3}$ M CaCl$_2$, 10$^{-4}$ M HNO$_3$. All measurements were done with 3 ISEs of the same composition.

**Pb$^{2+}$ measurements.** Calibration curves were taken from lower to higher concentrations at pH 4 (adjusted with 0.1 M HNO$_3$ using a Metrohm pH glass electrode (Metrohm AG, CH-9010 Herisau) in polyethylene beakers which
were pretreated with 0.1 M HNO₃ overnight. Potentials were measured with a custom made 16-channel electrode monitor at room temperature (20-21 °C) in stirred solutions. For each solution, potentials were measured until the drifts were smaller than 1 mV/min (10 min below 10⁻⁷ M, 40 min at 10⁻⁶ where ISE 3 has a super-Nernstian step), and 5 min at higher concentrations. The reference electrode (Metrohm Ag/AgCl in 3 M KCl, type 6.0729.100) had a 1 M NH₄NO₃ bridge electrolyte.
6. ION-SELECTIVE POLYMERIC MEMBRANE ELECTRODES FOR POTENTIOMETRIC MEASUREMENT OF Pb\textsuperscript{2+} IN ENVIRONMENTAL SAMPLES AT TRACE LEVELS

Potentiometric polymeric membrane electrodes based on electrically neutral ionophores can be used for the determination of heavy metal ions in natural water samples at nanomolar concentration. The obtained results are in good agreement with ICPMS measurements.

6.1 Lead in the environment and in the human body

Lead is a bluish-gray metal occurring in small amounts in the earth's crust. It can be found in all parts of our environment mainly as a result of human activities like mining, manufacturing and burning of fossil fuels. Lead is widely used for the production of batteries, ammunition, pipes and pigments, and for protection from gamma and X-ray radiations. In the last decades, the major origin of lead emission has been motor vehicle exhaust resulting from the use of tetraethyllead as "anti-knock" agent [81]. Only in very recent years the level of lead in the air has begun to decrease, due to the diminished use of lead in petrol.

Distribution of lead in the environment

Lead, the most abundant natural heavy metal element, is widely distributed in the earth's crust. Its primary forms in nature are insoluble salts or oxides of Pb(II) such as PbS, PbO\textsubscript{2}, PbCO\textsubscript{3} and PbSO\textsubscript{4} [82]. The presence of lead in the ecosystem is mainly due to anthropogenic sources like mining, industrial emission, automobile exhaust and burning of fossil fuels. This leads to the
strong deviations in the environmental lead concentrations between urban region of high industrial density and rural regions without industrial sites.

*Lead in aquatic systems*

With exception of lead chlorate, perchlorate, nitrate, and acetate, inorganic Pb(II) compounds are insoluble or only slightly soluble in water. Depending on pH and carbonate concentration, the main lead species found in environmental water are PbCO$_3^0$, PbOH$^+$, PbHCO$_3^+$, and PbSO$_4^0$. The free hydrated Pb$^{2+}$ ion is predominant only under acidic conditions (pH < 7) and in the absence of organic ligands. In systems with very high Cl$^-$ concentrations, such as estuaries and seawater, the dominant species is PbCl$^+$. The chemical speciation of Pb(II) in a typical organic-free natural water system is shown in Figure 6.1. In not seriously polluted water, the concentration of total lead is usually between 50 and 0.5 nmol/L [83], but it may be higher in tap water due to uptake from metallic tubes or lead-soldered joints in the plumbing system. Owing to the uptake from the air, the lead concentration can also be quite high in rainwater in regions of high industrial density.

*Uptake and distribution of lead in the human body*

The main sources of lead uptake by humans are inhaled air, drinking water, and food [83]. Lead can be inhaled from air in form of PbBrCl particles with a diameter below 0.5 mm or as alkyllead vapor, but it can also be rapidly absorbed through the skin by formation of lipid-soluble complexes with lactic acid and amino acids in sweat [84].
Figure 6.1. Distribution of the various Pb$^{2+}$ species as a function of pH in a sample solution containing only Pb$^{2+}$ and the inorganic anions CO$_3^{2-}$, HCO$_3^-$, Cl$^-$, OH$^-$, and SO$_4^{2-}$. The curves are obtained with the following formation constants [56, 57, 85] and concentrations: $\beta_{\text{PbCO}_3} = 10^{6.2}$, $\beta_{\text{PbHCO}_3} = 10^3$, $\beta_{\text{PbCl}} = 10^{1.6}$, $\beta_{\text{PbOH}} = 10^{6.3}$, $\beta_{\text{PbSO}_4} = 10^{2.7}$, $[\text{Pb}^{2+}]_{\text{tot}} = 1.0 \times 10^{-9}$, $[\text{CO}_3^{2-}]_{\text{tot}} = 2.9 \times 10^{-3}$, $[\text{SO}_4^{2-}]_{\text{tot}} = 1.55 \times 10^{-4}$, $[\text{Cl}]_{\text{tot}} = 1.25 \times 10^{-4}$. The pK$_a$ values for the dissociation of carbonic acid are pK$_{a1} = 6.35$ and pK$_{a2} = 10.25$.

The average lead content in ambient air can vary from less than 0.04 nmol m$^{-3}$ in remote areas to more than 50 nmol m$^{-3}$ in industrial regions and working environments [83] such as mines, storage battery factories, glassworks, etc. Tobacco smoking can also slightly contribute to the daily lead intake. The uptake of lead from the diet (drinking water and food) occurs in the gastrointestinal tract and can be influenced by other dietary substances. For example, iron, calcium and phosphate reduce the lead
absorption, whereas a lack of zinc, ascorbic acid or citric acid in diet enhance it. The *provisional tolerable weekly intake* (PTWI) of lead from drinking water and diets by an adult human being has been established by WHO (World Health Organisation) to be 0.24 mmol kg⁻¹ week⁻¹ [83].

After intake, lead is transported to the blood plasma and within few minutes it is transferred to the erythrocytes. Only a small amount (6%) remains dissolved in the plasma, from where it is slowly (with a half-time of 2-3 weeks) distributed to the soft tissues, teeth and skeleton [81, 83]. More than 90% of the total body lead is stored in the skeleton, from where it can be slowly released back into the blood with the consequence of chronic intoxication [86]. Indicators of human exposure are the lead levels in blood and urine (recent exposure), and in bones and teeth (long-term exposure).

*Toxicity*

Lead is a non-essential element with toxic effects on the human health. At biochemical level lead is capable to inhibit enzymes at several steps in the haem pathway and disturb the porphyrin metabolism. Moreover, it is an effective substitute for calcium and influences many calcium-dependent processes [81].

The toxic effects on humans can be both acute or chronic. Acute lead poisoning, which usually arises from the accidental ingestion of a soluble lead salt, is initially associated with anorexia, abdominal pain and vomiting, eventually followed by renal damage, encephalitis and coma. Chronic poisoning, usually due to accumulation of small quantities by inhalation, ingestion or skin absorption, may cause anemia, weakness, headache and the development of a black line in the gums. Long-term exposure can affect the peripheral nervous system with paralysis in the more advanced cases. Irreversible renal damages are also possible, even a long time after the abnormal lead exposure has terminated [81, 83].
6.2 Requirements for low detection limit and selectivity

The U.S. Environmental Protection Agency (EPA) demands that total lead in drinking water should not exceed 15 ppb (7.2 x 10^{-8} M [87]; action limit). Since the detection limit of a reliable analytical technique is expected to be at least ten times lower than the action limit, conventional ISE with a detection limit of about 10^{-6} M have never been considered to be suitable for the analysis of heavy metal ions in drinking and environmental water samples. Until now, the most widely used techniques for the determination of lead in liquid samples were graphite furnace or hydride generation atomic absorption spectroscopy (GF-AAS and HG-AAS), inductively coupled plasma combined with mass spectrometry (ICP-MS) or atomic emission spectrometry (ICP-AES) [88-91], and several voltammetric methods as polarography, anodic stripping voltammetry (ASV), and square wave anodic stripping voltammetry (SWASV) [85, 92-95]. Only recently, the lower detection limit of ISEs have been lowered by many orders of magnitude [30, 39], so that they may become an interesting alternative to other analytical methods. In fact, with the exception of the detection limit, ISEs have various advantages in comparison with other methods. For example, they allow in situ measurements, are not expensive, can be easily miniaturized, do not require sample preparation and are not redox-sensitive. Moreover, they allow in many cases the speciation of the investigated ions.

In absence of the ion fluxes through the membrane and the sensed aqueous layer, the thermodynamically defined lower detection limit of an ISE membrane is solely dictated by the selectivity coefficients and the concentration of interfering ions in the sample. In this case, the sensitivity to an ion is lost when it is completely replaced by an interfering ion in the phase boundary region of the membrane. The corresponding so-called static detection limit, $a_i(DL,\text{static})$, is given by equation 3.16. It is obvious that
trace level determinations in the presence of millimolar ionic backgrounds are only possible with highly selective membranes. This is the case of environmental samples (lake, river and drinking water) where the Ca\(^{2+}\), Na\(^{+}\), Mg\(^{+}\), and K\(^{+}\) concentrations are normally between 10\(^{-4}\) and 10\(^{-3}\) M, while the Pb\(^{2+}\) concentrations are at least five or six orders of magnitude lower.

Unfortunately, at submicromolar analyte concentrations, the static lower detection limits predicted by equation 3.16 are too optimistic, since in dilute solutions an exchange of just a small fraction of membrane analyte ions with sample interfering ions leads to a slight concentration polarization that induces a flux of primary ions from the membrane into the sample. As a consequence, the analyte concentration in the sensed aqueous surface layer becomes higher than in the sample bulk. Recently, a rather simple implicit equation has been derived by Prof. E. Bakker [78, 96] to predict the required selectivities in such situations (dynamic detection limits). The equation can be used to calculate the expected detection limit as a function of the concentration of any number of monovalent and divalent interfering ions in the sample. Here, the only simplified explicit relationships between the detection limit and the selectivity coefficient are given for systems containing just one interfering ion of chosen valency:

- primary and interfering ion have the same charge:

\[
\log c_{i}(DL) = \frac{1}{2} \log \left( \frac{1}{z_i} K_{i,j}^{\text{pot}} c_{j,S} q R_{i} \right)
\]  

(6.1)

- divalent primary and monovalent interfering ion:

\[
\log c_{i}(DL) = \frac{1}{3} \log \left( K_{i,j}^{\text{pot}} \left[ \frac{1}{2} c_{j,S} q R_{i} \right]^2 \right)
\]  

(6.2)
The coefficient $q$ incorporates the ratio of the diffusion coefficients of the analyte ion in both membrane and sample phases ($D_M$ and $D_S$) and the respective diffusion layer thicknesses ($\delta_M$ and $\delta_S$)

$$q = \frac{D_M \delta_S}{D_S \delta_M} \quad (6.3)$$

From equations 6.1 and 6.2 it is immediately evident that the concentration of ionic sites should be preferably low, ions should diffuse slowly in the organic phase, and selectivity coefficients must be small to reach low detection limits. Table 6.1 shows the measured and the required selectivity coefficients according to both the static and the dynamic model in order to obtain a lower detection limit of $7.2 \times 10^{-9}$ M Pb$^{2+}$. A value of $q = 0.001$ is thereby used, which is obtained with the diffusion coefficient of Pb$^{2+}$ in water ($D_S = 9.5 \times 10^{-6}$ cm$^2$ s$^{-1}$ [73]) and the estimated one of the ionophore complex in the membrane ($D_M = 10^{-8}$ cm$^2$ s$^{-1}$ [74]), and equal thickness of the diffusion layer in the aqueous and organic phases. Evidently, according to the new dynamic model, the required selectivity coefficients are much smaller than those obtained on the basis of the static model. It is important to remember that equations 6.1 and 6.2 are based on the assumption that the ion-exchange process is incomplete; i.e. only a few percent of the analyte ions are displaced by interfering ions in the membrane. Therefore, they are only valid when the predicted lower detection limit is at least one order of magnitude higher than the value calculated on the basis of equation 3.16.
Table 6.1. Required Selectivity coefficients for a lead ISE to achieve a 1.5 ppb (7.2 nM) detection limit in Zurich drinking water and comparison with experimental values.

<table>
<thead>
<tr>
<th>Ion J</th>
<th>conc. / M</th>
<th>( \log K_{p_{b,j}}^{\text{pot}} ) required</th>
<th>measured ( \log K_{p_{b,j}}^{\text{pot}} )</th>
<th>Expected ( c_{p_{b}}(DL) / M )</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Na}^+ )</td>
<td>( 3.5 \times 10^{-4} )</td>
<td>-1.2 ( \text{eq. 3.16} )</td>
<td>-3.9 ( \text{eq. 6.1-6.2} )</td>
<td>-6.3 ± 0.2</td>
</tr>
<tr>
<td>( \text{K}^+ )</td>
<td>( 3.8 \times 10^{-5} )</td>
<td>+0.7</td>
<td>-1.9</td>
<td>-6.3 ± 0.1</td>
</tr>
<tr>
<td>( \text{H}^+ )</td>
<td>( 1.8 \times 10^{-8} )</td>
<td>+7.3</td>
<td>+4.7</td>
<td>-3.5 ± 0.3</td>
</tr>
<tr>
<td>( \text{Ca}^{2+} )</td>
<td>( 1.6 \times 10^{-3} )</td>
<td>-5.3</td>
<td>-6.6</td>
<td>-12.3 ± 0.1</td>
</tr>
<tr>
<td>( \text{Mg}^{2+} )</td>
<td>( 4.9 \times 10^{-4} )</td>
<td>-4.8</td>
<td>-6.1</td>
<td>-12.4 ± 0.2</td>
</tr>
<tr>
<td>( \text{Cu}^{2+} )</td>
<td>( 4.7 \times 10^{-8} )</td>
<td>-0.8</td>
<td>-2.1</td>
<td>-3.7 ± 0.2</td>
</tr>
<tr>
<td>( \text{Cd}^{2+} )</td>
<td>( 8.9 \times 10^{-10} )</td>
<td>+0.9</td>
<td>-0.4</td>
<td>-5.4 ± 0.1</td>
</tr>
</tbody>
</table>

6.3 Optimization of the experimental technique for trace level measurements of \( \text{Pb}^{2+} \)

In the previous two chapters, the relevant effects that dictate the lower detection limits and the selectivity of polymeric membrane ISEs have been considered. It has been shown that potentiometric measurements at trace levels are possible if very selective ionophores are available and if the inner filling solution and the membrane parameters are carefully adjusted. Now it is demonstrated that ISEs based on solvent polymeric membrane can be used for the routine determination of nanomolar concentrations of \( \text{Pb}^{2+} \), even in samples containing a high background concentration of interfering ions, such as \( \text{H}^+ \), \( \text{Na}^+ \) or \( \text{Ca}^{2+} \).
Membrane and inner filling solution

First, the compositions of the ISE membrane and the inner electrolyte were optimized in order to minimize ion fluxes across the membrane. Ideally, for a perfect performance of the ISE any change in the sample composition should only induce an insignificant flux of lead ions across the sample/membrane interface. Generally, this could only be achieved with an inner electrolyte composition that in each case perfectly matches the sample, but unfortunately the sample composition is variable and not known before the measurement. Here, for practical reasons, the same inner filling solution is used for all sample concentrations (cf. Section 6.8). Its activity was chosen to be close to the required detection limit of $7.2 \times 10^{-9}$ M Pb$^{2+}$. To increase the stability of membrane potential at the inner side, the low activity of the free Pb$^{2+}$ in the inner compartment (ca. $1.3 \times 10^{-9}$ M) was kept constant by using an NTA buffer. With the equations given in Section 5.2, the membrane weighing parameters, and the selectivity coefficients (Table 6.1), it can be calculated that with this inner filling solution less than 5% of Pb$^{2+}$ is replaced by Na$^+$ at the inner membrane surface, so that the concentration gradients in the membrane remain small. Additionally, according to equations 6.1 and 6.2, it is possible to minimize the ion fluxes by preparing membranes that contain a small concentration of ionic sites, have a larger thickness or exhibit a smaller diffusion coefficient of the lead ionophore. In the present application, only the amount of lipophilic anionic sites was reduced but the other parameters were not changed in order to not increase the conditioning times of the membrane. The reduction of the anionic concentration sites yielded membrane selectivities that were not as good as with the unmodified composition (cf. Chapter 4) but still adequate for the present application. If necessary, it is possible to improve the selectivity without enhancing ion fluxes by increasing only the concentration of the ionophore but not of the sites.
**Conditioning solution**

Before use, the ISE membranes were conditioned overnight in a $5 \times 10^{-9}$ M Pb$^{2+}$ solution of pH 4 that contained a background of $10^{-3}$ M CaCl$_2$. A conditioning with such a dilute Pb$^{2+}$ solution, which is close to the target level, was only possible because the membrane cocktails contained the right amount of lead ions already before the conditioning. Otherwise, with this solution, too long conditioning times would be required to quantitatively exchange the counter ions (Na$^+$) of the lipophilic anion in the membrane. The conventional conditioning procedure (overnight in $10^{-2}$ M Pb$^{2+}$ solutions) was avoided in order to minimize the coextraction of Pb(NO$_3$)$_2$ from the conditioning solution into the membrane and to prevent the formation of PbCO$_3$ and Pb(OH)$_2$ crystals on the membrane surface. Calculations based on equations 5.3-5.6 and on the selectivity coefficients of Table 6.1, indicate that with this conditioning solution less than 3% of the lead ions in the membrane are exchanged by the H$^+$ or Ca$^{2+}$ present in the aqueous phase. Since after such a conditioning procedure no reequilibration of the membrane bulk is required, the ISE has a fast and Nernstian response in sample having a higher Pb$^{2+}$ concentration (cf. Figures 5.10 and 5.11).

**Measuring procedure**

To keep the overall measuring times short, calibrations were made at relatively high concentrations ($7 \times 10^{-6}$-1 x $10^{-4}$ M). In fact, as demonstrated in Figure 5.10, the response of ISEs is much faster at these concentrations than at lower ones. As shown in Figure 6.3, after achieving a stable EMF in the sample, the five-point calibration was taken within few minutes. In addition to shortening the measurement cycle, a fast calibration has the advantage that the contact between membrane and calibration solution is short, so that the coextraction of Pb(NO$_3$)$_2$ into the membrane is minimized. If the membranes had a longer contact with these relatively concentrated
calibration solutions, the \textit{EMF} recovery with the next dilute sample would take a much longer time due to the release of the coextracted lead ions. The recovery time after changing from the last calibration solution to a new sample is shown in Figure 6.3. It is evident that even during the short contact with the calibration solutions some changes in the membrane or at the membrane surface occur, so that the membrane needs several minutes to reequilibrate again with a dilute sample.

\textbf{Figure 6.2.} Determination of the \( \text{Pb}^{2+} \) activities in samples containing \( 10^{-3} \) M \( \text{CaCl}_2 \) and \( 10^{-4} \) M \( \text{HNO}_3 \) as ionic background by calibrating with a series of lead ion solutions of higher concentrations. The known logarithmic lead activities of the solutions are indicated by the numbers and by the vertical arrows. The measured values correspond to the intersection of calibration line with the horizontal line indicating the \textit{EMF} obtained in the sample solution.
To minimize measurement errors, the concentration of the calibration solutions is usually close to that in the sample. In this work, for the reasons explained above, the concentration of the sample and calibration solutions can differ by many orders of magnitude. However, the results in Figure 6.2 show that the slope of the calibration curves is highly reproducible, and therefore, the results are accurate despite the large difference in the activities between the samples and the calibration solutions.

**Figure 6.3.** Time traces for two of the experiments shown in Figure 6.2. The concentrations of the sample and calibration solutions are indicated by the numbers. The recovery time of the ISE membrane after calibration is longer for the more dilute sample.
Table 6.2: Known and found Pb\(^{2+}\) activities for the experiments shown in Figure 6.2. The standard deviations are higher for more dilute samples probably because of the larger concentration difference between sample and calibration solutions.

<table>
<thead>
<tr>
<th>spiked Pb(^{2+}) amount</th>
<th>found Pb(^{2+}) amount</th>
</tr>
</thead>
<tbody>
<tr>
<td>log activity</td>
<td>activity</td>
</tr>
<tr>
<td>-9.13</td>
<td>7.5 x 10(^{-10})</td>
</tr>
<tr>
<td>-8.43</td>
<td>3.7 x 10(^{-9})</td>
</tr>
<tr>
<td>-8.13</td>
<td>7.5 x 10(^{-9})</td>
</tr>
<tr>
<td>-7.43</td>
<td>3.7 x 10(^{-8})</td>
</tr>
<tr>
<td>-7.13</td>
<td>7.5 x 10(^{-8})</td>
</tr>
<tr>
<td>-6.43</td>
<td>3.7 x 10(^{-7})</td>
</tr>
</tbody>
</table>

6.4 Proton interference at low Pb\(^{2+}\) concentrations

In untreated natural water samples of about pH 8 Pb\(^{2+}\) is predominantly complexed by carbonate or hydroxide. Consequently, the potentiometric assessment of total lead is only possible after acidification of the sample to pH 4 or lower (cf. Figure 6.1).

The data in Table 6.1 indicate that at this pH proton is the main interferent, and that the practical (dynamic) lower detection limit is about 4 x 10\(^{-9}\) M, due to a minute Pb\(^{2+}\)/H\(^{+}\) ion exchange at the outer membrane surface. Experimentally, the pH dependence of the Pb\(^{2+}\)-ISE response was investigated by varying the HNO\(_3\) concentration of a solution containing 5 x 10\(^{-8}\) M Pb(NO\(_3\))\(_2\) (EPA action level) and a background of 10\(^{-3}\) M CaCl\(_2\). From Figure 6.4, it is evident that the EMF of the Pb\(^{2+}\)-ISE is not influenced by pH changes between 5.0 and 3.9. At lower pH, the potential increases just after addition of HNO\(_3\) and then slowly decreases to the initial potential. Only at pH < 3 the potential recovery is incomplete and a net potential shift can be observed. At pH 2.3, the potential increases by about 8 mV, which is about 6
mV higher than predicted with equation 6.3. However, it is not possible to exactly compare the two values as the measured potential had still a slow negative drift.

![Graph showing EMF response](image)

**Figure 6.4.** Time traces of the EMF response of a lead selective electrode in contact with a $5 \times 10^{-8}$ M Pb$^{2+}$ sample by lowering the sample pH. The respective time trace of a commercial pH glass electrode is shown as reference.

Figure 6.4 confirms that the H$^+$ selectivity of the ISE-membrane is sufficient to assess Pb$^{2+}$ at the action limit concentration of 15 ppb, and that small variations of the pH do not influence the Pb$^{2+}$ measurements. This result is important, because it is very difficult to precisely adjust the pH of environmental water samples without using highly concentrated buffers.
6.5 Measurements of Pb\textsuperscript{2+} in environmental samples

The applicability of ISEs and the proposed measuring and conditioning procedures for the determination of total lead concentrations in environmental water samples were evaluated by comparing the results obtained with ISE and ICPMS. The measurement procedure was carried out as described above, but Pb\textsuperscript{2+} concentrations instead of activities were used for calculations. The obtained results are not biased since the ionic strength of the sample during the calibration did not change, so that all the Pb\textsuperscript{2+} concentrations were proportional to the activities. The investigated samples ranged from tap, lake, river, and rainwater of Zurich (Switzerland), tap water of Bellinzona, (Switzerland), two bottled water sources (of the commercial brands Evian and Cristalp), and tap water of different sources in Budapest. These last two samples were obtained from old private homes with original lead plumbing still in place, and do not represent the average drinking water quality of Budapest. The results are presented in Table 6.3.

According to ICPMS analysis, the tap water of Zurich containes about 5 ppb lead, a value that comes very close to the concentration assessed with ISEs. Especially noteworthy is the fact that this sample contains more than 50-fold molar excess of copper over lead, which appear to have little influence on the potentiometric results. According to equations 6.2 and 6.3 and Table 6.1, H\textsuperscript{+} and Cu\textsuperscript{2+} are indeed the major interfering agents in this measurement. This sample was subsequently spiked with known concentration of lead, and ISEs were capable of accurately recovering the added lead concentrations in the range of $10^-8$ and $10^-7$ (see Table 6.5, first three lines). Clearly the ISE can be used as an adequate analytical technique to detect unacceptable concentrations of Pb\textsuperscript{2+} (over $7.2 \times 10^-8$ M; EPA action value).
Table 6.3. Total lead concentrations determined by direct potentiometry in environmental samples and comparison to ICPMS analysis.

<table>
<thead>
<tr>
<th></th>
<th>Pb$^{2+}$-ISE</th>
<th>Pb-ICPMS</th>
<th>Cu-ICPMS</th>
<th>Cd-ICPMS</th>
</tr>
</thead>
<tbody>
<tr>
<td>Zurich tap</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>0.71 ± 0.04</td>
<td>0.54 ± 0.07</td>
<td>6.70 ± 0.08</td>
<td>0.07 ± 0.01</td>
</tr>
<tr>
<td>nM</td>
<td>3.4 ± 0.2</td>
<td>2.6 ± 0.3</td>
<td>104 ± 1</td>
<td>0.7 ± 0.1</td>
</tr>
<tr>
<td>Zurich lake</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>0.78 ± 0.05</td>
<td>0.23 ± 0.01</td>
<td>4.35 ± 0.07</td>
<td>0.02 ± 0.01</td>
</tr>
<tr>
<td>nM</td>
<td>3.8 ± 0.2</td>
<td>1.1 ± 0.1</td>
<td>69 ± 1</td>
<td>0.2 ± 0.1</td>
</tr>
<tr>
<td>Glatt river (ZH)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>0.96 ± 0.08</td>
<td>0.62 ± 0.01</td>
<td>5.93 ± 0.07</td>
<td>0.02 ± 0.01</td>
</tr>
<tr>
<td>nM</td>
<td>4.6 ± 0.4</td>
<td>3.0 ± 0.1</td>
<td>93 ± 1</td>
<td>0.2 ± 0.1</td>
</tr>
<tr>
<td>Zurich rain</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>5.02 ± 0.20</td>
<td>5.12 ± 0.05</td>
<td>40.4 ± 0.02</td>
<td>0.10 ± 0.01</td>
</tr>
<tr>
<td>nM</td>
<td>24.0 ± 1.0</td>
<td>24.7 ± 0.2</td>
<td>636 ± 1</td>
<td>0.9 ± 0.03</td>
</tr>
<tr>
<td>Bellinzona tap</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>0.70 ± 0.06</td>
<td>0.09 ± 0.01</td>
<td>8.66 ± 0.06</td>
<td>&lt;0.02</td>
</tr>
<tr>
<td>nM</td>
<td>3.4 ± 0.3</td>
<td>0.72 ± 0.08</td>
<td>136 ± 1</td>
<td>&lt;0.2</td>
</tr>
<tr>
<td>Cristalp (bottle)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>1.06 ± 0.09</td>
<td>0.01 ± 0.01</td>
<td>0.34 ± 0.01</td>
<td>&lt;0.02</td>
</tr>
<tr>
<td>nM</td>
<td>5.1 ± 0.4</td>
<td>&lt;0.1 ± 0.1</td>
<td>5.3 ± 0.2</td>
<td>&lt;0.2</td>
</tr>
<tr>
<td>Evian (bottle)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>0.74 ± 0.05</td>
<td>0.02 ± 0.01</td>
<td>1.00 ± 0.03</td>
<td>&lt;0.02</td>
</tr>
<tr>
<td>nM</td>
<td>3.6 ± 0.4</td>
<td>&lt;0.1 ± 0.1</td>
<td>15.7 ± 0.5</td>
<td>&lt;0.2</td>
</tr>
<tr>
<td>Budapest tap 1</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>2.30 ± 0.8</td>
<td>0.92 ± 0.01</td>
<td>45.2 ± 0.1</td>
<td>&lt;0.02</td>
</tr>
<tr>
<td>nM</td>
<td>11 ± 4.0</td>
<td>4.4 ± 0.1</td>
<td>712 ± 1</td>
<td>&lt;0.2</td>
</tr>
<tr>
<td>Budapest tap 2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ppb</td>
<td>10.0 ± 1.0</td>
<td>2.59 ± 0.02</td>
<td>1300 ± 20</td>
<td>&lt;0.02</td>
</tr>
<tr>
<td>nM</td>
<td>49 ± 5.0</td>
<td>12.5 ± 0.1</td>
<td>20300 ± 300</td>
<td>&lt;0.2</td>
</tr>
</tbody>
</table>

Other water samples from Zurich included lake, river and rainwater. The ISEs showed acceptable performances also in these cases. The assessed lead concentrations never exceeded 5 x 10^{-9} M and were in reasonable agreement with ICPMS values. The next three samples shown in Table 6.3 (Bellinzona tap water and the bottled water samples) according to the ISE again do not exceed 5 x 10^{-9} M Pb$^{2+}$. However, ICPMS values were one or two order of magnitude lower, showing that the ISE results are biased by the proton interference, as expected from equation 6.2.
Table 6.4. Total lead concentrations determined by direct potentiometry in natural and spiked Zurich tap water spiked with Pb\(^{2+}\) and/or Cu\(^{2+}\) and comparison to ICPMS analysis.

<table>
<thead>
<tr>
<th>spiked Pb and Cu amount</th>
<th>found Pb and Cu amount</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Pb(^{2+})-ISE</td>
</tr>
<tr>
<td>natural sample</td>
<td>0.71 ± 0.04 ppb</td>
</tr>
<tr>
<td></td>
<td>3.4 ± 0.2 nM</td>
</tr>
<tr>
<td>2.5 ppb 6.70 ppb</td>
<td>2.61 ± 0.06 ppb</td>
</tr>
<tr>
<td>12.0 nM 104 nM</td>
<td>12.6 ± 0.3 nM</td>
</tr>
<tr>
<td>10.5 ppb 6.70 ppb</td>
<td>11.5 ± 0.3 ppb</td>
</tr>
<tr>
<td>51.0 nM 104 nM</td>
<td>55.6 ± 1.5 nM</td>
</tr>
<tr>
<td>20.5 ppb 6.70 ppb</td>
<td>22.0 ± 1 ppb</td>
</tr>
<tr>
<td>99.0 nM 104 nM</td>
<td>107 ± 12 nM</td>
</tr>
<tr>
<td>2.5 ppb 507 ppb</td>
<td>6.3 ± 0.4 ppb</td>
</tr>
<tr>
<td>12.0 nM 7.9 nM</td>
<td>30 ± 2 nM</td>
</tr>
<tr>
<td>2.5 ppb 1007 ppb</td>
<td>9.8 ± 0.8 ppb</td>
</tr>
<tr>
<td>12.0 nM 15.8 nM</td>
<td>47 ± 4 nM</td>
</tr>
<tr>
<td>2.5 ppb 1507 ppb</td>
<td>13.9 ± 1.5 ppb</td>
</tr>
<tr>
<td>12.0 nM 23.7 nM</td>
<td>67 ± 7 nM</td>
</tr>
</tbody>
</table>

The two tap water samples from Budapest contained higher lead concentrations than the other drinking water samples, even though none of them exceeded the action limit imposed by EPA. Especially interesting is the sample Budapest tap 2 containing a very high copper concentration (2 x 10\(^{-5}\) M), which appears to be the main interferent in the ISE measurement. Indeed, for samples containing this Cu\(^{2+}\) concentration, equation 6.1 predicts a lower lead detection limit of 2.3 x 10\(^{-8}\) M, which is twice as high than the value obtained with ICPMS. By adding the concentration predicted by the detection limit to the value determined with ICPMS one obtains a lead concentration close to the value suggested by ISE (4.9 x 10\(^{-8}\) M). The copper interference was further evaluated by spiking Zürich tap water with 2 ppb.
lead and various levels of copper, in order to mimic the Budapest sample (see Table 6.4, last three lines). The bias in the ISE measurements was very similar to that observed with the native Budapest sample (see Table 6.3, last line) and suggests that the interference model employed here is adequate despite its simplicity. Although the copper concentration is here exactly at the current action limit (1.3 ppm) [Agency, 1999 #1342], the bias in the ISE measurement is not large enough to indicate a lead concentration above the lead action limit (7.2 \times 10^{-8} \text{ M}). It means that copper may interfere substantially to give a false lead reading above the lead action limit only if copper itself has a higher concentration than the copper action limit. In practice, therefore, a simple potentiometric screening would reliably reveal a real quality problem with drinking water.

6.6 Speciation of low total Pb^{2+} concentrations in drinking water

In untreated natural water samples of about pH 8, Pb^{2+} is predominantly complexed by carbonate (cf. Figure 6.1). The measured pH dependence of the free Pb^{2+} activity in the unspiked drinking water of Zurich is shown in Figure 6.5. It clearly demonstrates that at pH 7.9 (native water) a few percent of the total Pb^{2+} concentration, i.e., about 2 \times 10^{-10} \text{ M}, could be accurately determined. This result does not contradict the previous considerations of dynamic lower detection limits, since at high pH the H^{+} interference is eliminated, so that the lower detection limit is caused by other interfering ions. In addition, due to the presence of carbonate, the effective concentration of Pb^{2+} and Cu^{2+} in the sensed aqueous layer is buffered, and the fluxes of complexed Pb^{2+} are bigger because of the high carbonate concentration [97]. The total Pb^{2+} concentration was measured with the same ISE, but at pH 4. The vertical arrow in the plot shows the lead concentration determined by ICPMS.
Figure 6.5. Potentiometric determination of Pb\(^{2+}\) in unspiked Zürich tap water at two pH values: pH 7.9 (native pH) and pH 4.0 (acidified sample). Measured potentials are indicated by horizontal dotted lines for the sample and by dots for the calibration solutions. The vertical arrow shows the lead concentration obtained with ICPMS.

The fact that low lead activities can be reliably specified is corroborated by Figure 6.6, which shows the calculated pH dependence of the free Pb\(^{2+}\) activity for two different carbonate concentrations together with the experimental values. For this measurement the Zürich tap water was spiked with 10 ppb Pb\(^{2+}\). Some deviations between the calculated and measured speciation curves can be caused, in addition to the experimental errors, by the magnitude of the used formation constants. Depending on the ionic strength and the literature source, the published complex formation constants of Pb\(^{2+}\)
with carbonate and bicarbonate deviate from each other up to one order of magnitude [56, 57, 85, 98]. Additionally, the total concentration of carbonate varies at different pHs and, therefore, during the titration with HNO₃ does not remain constant.

\[ \text{calculated for} \quad [\text{CO}_3^-]_{\text{tot}} = 2.48 \text{ mmol/l} \]

\[ \text{calculated for} \quad [\text{CO}_3^-]_{\text{tot}} = 4.14 \text{ mmol/l} \]

**Figure 6.6.** Direct potentiometric speciation of Pb²⁺ in Zürich tap water spiked with 10 ppb Pb²⁺ as function of the pH. Solid lines: calculated values considering the complex equilibria of Pb²⁺ given in the experimental part (Section 6.8, EMF measurements). The theoretical curves are given for a maximal and minimal CO₃ concentration in the water (data from web sources: www.WVZ.ch, Wasserversorgung Zürich). Errors bar are the standard deviation of measurements with 6 ISE.

The bias in the upper part of the calculated curves strongly depends on the total SO₄²⁻ concentration. Unfortunately, an accurate simulation of SO₄²⁻...
influence on the Pb\(^{2+}\) speciation is very difficult, since the free SO\(_4^{2-}\) amount depends on the concentration of other ions, i.e., Ca\(^{2+}\), Mg\(^{2+}\), and H\(^+\). Moreover, comparing the errors, one must remember that the ISE response is linear to the logarithm of the activity. Therefore, the potential difference between 1% and 10% of free Pb\(^{2+}\) (high pH) is 29.6 mV, while at low pH, between 50% and 100% of free lead, it is only 9 mV. For these measurements the inner filling solution was exactly the same as the conditioning solution, i.e., it contained 5 \(\times\) 10\(^{-9}\) M Pb(NO\(_3\))\(_2\), (without NTA buffer, see Section 6.7) at pH 4 in a background of 10\(^{-3}\) M CaCl\(_2\). The disadvantage of such an inner solution is that the Pb\(^{2+}\) concentration in it can change with time and may cause potential drifts.

### 6.7 Diffusion of NTA complexes through the ISE membrane

Surprisingly, the presence of ion buffers such as NTA and EDTA in the inner solution can induce potential shifts and drifts when the pH or the ionic strength of the sample is changed. This effect is further investigated in the present section. Figure 6.7 shows time traces of the responses of two ISEs differing only in their inner filling solutions. In the inner solution of ISE I (solid lines) the total Pb\(^{2+}\) concentration is 10\(^{-4}\) M and the free ion activity buffered with 10\(^{-3}\) M NTA is ca. 2 \(\times\) 10\(^{-9}\) M, while ISE II (dotted lines) contains a total Pb\(^{2+}\) concentration of 5 \(\times\) 10\(^{-9}\) M and no ion buffer.

The electrode potentials are first measured in a sample containing 5 \(\times\) 10\(^{-8}\) M Pb\(^{2+}\) and a background of 10\(^{-3}\) CaCl\(_2\) at pH 4 (sample A). Both ISEs exhibit good EMF stabilities. After 30 min sample A is replaced by B, which is tap water of Zürich acidified to pH 4 and spiked with 5 \(\times\) 10\(^{-8}\) M Pb\(^{2+}\). No EMF difference is expected between the two samples because they have the same Pb\(^{2+}\) activities and the interfering ions in tap water are sufficiently discriminated (cf. Chapter 6.2). However, ISE II that has NTA in the inner solution shows a negative potential shift of about 30 mV. It is evident that
the potential of ISE II does not only depend on the Pb\(^{2+}\) activity, but also on the sample composition so that an external calibration of the sensor is not feasible.

![Graph showing time traces of ISE responses](image)

**Figure 6.7.** Time traces of the response of two ISEs that differ only in the composition of the inner filling solution.

After reaching constant potentials in sample B, the measuring solution is changed to sample C, which is also tap water of Zürich spiked with 5 \(\times\) 10\(^{-8}\) M Pb\(^{2+}\), but its pH is 8 instead of 4. During the first two minutes both ISE show the same behavior, but then ISE II (with NTA) begins to drift toward a more positive potential. After 30 min both ISE responses were stable, but the potential difference between them was higher than in sample B. These
experiments indicate that the response of ISE II is somehow influenced by the pH of the sample. Finally, the \( \text{EMF} \) is recorded in acidified (pH 4) tap water of Zürich, spiked with \( 5 \times 10^{-5} \text{ M Pb}^{2+} \). The \( \text{EMF} \) difference between samples B and D is similar with both electrodes (83.7 for ISE I, 81.1 for ISE II), which are close to the theoretically expected value. It can be concluded that ISE containing NTA in the inner solution can be used if sample and calibration solutions have the same pH and ionic background. If this is not the case, the measured values can be biased by potential drifts. These drifts are likely to be caused by the diffusion of NTA species (e.g. PbHNTA, \( \log K = 15.2 \) [99]) through the membrane. The use of ion-exchanger resins instead of organic molecules as buffering agents may eliminate this effect and improve the performance of the ISEs.

6.8 Silver interference

The selectivity values measured in Chapter 4 indicate that ISEs based on the thioamide functionalized calix[4]arene strongly prefer \( \text{Ag}^+ \) over \( \text{Pb}^{2+} \). However, ion selectivities are determined at relatively high salt concentrations and, consequently, under steady-state conditions. Here it is shown that in ISEs conditioned with \( \text{Pb}^{2+} \) the interference of small \( \text{Ag}^+ \) concentrations occurs much later than predicted with the steady-state selectivity model (Equation 3.16). In the experiment, the \( \text{Pb}^{2+} \) conditioned ISE was first calibrated with \( \text{Pb}^{2+} \) solutions (Figure 6.8, right) to check its operation. Then, the same ISE was used to monitor the potential of a sample containing \( 5 \times 10^{-8} \text{ M Pb}^{2+} \) to which \( \text{Ag}^+ \) was stepwise added. According to the steady-state model [15] (dotted line), for a selectivity coefficient \( K_{\text{Pb,Ag}}^{\text{pot}} = +13.2 \) a response to \( \text{Ag}^+ \) was expected above \( 10^{-11} \text{ M Ag}^+ \). However, as seen in Figure 6.8, no important \( \text{EMF} \) changes are observed before a \( 10^{-8} \text{ M Ag}^+ \) concentration was reached.
The experimental observations can be explained if ion fluxes through the membrane and the aqueous layer are considered. At the sensor interface silver ions enter the membrane by exchanging with lead ions and diffuse into the bulk of the membrane. As a consequence the concentration of the aqueous \(\text{Ag}^+\) in the proximity of the membrane decreases below the values given by \(K_{\text{Ag},\text{Pb}}^{\text{bot}} a_{\text{Pb}}^{1/2}\), so that the ISE continues responding to \(\text{Pb}^{2+}\). Only when the \(\text{Ag}^+\) concentration in the bulk of the sample is high enough (>\(10^{-8}\) M) depleation at the phase boundary become neglectable, and the potential increases abruptly. This behaviour is related to the so-called Hulanicki effect [100], earlier described for anion ISEs.
Differently from the situation described in the previous part of the work (cf. Table 6.1), where concentration gradients biased the lower detection limit of the Pb$^{2+}$-ISE, in the Pb$^{2+}$/Ag$^+$ system investigated here ion fluxes improve the performance of the Pb$^{2+}$-ISE. In fact, in absence of ion fluxes the response to Ag$^+$ or Pb$^{2+}$ would be solely dictated by the selectivity coefficient, that predicts Ag$^+$-interference at much lower concentrations than those measured in the experiment. Unexpected selectivity problems could arise when trying to miniaturise sensors based on carrier with the same selectivity behaviour of the lead ionophor. Miniaturised sensors can have the advantage that they need smaller amounts of analyte and, as a consequence, smaller samples volumes and shorter equilibration times with the sample. However, they work at equilibrium conditions and static interference (cf. Section 6.2) is expected. Consequently, in a miniaturised membrane, the ionophore used in this work would be useful only for the detection Ag$^+$, but not for low concentrations of Pb$^{2+}$.

6.9 Long term potential stability

The EMF of the Pb$^{2+}$ ISE have been monitored during 4 days in a sample containing 5 x 10$^{-8}$ M Pb$^{2+}$, 10$^{-3}$ M CaCl$_2$ and 10$^{-4}$ m HNO$_3$. As shown in Figure 6.9 the potential changes were lower than 5 mV, indicating a good stability of the membrane and of the inner filling solution.
Figure 6.9. Potential of 3 ISEs with membranes of the same composition measured during 4 days in a 500 ml solution containing \(5 \times 10^{-8} \text{ M} \ \text{Pb(NO}_3\text{)}_2, 10^{-3} \text{ M CaCl}_2, 10^{-4} \text{ M HNO}_3\). Time = 0 corresponds to the unconditioned dry membrane.
6.10 Experimental Section

**Reagents.** Poly(vinyl chloride) (PVC), bis(2-ethylhexyl)sebacate (DOS), sodium tetrakis[3,5-bis(trifluoromethyl)phenyl]borate (NaTFPB), the ionophore 4-<i>tert</i>-butylcalix[4]arene-tetrakis(thioacetic acid dimethylamide) (lead ionophore IV), the lipophilic salt tetradecylammonium tetrakis(4-chlorophenyl)borate (ETH 500) and tetrahydrofuran (THF) were from Fluka AG (CH-8071 Buchs, Switzerland). Aqueous solutions were prepared with freshly deionized water (18.0 MΩ cm specific resistance) obtained with a NANOpure® reagent-grade water system (Barnstead, CH-4009 Basel, Switzerland). HNO₃, NaCl, KCl, CaCl₂ and MgCl₂ were of Suprapur grade from Merck (Darmstadt, Germany). Nitrilotriacetic acid (NTA) was Microselect purity, and the other used salts were of puriss. p.a. quality of Fluka. The Pb²⁺ standard solution with a concentration of 1000 mg l⁻¹ in 0.001 M HNO₃ was obtained from Fluka Trace analysis.

**Membranes.** The membranes contained 0.07 wt % (0.68 mmol kg⁻¹) lead ionophore IV, 0.03 wt % (0.30 mmol kg⁻¹) NaTFPB, 1.32 wt % (11.50 mmol kg⁻¹) ETH 500, 62.06 wt % DOS, 36.52 wt % PVC, and 0.15 mmol kg⁻¹ Pb(NO₃)₂. The membrane was prepared by dissolving 417 mg of membrane components (except Pb(NO₃)₂) in 5 mL THF. The lead salt was successively added to the membrane cocktail as 62 µL of an aqueous 10⁻³ M Pb(NO₃)₂ solution. This membrane solution was stirred for about 1 h. A membrane of ca. 200 µm was obtained by casting the solution into a glass ring (7.0 cm i.d.) fixed on a glass plate. The membrane used for the measurements of the selectivity coefficients had the same composition, but without Pb(NO₃)₂.

**Electrodes.** A disk of 4 mm diameter was punched from the above ion-selective membrane and glued to a plasticized PVC tubing with a THF/PVC slurry. The internal filling solution (if not specified otherwise) was 10⁻³ M
NTA, $2.1 \times 10^{-3}$ M NaOH, $10^{-4}$ M Pb(NO$_3$)$_2$, adjusted to pH 7 with 0.1 M NaOH. The calculated activity of uncomplexed Pb$^{2+}$ was $1.3 \times 10^{-9}$ M, with log $\beta_{\text{Pb-NTA}} = 11.39$, $\text{pK}_a^1 = 1.89$, $\text{pK}_a^2 = 2.49$, $\text{pK}_a^3 = 9.73$ [56, 57]. A diaphragm separated the internal filling solutions from the reference half cell (Ag/AgCl in $2 \times 10^{-3}$ M NaCl). The conditioning solution was always $4.8 \times 10^{-9}$ M Pb(NO$_3$)$_2$ (1 ppb), $10^{-3}$ M CaCl$_2$ and $10^{-4}$ M HNO$_3$. All ISEs were conditioned for at least 24 h and used for not more than one week. Each measurement was done with 4 ISEs of the same composition and repeated at least two times.

**EMF measurements.** Experiments were performed at pH 4 in polypropylene beakers, which were pretreated with 0.1 M HNO$_3$ overnight. The samples were collected in polyethylene bottles and the pH values were adjusted immediately after sampling by adding the required amount of 1 M HNO$_3$ (0.8 to 3.1 mL to a 500 mL sample in the case of natural water); no pH buffers were used. Potentials were measured with a custom-made 16-channel electrode monitor at room temperature (20-21 °C) in stirred solutions. The specific measurement protocol was the following: the ISEs were equilibrated in a 500 mL beaker containing the target sample and the potential was observed until the drift was smaller than 1 mV for a 10 min time period (ca. 20 min for samples with a Pb$^{2+}$ concentration higher than the conditioning solution). Subsequently, the ISE was immersed into 50 mL of the same sample, but adding 80 µL of Pb$^{2+}$ standard solution, and the potential was acquired for 5 min. Four more Pb$^{2+}$ additions were performed on the same sample (80, 150, 250, 500 µL of standard) and the potential was monitored for 1 min each time. For the measurements in Figures 6.4-6.6 the pH of the sample was successively changed from 8 to 4 by adding 0.1 or 1 M HNO$_3$ and monitored with a pH-electrode. The ISEs were then calibrated as described above. For this measurement the concentration of Pb$^{2+}$ in the inner filling solution was not buffered with NTA, but it was identical to the conditioning solution (see Section 6.7). The theoretical curves were
calculated with Mathcad PLUS 6 by taking into account the equilibria of Pb$^{2+}$ with the following species: CO$_3^{2-}$ (log $\beta_1 = 6.20$; log $\beta_2 = 8.85$; [CO$_3^{2-}$]$_{tot, min} = 2.48 \times 10^{-3}$ M; [CO$_3^{2-}$]$_{tot, max} = 4.14 \times 10^{-3}$ M); OH$^-$ (log $\beta_1 = 6.29$, log $\beta_2 = 10.88$, log $\beta_3 = 13.94$), SO$_4^{2-}$ (log $\beta_1 = 2.75$; log $\beta_2 = 3.47$; [SO$_4^{2-}$]$_{tot} = 1.55 \times 10^{-4}$ M); Cl$^-$ (log $\beta_1 = 1.60$; log $\beta_2 = 1.80$; log $\beta_3 = 1.70$; log $\beta_2 = 1.38$; [Cl$^-$]$_{tot} = 1.25 \times 10^{-4}$ M); NO$_3^-$ (log $\beta_1 = 1.17$; log $\beta_2 = 1.40$; [NO$_3^-$]$_{tot} = 7.20 \times 10^{-5}$ M); HCO$_3^-$ (log $\beta_1 = 3.20$); HSO$_4^-$ (log $\beta_1 = 2.10$). The logarithmic equilibrium constant for the protonation of CO$_3^{2-}$ were log $\beta_1 = 10.33$ and log $\beta_2 = 16.35$. The reference electrode (Metrohm Ag/AgCl in 3 M KCl, type 6.0729.100) had a 1 M NH$_4$NO$_3$ bridge electrolyte. The pH glass electrode was also from Metrohm. Since the ionic strength of the solutions during the measurements was kept constant, the EMF values were not corrected for the liquid junction potential, and Pb$^{2+}$ activities were assumed to be proportional to the concentrations. The selectivity coefficients were measured with the separate solution method with membranes initially not containing any Pb$^{2+}$. Calibration curves were taken first for the interfering ions and then for Pb$^{2+}$ [29]. A new set of three electrodes was used for each interfering ion. Selectivity coefficients were calculated from the EMF obtained with 10$^{-3}$ M solutions using the experimental slope. All EMF data were corrected for the liquid junction potential according to the Henderson equation. For selectivity determinations, activity coefficients were calculated according to the reference [9].

**ICPMS Measurements.** A last generation quadrupole-based ICPMS (ELAN 6100 DRC, PE/SCIEX, Concord, Canada) was used for heavy metal quantification in the water samples. The instrument was operated at a low sample uptake rate and at robust plasma conditions to minimize non-spectroscopic interferences. The major interfering elements for the ISE-measurements, Cu and Cd, were monitored additionally to Pb. The isotopes measured were $^{65}$Cu, $^{111}$Cd and $^{208}$Pb, which were chosen with respect to the highest abundance but to the least potential spectroscopic interferences (e.g.
Variations in the natural isotopic composition of Pb were not considered to be of relevance for the accuracy required in these measurements. Due to this fact only the most abundant isotope was used for analysis. 10 μg/L of Tl and Bi were added prior to the ICPMS measurements and included in the measurements to examine signal drift, which was found to be virtually absent. The sample related signal variation was 2.5% with a variation for periodic repetition of one sample below 0.5%.

The instrumental parameters under standard operation mode were the following: Rf-Power: 1350 W; Carrier gas flow: 0.95 L/min Ar; Auxiliary gas flow: 0.9 L/min Ar; Plasma gas flow: 14.5 L/min Ar; Nebulizer: Quartz, concentric; Spray chamber: Quartz, cyclonic; Sample uptake rate: 0.7 mL/min; Autolens: on; Measurement mode: peak hop; dwell time per isotope: 200 ms; no. of sweeps: 5; no. of readings: 1; no. of replicates: 10. The limits of detection achieved under these conditions were 0.002 μg/L (9.7 x 10^{-12} M) for 208Pb, 0.03 μg/L (4.7 x 10^{-10} M) for 65Cu and 0.02 μg/L (1.8 x 10^{-10} M) for 111Cd.
7. EXPERIMENTAL PART

7.1 Reagents

Poly(vinyl chloride) (PVC), the plasticizers bis(2-ethylhexyl) sebacate (DOS) and 2-nitrophenyl octyl ether (o-NPOE), the lipophilic sites sodium and potassium tetrakis-[3,5-bis(trifluoromethyl)phenyl]borate (NaTFPB, KTFPB), the lipophilic salt tetradodecyl-ammonium tetrakis(4-chlorophenyl)borate (ETH 500), the Pb²⁺-ionophores Methylene-bis-N,N-diisobutyldithiocarbamate (MBDiBDTC, ionophore VII), 4-tert.-butyl-calix[4]arene-tetrakis(thioacetic acid dimethylamide) (ionophore VI), the Ca²⁺-ionophore N,N-dicyclohexyl-N,N'-dioctadecyl-3-oxapentanediamide (ETH 5234), and tetrahydrofuran (THF) were from Fluka AG (CH-8071 Buchs).

The Pb²⁺-ionophores N,N'-diheptyl-N,N',6,6-tetramethyl-4,8-dioxo-undecanedi-amide) (ETH 295, ionophore V), N,N'-dioctadecyl-N,N'-dipropyl-3,6-dioxoactanediamide) (ETH 322, ionophore I), N,N,N',N'-tetradodecyl-3,6-dioxoactanediamide) (ETH 5428, ionophore II), N,N,N',N'-tetradodecyl-3,6-dioxoactane-1-thio-8-oxodiamide) (ETH 5493, ionophore III), and N,N,N',N'-tetradodecyl-3,6-dioxoactanedithioamide) (ETH 5435, ionophore IV), as well as the chromoionophores, 4-(octadecylamino)-azobenzene (ETH 5315) and 9-dimethylamino-5-[4-(15-butyl-1,13-dioxo-2,14-dioxanonadecyl) phenylimino]-5H-benzo[α]phenoazazine (ETH 5418) were synthetized in our laboratory as described [37, 38, 52].

Aqueous solutions were prepared with freshly deionized water (specific resistance, 18.1 MΩ cm) obtained with a NANOpure® reagent-grade water system (Barnstead, CH-4009 Basel, Switzerland).

The salts NaCl, KCl, CaCl₂ and MgCl₂ and the acid HNO₃ were of Suprapur grade from Merck (Darmstadt, Germany). Nitrilotriacetic acid (NTA), disodium ethylenediamine tetraacetic acid (Na₂EDTA) and all other salts were
of puriss p.a. or Microselect quality from Fluka or p.a. quality from Merck. The Pb\(^{2+}\) standard solution with a concentration of 1000 mg l\(^{-1}\) in 0.001 M HNO\(_3\) was obtained from Fluka Trace analysis.

### 7.2 Structures

**Plasticizers**

o-NPOE, C\(_{14}\)H\(_{21}\)NO\(_3\), Mr 251.33

![o-NPOE structure]

DOS, C\(_{26}\)H\(_{50}\)O\(_4\), Mr 426.69 g

![DOS structure]

**Lipophilic anionic sites**

NaTFPB, C\(_{32}\)H\(_{12}\)BF\(_{24}\)Na, Mr 886.21 g
KTFPB, C\(_{32}\)H\(_{12}\)BF\(_{24}\)K, Mr 902.32 g
Lipophilic salt

ETH 500, $C_{72}H_{116}BCl_4N$, Mr 1148.29

Ionophores

ETH 322, $C_{48}H_{96}N_2O_4$, Mr 793.34
ETH 5428, C\textsubscript{54}H\textsubscript{108}N\textsubscript{2}O\textsubscript{4}, Mr 849.45

ETH 5493, C\textsubscript{54}H\textsubscript{108}N\textsubscript{2}O\textsubscript{3}S, Mr 865.51

ETH 5435, C\textsubscript{54}H\textsubscript{108}N\textsubscript{2}O\textsubscript{2}S\textsubscript{2}, Mr 881.58

ETH 295, C\textsubscript{27}H\textsubscript{54}N\textsubscript{2}O\textsubscript{4}, Mr 470.73
MBDiBDTC, $\text{C}_{19}\text{H}_{38}\text{N}_{2}\text{S}_{4}$, Mr 422.77

\[
\begin{array}{c}
\text{N} \\
\text{S} \\
\text{S} \\
\text{S} \\
\text{N}
\end{array}
\]

$\text{Pb}^{2+}$-selective calix[4]arene derivate, $\text{C}_{60}\text{H}_{84}\text{N}_{4}\text{O}_{4}\text{S}_{4}$, Mr 1054.59

\[
\begin{array}{c}
\text{N} \\
\text{S} \\
\text{O} \\
\text{S}
\end{array}
\]

ETH 5234, $\text{C}_{32}\text{H}_{100}\text{N}_{2}\text{O}_{3}$, Mr 801.37

\[
\begin{array}{c}
\text{N} \\
\text{O} \\
\text{O} \\
\text{N}
\end{array}
\]

Chromoionophores

ETH 5315, $\text{C}_{30}\text{H}_{47}\text{N}_{3}$, 449.71, $\text{pK}_a$ 5.2 [51]
ETH 5418, C_{45}H_{57}N_{3}O_{5}, Mr 719.95, pKa 9.0 [51]

7.3 Compositions of the membranes

The membrane compositions described in the experimental sections of chapters 4, 5, and 6 are summarised in the tables below.

*Membranes of chapter 4*

**Table 7.1.** Membrane composition of the ISE-membranes used for the determination of the ion-ionophore formation constants.

<table>
<thead>
<tr>
<th>Ionophore name</th>
<th>NaTFPB (mmol/kg)</th>
<th>DOS (wt%)</th>
<th>PVC (wt%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>ETH 322</td>
<td>15.5</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5428</td>
<td>15.3</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5493</td>
<td>16.3</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5435</td>
<td>16.2</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 295</td>
<td>28.6</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
<tr>
<td>MBDiBDTC</td>
<td>29.3</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
<tr>
<td>Ionophore VI</td>
<td>14.9</td>
<td>5.0</td>
<td>32-33 %</td>
</tr>
</tbody>
</table>
Table 7.2. Membrane composition of the ISE-membranes used to determine the stoichiometries of the ion-ionophore complexes.

<table>
<thead>
<tr>
<th>Ionophores</th>
<th>NaTFPB (mmol/kg)</th>
<th>DOS (wt%)</th>
<th>PVC (wt%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>ETH 5428</td>
<td>10.7</td>
<td>3.3</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5428</td>
<td>31.0</td>
<td>3.2</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5493</td>
<td>10.2</td>
<td>3.2</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5493</td>
<td>29.3</td>
<td>3.1</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5435</td>
<td>10.2</td>
<td>3.2</td>
<td>32-33 %</td>
</tr>
<tr>
<td>ETH 5435</td>
<td>29.3</td>
<td>3.2</td>
<td>32-33 %</td>
</tr>
<tr>
<td>Ionophore VI</td>
<td>1.6</td>
<td>0.5</td>
<td>32-33 %</td>
</tr>
<tr>
<td>Ionophore VI</td>
<td>4.3</td>
<td>0.5</td>
<td>32-33 %</td>
</tr>
</tbody>
</table>

Table 7.3. Membrane composition of the optode-membranes used for the determination of the ion-ionophore formation constants. The DOS and PVC amounts (wt%) were 65-66% and 32-33%, respectively, for all membranes.

<table>
<thead>
<tr>
<th>Ionophores</th>
<th>Chromoionophores</th>
<th>KTFPB (mmol/kg)</th>
</tr>
</thead>
<tbody>
<tr>
<td>name</td>
<td>name (mmol/kg)</td>
<td>(mmol/kg)</td>
</tr>
<tr>
<td>ETH 322</td>
<td>ETH 5315</td>
<td>9.4</td>
</tr>
<tr>
<td>ETH 5428</td>
<td>ETH 5418</td>
<td>9.1</td>
</tr>
<tr>
<td>ETH 5493</td>
<td>ETH 5418</td>
<td>9.2</td>
</tr>
<tr>
<td>ETH 5435</td>
<td>ETH 5418</td>
<td>9.3</td>
</tr>
<tr>
<td>ETH 295</td>
<td>ETH 5418</td>
<td>9.2</td>
</tr>
<tr>
<td>MBDiBDTC</td>
<td>ETH 5418</td>
<td>9.2</td>
</tr>
<tr>
<td>Ionophore VI</td>
<td>ETH 5418</td>
<td>9.2</td>
</tr>
<tr>
<td></td>
<td></td>
<td>10.8</td>
</tr>
</tbody>
</table>
Membranes of Chapter 5

Table 7.4. Membrane composition Ca$^{2+}$-ISEs.

<table>
<thead>
<tr>
<th>ETH 5234 (mmol/kg)</th>
<th>NaTFPB (mmol/kg)</th>
<th>ETH 500 (mmol/kg)</th>
<th>DOS (wt%)</th>
<th>PVC (wt%)</th>
<th>Thickness (μm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>13.0</td>
<td>4.6</td>
<td>-</td>
<td>32-33 %</td>
<td>65-66 %</td>
<td>150</td>
</tr>
<tr>
<td>12.8</td>
<td>4.5</td>
<td>-</td>
<td>32-33 %</td>
<td>65-66 %</td>
<td>75</td>
</tr>
<tr>
<td>13.1</td>
<td>4.8</td>
<td>-</td>
<td>32-33 %</td>
<td>65-66 %</td>
<td>600</td>
</tr>
<tr>
<td>13.2</td>
<td>4.5</td>
<td>-</td>
<td>49-50 %</td>
<td>49-50 %</td>
<td>160</td>
</tr>
<tr>
<td>13.3</td>
<td>4.5</td>
<td>-</td>
<td>20-21 %</td>
<td>78-79 %</td>
<td>150</td>
</tr>
<tr>
<td>1.38</td>
<td>0.45</td>
<td>5.6</td>
<td>32-33 %</td>
<td>65-66 %</td>
<td>150</td>
</tr>
<tr>
<td>0.14</td>
<td>0.05</td>
<td>5.6</td>
<td>32-33 %</td>
<td>65-66 %</td>
<td>150</td>
</tr>
</tbody>
</table>

Membranes of Chapter 6

Table 7.5. Membrane composition Pb$^{2+}$-ISEs. To the membrane cocktail were added 50 mol% (relative to NaTFPB) of Pb(NO$_3$)$_2$.

<table>
<thead>
<tr>
<th>Pb$^{2+}$ Ionophore VI (mmol/kg)</th>
<th>NaTFPB (mmol/kg)</th>
<th>ETH 500 (mmol/kg)</th>
<th>DOS (wt%)</th>
<th>PVC (wt%)</th>
<th>Thickness (μm)</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.7</td>
<td>0.3</td>
<td>11.5</td>
<td>62 %</td>
<td>36.5 %</td>
<td>ca. 150</td>
</tr>
</tbody>
</table>

7.4 Electrodes and EMF-measurements

The ion-selective membranes were punched from the master membrane and glued to a plasticized PVC tubing with a THF/PVC slurry. If necessary, the internal filling solution of the ISE was separated from the reference half-cell (Ag/AgCl in a chloride solution) by a bridge electrolyte. The reference electrodes were from Methrom (Ag/AgCl in 3 M KCl, type 6.0729.100) or custom made of the type free flowing [53]. The Potential differences were measured with a custom-made 16-channel electrode monitor and corrected for
the liquid junction potential of the reference electrode according to the Henderson formalism [1, 70]. Activity coefficients were calculated using the Debye-Hückel approximation [9].
8. GLOSSARY

\( a_i \) activity [mol l\(^{-1}\)] of the ion \( I \)

\( a_i(I) \) activity [mol l\(^{-1}\)] of the ion \( I \) in an ideal sample containing only \( I \)

\( a_{I,S} \) activity [mol l\(^{-1}\)] of the cation \( I \) in the sample

\( a_{I,M} \) activity [mol l\(^{-1}\)] of the cation \( I \) in the membrane

\( a_{i,S} \) activity [mol l\(^{-1}\)] of the ion \( i \)

\( a_{i,\text{ref}} \) activity [mol l\(^{-1}\)] of the ion \( i \) in the bridge electrolyte of the reference electrode

\( a_{J,S} \) activity [mol l\(^{-1}\)] of the ion \( J \) in the sample

\( a_{J}(J) \) activity [mol l\(^{-1}\)] of the ion \( J \) in an ideal sample containing only \( J \)

\( a_{X,M} \) activity [mol l\(^{-1}\)] of the anion \( X \) in the membrane

\( a_{X,S} \) activity [mol l\(^{-1}\)] of the anion \( X \) in the sample

\( C \) chromoionophore

\( C_T \) total chromoionophore concentration in the membrane

\( c_i \) concentration of the diffusing species \( i \) [mol cm\(^{-3}\)]

\( c_{I,bulk} \) concentration of the ion \( I \) in the bulk of the sample

\( c_{I,PB} \) concentration of the ion \( I \) at the phase boundary

\( DL \) Detection Limit

\( D_i \) diffusion coefficient for the diffusing species \( i \) [cm\(^2\) s\(^{-1}\)]

\( EMF \) electromotive force

\( E_{\text{const}} \) sample independent potential contribution to \( EMF \)

\( E_{D,M} \) Diffusion potential within the membrane

\( E_{D,\text{ref}} \) liquid junction potential at the sample/reference electrode

\( E_{1} \) membrane-depending potential difference of the measuring cell consisting of a reference electrode and an ISE responding only to changes in \( a_{I,S} \)
\( E_{IJ} \) membrane-depending potential difference of the measuring cell consisting of a reference electrode and an ISE responding simultaneously to several ionic species

\( E_i^0 \) intercept of the linear response function of the ISE responding to the ion \( I \)

\( E_j \) membrane-depending potential difference of the measuring cell consisting of a reference electrode and an ISE responding only to changes in \( a_{j,s} \)

\( E_j^0 \) intercept of the linear response function of the ISE responding to the ion \( J \)

\( E_M \) membrane potential

\( E_{M,\text{const}} \) sum of all sample independent membrane potentials

bridge electrolyte interface

\( E_{PB}^- \) phase boundary potential at the ISE membrane/inner filling solution interface

\( E_{PB}^+ \) phase boundary potential at the ISE membrane/sample interface

EDTA ethylenediamine tetraacetic acid

EPA Environmental Protection Agency

\( F \) Faraday constant \((96487 \text{ C mol}^{-1})\)

\( K_{\text{coex}} \) overall coextraction constant of a membrane with ionophore

\( K_{\text{exch}} \) overall ion-exchange equilibrium constant of a membrane with ionophore

\( K_{IJ} \) ion-exchange constant of the uncomplexed ion \( I \) and \( J \) between sample and membrane

\( K_{IX} \) coextraction constant

\( K_{i,j}^\text{pot} \) potentiometric selectivity coefficient of the ISE for the ions \( I \) and \( J \)

\( k_i \) "single ion distribution coefficient" of the cation \( I \)

\( k_j \) "single ion distribution coefficient" of the cation \( J \)
\( k_X \) “single ion distribution coefficient” of the anion \( X \)
\( I \) primary cation
\( (IE) \) ion-exchanger membrane
\( IL \) ion-ionophore complex of the primary ion
\( \text{ICPMS} \) Inductive Coupled Plasma Mass Spectrometry
\( \text{ISE} \) ion-selective electrode
\( J \) interfering cation
\( JL \) ion-ionophore complex of the interfering ion
\( J_i(x) \) rate of transfer of the species \( i \) per unit area of the section in the \( x \)-coordinate, flux \([\text{mol cm}^{-2} \text{s}^{-1}]\)
\( L \) ionophore (ligand)
\( (L) \) ionophore-based membrane
\( L_T \) total ionophore concentration in the membrane
\( M \) membrane, if used as index
\( M_\text{M} \) molar \((\text{mol/l})\)
\( \text{mV} \) \(10^{-3}\) Volt
\( \text{NTA} \) nitrilotriacetic acid
\( \text{nM} \) nanomolar \((10^{-9}\ \text{mol/l})\)
\( n_i \) stoichiometry of the primary ion-ionophore complex
\( n_j \) stoichiometry of the interfering ion-ionophore complex
\( \text{PVC} \) poly(vinyl chloride)
\( \text{ppm} \) part per milion
\( \text{ppb} \) part per bilion
\( R \) molar gas constant \((8.314 \ \text{J K}^{-1} \text{mol}^{-1})\)
\( R_T \) total concentration of lipophilic anionic sites in the membrane
\( S \) sample, if used as index
\( s_i \) slope of the \( I \)-selective ISE
\( s_j \) slope of the \( J \)-selective ISE
\( T \) absolute temperature \([\text{K}]\)
\( \text{THF} \) tetrahydrofuran
TMA tetramethylammonium

$u_i$ absolute mobility of the ion $i$ \([\text{cm}^2 \text{ mol} \text s^{-1} \text J^{-1}]\)

WHO World Health Organisation

$x$ counter anion

$x$ space coordinate measured perpendicularly to the section \([\text{cm}]\)

$z_i$ charge of the ion $i$

$z_I$ charge of the cation $I$

$z_J$ charge of the cation $J$

$z_X$ charge of the anion $X$

$\beta$ formal complex formation constant

$\phi$ electrical potential

$\phi_S$ electrical potential in the sample

$\phi_M$ electrical potential in the membrane

$\phi(x)$ electrical potential in $x$-coordinate

$\tilde{\mu}_I$ electrochemical potential of the ion $I$

$\mu_I$ chemical potential of the ion $I$

$\mu_I^0$ standard free energy of the ion $I$

$\tilde{\mu}_{I,S}$ electrochemical potential of the ion $I$ in the sample

$\tilde{\mu}_{I,M}$ electrochemical potential of the ion $I$ in the sample

$\delta_M$ thickness of the potential determining diffusion layer in the membrane

$\delta_S$ thickness of the potential determining diffusion layer in the sample (Nernstian layer)

$\tilde{\nu}(x)$ hydrodynamic velocity in $x$-coordinate \([\text{cm s}^{-1}]\)
9. REFERENCES


(60) Hulanicki, A.; Lewenstam, A. Talanta 1977, 24, 171-175, Interpretation of selectivity coefficients of solid-state ion-selective electrodes by means of the diffusion layer model.


Injection On-Line Microcolumn Preconcentration and Separation Coupled with Flame Atomic Absorption Spectrometry for Interference-Free Determination of Trace Lead in Biological and Environmental Samples.


**Curriculum Vitae**

1973 Born on the 8th of November at Bellinzona (TI), Switzerland
1974-1984 Primary school education at Bellinzona
1984-1988 Secondary school education at Bellinzona
1988-1992 Post secondary education at Liceo Cantonale, Bellinzona, passing Matrick of scientific type
1992-1997 Studies at the Chemistry Departement of the Swiss Federal Institute of Technology (ETH) Zürich
1997 Chemistry degree of the Swiss Federal Institute of Technology (ETH) Zürich, diploma thesis in the group of Prof. Dr. Ernö Pretsch
1997-2001 Ph. D. studies in the group of Prof. Dr. Ernö Pretsch on chemical sensors
1998-2000 Laboratory assistant for gas-chromatography and GC-MS in analytical chemistry

Languages: Italian, German, French, English

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