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Review

The Design of Ligands for the Transport of Metal Salts in Extractive Metallurgy

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Anion recognition is key to the development of reagents that can selectively transport metal salts in extractive metallurgy. This review outlines flowsheets that require such reagents and considers how anion receptors could be used to transfer base metal salts into water-immiscible media in solvent extraction processes. Particular attention is paid to the extraction of sulfate because there is a need to develop systems for the unit operations of concentration and separation to process sulfate streams.

Keywords: Hydrometallurgy; Sulfate; Extraction; Ditopic

INTRODUCTION

The recovery and purification of metals from their ores can be accomplished by using a variety of techniques such as pyrometallurgy, hydrometallurgy and froth flotation. The increasing number of environmental regulations has heightened the need for new, 'greener' methods. One possible solution is to develop extractants that can transport metal salts in a hydrometallurgical solvent extraction process to effect the concentration and separation of the desired metal. Such an approach could be based on ditopic ligands that show high selectivity for the targeted metal cations and anion(s). This is particularly challenging with respect to the binding and co-extraction of the sulfate anion, often the dominant anion in pregnant leach solutions, into a water-immiscible solvent because of its high affinity for aqueous solution. However, recent developments in the field of anion recognition chemistry have led to an increase in the number and variety of sulfate-selective receptors, which may open up routes to the design of selective and efficient reagents to transport metal sulfates.

METALLURGY

Metallurgy encompasses the extraction of metals from their ores and the purification or modification of the metals for a more useful purpose. The term, in practice, is more often used to refer to an industrial process than a laboratory technique. Metallurgy is a technology that dates back to 4000 BC [1] or earlier and deals with the principles whereby metals are combined to form alloys such as brass or bronze. The main techniques currently used in industry to process metal ores are pyrometallurgy and hydrometallurgy.

PYROMETALLURGY

Initially, the earth's atmosphere was mainly composed of SO_2 and CO_2, which meant that the metal ores formed were metal sulfide ores. As plants emerged and began the process of photosynthesis, oxygen started to become a major component of the atmosphere. Exposed layers of metal sulfide ores were therefore subject to oxidation to produce oxidic ores. Consequently, metal ore deposits commonly contain a sulfidic lower layer, a mixed sulfidic/oxidic ('transition ore') middle layer and a superficial oxidic layer [1].

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Pyrometallurgy is used to extract metals from their ores [2] by heating the latter to high temperatures, often with a reducing agent such as carbon, to produce the metal. A reductant is used when the reduction product is thermodynamically stable, as this favours the production of the metal. For example, if carbon is used in the process of extraction from an oxidic metal ore, then carbon monoxide is formed. This method is not applicable to sulfidic ores because the carbon disulfide produced is barely thermodynamically stable and the analogue of carbon monoxide, carbon sulfide, is highly unstable [3]. Therefore, before the reduction process, metal sulfides are often roasted in air to convert them to metal oxides. This occurs because the oxidation produces sulfur dioxide, which is highly thermodynamically favourable, for example:

$$\text{(i)}\quad ZnS + \frac{3}{2}O_2 \rightarrow ZnO + SO_2$$

$$\text{(ii)}\quad ZnO + 2C \rightarrow Zn + 2CO$$

A reducing agent is not required in some cases where heating the ore to high temperatures can be sufficient to extract the metal. An example of this is copper smelting, in which copper can be recovered from its sulfidic ore at high temperatures of around 1300°C by a two-stage process in air [3]. This can be represented, with respect to copper, as:

$$\text{(i)}\quad 2CuS + O_2 \rightarrow Cu_2S + SO_2$$

$$\text{(ii)}\quad Cu_2S + O_2 \rightarrow 2Cu + SO_2$$

This is again favoured thermodynamically by the formation of sulfur dioxide.

The smelting of copper ores [1] has been used for metal extraction since as early as 4000 BC. This gives the method a significant advantage over many other techniques as the process has been refined and improved for many years and new installations can easily be constructed by a number of specialized companies. Pyrometallurgy can also be carried out on an extremely large scale and is very economical, making it much more profitable than other extraction techniques [4]. These methods were later adapted and applied to the recovery of other metals such as iron, nickel and zinc [2]. However, some of the metals produced by industry need complex recovery processes. Titanium, for example, is made by the Kroll process, which converts TiO_2 to $TiCl_4$ by coke production in the presence of Cl_2. The titanium tetrachloride produced is then reduced with molten magnesium in an inert argon atmosphere to give titanium metal [3].

Pyrometallurgical extraction does have some major disadvantages; for example, the emission of pollutants and toxic gases [2] such as SO_2 and intensely toxic heavy metals [4]. This problem is becoming increasingly significant as environmental regulations get stricter and more difficult to meet. Fossil fuels that are used during the extraction can also form gaseous pollutants. The 'off-gases' can contain fine dusts that need to be recovered to further reduce pollution [2]. In addition, when applied to low-grade ores, pyrometallurgical techniques are often not cost effective.

Pyrometallurgy is an important metal extraction technique, but other methods are needed to process the large stocks of low-grade ores and meet new environmental regulations.

GRAVITY AND FROTH FLOTATION METHODS

Gravity provides the simplest technique for metal ion concentration. Dense metal or ore particles can be separated in swirling water from significantly less dense silicate and other rock-based minerals, as the latter form a suspension and the metal/ore particles sink to the bottom [3].

Another technique used in metallurgy is froth flotation and this is typically used to obtain ore 'concentrates' and remove impurities. Metal sulfides are the ores most commonly treated in this way, but a number of other examples can be found such as the separation of KCl from NaCl [3]. The process is heavily dependent on converting the desired hydrophilic solids into hydrophobic ones and subsequently removing them as an oil-based froth. The unwanted materials (silicates, aluminosilicates) are removed as an aqueous slurry from the bottom of the separator.

Flotation agents are used that can be selectively attached to the surfaces of the metal ore particles [5]. For example, in the treatment of metal sulfide ores, suitable agents for the metal sulfide particles have polar 'head groups' that contain sulfur atoms. These tend to adsorb on a metal sulfide surface but not onto silicate or aluminosilicate surfaces. The attached alkyl 'tails' of these collectors render the complexed ore particle hydrophobic [5,6], which allows ore concentration by removal of the resultant hydrophobic froth.

HYDROMETALLURGY

Hydrometallurgical techniques have fewer environmental problems than the foregoing. As world supplies of high-grade ores are decreasing, techniques are required for the recovery of metals from low-grade ores, transition ores, dumpstocks, mixed metals and residues [2]. Hydrometallurgy is being developed as a more environmentally sound and diverse technique [4]. In addition, it has some economic benefits, as it avoids the high temperatures of pyrometallurgy. Transport costs can be reduced because plants can be situated near the ore bodies [4].

Hydrometallurgical methods use reactions in aqueous solutions to extract selectively (this can involve metal complex formation) and/or concentrate targeted metal ions. The hydrometallurgical process often involves leaching of the ore as the initial step. For example, gold particles in crushed rock can be recovered by cyanide leaching with aerated sodium cyanide solution to allow $[Au(CN)_2]^-$ to be brought into solution [5,6]. The undesired impurities (aluminosilicates) do not dissolve, which gives:

$$2Au(s) + O_2 + 4CN^-(aq) + 2H_2O \rightarrow$$

$$2[Au(CN)_2]^-(aq) + H_2O_2(aq) + 2OH^-(aq)$$

The reaction can take several days, after which the gold can be obtained by electrolysis [7]. Ammoniacal leaching can be used to separate base metal ions. The Sherritt Gordon process uses the power of aqueous ammonia to complex Cu^{2+}, Ni^{2+} and Co^{3+} but not Mn^{2+}, Fe^{2+} and Fe^{3+}. After leaching, the individual Cu^{2+}, Ni^{2+} and Co^{3+} ions can be separated by complicated aqueous chemistry. The ions are then reduced to the base metals using hydrogen under conditions of high pressure and temperature [3].

Other leaching processes can use very different methods; for example, bacteria for the oxidation of insoluble mineral sulfides to the commonly water-soluble metal sulfates. Thus *Thiobacillus ferrooxidans* can use the sulfide–sulfate redox cycle to drive its own metabolic processes [3]. This type of reaction requires oxygen:

$$MS(s) + 2O_2(g) \rightarrow M^{2+}(aq) + SO_4^{2-}(aq)$$

The metal can then be recovered by a hydrometallurgical process. This has the very significant advantage over the roasting in air of sulfidic ores performed in pyrometallurgy, that SO_2, an air pollutant, is not released to the atmosphere [5].

Liquid–liquid solvent extraction is an important technique in hydrometallurgy. It is capable of dealing with very difficult metal ion separations and generally involves the selective transport of metal ions from an aqueous phase into an immiscible organic phase. This technology has a number of possible applications, and one of the most important is the extraction of transition metals from low-grade ores. There are four steps in this form of metal recovery, consisting of leaching, extraction, stripping and electrowinning [4,5]. The first part of the process, leaching, is the selective dissolution of the metal and this is often carried out with sulfuric acid. There are two ways that leaching is carried out: simple leaching (at atmospheric temperature and pressure) and pressure leaching (pressure and temperature are increased to quicken the process) [2]. Extraction is the next stage in the process and involves the selective transport of metal ions into a non-polar organic phase by means of a hydrophobic ligand, often an organic acid, dissolved in the organic solvent (overbar indicates solvation in the organic phase), for example, as in Eq (1):

$$M^{2+} + \overline{2LH} \rightleftharpoons \overline{ML_2} + 2H^+ \qquad (A)$$

The stripping stage involves the removal of the metal ion from the reagent, for example by contacting the loaded organic phase with an aqueous phase at lower pH than the leaching solution:

$$\overline{ML_2} + 2H^+ \rightleftharpoons M^{2+} + \overline{2LH}$$

This produces a pure metal salt solution [2,4,5] that can be used in the electrowinning stage to give pure metal.

An example of such a process that is used in industry is the solvent extraction of copper by selective phenolic oxime ligands (Fig. 1) [4]. In this process the copper oxide is leached from its ore using sulfuric acid to give a pregnant leach solution containing a number of metal sulfates. The leached Cu(II) ions can be extracted into the organic phase by the hydrophobic ligand. The metal ion displaces two protons from two ligand molecules, which are released into the aqueous phase, regenerating the sulfuric acid used to leach the ore. The copper is stripped from the loaded organic phase back into an aqueous phase of lower pH than the pregnant leach solution to give a pure copper sulfate solution. The ligand is simultaneously reprotonated in the organic phase and is recycled. The pure copper sulfate solution then undergoes electrolysis to produce pure copper metal (product) and sulfuric acid (spent electrolyte), which is fed back into the stripping stage (Fig. 2). The pH gradient is the driving force for the transport of the copper ion across the circuit. The success of the process is based on the excellent overall materials balance (Fig. 2) and the charge neutrality of the copper ligand complex, which provides organic solubility. The copper produced is also of a very high quality (electrical conductivity grade) and the reagent can be recycled.

The leaching of metal sulfide ores requires different reactions and this can cause problems in

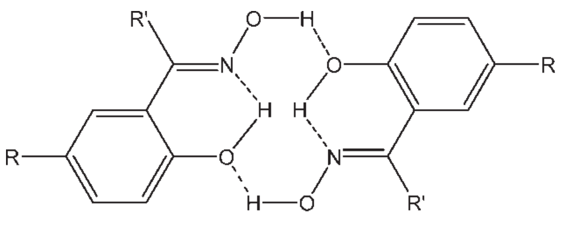


FIGURE 1 The commercial phenolic oxime copper extractants [4,5]. R = alkyl, R′ = H, alkyl.

$CuO + H_2SO_4$	$\rightleftharpoons$	$CuSO_4 + H_2O$	Leach
$CuSO_4 + \overline{2LH}$	$\rightleftharpoons$	$\overline{CuL_2} + H_2SO_4$	Extract
$\overline{CuL_2} + H_2SO_4$	$\rightleftharpoons$	$CuSO_4 + \overline{2LH}$	Strip
$CuSO_4 + H_2O$	$\rightleftharpoons$	$Cu + 1/2O_2 + H_2SO_4$	Electrowin
CuO	$\rightleftharpoons$	$Cu + 1/2O_2$	Net

FIGURE 2 Materials balance for heap leach/solvent extraction/electrowinning of copper based upon the process involving phenolic oxime extractants [4,5].

the materials balance of the process if the conventional extractants, such as the phenolic oximes, are used. The dilemma occurs because sulfidic ores cannot be leached efficiently with sulfuric acid and conversion to the oxidic ores by roasting in air releases SO_2, which must be recovered [3]. Oxidative leaching can be performed by roasting the sulfidic ore at a lower temperature in a limited air supply to produce the metal sulfate, which can subsequently be dissolved in water for extraction [3,8,9]. The problem with these types of leaching, when used in conjunction with commercial organic acid solvent extractants, LH, is caused by the release of two protons into the aqueous phase (Fig. 3), creating a build-up of acid. This presents no problem in the metal recovery from oxidic ores because the acid is reused in the leaching step.

However, for sulfidic ores, acid is not consumed in leaching and therefore builds up at the front end of the circuit. This requires either acid recovery or neutralization with a base to control pH. Acid recovery is very expensive and neutralization is not favoured because a salt waste is generated, which has to be removed from solution.

There have been a number of suggestions on how to tackle this problem [10]. One approach involves avoiding the release of sulfuric acid during extraction by using a ditopic ligand that can extract both a metal cation and a sulfate anion (see extract step, Fig. 4) [11]. For reagents based on zwitterionic binding sites (*see below*), the stripping of the cation and anion from the ligand could then be performed in two steps. First the cation can be stripped, as in previous procedures, by using sulfuric acid of low pH to produce a pure metal sulfate solution. Second, the anion can be removed by subsequent stripping with an aqueous base, B, to recycle the ligand (Fig. 4). With careful planning, the salt formed could be a saleable product. The pure metal sulfate solution would be fed to an electrowinning stage to give pure metal and to regenerate the sulfuric acid used in stripping the metal.

Extraction of metals by this method has potential applications for treating acid mine drainage and other effluent streams [12]. This could be done by a 'subtractive approach', whereby a metal sulfate is selectively removed without the pH being affected. A strategy such as this is also important in primary metal recovery because the cation exchange reagents currently used, such as the phenolic oximes, release protons into the aqueous phase upon complexing the metal cation, which in turn lowers the pH and therefore reduces the efficiency of loading metal ions in competition with protons (see equation (A)). This problem is exacerbated when the pregnant leach solution contains a high concentration of extractable metal and an efficient process demands a high recovery from each solvent extract step.

$MS + 2O_2$	$\rightleftharpoons$	MSO_4	Leach
$MSO_4 + \overline{2LH}$	$\rightleftharpoons$	$\overline{ML_2} + H_2SO_4$	Extract
$\overline{ML_2} + H_2SO_4$	$\rightleftharpoons$	$MSO_4 + \overline{2LH}$	Strip
$MSO_4 + H_2O$	$\rightleftharpoons$	$M + 1/2O_2 + H_2SO_4$	Electrowin
$MS + 3/2O_2 + H_2O$	$\rightleftharpoons$	$Cu + H_2SO_4$	Net

FIGURE 3 Sulfuric acid build-up.

$MS + 2O_2$	$\rightleftharpoons$	MSO_4	Leach
$MSO_4 + \overline{L}$	$\rightleftharpoons$	$\overline{MLSO_4}$	Extract
$\overline{MLSO_4} + H_2SO_4$	$\rightleftharpoons$	$MSO_4 + \overline{[LH_2]SO_4}$	Strip cation
$\overline{[LH_2]SO_4} + 2B$	$\rightleftharpoons$	$\overline{L} + [BH_2]SO_4$	Strip anion
$MSO_4 + H_2O$	$\rightleftharpoons$	$M + 1/2O_2 + H_2SO_4$	Electrowin
$MS + 3/2O_2 + H_2O + 2B$	$\rightleftharpoons$	$M + [BH_2]SO_4$	Net

FIGURE 4 Materials balance for metal recovery via metal sulfate solvent extraction.

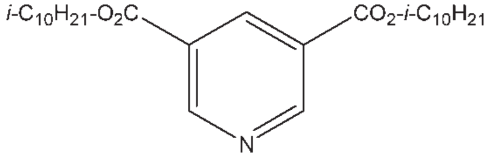


FIGURE 5 The commercial copper extractant CLX50 used in the CUPREX process [13].

There are already some examples of extractants that transport metal salts in hydrometallurgical processes but most of these use chloride as the accompanying anion, for example in the CUPREX process [13]. This uses oxidative leaching of sulfidic ores with ferric chloride to produce pregnant leach solutions (containing mainly $CuCl_2$, $FeCl_2$ and $FeCl_3$), liberating elemental sulfur. In the extract step, a neutral ligand such as CLX50 (Fig. 5) transports $CuCl_2$ with a good materials balance.

The extraction equilibrium in the process is highly dependent on the activity of the Cl^- anion in the feed solutions. This in turn is highly dependent on the stability of chlorometallate complexes $[MCl_x]^{2-x}$ and $[MCl_x]^{3-x}$ for the metals present [14,15], and the equations in Fig. 6 are a simplified representation of the solution speciation.

There are practical problems associated with this process as the Cl_2 generated during electrowinning causes engineering problems and is a safety hazard. Also the hydrochloric acid medium from which the $CuCl_2$ is extracted is highly corrosive and this again creates engineering difficulties [16]. The transport of SO_4^{2-} instead of Cl^- in a solvent extraction process would greatly reduce these problems, and the electrowinning of metal sulfates is well defined and sulfuric acid is less corrosive [5].

Hydrometallurgical methods can therefore be used to effectively recover metals in a variety of different ways. The increasing number of environmental regulations makes research into these types of processes of key importance. This is particularly apparent in the processing of metal sulfide ores because although pyrometallurgical methods are efficient, they release toxic gases such as SO_2. Commercial organic acid reagents, used in the hydrometallurgical processing of metal oxide ores, also generate an unwanted acid build-up at the front end of the circuit if used with metal sulfides. This problem could be addressed by the design of ditopic ligands that could extract selectively both metal cations and their attendant anion(s), as the acid could

$4FeCl_3 + Cu_2S$	$\rightleftharpoons$	$4FeCl_2 + 2CuCl_2 + S\downarrow$	Leach
$2CuCl_2 + 4\overline{L}$	$\rightleftharpoons$	$\overline{2CuL_2Cl_2}$	High Cl^- activity
$(\overline{2CuL_2Cl_2})$	$\rightleftharpoons$	$(2CuCl_2 + 4\overline{L})$	Low Cl^- activity
$2CuCl_2$	$\rightleftharpoons$	$2Cu + 2Cl_2$	Electrowinning
$4FeCl_2 + 2Cl_2$	$\rightleftharpoons$	$4FeCl_3$	Leach recovery
Cu_2S	$\rightleftharpoons$	$2Cu + S$	Overall

FIGURE 6 A simplified flow sheet and materials balance for the recovery of copper from sulfidic ores by oxidative chloride leaching, solvent extraction and electrowinning [5].

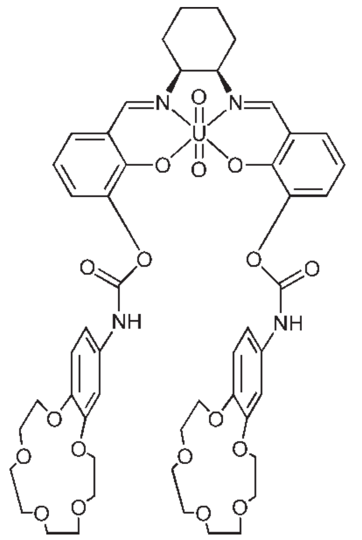


FIGURE 7 Ditopic ligand developed by Reinhoudt and co-workers [17] to bind $K^+/H_2PO_4^-$.

be transported with the targeted metal ion across the circuit.

LIGAND DESIGN FOR SOLVENT EXTRACTION OF METAL SALTS

To make solvent extraction of metal salts viable, highly selective ligands that can complex both the metal cation and the attendant anion(s) are required. Design of these ligands should focus on producing separate cation- and anion-binding sites or the anion could be bound directly to the metal centre.

To facilitate stripping procedures, it is desirable to have separated binding sites in the extractant. There are some examples of this type of ditopic ligand in the literature. Reinhoudt and co-workers constructed a ligand that contained two cation-binding crown ether groups and two anion-binding amide moieties, which were aided in their role by a Lewis acidic UO^{2+} group in relatively close proximity [17] (Fig. 7). The receptor successfully complexed both potassium and dihydrogenphosphate ions.

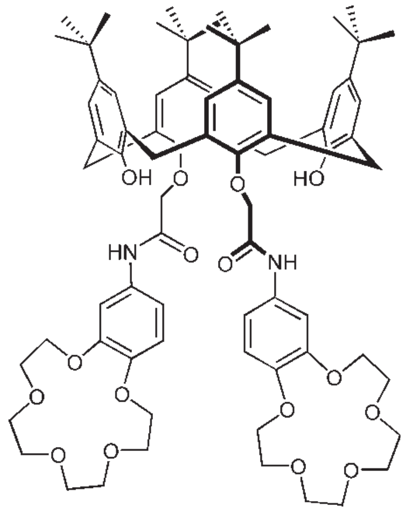


FIGURE 8 Calixarene-based ditopic receptor developed by Beer *et al.* [18].

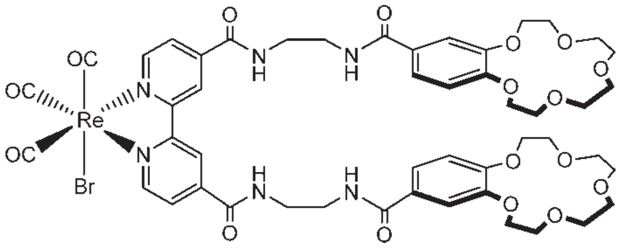


FIGURE 9 Ditopic receptor developed by Beer and Dent [19].

Another example of this type was reported by Beer *et al.* in which the receptor contained two benzo[15]-crown-5 moieties attached to the lower rim of calix[4]arene through an amide linker [18] (Fig. 8). In the absence of alkali metal cations, the amide groups bound no anions, but on addition of potassium ions a sandwich complex was formed between the two crown ether units. This templated the formation of an anion-binding site by bringing the amide moieties closer together. The system showed particular selectivity for the dihydrogenphosphate anion. This system demonstrates the benefit of using ditopic ligands with separate binding sites. Cooperative binding can be achieved by using cation complexation to facilitate anion binding.

In related work, Beer and Dent also showed that cation binding can change the anion selectivity of a ligand. They developed a heteroditopic rhenium(I) bipyridyl receptor (Fig. 9), which again used amide groups for anion binding [19]. The ligand, when no K^+ ions were present, was found to bind $H_2PO_4^-$ over Cl^- as 1:1 complexes ($K_{H_2PO_4} = 205\,M^{-1}$ and $K_{Cl} = 55\,M^{-1}$ in DMSO-d_6). However, after K^+ binding, an intramolecular sandwich complex forms and the association constant decreases for $H_2PO_4^-$ and significantly increases for Cl^- ($K_{H_2PO_4} = 35\,M^{-1}$ and $K_{Cl} = 300\,M^{-1}$ in DMSO-d_6). This demonstrates how careful ditopic ligand design can tune anion selectivity as the smaller Cl^- anion shows a greater affinity for the more compact binding cavity formed upon K^+ complexation.

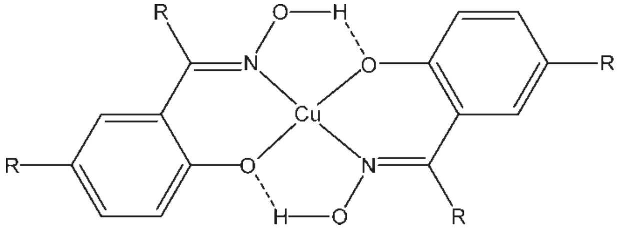


FIGURE 10 Phenolic-oxime *pseudo*-macrocyclic copper complex [20].

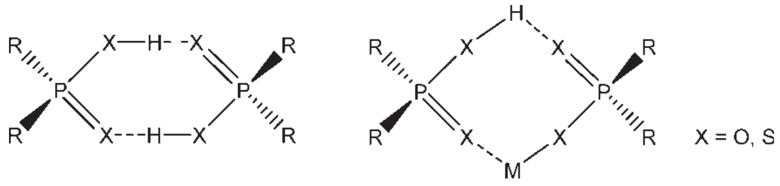


FIGURE 11 Organophosphorus acid extractants and their mode of metal binding [5].

It is important to consider the behaviour of ligands in the non-polar solvents used in industrial solvent extraction when designing extractants. Intra- and intermolecular interactions, such as hydrogen bonds, become increasingly significant and can be instrumental in controlling the efficiency of complexation and extraction.

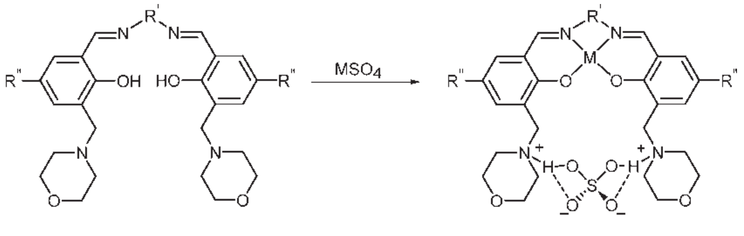
An example of this uses the well-known phenolic oxime extractants [4]. These are used in hydrocarbon solvents for the recovery of copper (see Section 5). Intermolecular hydrogen bonds between the oxime OH groups and the phenolate oxygen atoms lead to the formation of *pseudo*-macrocyclic copper complexes [20] (Fig. 10) of high stability, which contribute to the efficiency of the process.

In another example, the organophosphorus acid metal extractants show very strong interligand hydrogen bonds causing the formation of dimers in hydrocarbon solvents [6]. The eight-membered rings in the dimers are retained upon metal complexation, in the presence of excess extractant (Fig. 11). This has important consequences as the bite angle defined by the *pseudo*-chelate rings favours tetrahedral coordination and allows the selective extraction of metals such as zinc. For example, di(2-ethylhexyl)phosphoric acid (D2EHPA) extracts Zn^{2+} more strongly than other first transition series metal ions [5,6]. Ditopic ligands have been designed in the Tasker group [11] to transport metal salts. The metals that were targeted for extraction were copper and nickel and the anion chosen was sulfate, the anion present in many pregnant leach solutions. A 'salen'-based binding site (salen = *N,N'*-(salicylidene)ethylenediamine dianion) was used as a coordination site for the metal (Fig. 12) as it is known to bind strongly to nickel and copper [21,22].

The prototype ligands [11] (Fig. 12) also have two pendant tertiary amines that are capable of capturing the protons released from the phenolic oxygen atoms upon metal complexation giving the ligand a zwitterionic form (positive and negative sites are present but the molecule is overall charge-neutral). These arms are templated by the coordination of a metal ion to create a dicationic binding site for anions. This is important for the encapsulation of the sulfate anion in the solid state [11] and is also essential for the efficient solvent extraction of sulfate [23]. The anion-binding site also has the ability to hydrogen bond to the captured anion. These types of interactions are frequently seen in biological systems [24–26] and are important to the efficiency of transport in solvent extraction processes.

An advantage of the ligand's zwitterionic form is that no acid is lost to the aqueous phase upon formation of a complex with a metal salt so the pH of the raffinate is similar to that of the feed solution.

The overall charge neutrality of the complex is also important for solubility in hydrocarbons or other solvents of low polarity. These design features open up the possibility for the ligand to be used in hydrometallurgical operations in sulfate media (Fig. 13). A practicable metal salt extractant must bind the targeted metal cation and anion(s) strongly and selectively. This could be performed by a variety of methods such as using the coordination of a metal cation to template the formation of the anion-binding site or by taking advantage of the importance of hydrogen-bonding interactions in non-polar solvents. To develop highly selective systems for the desired metal cation and anion(s), convenient and efficient methods to assess selectivity are required.



R' = $-(CH_2)_2-$(A), *o*-C_6H_4(B), (±)-*trans*-1,2-cyclohexane (C)

R'' = Bu^t, branched nonyl (used for solvent extraction to increase organic solubility)

FIGURE 12 Prototype ditopic ligands for the transport of metal sulfates [11].

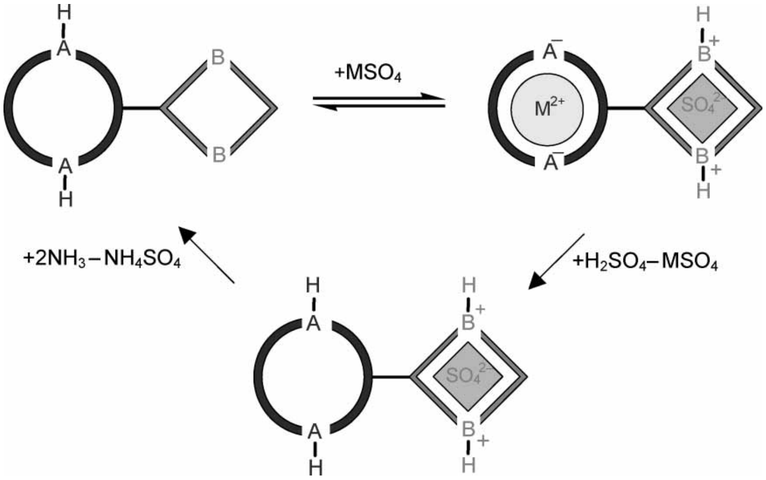


FIGURE 13 Application of zwitterionic ditopic ligands with dicationic/dianionic sites for the transport of metal sulfates.

METAL AND ANION DETECTION METHODS FOR ASSESSING SELECTIVITY

Reliable analytical techniques are needed to obtain accurate data for the amount of metal and anion extracted or the relative strength of their binding. Metal ions can be detected by a variety of methods [27] (e.g. inductively coupled plasma optical emission spectroscopy (ICP-OES), atomic absorption spectroscopy, UV–Vis spectroscopy), which can be used to determine the selectivity of their extraction. ICP-OES is very good for accurately analysing samples for metal content (less than 10 ppb detection limits) [27] and in a solvent extraction experiment it can be used to analyse both the organic and the aqueous phase to check for materials balances.

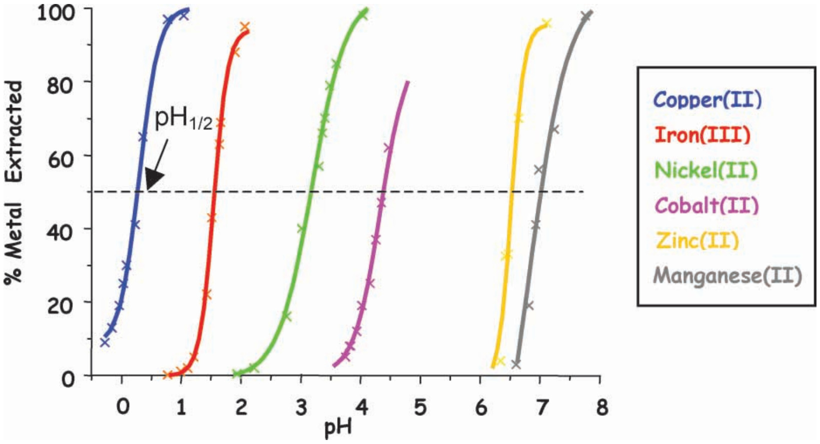


FIGURE 14 The $pH_{1/2}$ of a number of metals with P50-oxime (a phenolic oxime) copper extractant [4].

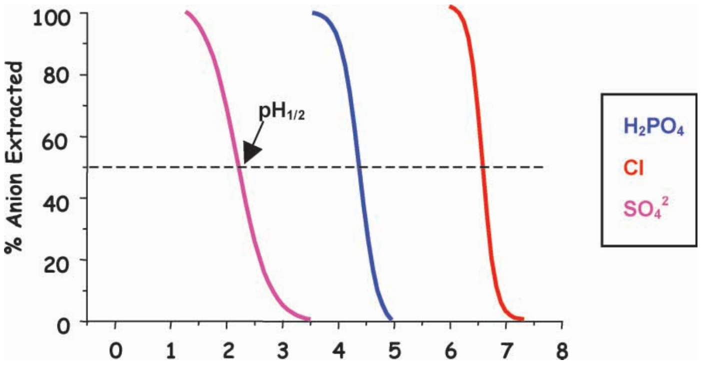


FIGURE 15 Schematic representation of how anion selectivity could be measured in an SX experiment.

In hydrometallurgy, a common way of comparing the 'strengths' of organic acid (LH) extractants is by measuring the $pH_{1/2}$, that is the pH associated with 50% loading of the extractant at a stated concentration and for a defined composition of aqueous feed [5]. Conventionally, for the extraction of metal ions (equation (A), Section 5), the lower the $pH_{1/2}$ the 'stronger' the metal complexation. If the metal content of the organic phase is determined over a range of pH, then $pH_{1/2}$ values can be obtained and selectivities assigned (Fig. 14).

The stability of anion complexation is examined in a number of different ways in the literature but most have involved a single-phase system. One popular method is based on potentiometry, or more specifically pH-metry [28,29]. This technique is nearly always carried out using a glass pH electrode and, to be effective, it needs to relate the consumption or release of protons in the complexation process to the concentration of the anion-receptor complex formed. However, as ligands and anions can often capture protons independently, the first step in the procedure is usually to determine the basicity constants of the ligand and anion separately [30]. Once these are known, an alkali or mineral acid is titrated into a solution of the ligand, anion and protons, and the variation in the free hydrogen ion concentration is monitored. The stability constants of anion complexation can then be determined by computer programs [30]. This technique has been used successfully to generate anion selectivity data. To date, most of the studies have been carried out in water alone.

NMR is probably the most widely used technique for the determination of anion binding constants in non-aqueous solvents [31]. It has also provided some of the most important structural information on the solution behaviour of anion receptors. The experimentally measurable parameters (chemical shifts, coupling constants, etc.) in NMR can be used to determine complex stability constants. These parameters depend on the NMR-sensitive nuclei present and their chemical environments as well as the reaction rates of the exchange taking place between them. Slow or fast exchange gives different information: in the case of slow exchange, the amount of free and complexed species present in equilibrium can be determined from integration of signal area, which allows an estimate of the stability constants. Under fast exchange, averaged signals of chemical shift are obtained and these can be used to calculate equilibrium constants. NMR can also give information on solution structure as the NMR-sensitive nuclei most heavily involved in anion binding will exhibit the largest changes in chemical shift and relaxation times. The disadvantages of NMR are that it can suffer from low sensitivity resulting from the high association, and relatively high concentrations (10^{-2}–10^{-3} M) of both host and guest are required to obtain appreciable NMR signals [32]. NMR also cannot be used in the presence of some paramagnetic nuclei (e.g. Cu^{2+}).

Anion coordination can also affect the electromagnetic spectrum of the anion and/or the receptor in the 10^{-8}–10^{-4} m wavelength region, which allows the use of UV–Vis spectrophotometry and IR spectroscopy. Both of these techniques use the Beer–Lambert–Bouger law,

$$A_p = \sum \varepsilon_{ip} l C_i,$$

where A_p and ε_{ip} are, respectively, the absorbance of the solution and the molar absorption coefficient of the *i*th species at the wavelength denoted by the index *p*, C_i is the concentration of the *i*th species, and *l* is the optical path length [30]. This relates the absorption data obtained to concentrations, which can subsequently be used to calculate equilibrium constants for anion binding. These techniques are normally used to complement other methods [33], as they are less accurate, but they have also been used as the dominant technique in some cases [32,34] because they have the advantage that studies can take place at low and high pH and there is no restriction to aqueous solution. Where ε_{ip} is high, they are applicable to low concentrations. Also, as in NMR, structural information about the anion-receptor complexes can be discovered [33].

A rapidly growing and popular way of monitoring anion concentrations is the use of anion selective electrodes [30]. There are several different metal ion selective electrodes available but the range of anion selective electrodes is still relatively small. Those for halides are now widely available and successfully used, often in combination with a glass electrode, to give information on the free concentration of the anion under study and the hydrogen ion concentration [35]. They can operate in water and in polar solvents such as acetonitrile.

For analysis by ICP-OES, elements generally require one or more atomic or ionic emission lines in the 160 to 900 nm region to be detected accurately. This technique is not as effective for non-metallic elements (e.g. carbon, halogens) [27] as either their emission lines do not fall in the correct range or they have poor detection limits (10–100 ppb). Sulfur is an exception as it has emission lines at low wavelengths (180–190 nm) capable of giving reliable results. The sulfate anion can therefore be analysed by ICP-OES. In metal salt extraction this allows the amount of metal and sulfate extracted by a ligand to be

TABLE I Radii of selected isoelectronic cations and anions [37]

Cation	*r* (Å)	Anion	*r* (Å)
Na^+	1.16	F^-	1.19
K^+	1.52	Cl^-	1.67
Rb^+	1.66	Br^-	1.82
Cs^+	1.81	I^-	2.06

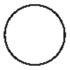
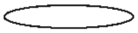
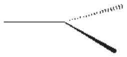
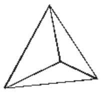
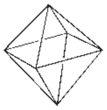
Anion Structure	Shape	Anions
	Spherical	F^-, Cl^-, Br^-, I^-
	Linear	N_3^-, SCN^-, OH^-, CN^-
	Trigonal Planar	CO_3^{2-}, NO_3^-
	Tetrahedral	$H_2PO_4^{2-}$, ClO_4^-, SO_4^{2-}, BF_4^-
	Octahedral	$[Fe(CN)_6]^{4-}$, PF_6^-

FIGURE 16 The different structures of anions [38].

analysed simultaneously as the ICP-OES can perform multi-element analysis.

The selectivity of metal ion extraction can be assessed by the determination of differences in $pH_{1/2}$ values (Fig. 14) [4]. If $pH_{1/2}$ values of sulfate complexation can be obtained by monitoring the extraction of sulfate over a range of pH, then they could be used to assess the 'strength' of sulfate binding in a two-phase solvent extraction experiment. In this case, the higher the $pH_{1/2}$, the 'stronger' the binding as complexation will only occur when protonation of the anion-binding site (e.g. tertiary amine groups) is favourable. If the concentrations of other anions in the organic phase can also be determined over a range of pH by using some of the anion analysis techniques used in the literature, then anion selectivity in a solvent extraction experiment could be determined using the same principles as metal cation selectivity. This is shown hypothetically in Fig. 15 for a system in which Cl^- is extracted at a higher pH than $H_2PO_4^-$ and SO_4^{2-} so the 'strength' of Cl^- binding is greater. This novel approach to assessing anion selectivity should allow rapid and efficient determinations of the 'strength' of complexation in a two-phase solvent extraction system.

DEVELOPING SULFATE-SELECTIVE ANION RECEPTORS

New methods are required for the treatment of metal sulfide ores and one potential strategy involves the development of extractants for metal sulfates (see Section 5). This is complicated because selective binding is needed for both the targeted metal cation and the sulfate anion to maintain good material balances. The more challenging of these targets is to design sulfate-selective anion receptors.

Anion coordination chemistry is an emerging area of supramolecular chemistry. Cation recognition is already very well developed and it has taken time for the coordination chemistry of anions to receive a significant level of interest [30]. This is surprising when the important part anions play in biology, medicine, catalysis and the environment is considered. For example, in biological systems, anions carry genetic information (as DNA is a polyanion) and most of the enzyme substrates and cofactors are anionic in nature [36]. The main reason that less attention has been given to the design of anion receptors is its challenging nature. Anions are larger than their isoelectronic cations (Table I), resulting in a lower charge-to-radius ratio [37]. This means that they are more difficult to bind electrostatically (cations are much smaller so their charge is more concentrated). Another difficulty associated with

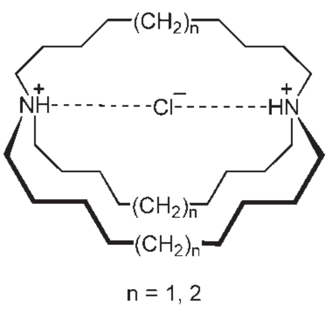


FIGURE 17 Example of chloride katapinates [39].

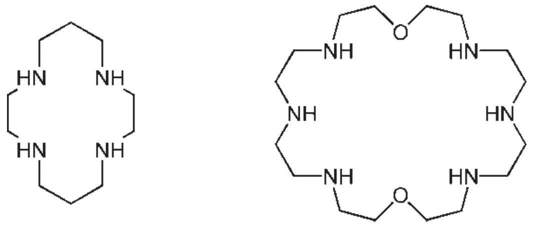


FIGURE 18 Early macrocyclic anion receptors [40,41].

anions is that they are subject to protonation as the pH is lowered. Receptors must therefore be designed to work in a specific pH range as the anion's charge will decrease with pH.

Anionic species also have a variety of different structures (some common ones are shown in Fig. 16) [38]. The design of receptors is therefore specific for the anion's particular geometry and is far more complicated than for the inherently spherical cations.

The first example of a synthetic anion inclusion complex was reported in 1968 by Park and Simmons [39]. They discovered a new type of ion-pairing in which a halide guest could occupy the cavity of an organic cage. These complexes became known as the 'halide katapinates' (Fig. 17) and were the basis for most of the early work on anion complexation. This work mainly consisted of synthesis of a large number of macrocyclic oligoamines (two are shown in Fig. 18) [40,41]. The designs used ammonium groups to facilitate anion binding.

Work has also been carried out on a variety of macro-oligocyclic ligands (Fig. 19) [42,43]. These created a more three-dimensional cavity for the anion. The tetrahedral ligand (shown on the right of Fig. 19) uses purely electrostatic anion binding, in contrast to ammonium groups, which have a combination of electrostatic and hydrogen bonding. Several anions, with very different shapes (spherical, tetrahedral, etc.) and charges, were found to bind to these macrocyclic ligands with a range of selectivities dependent on cavity size and protonation state [42,43].

There are now many different types of anion receptors in the literature [44]. Many attempts to bind anions have involved the use of amides because of their ability to function as hydrogen-bond donors to anions. The hydrogen bonding is directional and is therefore very useful when designing receptors for anions of specific shapes and geometries. The first purely amide-based anion receptor was developed in 1986 by Pascal *et al.* [45] and it was found to bind fluoride anions in deuterated DMSO (Fig. 20).

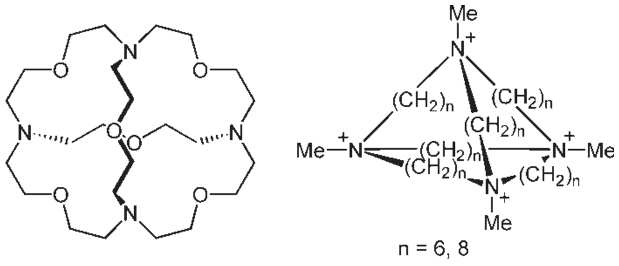


FIGURE 19 Early anion receptors with three-dimensional cavities [42,43].

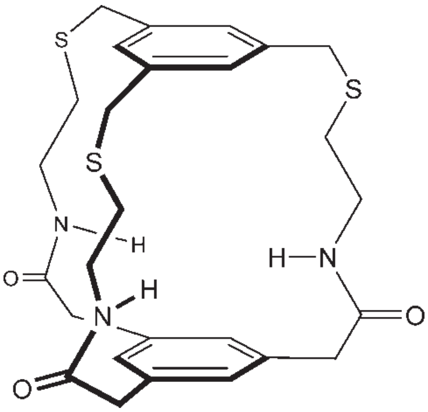


FIGURE 20 The first purely amide-based anion receptor developed by Pascal *et al.* [45].

In 1993, Reinhoudt and co-workers developed the design possibilities further by synthesizing a variety of acyclic tripodal anion receptors containing amide groups [46]. The ligands were all C_3 symmetric and therefore set up to bind tetrahedral anions (Fig. 21).

Urea and thiourea are also very good hydrogen-bond donors and are particularly effective because of their excellent directed hydrogen bonding (Fig. 22), which facilitates the binding of bidentate anions with high stability constants. Y-shaped anions such as carboxylate are especially strongly bound [44].

Derivatives of simple (thio)urea receptors were developed by Umezama and co-workers [47] to take advantage of this ability to bind more complex anions selectively. In one of the ligands, two thiourea groups were used to create a rigid structure in which dihydrogen phosphate could be bound. High stability constants were obtained and this was credited to the pre-organization in the receptor (Fig. 23).

Pyrrole NH groups have become one of the most useful anion-binding motifs [48]. Unlike amides, they contain no hydrogen-bond acceptors, which can often cause intra- or intermolecular hydrogen bonding that weakens a receptor as anion-binding competes with self-association.

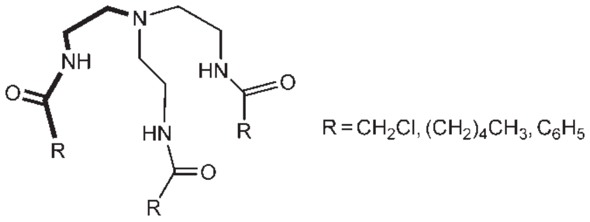


FIGURE 21 Tripodal amide receptors that show selective binding of H_2PO_4 [46].

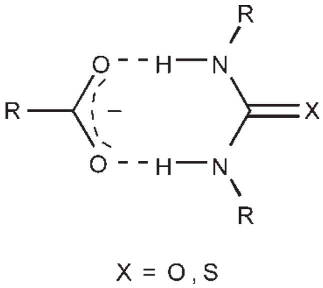


FIGURE 22 The ideal two-point interaction between (thio)urea and carboxylate anions [44].

Another advantage is that pyrroles are neither very basic nor acidic so they can sustain an anion-receptor complex under a number of conditions. They are also very versatile as they can be easily functionalized and incorporated into other structures [48]. They were first discovered to be useful in the area of anion recognition in 1990, when a diprotonated pentapyrrolic macrocycle (Fig. 24), known as a sapphyrin, was found by chance to bind a fluoride anion within hydrogen-bonding distance of all five NH groups in the solid state [49]. A crystal structure was later obtained of the same macrocycle binding chloride by hydrogen bonds; however, this time one chloride was bound above the cavity and the other was bound below. This suggested that fluoride would be more strongly bound in solution as it fitted better into the cavity. This was confirmed by fluorescence titration experiments, which showed a selectivity of three orders of magnitude for fluoride over chloride [50].

Pyrrole itself has recently been found to stabilize an anion complex in the solid state, but these interactions were not found to persist significantly in solution [51] and highlighted the fact that pyrrole may work most efficiently as an important component of a larger anion receptor.

This concept has recently been developed by Gale and co-workers [52–54], who have synthesized a series of 2,5-diamidopyrroles. Many of those developed have shown particular selectivity for oxo-anions. The receptor (a) in Fig. 25 was found to bind the benzoate anion with an association constant of $2.5 \times 10^3\,M^{-1}$, in contrast to the monofunctionalized receptor (b), which has an association constant of only $202\,M^{-1}$. The higher affinity was attributed to the ability of the 2,5-diamidopyrrole to form three hydrogen bonds to the benzoate [55] rather than just the two possible in (b). This was later confirmed by a crystal structure of the complex showing how well-positioned anion-binding groups can be used in combination to create a more effective receptor framework.

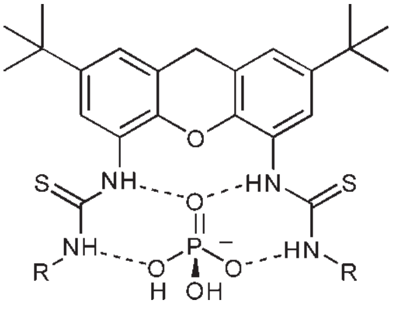


FIGURE 23 Preorganized thiourea receptor for complex anions [47].

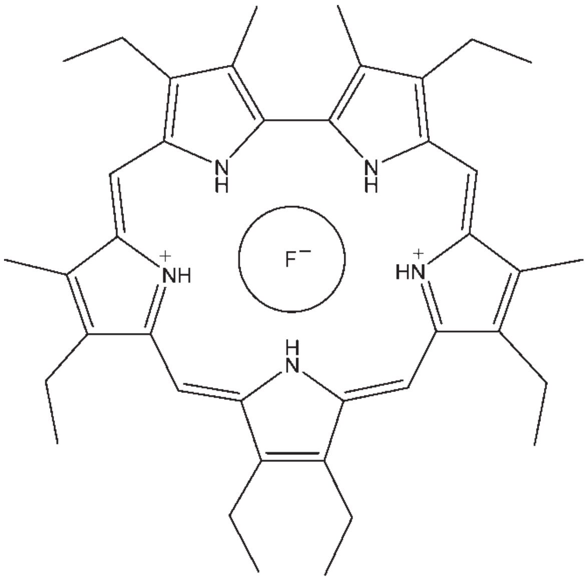


FIGURE 24 Diprotonated pentapyrrolic macrocycle with coordinated fluoride [49].

If a metal sulfate extractant is to be developed for the treatment of metal sulfide ores it is desirable to have an anion-binding site that is preferentially selective for sulfate over other anions. The literature has been lacking in synthetic sulfate-selective receptors and only a handful of examples had been published until very recently.

One of the most successful designs has involved the use of guanidinium groups [56]. The guanidinium moiety is set up very much like ureas in that it can form two hydrogen bonds to anions such as carboxylate. In addition, it is also positively charged resulting in the formation of more strongly bound complexes with anions. It is protonated over a wide pH range ($pK_a \sim 13.5$) and this obviously aids anion binding. In an example in Fig. 26, two bicyclic guanidine units are used in a rigid ligand backbone containing an aromatic spacer to induce the selectivity [56]. The receptor is designed for tetrahedral anions as the two bicyclic guanidines align perpendicular to each other. The system shows a remarkable selectivity for sulfate over hydrogen phosphate, indicating that this particular disposition of the guanidinium groups is favoured by the sulfate anion.

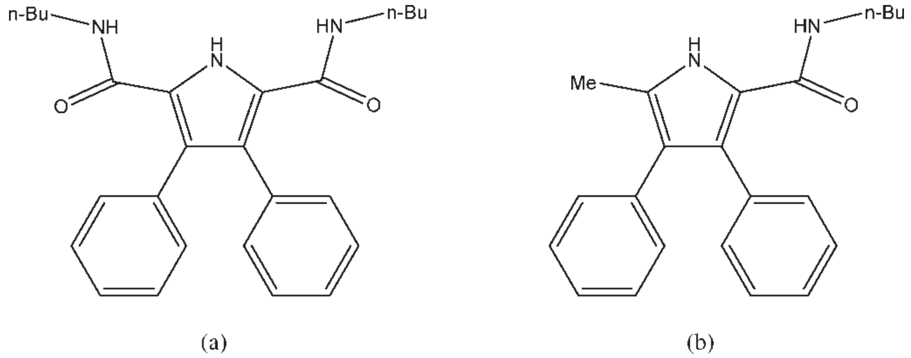


FIGURE 25 A 2,5-diamidopyrrole (a) and a 2-amido-5-methylpyrrole (b) for anion binding [54].

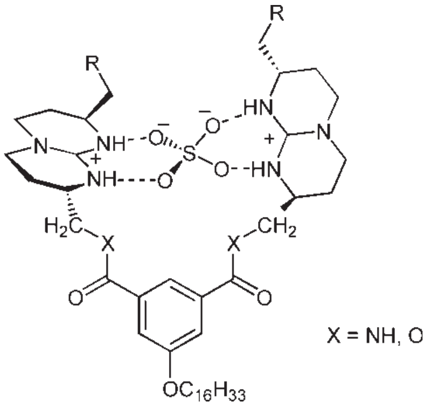


FIGURE 26 Ditopic guanidinium receptors for sulfate extraction [56].

The efficacy of guanidinium units to bind sulfate was also shown by de Mendoza and co-workers [57], who demonstrated how two strands of a guanidinium-based compound (Fig. 27) could fold in a double helical structure around a sulfate anion. Unfortunately, guanidinium groups would not be suitable components of reagents for most hydrometallurgical processes as they are difficult and expensive to synthesize, and the high pK_a of the unit would make it very difficult to strip the sulfate and recycle the ligand.

Sessler *et al.* have synthesized a diamidodipyrromethane macrocycle that shows an extraordinarily high sulfate over nitrate selectivity (Fig. 28) [58], which is needed for efficient removal of the sulfate anion from nitrate-rich radioactive waste streams [59]. This is required because, unlike nitrate anions, they can adversely affect the treatment process. The macrocycle is a hybrid of the well-known anion-binding motif pyridine-2,6-dicarboxamide [60] and dipyrromethane fragments, which are key anion-binding components in calixpyrroles (see Fig. 32) [51]. Development of the system demonstrates how anion-binding units can be combined in novel ways to alter selectivity patterns.

Two very different but very successful receptors for sulfate have been reported by Gale and co-workers. The first make use of his well-developed 2,5-diamidopyrrole backbone, which has been used to create anion receptors with varying properties [52–54]. In this case, amine pendant arms were attached to the 2,5-diamidopyrrole unit (Fig. 29(a)) and the resultant receptor showed a particularly high affinity for the HSO_4^- anion [61]. This was attributed to a proton transfer from the HSO_4^- anion to the receptor to produce a more highly charged receptor and the dianion SO_4^{2-}. Indeed, when the analogous ammonium receptor was prepared, only a very weak association was observed, probably due to the absence of proton exchange. This demonstrates how the nature of a receptor's pendant groups can have a profound effect on its anion-binding properties. The second design is based on a Pt(II) complex of four urea-functionalized isoquinoline ligands (Fig. 29(b)), which shows selectivity for sulfate over chloride, bromide and iodide [62]. The complex selectively encapsulates sulfate in a 'cone' formation in both solution and the solid state by a combination of electrostatic attraction from the Pt centre and

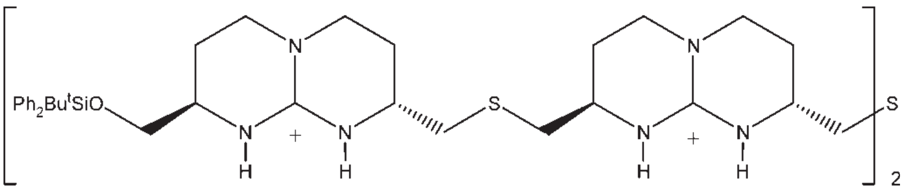


FIGURE 27 A guanidinium receptor for sulfate encapsulation [57].

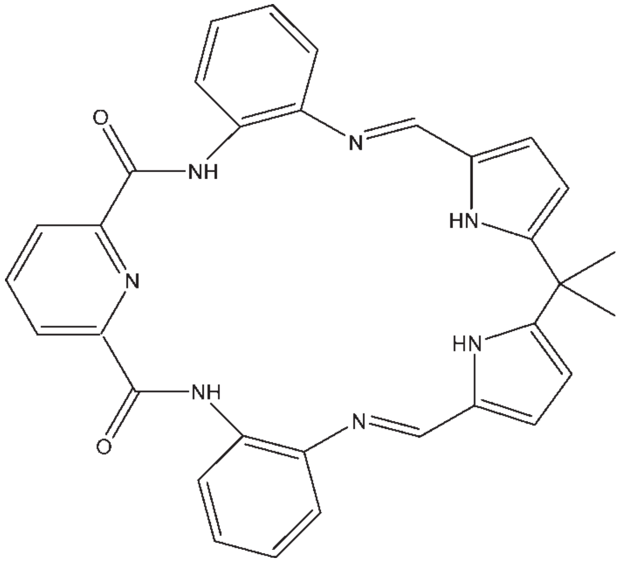


FIGURE 28 A new diamidodipyrromethane macrocycle with high SO_4^{2-} vs. NO_3^- selectivity [58].

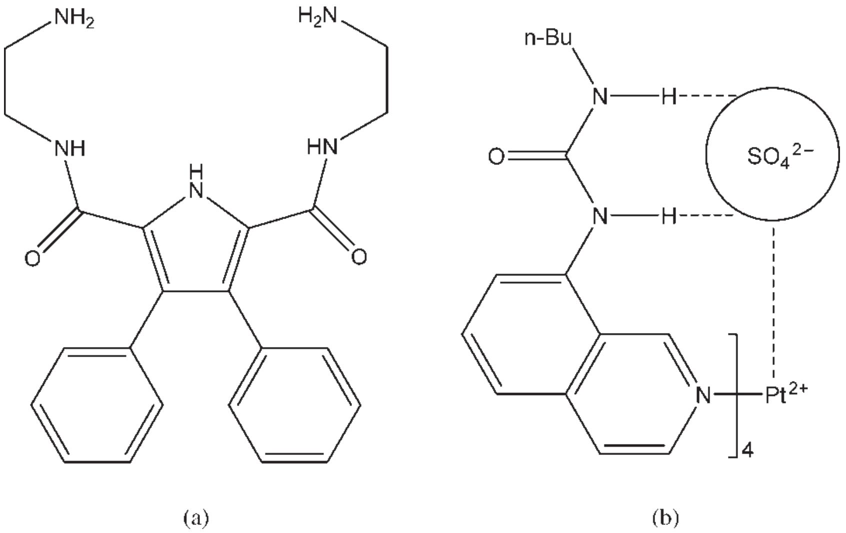
hydrogen bonding from the ureas. In fact, the urea groups contribute a total of eight hydrogen bonds to three of the four sulfate oxygens, and this is made possible by the flexibility of the receptor.

A similar number of hydrogen-bond donors are present in the sulfate-binding protein (SBP), which has been developed by nature for the selective transport of the sulfate anion [63]. This natural anion receptor is a specialized protein that binds sulfate in a solvent-inaccessible cavity by means of seven hydrogen bonds. These are received from the peptide backbone NH, serine OH and tryptophan NH groups. The SBP, in contrast to the phosphate-binding protein (PBP), only binds fully deprotonated anions, and this is facilitated by the absence of any of the hydrogen-bond acceptors that are important to the operation and selectivity of the PBP [30]. This illustrates how very small differences in receptor design can profoundly affect selectivity.

Hydrogen bonds are very important for selective anion encapsulation and an electrostatic interaction is not always necessary. This has been especially well proven by Kubik *et al.*, who have developed a neutral anion receptor that uses only hydrogen bonds for recognition but still shows appreciable anion selectivity in water/methanol solution [64]. This is particularly notable because neutral receptors often fail to function in aqueous media because of its competitive nature [64]. The compound consists of two cyclohexapeptide subunits containing L-proline and 6-aminopicolinic acid subunits, in an alternating sequence, which are connected by an adipic acid spacer (Fig. 30).

The receptor shows a particularly high affinity and selectivity for sulfate in aqueous solution, forming 1:1 complexes. The encapsulation of the anion is likened to the storage of a pearl in an oyster as the two cyclohexapeptide subunits operate as two 'shells' that enclose the sulfate anion ('pearl') by means of an adipic acid spacer or 'hinge'. This has led to the use of the term 'molecular oyster'. The anion is bound exclusively by hydrogen bonds in the cavity, sufficiently strongly to prevent aqueous solvation, and is therefore a useful model for the SBP.

The design of selective sulfate receptors for incorporation into reagents for extractive metallurgy is clearly a challenging target as there are only a few literature examples. However, the successful systems have very varied structures, opening up a number of design possibilities, particularly because very subtle changes to the receptor framework can have a significant effect on the selectivity. In addition to the challenge of *selectivity* of transport of sulfate into



(a) (b)

FIGURE 29 Two different but effective sulfate receptors developed by Gale and co-workers [61,62].

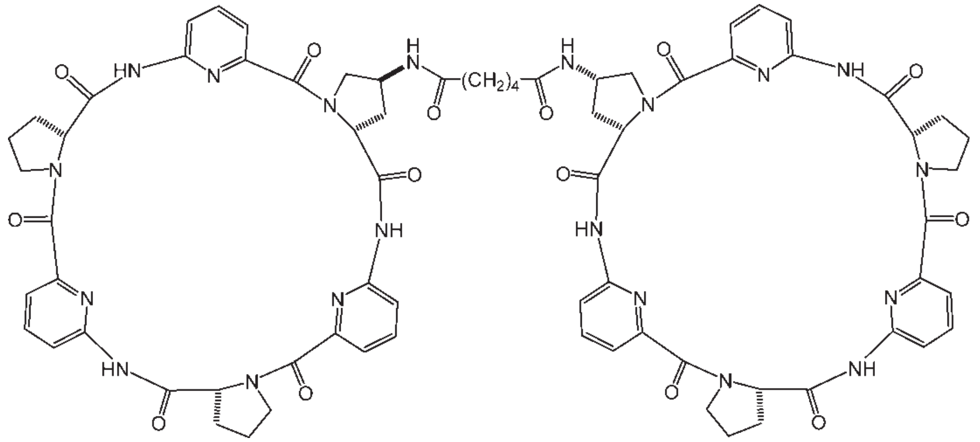


FIGURE 30 A neutral receptor with a remarkable sulfate affinity in aqueous solution [64].

a water-immiscible solvent, we face another problem, the design of reagents with sufficient '*strength*' to effect transfer from water, given the very high hydration energies of sulfate. This problem is related to the Hofmeister series (or bias).

THE HOFMEISTER SERIES

The limited number of selective sulfate receptors reported in the literature is surprising, considering the important role the sulfate anion plays in biology [63] and the problems it causes in radioactive waste treatment processes [59]. The difficulties associated with generating an efficient sulfate receptor are strongly linked to a phenomenon known as the Hofmeister series, which was first noted in 1888 when Franz Hofmeister was conducting experiments to rank the ability of various ions to precipitate a mixture of hen egg white proteins (Fig. 31) [65]. The series was found to have many applications and has been used to explain a body of data on the behaviour of water at interfaces and a number of associated phenomena [66].

In the specific area of solvent extraction of ionic materials, it has been suggested that the term 'Hofmeister bias' is used [67]. This can be used to predict why an anion such as ClO_4^- will generally be extracted from aqueous media into a water-immiscible solvent in preference to an anion such as SO_4^{2-}. The reason is that SO_4^{2-} is much more heavily hydrated than ClO_4^-, and therefore more hydrophilic, so it prefers to stay in an aqueous environment. In general, anion extraction is more favourable for large, charge-diffuse anions and distribution coefficients can be predicted from the Gibbs free energies of solvation/hydration [30,67]. To reverse the Hofmeister bias, an anion-binding solvent extractant is required to bind sufficiently strongly to compensate for dehydration associated with transfer to the organic phase. This is often carried out by using hydrogen-bond donors [68] or Lewis acid groups [30]. Numerous attempts have been made but, as yet, only limited success has been achieved in reducing or reversing the Hofmeister bias.

The selective solvation of hydrophilic phosphate anions is achieved in nature by the PBP by means of 12 hydrogen bonds [30]. This level of specificity is very difficult to achieve when designing receptors, but some systems, such as the fluorinated calixpyrroles developed by Sessler *et al.* [69], have had some success. The neutral receptors were found to extract caesium salts of smaller anions as effectively as the large charge-diffuse iodide. For example, β-octafluoro-*meso*-octamethylcalix[4]pyrrole (Fig. 32) extracts chloride as effectively as iodide into nitrobenzene.

An alternative strategy for the recovery of metal salts with anions that have very high hydration energies may be to design receptors that can transport partially hydrated anions. It has recently been shown that, in the solid state, the copper sulfate complex of a macrocyclic ligand (shown in Fig. 33) has sulfate encapsulated as a water-bridged dimer [70]. The macrocyclic ligand shows a greater 'strength' of sulfate transfer in solvent extraction than closely related ligands, indicating that this effect may be important for

strongly hydrated anions	**weakly hydrated anions**

Citrate^{3-} > SO_4^{2-} = tartrate^{2-} > HPO_4^{2-} > CrO_4^- > acetate$^-$ > HCO_3^- > Cl^- > NO_3^- > ClO_4^-

FIGURE 31 The original Hofmeister series [65].

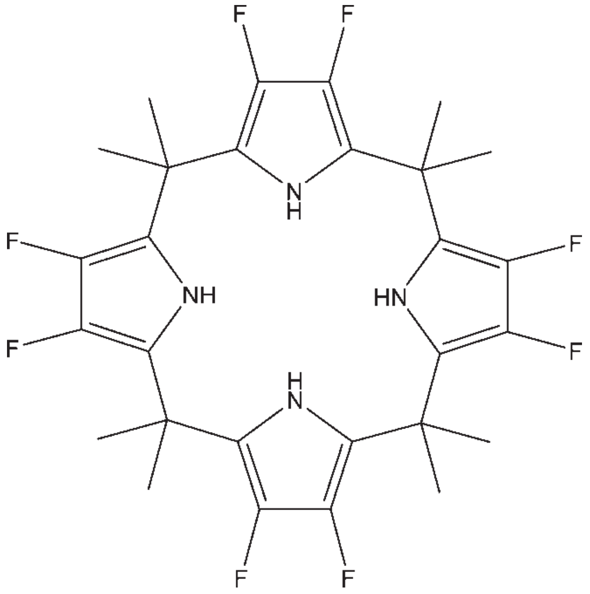


FIGURE 32 A β-fluorinated calyx [4]pyrrole that extracts Cl^- and Br^- as effectively as I^- [69].

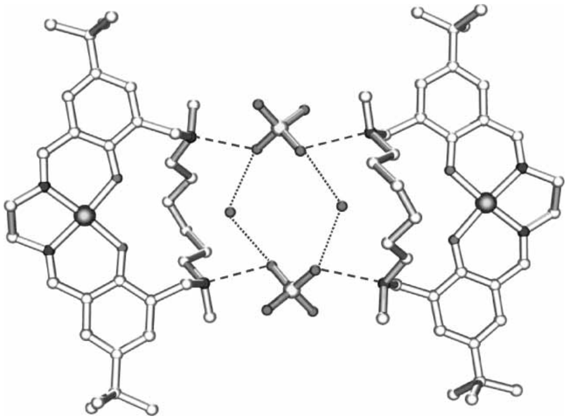


FIGURE 33 The tetra-anionic water-bridged sulfate dimer formed in the solid state between two macrocyclic copper complexes [70].

the efficient transport of strongly hydrated anions [70].

The design of receptors that can reverse the Hofmeister bias is a key requirement for effecting the separation of hydrophilic anions in radioactive waste remediation and various other applications such as new base-metal recovery processes, which will exploit new leaching technologies for sulfidic ores and generate sulfate pregnant leach solutions.

CONCLUSION

Although efficient metal recovery can be facilitated by several different techniques such as pyrometallurgy, froth flotation and hydrometallurgy, it is the latter that shows most promise for the development of greener processes using a 'systems engineering' approach to minimize environmental impact [71]. To open up new flowsheets to process sulfidic ores of base metals, reagents that transport metal sulfates with high selectivity are needed to effect the unit operations of *concentration* and *separation*. Recent advances in the development of anion-selective reagents described in this review provide a variety of design possibilities that might lead to practicable extractants and should encourage the collaboration of supramolecular chemists and metallurgists.

References

[1] Shriver, D. F.; Atkins, P. W.; Langford, C. H. *Inorganic Chemistry*; 2nd Ed. W.H. Freeman and Co. New York, 1994.

[2] Habashi, F. *Miner. Eng.* **1994**, 7.

[3] Swaddle, T. W. *Inorganic Chemistry: An Industrial and Environmental Perspective*; Academic Press: Toronto, 1997.

[4] Szymanowski, J. *Hydroxyoximes and Copper Hydrometallurgy*; CRC Press, Inc. London, 1993.

[5] Tasker, P. A.; Plieger, P. G.; West, L. C. In *Comprehensive Coordination Chemistry*; McCleverty, J. A, Meyer, T. B., Eds.; Vol. 9, Elsevier: Oxford, 2004; p 759.

[6] Nicol, M. J.; Fleming, C. A.; Preston, J. S. In *Comprehensive Coordination Chemistry*; Wilkinson, G., Ed.; Vol. 6, Pergamon Press: Oxford, 1987; p 779.

[7] Pletcher, D.; Walsh, F. C. *Industrial Electrochemistry*; Chapman and Hall: London, 1990.

[8] Rubisov, D. H.; Krowinkel, J. M.; Papangelakis, V. G. *Hydrometallurgy* **2000**, *58*, 1.

[9] Rubisov, D. H.; Papangelakis, V. G. *Hydrometallurgy* **2000**, *58*, 89.

[10] Sole, K. C. *International Solvent Extraction Conference Cape Town, South Africa, March 17–21*, **2002**, 1033.

[11] Miller, H. A.; Laing, N.; Parsons, S.; Parkin, A.; Tasker, P. A.; White, D. J. *J. Chem. Soc., Dalton Trans.* **2000**, 3773.

[12] White, D. J.; Laing, N.; Miller, H. A.; Parsons, S.; Tasker, P. A.; Coles, S. *J. Chem. Soc., Chem. Commun.* **1999**, 2077.

[13] Dalton, R. F.; Diaz, G.; Price, R.; Zunkel, A. D. *J. Miner. Met. Mater. Soc.* **1991**, *43*, 51.

[14] Szymanowski, J. *J. Radioanal. Nucl. Chem.* **1996**, *208*, 183.

[15] Cote, G.; Jakubiak, A.; Bauer, D.; Szymanowski, J.; Mokili, B.; Poitrenaud, C. *Solvent Extr. Ion Exch.* **1994**, *12*, 99.

[16] Akkus, N.; Campbell, J. C.; Davidson, J.; Henderson, D. K.; Miller, H. A.; Parkin, A.; Parsons, S.; Plieger, P. G.; Swart, R. M.; Tasker, P. A.; West, L. C. *J. Chem. Soc., Dalton Trans.* **2003**, 1932.

[17] Rudkevich, D. M.; Brzozka, Z.; Palys, M.; Visser, H. C.; Verboom, W.; Reinhoudt, D. N. *Angew. Chem. 106* **1994**, 480.

[18] Beer, P. D.; Chen, Z.; Gale, P. A.; Heath, J. A.; Knubley, R. J.; Ogden, M. I.; Drew, M. G. B. *J. Inclusion Phenom. Mol. Recognit. Chem.* **1994**, *19*, 343.

[19] Beer, P. D.; Dent, S. W. *J. Chem. Soc., Chem. Commun.* **1998**, 825.

[20] Smith, A. G.; Tasker, P. A.; White, D. J. *J. Coord. Chem. Rev.* **2003**, *241*, 61.

[21] Di Bella, S.; Fragala, I.; Ledoux, I.; Diaz-Garcia, M. A.; Marks, T. J. *J. Am. Chem. Soc.* **1997**, *119*, 9550.

[22] Suresh, E.; Bhadbhade, M. M.; Srinivas, D. *Polyhedron* **1996**, *15*, 4133.

[23] Galbraith, S. G.; Plieger, P. G.; Tasker, P. A. *J. Chem. Soc., Chem. Commun.* **2002**, 2662.

[24] Cheng, Y.; Mason, A. B.; Woodworth, R. C. *Biochemistry* **1995**, *34*, 14879.

[25] Jennings, M. L.; Smith, J. S. *J. Biol. Chem.* **1992**, *267*, 13964.

[26] Kanyo, Z. F.; Christianson, D. W. *J. Biol. Chem.* **1991**, *266*, 4264.

[27] Settle, A. *Handbook of Instrumental Techniques for Analytical Chemistry*; Prentice Hall PTR: NJ, 1997.

[28] Bianchi, A.; Micheloni, M.; Paoletti, P. *Coord. Chem. Rev.* **1991**, *110*, 17.

[29] Izatt, R. M.; Pawlak, K.; Bradshaw, J. S.; Bruening, R. L. *Chem. Rev. 91* **1991**, 1721.

[30] Bianchi, A., Bowman-James, K., Garcia-Espana, E., Eds.; *Supramolecular Chemistry of Anions*; Wiley-VCH: New York, 1997.

[31] Connors, K. A. *Binding Constants: The Measurements of Molecular Complex Stability*; Wiley-VCH: New York, 1987.

[32] Xie, H.; Yi, S.; Wu, S. *J. Chem. Soc., Perkin Trans. 2*, **1999**, 2751.

[33] Jagessar, R. C.; Burns, D. H. *J. Chem. Soc., Chem. Commun.* **1997**, 1685.

[34] Xie, H.; Yi, S.; Yang, X.; Wu, S. *New J. Chem.* **1999**, *23*, 1105.

[35] Beck, M. T.; Nagypal, I. *Chemistry of Complex Equilibria*; Akedemiai Kiado: Budapest, 1990.

[36] Christianson, D. W.; Lipscomb, W. N. *Acc. Chem. Res.* **1989**, *22*, 62.

[37] Shannon, R. D. *Acta Crystallogr., Sect. A* **1976**, *A32*, 751.

[38] Beer, P. D.; Gale, P. A. *Angew. Chem., Int. Ed. Engl.* **2001**, *40*, 486.

[39] Park, C. H.; Simmons, H. E. *J. Am. Chem. Soc.* **1968**, *90*, 2431.

[40] Dietrich, B.; Hosseini, M. W.; Lehn, J. M.; Sessions, R. B. *J. Am. Chem. Soc.* **1981**, *103*, 1282.

[41] Kodama, M.; Kimura, E. *J. Chem. Soc., Dalton Trans.* **1977**, 1473.

[42] Graf, E.; Lehn, J. M. *J. Am. Chem. Soc.* **1975**, *97*, 5022.

[43] Schmidtchen, F. P. *Angew. Chem. 89* **1977**, 751.

[44] Gale, P. A. *Coord. Chem. Rev.* **2000**, *199*, 181.

[45] Pascal, R. A.; Spergel, J.; Van Engen, D. *Tetrahedron Lett.* **1986**, *27*, 4099.

[46] Valiyaveettil, S.; Engbersen, J. F. J.; Verboom, W.; Reinhoudt, D. N. *Angew. Chem.* **1993**, *105*, 942.

[47] Buhlmann, P.; Nishizawa, S.; Xiao, K. P.; Umezawa, Y. *Tetrahedron* **1997**, *53*, 1647.

[48] Sessler, J. L.; Camiolo, S.; Gale, P. A. *Coord. Chem. Rev.* **2003**, *240*, 17.

[49] Sessler, J. L.; Cyr, M. J.; Lynch, V.; McGhee, E.; Ibers, J. A. *J. Am. Chem. Soc.* **1990**, *112*, 2810.

[50] Sessler, J. L.; Cyr, M.; Furuta, H.; Kral, V.; Mody, T.; Morishima, T.; Shionoya, M.; Weghorn, S. *Pure Appl. Chem.* **1993**, *65*, 393.

[51] Gale, P. A.; Sessler, J. L.; Kral, V.; Lynch, V. *J. Am. Chem. Soc.* **1996**, *118*, 5140.

[52] Camiolo, S.; Gale, P. A.; Hursthouse, M. B.; Light, M. E.; Shi, A. J. *J. Chem. Soc., Chem. Commun.* **2002**, 758.

[53] Denuault, G.; Gale, P. A.; Hursthouse, M. B.; Light, M. E.; Warriner, C. N. *New J. Chem.* **2002**, *26*, 811.

[54] Gale, P. A.; Camiolo, S.; Chapman, C. P.; Light, M. E.; Hursthouse, M. B. *Tetrahedron Lett.* **2001**, *42*, 5095.

[55] Camiolo, S.; Gale, P. A.; Hursthouse, M. B.; Light, M. E. *Tetrahedron Lett.* **2002**, *43*, 6995.

[56] Stephan, H.; Gloe, K.; Schiessl, P.; Schmidtchen, F. P. *Supramol. Chem.* **1995**, *5*, 273.

[57] Sanchez-Quesada, J.; Seel, C.; Prados, P.; de Mendoza, J.; Dalcol, I.; Giralt, E. *J. Am. Chem. Soc.* **1996**, *118*, 277.

[58] Sessler, J. L.; Katayev, E.; Pantos, G. D.; Ustynyuk, Y. A. *J. Chem. Soc., Chem. Commun.* **2004**, 1276.

[59] Vienna, J. D.; Schweiger, M. J.; Smith, M. D.; Smith, D. E.; Crum, J. V.; Peeler, D. K.; Reamer, I. A.; Musick, C. A.; Tillotson, R. D.; Pacific Northwest National Laboratory: Washington, 1999.

[60] Kavallieratos, K.; Bertao, C. M.; Crabtree, R. H. *J. Org. Chem.* **1999**, *64*, 1675.

[61] Navakhun, K.; Gale, P. A.; Camiolo, S.; Light, M. E.; Hursthouse, M. B. *J. Chem. Soc., Chem. Commun.* **2002**, 2084.

[62] Bondy, C. R.; Gale, P. A.; Loeb, S. J. *J. Am. Chem. Soc.* **2004**, *126*, 5030.

[63] Pflugrath, J. W.; Quiocho, F. A. *Nature* **1985**, *314*, 257.

[64] Kubik, S.; Kirchner, R.; Nolting, D.; Seidel, J. *J. Am. Chem. Soc.* **2002**, *124*, 12752.

[65] Hofmeister, F. *Arch. Exp. Pathol. Pharmakol.* **1888**, 247.

[66] Collins, K. D.; Washabaugh, M. W. Q. *Rev. Biophys.* **1985**, *18*, 323.

[67] Moyer, B. A.; Bonnesen, P. V.; Delmau, L. H.; Haverlock, T. J.; Kavallieratos, K.; Levitskaia, T. G. *International Solvent Extraction Conference Cape Town, South Africa, March 17–21*, **2002**, 299.

[68] Kral, V.; Sessler, J. L.; Shishkanova, T. V.; Gale, P. A.; Volf, R. *J. Am. Chem. Soc.* **1999**, *121*, 8771.

[69] Levitskaia, T. G.; Marquez, M.; Sessler, J. L.; Shriver, J. A.; Vercouter, T.; Moyer, B. A. *J. Chem. Soc., Chem. Commun.* **2003**, 2248.

[70] Plieger, P. G.; Tasker, P. A.; Galbraith, S. G. *J. Chem. Soc., Dalton Trans.* **2004**, 313.

[71] Richards, D. J. *The Industrial Green Game*; National Academy Press: Washington DC, 1997.