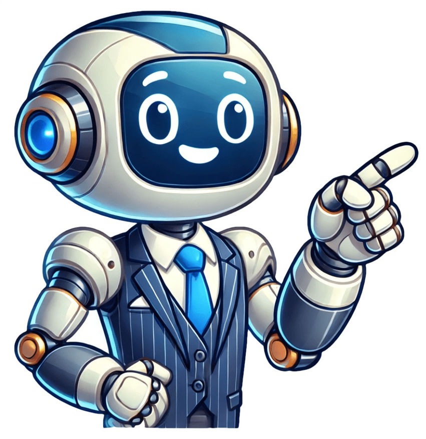
Continue



Ligands can be further characterized as monodentate, bidentate, tridentate etc. where the concept of teeth (dent) is introduced. Monodentate ligands bind through only one donor atom. Monodentate means "one-toothed." The halides, phosphines, ammonia and amines seen previously are monodentate ligands. Bidentate ligands bind through two donor sites. Bidentate means "two-toothed." An example of a bidentate ligand is ethylenediamine. It can bind to a metal via two donor atoms at once. Figure \(\PageIndex{1}\): The hypothetical "lobster ligand" binds to the \(Ni^{2+}\) via two donor sites. Bidentate binding allows a ligand to bind more tightly. Tridentate ligands, which bind through three donors, can bind even more tightly, and so on. This phenomenon is generally called the "chelate effect." This term comes from the Greek chelos, meaning "crab." A crab does not have any teeth at all, but it does have two claws for tightly holding onto something for a couple of reasons. A very simple analogy is that, if you are holding something with two hands rather than one, you are not as likely to drop it. In the examples previously disccussed, each ligand only forms one bond with the central metal ion to give the complex ion. Such a ligand is said to be unidentate. That means literally that it only has one tooth! It only has one pair of electrons that it can use to bond to the metal - any other lone pairs are pointing in the wrong direction. Some ligands, however, have rather more teeth! These are known generally as multidentate or polydentate ligands, but can be broken down into a number of different types. Bidentate ligands have two lone pairs, both of which can bond to the central metal ion. The two commonly used examples are 1,2-diaminoethane (old name: ethylenediamine - often given the abbreviation "en"), and the ethanedioate ion (old name: oxalate). In the ethanedioate ion, there are lots more lone pairs than the two shown, but these are the only ones we are interested in. You can think of these bidentate ligands rather as if they were a pair of headphones, carrying lone pairs on each of the "ear pieces". These will then fit snuggly around a metal ion. You might find this abbreviated to \([Ni(en)_3]^{2+}\). The structure of the ion looks like this: In this case, the "ear pieces" are the nitrogen atoms of the NH2 groups - and the "bit that goes over your head " is the \(-CH_2CH_2-\) group. If you were going to draw this in an exam, you would obviously want to draw it properly - but for learning purposes, drawing all the atoms makes the diagram look unduly complicated! Notice that the arrangement of the bonds around the central metal ion is exactly the same as it was with the ions with 6 water molecules attached. The only difference is that this time each ligand uses up two of the positions - at right angles to each other. Because the nickel is forming 6 co-ordinate bonds, the coordination number of this ion is 6, despite the fact that it is only joined to 3 ligands. Coordination number counts the number of bonds, not the number of ligands. This is the complex ion formed by attaching 3 ethanedioate (oxalate) ions to a chromium(III) ion. The shape is exactly the same as the previous nickel complex. The only real difference is the number of charges. The original chromium ion carried 3+ charges, and each ethanedioate ion carried -2, i.e., \[(+3) + (3 \times -2) = -3. onumber \] The structure of the ion looks like this: Again, if you drew this in an exam, you would want to show all the atoms properly. If you need to be able to do this, practice drawing it so that it looks clear and tidy! Refer back to the diagram of the ethanedioate ion further up the page to help you. A quadridentate ligand has four lone pairs, all of which can bond to the central metal ion. An example of this occurs in haemoglobin (American: hemoglobin). The functional part of this is an iron(II) ion surrounded by a complicated molecule called heme. This is a sort of hollow ring of carbon and hydrogen atoms, at the center of which are 4 nitrogen atoms with lone pairs on them. Heme is one of a group of similar compounds called porphyrins. They all have the same sort of ring system, but with different groups attached to the outside of the ring. You aren't going to need to know the exact structure of the haem at this level. We could simplify the heme with the trapped iron ion as: Each of the lone pairs on the nitrogen can form a co-ordinate bond with the iron(II) ion - holding it at the center of the complicated ring of atoms. The iron forms 4 co-ordinate bonds with the heme, but still has space to form two more - one above and one below the plane of the ring. The protein globin attaches to one of these positions using a lone pair on one of the nitrogens in one of its amino acids. The interesting bit is the other position. Figure \(\PageIndex{2}\): Overall, the complex ion has a co-ordination number of 6 because the central metal ion is forming 6 co-ordinate bonds. The water molecule which is bonded to the bottom position in the diagram is easily replaced by an oxygen molecule (again via a lone pair on one of the oxygens in \(O_2\)) - and this is how oxygen gets carried around the blood by the haemoglobin. When the oxygen gets to where it is needed, it breaks away from the haemoglobin which returns to the lungs to get some more. You probably know that carbon monoxide is poisonous because it reacts with hemeoglobin. It bonds to the same site that would otherwise be used by the oxygen - but it forms a very stable complex. The carbon monoxide doesn't break away again, and that makes that hemeoglobin molecule useless for any further oxygen transfer. A hexadentate ligand has 6 lone pairs of electrons - all of which can form co-ordinate bonds with the same metal ion. The best example is EDTA. The diagram shows the structure of the ion with the important atoms and lone pairs picked out. The EDTA ion entirely wraps up a metal ion using all 6 of the positions that we have seen before. The co-ordination number is again 6 because of the 6 co-ordinate bonds being formed by the central metal ion. The diagram below shows this happening with a copper(II) ion. Here is a simplified version. Make sure that you can see how this relates to the full structure above. The overall charge, of course, comes from the 2+ on the original copper(II) ion and the 4- on the \(EDTA^{4-}\) ion. Jim Clark (Chemguide.co.uk) A ligand is an atom or atom group that can donate a lone pair of electrons to a transition metal ion to form a complex through the formation of co-ordinate bonds. All ligands have lone pairs of electrons. Some common ligands are: water, ammonia, cyanide and chloride ions. Ligands in Complex Ions If ligands have one pair of electrons, they are called monodentate ligands. Examples include water, ammonia, chloride ions and cyanide ions. Water and ammonia are uncharged, whereas chloride and cyanide have a -1 charge. If ligands have two pairs of electrons, they are called bidentate ligands. Examples include 1,2-diaminoethane or the ethanedioate ion (C2O4^{2}-) Their structures are shown below: Ligands in Complex Ions Multidentate Ligands If ligands have many lone pairs of electrons, they are called multidentate ligands. The example you need to be familiar with is the ethylenediaminetetracetate ion or EDTA 4- ion for short. This has six pairs of electrons. Ligands in Complex Ions The co-ordination number is the number of co-ordinate bonds formed with the central transition metal atom or ion. Small ligands, like water and ammonia, have a co-ordination number of six, as six of these ligands can fit around the central transition metal atom or ion. Bidentate ligands also have a co-ordination number of six. This is because three of these molecules can fit around the central transition metal ion. EDTA 4- ion has six pairs of electrons. Each EDTA ion makes six co-ordinate bonds around the central transition metal atom. The coordination number is six. Ligands in Complex Ions Personalised tuition to meet your needs with flexible scheduling Larger ligands, like chloride ions, have a co-ordination number of four, as only four ligands can fit around the central transition metal atom or ion. Ligands in Complex Ions →What are transition metals? Transition metals are a group of elements located in the middle of the periodic table and are characterized by their ability to form multiple stable oxidation states. Some of the most well-known transition metals include iron, copper, and gold. →What are ligands? Ligands are chemical species that bond to transition metals to form coordination compounds. Ligands can be neutral molecules or ions, and they can bond to the metal through coordination bonds, which are covalent bonds between the metal and the ligand. →What is the role of ligands in transition metal chemistry? Ligands play a crucial role in transition metal chemistry by determining the oxidation state and coordination number of the metal, as well as the physical and chemical properties of the resulting coordination compound. →What are the different types of ligands? There are several types of ligands, including monodentate ligands, which bond to the metal through a single coordination bond, and polydentate ligands, which bond to the metal through multiple coordination bonds. Other types of ligands include chelating ligands, which form cyclic structures around the metal, and anionic ligands, which are negatively charged. →How do ligands affect the color of transition metal compounds? The color of transition metal compounds is often related to the energy required to promote an electron from a lower energy orbital to a higher energy orbital. Ligands can influence the energy levels of these orbitals, leading to changes in the color of the coordination compound. →What is the difference between coordination number and oxidation state? Coordination number refers to the number of coordination bonds between a transition metal and its ligands, while oxidation state refers to the charge on the metal ion in a coordination compound. The coordination number and oxidation state of a transition metal can be influenced by the type and number of ligands present. →Why is the study of transition metals and ligands important in A-Level Chemistry? The study of transition metals and ligands is important in A-Level Chemistry because it provides a fundamental understanding of the properties and behavior of these elements, which are important in many industrial and biological processes. This knowledge is essential for understanding topics such as catalysis, the coordination of metal ions in biological systems, and the synthesis of coordination compounds. →What are some common techniques used to study transition metals and ligands? Common techniques used to study transition metals and ligands include spectroscopy, such as ultraviolet-visible spectroscopy and infrared spectroscopy, and crystallography, which involves the study of the structure and properties of crystals. These techniques allow for the determination of the oxidation state and coordination number of transition metals and the identification of the ligands present in coordination compounds. Depending on the size of the ligands and the number of dative bonds to the central metal ion, transition element complexes have different geometriesDative bonds can also be referred to as coordinate bonds, especially when discussing the geometry of a complexCentral metal atoms or ions with two coordinate bonds form linear complexesThe bond angles in these complexes are 180oThe most common examples are a copper (I) ion, (Cu+), or a silver (I) ion, (Ag+), as the central metal ion with two coordinate bonds formed to two ammonia ligandsExamples of a linear complexThe second example is the diamminesilver(I) ion, $[Ag(NH_3)_2]^+$, which is present in Tollens' reagentTollens' reagent is used to test for the aldehyde functional group in organic moleculesIn the test, the silver(I) ion is reduced to silver atoms that produce a characteristic silver mirror on the test tube wallsTetrahedralWhen there are four coordinate bonds the complexes often have a tetrahedral shapeComplexes with four chloride ions most commonly adopt this geometryChloride ligands are large, so only four will fit around the central metal ionThe bond angles in tetrahedral complexes are 109.5oExample of a tetrahedral complexSquare planarSometimes, complexes with four coordinate bonds may adopt a square planar geometry instead of a tetrahedral oneCyanide ions (CN-) are the most common ligands to adopt this geometryAn example of a square planar complex is cisplatinThe bond angles in a square planar complex are 90oCisplatin is an example of a square planar complexOctahedralOctahedral complexes are formed when a central metal atom or ion forms six coordinate bondsThis could be six coordinate bonds with six small, monodentate ligandsExamples of such ligands are water and ammonia molecules and hydroxide and thiocyanate ionsIt could be six coordinate bonds with three bidentate ligandsEach bidentate ligand will form two coordinate bonds, meaning six coordinate bonds in totalExamples of these ligands are 1,2-diaminoethane and the ethanedioate ionIt could be six coordinate bonds with one multidentate ligandThe multidentate ligand, for example EDTA4-, forms all six coordinate bondsThe bond angles in an octahedral complex are 90oExamples of octahedral complexesTransition element complexes can exhibit stereoisomerismEven though transition element complexes do not have a double bond, they can still have geometrical isomersSquare planar and octahedral complexes with two pairs of different ligands exhibit cis-trans isomerism (this is a special case of E-Z isomerism)An example of a square planar complex with two pairs of ligands is the anti-cancer drug cis-platinWhereas cis-platin has beneficial medical effects by binding to DNA in cancer cells, trans-platin cannot be used in cancer treatmentCis-platin (the Z-isomer) and trans-platin (the E-isomer) is an example of a square planar transition element complex that exhibits geometrical isomerismAs long as a complex ion has two ligands attached to it that are different to the rest, then the complex can display geometric isomerismExamples of octahedral complexes that exhibit geometrical isomerism are the [Cu(NH3)4(H2O)2]2+ and [Ni(H2NCH2CH2NH2)2Cl2]2+ complexes[Ni(H2NCH2CH2NH2)2Cl2]2+ can also be written as [Ni(en)2Cl2]2+Like in the square planar complexes, if the two 'different' ligands are adjacent (next) to each other then that is the 'cis' isomer, and if the two 'different' ligands are opposite each other then this is the 'trans' isomerIn [Cu(NH3)4(H2O)2]2+, the two water ligands are adjacent to each other in the cis isomer and are opposite each other in the trans isomerOctahedral transition metal complexes exhibiting geometrical isomerismOptical isomerismOctahedral complexes with bidentate ligands also have optical isomersThis means that the two forms are non-superimposable mirror images of each otherThey have no plane of symmetry, and one image cannot be placed directly on top of the otherThe optical isomers only differ in their ability to rotate the plane of polarised light in opposite directionsExamples of octahedral complexes that have optical isomers are the [Ni(H2NCH2CH2NH2)3]2+and [Ni(H2NCH2CH2NH2)2(H2O)2]2+ complexesThe ligand H2NCH2CH2NH2 can also be written as 'en' insteadOctahedral transition metal complexes exhibiting optical isomerismDrawing stereochemical formulaeChemists use a convention of wedge drawings to represent three dimensional moleculesThe convention is thata solid line is a bond in the same plane as the papera dotted line is a bond receding behind the plane of the paper(this can also be hatched or shaded wedges)a solid wedge is a bond coming out of the paperFour main shapes of transition metal complexes using stereochemical formulaeDid this page help you? Ligands play a vital role in chemistry, particularly in coordination compounds. A ligand is a molecule or ion that can donate a pair of electrons to a central metal atom or ion. The donated electron pair forms coordinate bonds with the metal, forming coordination compounds. Ligands act as Lewis bases (electron pair donors), and the central atom acts as a Lewis acid (electron pair acceptor). The nature of metal-ligand bonding can range from covalent to ionic. Moreover, the metal-ligand bond order can range from one to three. [1-4] In nearly all situations, metals and metalloids form bonds with ligands. The ligand influences the reactivity of the central atom in a complex, ligand substitution rates, ligand reactivity, and redox properties. Selecting appropriate ligands is crucial in various practical fields, such as bioinorganic and medicinal chemistry, homogeneous catalysis, and environmental chemistry. The word ligand comes from Latin and means "tie or bind". Ligands can be classified into different types based on the number of binding sites with the central metal atom, charge, and size. [1-4] Monodentate ligands, also known as unidentate ligands, form only one bond with a metal ion. It means they have a single donor atom capable of coordinating with the metal center. Some examples of monodentate ligands include water (H2O), ammonia (NH3), chloride ion (Cl–), cyanide ion (CN–), and pyridine (C5H5N). Because they bind to the center through one atom only, they are called monodentate, meaning one tooth. These ligands possess specific properties that make them useful in coordination chemistry. Firstly, their ability to form a single bond allows for the formation of simple complexes, making them relatively easy to study and manipulate. Monodentate ligands also tend to be highly selective in their binding, as they have a specific site on the metal ion where coordination occurs. Another essential property of monodentate ligands is their ability to increase the solubility and stability of metal complexes. By coordinating with the metal ion, they can prevent unwanted precipitation or decomposition reactions, ensuring the longevity of the complex. It is worth noting that while monodentate ligands can only form one bond with a metal ion, they can still participate in multiple coordination sites if multiple ligands are present. It allows for forming more complex structures, such as polymeric or macrocyclic compounds. Bidentate ligands are molecules or ions with two donor atoms capable of forming two bonds with a metal ion. They are commonly used in coordination chemistry to form stable complexes with transition metals. Some examples of bidentate ligands include ethylenediamine (en), oxalate (ox), and acetylacetonate (acac). Bidentate ligands are often called chelating agents because they form a ring-like structure known as a chelate when coordinated with a metal ion. The term chelate comes from Greek, meaning claw. This term describes ligands that can grab the central ion in two or more places, like a claw. The binding affinity of a chelating system depends on the "bite angle" or chelating angle. The chelate structure provides increased stability to the complex. It has been found that the chelate effect is predominantly an effect of entropy. Bidentate ligands also tend to have a flexible structure, allowing them to adopt different conformations when coordinating with the metal ion. Polydentate ligands, also known as multidentate ligands, can donate multiple electron pairs to a central metal ion in a coordination complex. These ligands have multiple binding sites, which allow them to form multiple bonds with the metal ion simultaneously. Molecules with three donor atoms are called tridentate; four are called tetradentate, five are called pentadentate, and six are called hexadentate. The chelating effect is also observed in polydentate ligands. The resulting chelate structure enhances the stability of the coordination complex by reducing the chances of the ligand dissociating from the metal ion. Examples of polydentate ligands include diethylenetriamine (dien) and ethylenediaminetetraacetic ion (ETDA4-). Diethylenetriamine has three nitrogen atoms, while EDTA4- is a complex ion with six coordination sites – two nitrogen and four oxygen atoms. The chelating effect and the ability of polydentate ligands to form stable coordination complexes make them valuable in various applications. For instance, in biochemistry, EDTA is commonly used for metal ion chelation in medical treatments and diagnostic tests. In industry, polydentate ligands are utilized for catalysis and in pharmaceuticals and other chemicals. Ligands having more than one potential donor atom are known as ambidentate ligands. For example, thiocyanate ion (NCS–) can bind to the central metal ion with either nitrogen or sulfur atoms. Also, the nitrate ion (NO2-) can bind at either the nitrogen atom or one of the oxygen atoms. Other ligands include bridging ligands, which bind to two or more metal ions, and organic ligands, which are the organic derivatives of the abovementioned inorganic ligands. References Study.comThoughtco.comChem.libretexts.orgChemed.chem.purdue.edu A ligand is an ion or molecule that can donate a pair of electrons to a central metal atom or ion to form a coordinate covalent bond, resulting in the formation of a coordination compound. Ligands play a crucial role in determining the structure, reactivity, and properties of coordination compounds.1.0What is Ligand?A ligand is an ion or molecule that donates a pair of electrons to a central metal atom or ion, forming a coordination complex through a coordinate covalent bond.The term "ligand" originates from the Latin word "ligare," meaning "to tie" or "to bind," reflecting its role in binding to metal centres. Ligands can be classified as anions (negatively charged), cations (positively charged), or neutral molecules.As Lewis bases, ligands donate electron pairs to the central metal atom, which acts as a Lewis acid by accepting these electron pairs.The nature of bonding between a metal and a ligand can range from covalent (electron sharing) to ionic (electrostatic attraction), influencing the stability and properties of the coordination complex.2.0Characteristics of LigandsLigands act as Lewis bases, which means they are electron pair donors. They provide a lone pair of electrons to the central metal atom or ion (which acts as a Lewis acid, or electron pair acceptor).The bond formed between a ligand and the central metal ion is a coordinate covalent bond (also known as a dative bond), where both electrons in the bond originate from the ligand.Ligands can be neutral molecules (such as H_2O, NH_3) or anions (such as Cl^-, CN^-). Some ligands can even be cations (though less common, like NO^+ in certain contexts).3.0Classification of LigandsLigands are classified based on several criteria, including their charge, denticity (the number of donor atoms), and the type of donor atom(s) involved.1. Based on Charge:Neutral Ligands: These ligands have no charge. Common examples include:Ammine (NH_3)Aqua (H_2O)Carbonyl (CO)Phosphine (PR_3)Anionic Ligands: These ligands carry a negative charge. Common examples include:Chloro (Cl^-)Cyano (CN^-)Hydroxo (OH^-)Nitrate (NO_3^-)Cationic Ligands: These ligands are positively charged, though less common. An example is:Nitrosyl (NO^+)2. Based on Denticity (Number of Donor Atoms):Monodentate (Unidentate) Ligands: These ligands have a single donor atom that coordinates to the central metal ion. Each monodentate ligand forms one coordinate bond with the metal.Examples: Water (H_2O), Ammonia (NH_3), Chloride (Cl^-), Cyanide (CN^-).Bidentate Ligands: These ligands have two donor atoms that can coordinate to a single metal ion simultaneously, forming two bonds. This often results in the formation of a five- or six-membered ring with the metal.Examples: Img 3Polydentate (Multidentate) Ligands: These ligands have multiple donor atoms that can coordinate to the same metal ion, forming more than two bonds. Polydentate ligands are often referred to as chelating agents because they can "chelate" the metal ion, forming very stable ring-like structures.Examples: Img 43. Based on the Type of Donor Atom:Anionic Ligands: Ligands that donate electrons from negatively charged donor atoms.Examples: Chloride (Cl^-), Hydroxide (OH^-), Sulfate (SO_4^{2-}).Neutral Ligands: Ligands that donate electrons from neutral donor atoms.Examples: Ammonia (NH_3), Water (H_2O), Phosphine (PH_3).Ambidentate Ligands: Ligands that have two different donor atoms but can only coordinate through one donor atom at a time.Examples: Nitro (NO_2^-), which can coordinate through nitrogen (as in nitro, $-NO_2$) or through oxygen (as in nitrito, -ONO); Thiocyanate (SCN^-), which can coordinate through sulfur (as thiocyanato, -SCN) or nitrogen (as iso-thiocyanato, -NCS). Img 54.0Types of Ligand CoordinationThese ligands are bonded to only one metal atom. Most common ligands in coordination chemistry act as terminal ligands.Examples: Water (H_2O), Ammonia (NH_3), Chloride (Cl^-).These ligands can coordinate two or more metal atoms simultaneously, acting as a "bridge" between them.Examples: Hydroxo (OH^-), Carbonyl (CO), Nitrosyl (NO).5.0Importance of Ligands in Coordination CompoundsLigands play a crucial role in the stability and formation of coordination compounds. Polydentate ligands, due to their ability to form multiple bonds with the central metal ion, typically result in more stable complexes (chelate effect).The type and number of ligands directly influence the geometry of the coordination complex. For example, six monodentate ligands often lead to an octahedral geometry, while four ligands can result in either tetrahedral or square planar geometry depending on the metal and ligands involved.Ligands affect the electronic properties of the central metal ion, including its oxidation state, electron configuration, and magnetic properties. For example, strong field ligands like CN^- can cause a large splitting of the d-orbitals in transition metals, resulting in low-spin complexes.Reactivity and Catalysis:The nature of the ligand can greatly influence the reactivity of the coordination compound. In catalysis, ligands are crucial in modulating the activity, selectivity, and stability of metal catalysts.Spectroscopic Properties:Ligands affect the absorption spectra of coordination compounds. Changes in the ligand field, induced by different ligands, can lead to varying colors in transition metal complexes due to d-d electronic transitions.Ligands are vital in biological systems. For instance, in hemoglobin, the heme group (a multidentate ligand) binds to the iron ion, facilitating oxygen transport in the blood. Similarly, ligands in chlorophyll play a critical role in photosynthesis.6.0Examples of Ligands and Their Complex Type The structure of the ion looks like this: In this case, the "ear pieces" are the nitrogen atoms of the NH2 groups - and the "bit that goes over your head " is the -CH2CH2- group. If you were going to draw this in an exam, you would obviously want to draw it properly - but for learning purposes, drawing all the atoms makes the diagram look unduly complicated! Notice that the arrangement of the bonds around the central metal ion is exactly the same as it was with the ions with 6 water molecules attached. The only difference is that this time each ligand uses up two of the positions - at right angles to each other. Because the nickel is forming 6 co-ordinate bonds, the co-ordination number of this ion is 6, despite the fact that it is only joined to 3 ligands. Co-ordination number counts the number of bonds, not the number of ligands. Cr (C2O4)3 3- This is the complex ion formed by attaching 3 ethanedioate (oxalate) ions to a chromium(III) ion. The shape is exactly the same as the previous nickel complex. The only real difference is the number of charges. The original chromium ion carried 3+ charges, and each ethanedioate ion carried 2-. (3+) + (3 x 2-) = 3-. The structure of the ion looks like this: Again, if you drew this in an exam, you would want to show all the atoms properly. If you need to be able to do this, practice drawing it so that it looks clear and tidy! Refer back to the diagram of the ethanedioate ion further up the page to help you. A quadridentate ligand A quadridentate ligand has four lone pairs, all of which can bond to the central metal ion. An example of this occurs in haemoglobin (American: hemoglobin). The functional part of this is an iron(II) ion surrounded by a complicated molecule called haem (heme). This is a sort of hollow ring of carbon and hydrogen atoms, at the centre of which are 4 nitrogen atoms with lone pairs on them. Haem is one of a group of similar compounds called porphyrins. They all have the same sort of ring system, but with different groups attached to the outside of the ring. You aren't going to need to know the exact structure of the haem at this level. We could simplify the haem with the trapped iron ion as: Each of the lone pairs on the nitrogen can form a co-ordinate bond with the iron(II) ion - holding it at the centre of the complicated ring of atoms. The iron forms 4 co-ordinate bonds with the haem, but still has space to form two more - one above and one below the plane of the ring. The protein globin attaches to one of these positions using a lone pair on one of the nitrogens in one of its amino acids. The interesting bit is the other position. Overall, the complex ion has a co-ordination number of 6 because the central metal ion is forming 6 co-ordinate bonds. The water molecule which is bonded to the bottom position in the diagram is easily replaced by an oxygen molecule (again via a lone pair on one of the oxygens in O2) - and this is how oxygen gets carried around the blood by the haemoglobin. When the oxygen gets to where it is needed, it breaks away from the haemoglobin which returns to the lungs to get some more. You probably know that carbon monoxide is poisonous because it reacts with haemoglobin. It bonds to the same site that would otherwise be used by the oxygen - but it forms a very stable complex. The carbon monoxide doesn't break away again, and that makes that haemoglobin molecule useless for any further oxygen transfer. A hexadentate ligand A hexadentate ligand has 6 lone pairs of electrons - all of which can form co-ordinate bonds with the same metal ion. The best example is EDTA. EDTA is used as a negative ion - EDTA4-. The diagram shows the structure of the ion with the important atoms and lone pairs picked out.