

Explain the following terms or concepts.	Marks <b>1</b>
Lewis base  <b>A species that can form a bond by donating a lone pair of electrons, <i>e.g.</i> H<sub>2</sub>O.</b>	

- Give the oxidation number of the indicated atom in the following compounds.

2

Compound	Atom	Oxidation number
$\text{HNO}_2$	N	
$\text{NI}_3$	N	
$[\text{Co}(\text{NH}_3)_5\text{Cl}]\text{SO}_4$	Co	
$\text{K}_3[\text{CrCl}_6]$	Cr	

- Write down the formulas for the following compounds.

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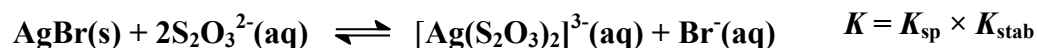
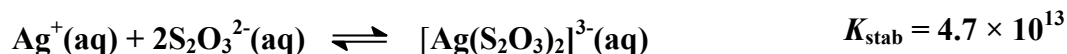
Compound	Formula
hexaaquacobalt(II) carbonate	$[\text{Co}(\text{OH}_2)_6]\text{CO}_3$
tetraamminecopper(II) sulfate	$[\text{Cu}(\text{NH}_3)_4]\text{SO}_4$
ammonium hexafluoridoferrate(III)	$(\text{NH}_4)_2[\text{FeF}_6]$
potassium hexacyanomanganate(II)	$\text{K}_4[\text{Mn}(\text{CN})_6]$

**THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.**

- The  $K_{sp}$  of AgBr is  $5.0 \times 10^{-13}$ . The  $K_{stab}$  of  $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$  is  $4.7 \times 10^{13}$ . Calculate the value of the equilibrium constant for the dissolution of AgBr in  $\text{Na}_2\text{S}_2\text{O}_3$  solution.

Marks  
7

The reactions corresponds to  $K_{sp}$  and  $K_{stab}$  can be added together to give the reaction for the dissolution of AgBr in  $\text{Na}_2\text{S}_2\text{O}_3$  solution:



The equilibrium constant for the overall reaction is the product of the equilibrium constants for the individual reactions:

$$K = K_{sp} \times K_{stab} = (5.0 \times 10^{-13}) \times (4.7 \times 10^{13}) = 24$$

Answer: 24

Calculate the solubility of AgBr in 2.0 M  $\text{Na}_2\text{S}_2\text{O}_3$ .

The solubility can be calculated using a reaction table, assuming  $x$  mol dissolves:

	AgBr(s)	$2\text{S}_2\text{O}_3^{2-}(\text{aq})$	$\rightleftharpoons$	$[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}(\text{aq})$	$\text{Br}^-(\text{aq})$
initial	excess	2.0		0	0
change	-x	-2x		+x	+x
final	excess	$2.0 - 2x$		x	x

$$K = \frac{[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}(\text{aq})[\text{Br}^-(\text{aq})]}{[\text{S}_2\text{O}_3^{2-}(\text{aq})]^2}$$

$$= \frac{(x)(x)}{(2.0 - 2x)^2} = \frac{x^2}{(2.0 - 2x)^2} = 24$$

Taking square roots of both sides gives:

$$\frac{x}{(2.0 - 2x)} = (24)^{1/2} \quad x = 0.91 \text{ mol L}^{-1}$$

Answer: 0.91 mol L<sup>-1</sup>

The  $K_{stab}$  for  $[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$  is much greater than the  $K_{stab}$  for  $[\text{Ag}(\text{NH}_3)_2]^+$ . Explain why this is so.

$\text{S}_2\text{O}_3^{2-}$  is a stronger ligand than  $\text{NH}_3$ , presumably because of its negative charge.

**Marks**  
**9**

- Complete the following table.

Coordination compound	Oxidation number of transition metal	Number of $d$ electrons around transition metal	Arrangement of $d$ electrons					
$\text{K}_2[\text{PtCl}_4]$	+2	$d^8$	<table border="1" style="display: inline-table; vertical-align: middle;"><tr><td>↑↓</td><td>↑↓</td><td>↑↓</td><td>↑↓</td><td> </td></tr></table>	↑↓	↑↓	↑↓	↑↓	
↑↓	↑↓	↑↓	↑↓					
$\text{Na}[\text{MnO}_4]$	+7	$d^0$	<table border="1" style="display: inline-table; vertical-align: middle;"><tr><td> </td><td> </td><td> </td><td> </td><td> </td></tr></table>					
$(\text{NH}_4)_2[\text{CoCl}_4]$	+2	$d^7$	<table border="1" style="display: inline-table; vertical-align: middle;"><tr><td>↑↓</td><td>↑↓</td><td>↑</td><td>↑</td><td>↑</td></tr></table>	↑↓	↑↓	↑	↑	↑
↑↓	↑↓	↑	↑	↑				
$[\text{Cr}(\text{NH}_3)_5(\text{OH}_2)]\text{Cl}_3$	+3	$d^3$	<table border="1" style="display: inline-table; vertical-align: middle;"><tr><td>↑</td><td>↑</td><td>↑</td><td> </td><td> </td></tr></table>	↑	↑	↑		
↑	↑	↑						

Identify one paramagnetic and one diamagnetic species from the above table.

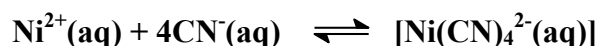
Paramagnetic: $(\text{NH}_4)_2[\text{CoCl}_4]$ and $[\text{Cr}(\text{NH}_3)_5(\text{OH}_2)]\text{Cl}_3$	Diamagnetic: $\text{K}_2[\text{PtCl}_4]$ and $\text{Na}[\text{MnO}_4]$
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- Give the equilibrium concentration of  $\text{Ni}^{2+}(\text{aq})$  ions in a solution formed by dissolving 0.15 mol of  $\text{NiCl}_2$  in 0.500 L of 2.00 M KCN solution. The  $K_{\text{stab}}$  of  $[\text{Ni}(\text{CN})_4]^{2-} = 1.7 \times 10^{30}$ .

When the  $\text{NiCl}_2$  is added to 0.500 L, the concentration of  $\text{Ni}^{2+}(\text{aq})$ :

$$\begin{aligned} \text{concentration} &= [\text{Ni}^{2+}(\text{aq})] = \text{number of moles} / \text{volume} \\ &= 0.15 \text{ mol} / 0.500 \text{ L} = 0.30 \text{ M} \end{aligned}$$

As  $K_{\text{stab}}$  is so large, essentially *all* of this will be complexed by the excess  $\text{CN}^-(\text{aq})$  ions:



As essentially all of the  $\text{Ni}^{2+}(\text{aq})$  becomes  $[\text{Ni}(\text{CN})_4]^{2-}(\text{aq})$ :

$$\begin{aligned} [\text{Ni}(\text{CN})_4]^{2-}(\text{aq}) &= 0.30 \text{ M} \\ [\text{CN}^-(\text{aq})] &= (2.00 - 4 \times 0.30) \text{ M} = 0.80 \text{ M} \end{aligned}$$

$K_{\text{stab}}$  is the equilibrium constant for the reaction so:

$$K_{\text{stab}} = \frac{[\text{Ni}(\text{CN})_4]^{2-}(\text{aq})}{[\text{Ni}^{2+}(\text{aq})][\text{CN}^-(\text{aq})]^4}$$

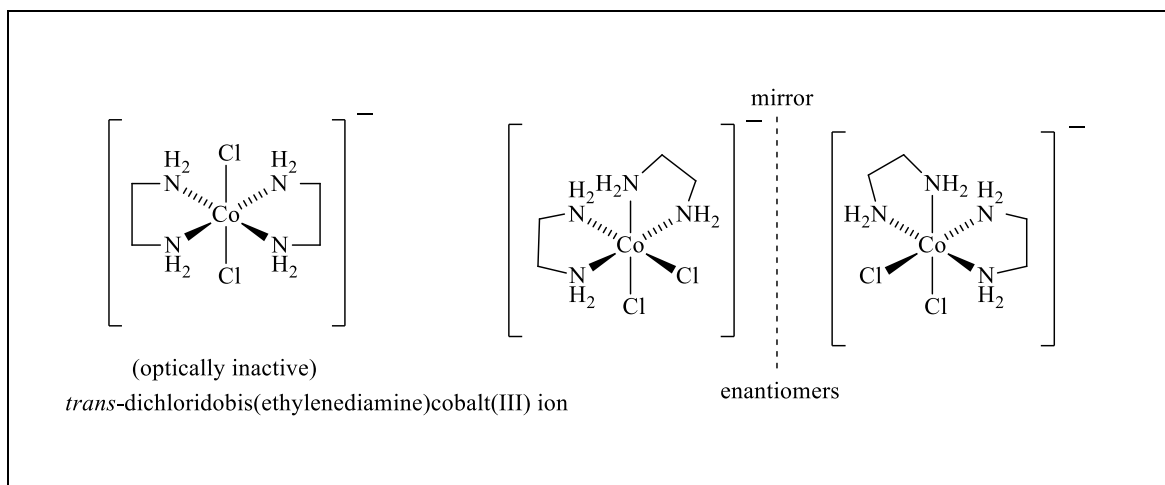
If the *tiny* amount of uncomplexed  $\text{Ni}^{2+}(\text{aq})$  has a concentration of  $x$  M then:

$$K_{\text{stab}} = \frac{(0.30)}{(x)(0.80)^4} = 1.7 \times 10^{30} \quad \text{so } x = 4.3 \times 10^{-31} \text{ M}$$

Answer:  $4.3 \times 10^{-31} \text{ M}$

- Draw all stereoisomers of the complex ion of  $[\text{CoCl}_2(\text{en})_2]\text{Cl}$ . Label the non-optically active isomer with its systematic name.  
(en = ethylenediamine = 1,2-ethanediamine =  $\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2$ )

Marks  
4

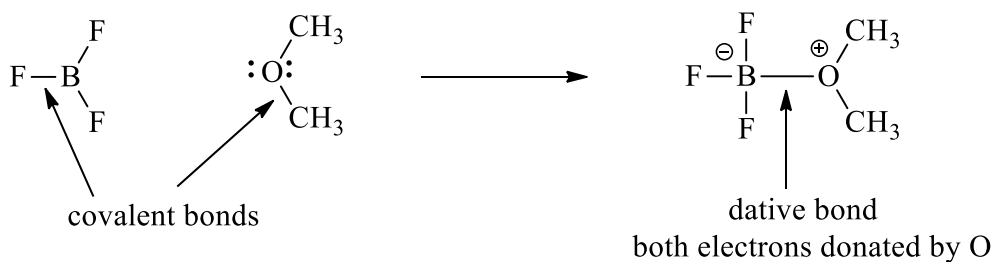


- What is a dative bond and does it differ from a covalent bond? Use examples from coordination chemistry and elsewhere to illustrate your answer.

2

**Dative bonds are similar to covalent bonds, but both electrons in the bond are donated by one atom.**

**They are typically found in coordination complexes where the lone pair on the ligand forms a bond with the metal ion. See the above Co(III) complexes for examples and the  $\text{BF}_3$ -ether adduct below for a different type of example. In general, dative bonds are weaker, longer and less directional than covalent bonds.**



- Explain why iron storage proteins are necessary for the transport of iron both intracellularly and extracellularly within the bloodstream at a pH of 7.4.

2

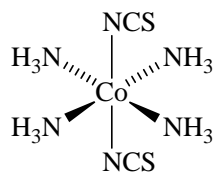
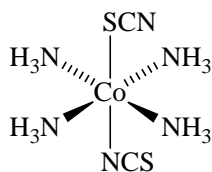
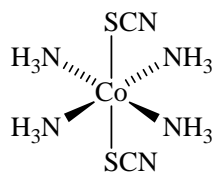
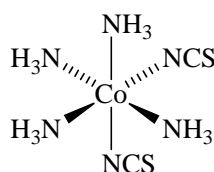
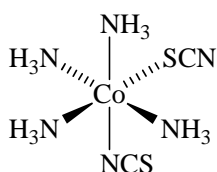
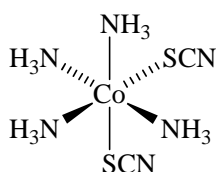
**The  $K_{sp}$  of  $\text{Fe}(\text{OH})_3$  is so low, that even at pH 7.4 there are sufficient  $\text{OH}^-$  ions present to precipitate the  $\text{Fe}^{3+}$  ions as  $\text{Fe}(\text{OH})_3$ .**

**To avoid precipitation and to allow a higher concentration of  $\text{Fe}^{3+}$  to be circulated,  $\text{Fe}^{3+}$  is complexed by *transferrin* in the bloodstream and iron is stored within *ferritin* within the cell.**



**Marks**  
**7**

- Write out the full name in standard notation of  $[\text{Co}(\text{NH}_3)_4(\text{SCN})_2]\text{Cl}$  and draw all the possible isomers of the complex ion.

**tetraamminedithiocyanatocobalt(III) chloride**

*trans-isomers*

*cis-isomers*

2 thiocyanato ligands

 1 thiocyanato ligand &  
 1 isothiocyanato ligand

2 isothiocyanato ligands

Describe and contrast the nature of the chemical bonds:

- between N and H in  $\text{NH}_3$ ;
- between Co and  $\text{NH}_3$ ; and
- between  $[\text{Co}(\text{NH}_3)_4(\text{SCN})_2]^+$  and  $\text{Cl}^-$  in this compound.

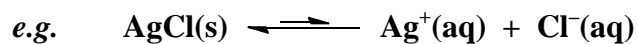
**N–H bonds are covalent in  $\text{NH}_3$ . These bonds are relatively short, strong and highly directional. They involve the sharing of electrons from both atoms involved in the bond.**

**$\text{Co}:\text{NH}_3$  coordination bonds are due to the donation of the lone pair of electrons on N to the  $\text{Co}^{3+}$ . These bonds are highly polar and are generally weaker, longer and less directional than covalent bonds.**

**$[\text{Co}(\text{NH}_3)_4(\text{SCN})_2]^+$  and  $\text{Cl}^-$  are ionically bonded in the solid state due to Coulombic attraction between the oppositely charged ions. These bonds are strong but not directional (i.e. they occur between every pair of ions in the solid with a strength that decreases with their separation).**

- During lectures a demonstration was performed called the “One pot experiment”. In this experiment, silver ions reacted with an alternating series of anions and ligands to form insoluble precipitates and soluble complexes. Explain how an insoluble precipitate can possibly be “dissolved” by the addition of ligands to the solution.

**The insoluble precipitate is actually in equilibrium with its ions, but the equilibrium lies heavily to the left.**



**When a ligand is added, the  $\text{Ag}^{\text{+}}(\text{aq})$  ions form a complex and are removed from the above equilibrium.**



**Due to Le Chatelier’s principle, more  $\text{AgCl(s)}$  must dissolve to try and re-establish the equilibrium and eventually all the “insoluble“ precipitate will dissolve.**

- The general formula for a nickel(II) chloride compound complexed with ammonia is  $[\text{Ni}(\text{NH}_3)_x]\text{Cl}_2$ . A 0.59 g sample of the salt was dissolved in water and the ammonia from it was titrated with 153 mL of 0.100 M HCl. What is the value of the coefficient  $x$ ?

Marks  
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The molar mass of  $[\text{Ni}(\text{NH}_3)_x]\text{Cl}_2$  is:

$$\begin{aligned} & (58.69 \text{ (Ni)} + x (14.01 \text{ (N)} + 3 \times 1.008 \text{ (H)}) + 2 \times 35.45 \text{ (Cl)}) \text{ g mol}^{-1} \\ & = (129.59 + 17.034x) \text{ g mol}^{-1} \end{aligned}$$

A 0.59 g sample therefore corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.59}{(129.59 + 17.034x)} \text{ mol} \quad (1)$$

The number of moles in 153 mL of 0.100 M HCl is:

$$\begin{aligned} \text{number of moles} &= \text{concentration} \times \text{volume} \\ &= 0.100 \text{ mol L}^{-1} \times 0.153 \text{ L} = 0.0153 \text{ mol} \end{aligned}$$

Ammonia reacts with HCl according to the reaction  $\text{NH}_3 + \text{HCl} \rightleftharpoons \text{NH}_4\text{Cl}$  and so this is equal to the number of moles of  $\text{NH}_3$  present. Each mol of  $[\text{Ni}(\text{NH}_3)_x]\text{Cl}_2$  contains  $x$  mol of  $\text{NH}_3$  so the number of moles of  $[\text{Ni}(\text{NH}_3)_x]\text{Cl}_2$  is:

$$\text{number of moles} = 0.0153 / x \text{ mol} \quad (2)$$

The value of  $x$  is calculated by equating (1) and (2). This is easiest to achieve by trial and error.

$x$	(1) / mol	(2) / mol
1	0.0040	0.015
2	0.0036	0.0077
3	0.0033	0.0051
4	0.0030	0.0038
5	0.0027	0.0031
6	0.0026	0.0026
7	0.0024	0.0022

The best agreement is for  $x = 6$  - a common coordination number for Ni(II).

Answer: 6

- Can methane act as a ligand? Explain your answer.

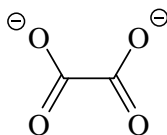
Marks

2

**No. Ligands are, by definition, electron pair donors and so require at least one lone pair of electrons that can be donated to a metal ion to form a covalent bond. Methane, CH<sub>4</sub>, has no lone pairs.**

- Fe(II) generally forms octahedral complexes. How many different complex ions can be formed when Fe(NO<sub>3</sub>)<sub>2</sub> is dissolved in an aqueous solution of sodium oxalate? The structure of the oxalate ligand is shown below.

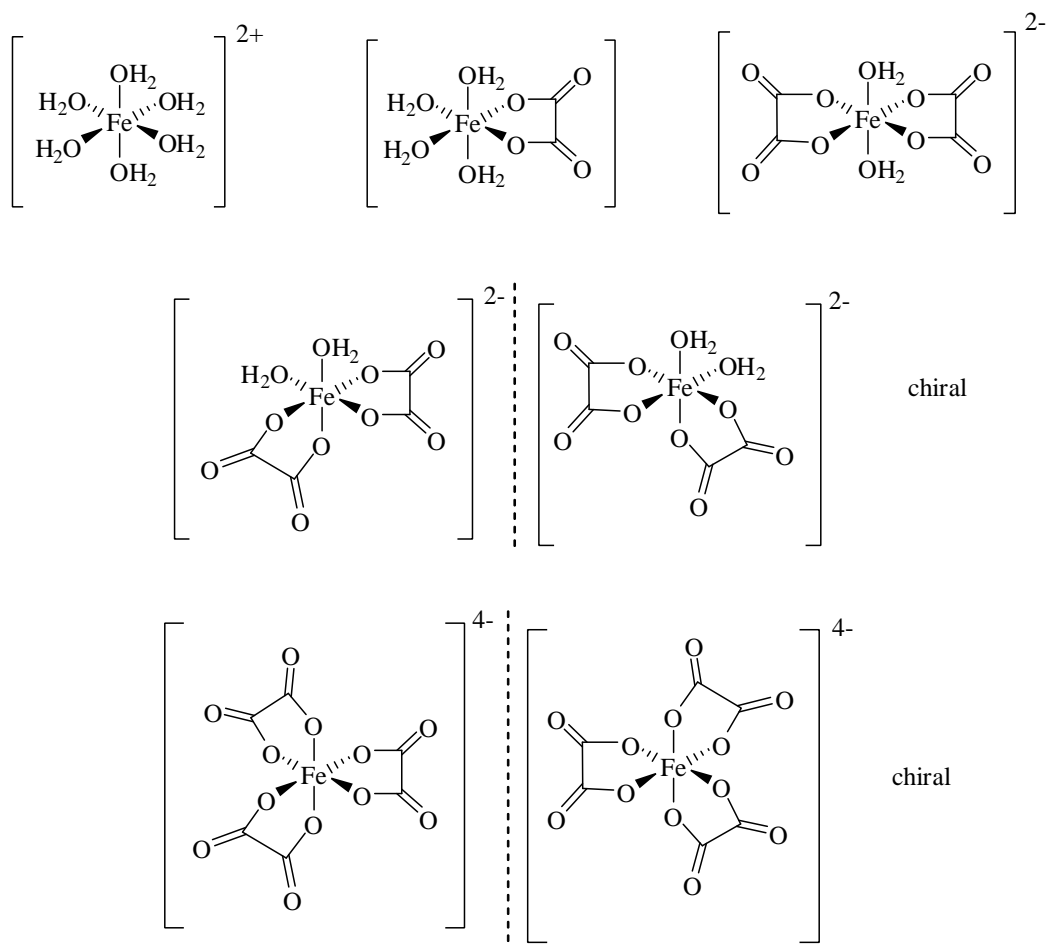
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**7 complex ions are possible – see below.**

Answer: 7

Draw diagrams of any 3 of these complexes, including at least one that is chiral.



- Complete the following table.

4

Formula	Systematic name
$[\text{CrCl}(\text{NH}_3)_5]\text{Cl}_2$	<b>pentaamminechloridochromium(III) chloride</b>
<b><math>[\text{PtBr}_2(\text{CO})_4](\text{NO}_2)_2</math></b>	dibromidotetracarbonylplatinum(IV) nitrite
$\text{K}_3[\text{CrF}_6]$	<b>potassium hexafluoridochromate(III)</b>
$(\text{NH}_4)_3[\text{CuF}_5(\text{OH}_2)]$	<b>ammonium aquapentafluoridocuprate(II)</b>

- State two chemical factors that contribute to the bioavailability of a heavy metal in the human body.

2

**The solubility of the metal ion in the body and the transport properties of the ion.**

**Marks**  
**5**

- Cisplatin,  $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$ , is a particularly effective chemotherapy agent against certain types of cancer. Calculate the concentration of  $\text{Pt}^{2+}(\text{aq})$  ions in solution when 0.075 mol of cisplatin is dissolved in 1.00 L of a 1.00 M solution of  $\text{NH}_3$ .  
 $K_{\text{stab}}$  of  $[\text{Pt}(\text{NH}_3)_2\text{Cl}_2] = 3.4 \times 10^{12}$ .

The initial concentration of cisplatin when 0.075 mol is dissolved in 1.00 L is 0.075 M.

As  $K_{\text{stab}}$  refers to the formation of the complex, the reaction table is:

	$\text{Pt}^{2+}(\text{aq})$	$2\text{Cl}^-(\text{aq})$	$2\text{NH}_3(\text{aq})$	$\rightleftharpoons$	$[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]$
<b>Initial</b>	<b>0</b>	<b>0</b>	<b>1.00</b>		<b>0.075</b>
<b>Change</b>	<b>+x</b>	<b>+2x</b>	<b>+2x</b>		<b>-x</b>
<b>Equilibrium</b>	<b>x</b>	<b>2x</b>	<b>1.00 + 2x</b>		<b>0.075 - x</b>

Hence:

$$K_{\text{stab}} = \frac{[\text{Pt}(\text{NH}_3)_2\text{Cl}_2]}{[\text{Pt}^{2+}][\text{Cl}^-]^2[\text{NH}_3]^2} = \frac{(0.075-x)}{(x)(2x)^2(1.00+2x)^2} = 3.4 \times 10^{12}$$

As  $K_{\text{stab}}$  is so large,  $x$  will be very, very small and so  $(1.00 + 2x) \sim 1.00$  and  $(0.075 - x) \sim 0.075$ . With this:

$$K_{\text{stab}} \sim \frac{(0.075)}{(x)(2x)^2(1.00)^2} = \frac{(0.075)}{(4x)^3} = 3.4 \times 10^{12}$$

$$x = [\text{Pt}^{2+}(\text{aq})] = 1.8 \times 10^{-5} \text{ M}$$

Answer:  $1.8 \times 10^{-5} \text{ M}$

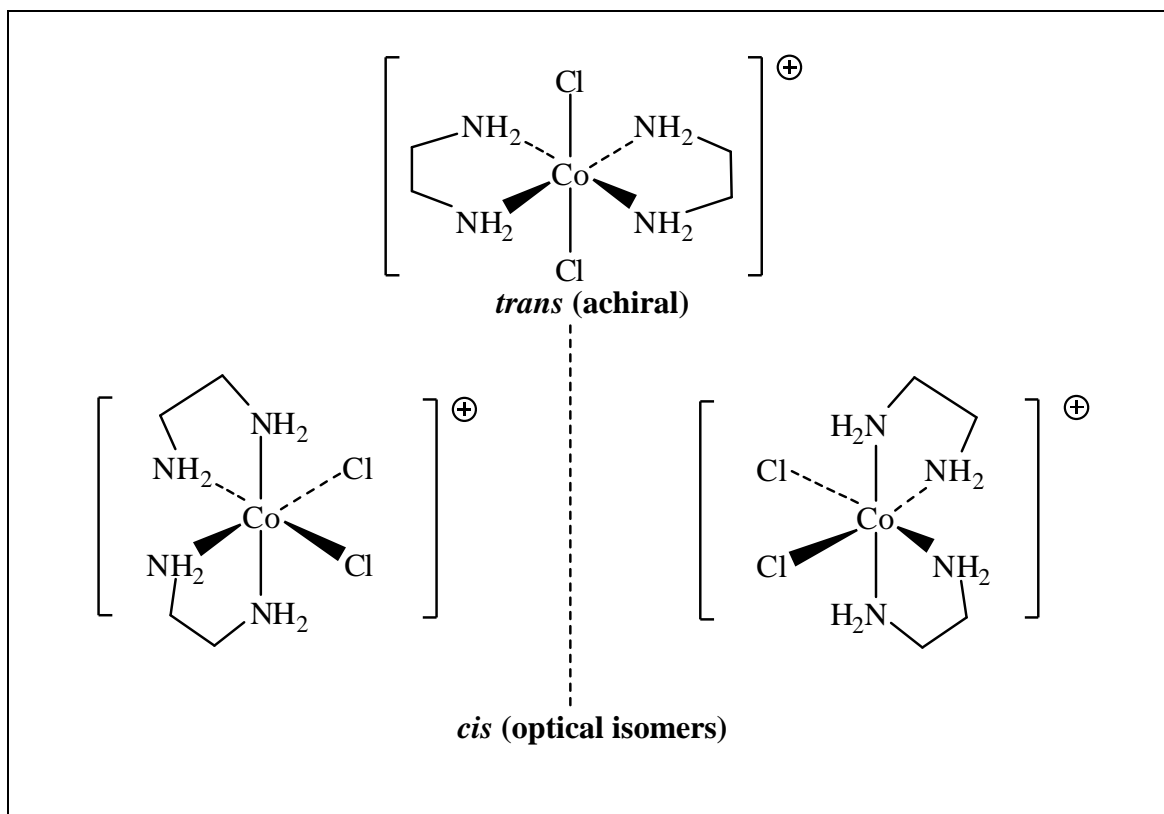
What changes would occur to the values of  $K_{\text{stab}}$  for cisplatin and the concentration of  $\text{Pt}^{2+}(\text{aq})$  ions if solid KCl were dissolved in the above solution?

$K_{\text{stab}}$	increase	<u>no change</u>	decrease
$[\text{Pt}^{2+}(\text{aq})]$	increase	no change	<u>decrease</u>

- $K_{\text{stab}}$  is the stability constant – it is a constant at any given temperature.
- From Le Chatelier's principle, if  $[\text{Cl}^-(\text{aq})]$  is increased then the equilibrium will shift to the right and so  $[\text{Pt}^{2+}(\text{aq})]$  will decrease.

- Draw all possible stereoisomers of the complex ion  $[\text{CoCl}_2(\text{en})_2]^+$ . Label each as *cis* or *trans*. en = ethylenediamine =  $\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2$

3



- Name the following complexes.

2



**sodium hexafluoroaluminate(III)**



**pentaamminebromidocobalt(III) sulfate**

- Zinc sulfate (0.50 g) is dissolved in 1.0 L of a 1.0 M solution of KCN. Write the chemical equation for the formation of the complex ion  $[\text{Zn}(\text{CN})_4]^{2-}$ .

**Marks**  
**3**

Calculate the concentration of  $\text{Zn}^{2+}(\text{aq})$  in solution at equilibrium. Ignore any change in volume upon addition of the salt.  $K_{\text{stab}}$  of  $[\text{Zn}(\text{CN})_4]^{2-} = 4.2 \times 10^{19} \text{ M}^{-4}$ .

The formula mass of  $\text{ZnSO}_4$  is  $65.39 (\text{Zn}) + 32.07 (\text{S}) + 4 \times 16.00 (\text{O}) = 161.46$ .

0.50 g therefore corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{formula mass}} = \frac{0.50 \text{ g}}{161.46 \text{ g mol}^{-1}} = 0.0031 \text{ mol}$$

As  $K_{\text{stab}} = 4.2 \times 10^{19}$  and is *very* large, the reaction essentially goes to completion. The reaction requires a 4:1 ratio  $\text{CN}^- : \text{Zn}^{2+}(\text{aq})$  ions and as 0.0031 mol of  $\text{Zn}^{2+}$  and 1.0 mol of  $\text{CN}^-$  are present,  $\text{CN}^-$  is in excess.

Let the tiny amount of uncomplexed  $\text{Zn}^{2+}(\text{aq})$  and its concentration in 1.0 L be:

$$\text{amount of } \text{Zn}^{2+}(\text{aq}) = x \text{ mol and } [\text{Zn}^{2+}(\text{aq})] = \frac{\text{number of moles}}{\text{volume}} = \frac{x}{1.0} \text{ M}$$

The amount of  $[\text{Zn}(\text{CN})_4]^{2-}(\text{aq})$  formed is therefore:

$$\text{amount of } [\text{Zn}(\text{CN})_4]^{2-}(\text{aq}) = (0.0031 - x) \sim 0.0031 \text{ mol as } x \text{ is so small.}$$

Hence,

$$[[\text{Zn}(\text{CN})_4]^{2-}(\text{aq})] \sim \frac{0.0031}{1.0} = 0.0031 \text{ M}$$

Formation of 0.0031 mol of  $[\text{Zn}(\text{CN})_4]^{2-}(\text{aq})$  requires  $(4 \times 0.0031) = 0.012$  mol of cyanide, leaving:

$$\text{amount of } \text{CN}^- = (1.0 - 0.012) = 0.99 \text{ mol and } [\text{CN}^-(\text{aq})] = \frac{0.99}{1.0} \text{ M}$$

Hence,

$$K_{\text{stab}} = \frac{[[\text{Zn}(\text{CN})_4]^{2-}(\text{aq})]}{[\text{Zn}^{2+}(\text{aq})][\text{CN}^-(\text{aq})]^4} = \frac{(0.0031)}{(x)(0.99)^4} = 4.2 \times 10^{19}$$

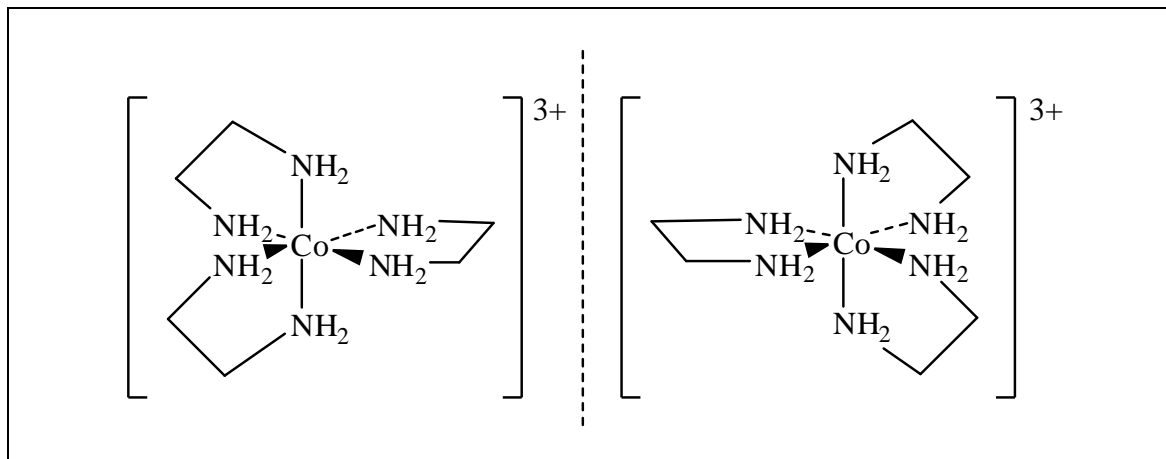
$$x = 7.7 \times 10^{-23} \text{ mol and so } [\text{Zn}^{2+}(\text{aq})] = 7.7 \times 10^{-23} \text{ M}$$

Answer:  $7.8 \times 10^{-23}$



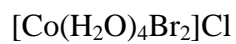
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- Draw all stereoisomers of the complex ion of  $[\text{Co}(\text{en})_3]\text{Br}_3$ .  
(en = ethylenediamine =  $\text{NH}_2\text{CH}_2\text{CH}_2\text{NH}_2$ )

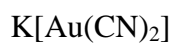


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- Name the following complexes.



**tetraaquadibromidocobalt(III) chloride**



**potassium dicyanoaurate(I)**

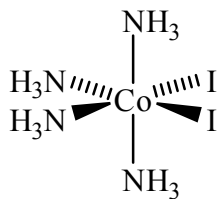
- Write the chemical equation for the formation of the complex ion  $[\text{Cd}(\text{NH}_3)_4]^{2+}$ .

**Marks**  
**2**

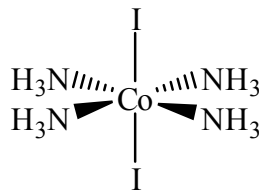
Write the associated stability constant expression ( $K_{\text{stab}}$ ).

$$K_{\text{stab}} = \frac{[[\text{Cd}(\text{NH}_3)_4]^{2+}]}{[\text{Cd}^{2+}][\text{NH}_3]^4}$$

- Draw all of the geometric isomers for the complex ion  $[\text{CoI}_2(\text{NH}_3)_4]^+$ . Label each isomer with its systematic name.



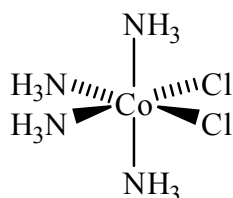
*cis*-tetraamminediiodidocobalt(II)



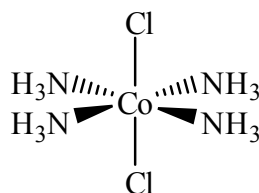
*trans*-tetraamminediiodidocobalt(II)

- Coordination complexes can display a number of types of isomerism. Draw a simple diagram showing a pair of geometric isomers. Label your diagram with the systematic name of each isomer.

The compound  $[\text{CoCl}_2(\text{NH}_3)_4]$  can exist as in two isomeric forms which differ in the arrangement of the ligands in space.

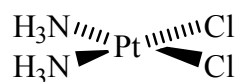


*cis*-tetraamminedichloridocobalt(II)



*trans*-  
tetraamminedichloridocobalt(II)

$\text{Pt}(\text{NH}_3)_2\text{Cl}_2$  can also exist as *cis* and *trans* isomers:



*cis*-diamminedichloridoplatinum(II)



*trans*-  
diamminedichloridoplatinum(II)

- Complete the following table.

<b>Formula</b>	<b>Systematic name</b>
$[\text{CrCl}(\text{OH}_2)_5]\text{Cl}_2$	<b>pentaaquachloridochromium(III) chloride</b>
$\text{NH}_4\text{Fe}(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O}$	<b>ammonium iron(III) sulfate-6-water</b>
$\text{Na}_3[\text{CoCl}_6]$	sodium hexachloridocobaltate(III)
$\text{Al}(\text{NO}_3)_3 \cdot 9\text{H}_2\text{O}$	aluminium nitrate-9-water
$\text{KNO}_2$	potassium nitrite
$\text{KHSO}_4$	<b>potassium hydrogensulfate</b>

**Marks**  
**3**

- Classify each of the following as either “soluble” or “insoluble” in water at 298 K.

Compound	Solubility	Compound	Solubility
$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$	<b>soluble</b>	$\text{HgCl}_2$	<b>soluble</b>
$\text{Li}_2\text{CO}_3$	<b>insoluble</b>	$\text{Zn}(\text{CH}_3\text{CO}_2)_2$	<b>soluble</b>
$\text{MnO}_2$	<b>insoluble</b>	$\text{SrSO}_4$	<b>insoluble</b>

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- Complete the following table.

FORMULA	SYSTEMATIC NAME
$\text{NH}_4\text{Fe}(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O}$	<b>ammonium iron(III) sulfate-6-water</b>
$[\text{Cr}(\text{OH}_2)_5\text{Cl}]\text{SO}_4$	<b>pentaaquachloridochromium(III) sulfate</b>
$\text{NaH}_2\text{PO}_4$	<b>sodium dihydrogenphosphate</b>
$\text{HClO}_4$	<b>perchloric acid</b>
$\text{As}_2\text{O}_3$	<b>diarsenic trioxide or arsenic(III) oxide</b>
$[\text{PdCl}_2(\text{NH}_3)_2]$	<b>diamminedichloridopalladium(II)</b>
$\text{SO}_2$	<b>sulfur dioxide</b>
<b>KNCS</b>	potassium thiocyanate
<b><math>\text{NaNO}_2</math></b>	sodium nitrite
<b><math>[\text{CoBr}_2(\text{OH}_2)_4]\text{Cl}</math></b>	tetraaquadibromidocobalt(III) chloride
<b><math>\text{Na}_3[\text{Fe}(\text{CN})_6]</math></b>	sodium hexacyanidoferrate(III)
<b><math>\text{PbO}_2</math></b>	lead(IV) oxide
<b><math>\text{O}_2^{2-}</math></b>	peroxide ion
<b><math>\text{Ni}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O}</math></b>	nickel(II) nitrate-6-water