28.5 STABILITY CONSTANT

Learning outcomes

Candidates should be able to:

1 define the stability constant, K_{stab}, of a complex as the equilibrium constant for the formation of the complex ion in a solvent (from its constituent ions or molecules)

2 write an expression for a K_{stab} of a complex ($[H_2O]$ should not be included)

3 use K_{stab} expressions to perform calculations

4 describe and explain ligand exchanges in terms of K_{stab} values and understand that a large K_{stab} is due to the formation of a stable complex ion

Initial assessment

1. Work with a partner to write a general set of rules to derive the expression for the equilibrium constant, *K*c, for a reaction such as:
   $m\,A(aq) + n\,B(aq) \rightleftharpoons p\,C(aq) + q\,D(aq)$

Introduction

Aqueous solutions of transition element ions are hydrated. They are complex ions with water molecules acting as ligands, forming dative (coordinate) bonds to the central metal ion.

Different ligands form complexes with different stabilities. For example, when we add concentrated aqueous ammonia to an aqueous solution of copper(II) sulfate, the ammonia ligands displace water ligands in a stepwise process.

$[Cu(H_2O)_6]2+(aq) + NH_3(aq) \rightleftharpoons [Cu(NH_3)(H_2O)5]2+(aq) + H_2O(l)$

$[Cu(NH_3)(H_2O)5]2+(aq) + NH_3(aq) \rightleftharpoons [Cu(NH_3)_2(H_2O)4]2+(aq) + H_2O(l)$

As we increase the concentration of ammonia, this process continues until four of the water molecules are replaced by ammonia molecules. The solution formed is a deep blue colour. The overall ligand substitution reaction is:

$$\underset{\text{blue solution}}{[Cu(H_2O)_6]^{2+}} + 4NH_3(aq) \rightleftharpoons \underset{\text{deeper blue solution}}{[Cu(NH_3)_4(H_2O)_2]^{2+}}(aq) + 4H_2O(l)$$

We can think of this exchange of ligands in terms of competing equilibria of the forward and backward reactions. The position of equilibrium lies in the direction of the more stable complex. In this case, the complex with ammonia as a ligand is more stable than the complex with just water as a ligand. If we dilute the complex with water, the position of equilibrium shifts to the left, and a complex with more water molecules as ligands forms.

The stability of the complex is expressed in terms of the equilibrium constants for ligand displacement. This is called the stability constant. Usually, an overall stability constant, *K*stab, is given rather than the stepwise constants. The method for writing equilibrium expressions for stability constants is similar to the one we used for writing *K*c. So for the equilibrium:

$[Cu(H_2O)_6]^{2+} + 4Cl^-(aq) \rightleftharpoons [CuCl_4]^{2-}(aq) + 6H_2O(l)$

the expression for the stability constant is:

$$K_{stab} = \frac{\left[[CuCl_4]^{2-}(aq)\right]}{\left[Cu(H_2O)_6\right]^{2+}\ [Cl-(aq)]^4}$$

Note:

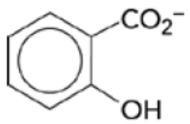
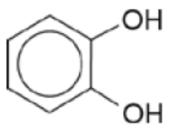
- water does not appear in the equilibrium expression because it is in such a large excess that its concentration is regarded as being constant

- the units for the stability constant are worked out in the same way as for the units of *K*c. For example, in the above case:

$$K_{\text{stab}} = \frac{\left[[\mathrm{CuCl_4}]^{2-}(\mathrm{aq})\right]}{\left[\mathrm{Cu(H_2O)_6}\right]^{2+}[\mathrm{Cl^-(aq)}]^4} \quad \frac{\cancel{(\mathrm{mol\ dm^{-3}})}}{\cancel{(\mathrm{mol\ dm^{-3}})} \times (\mathrm{mol\ dm^{-3}})^4} = \mathrm{dm^{12}\ mol^{-4}}$$

Stability constants are often given on a $\log_{10}$ scale. When expressed on a $\log_{10}$ scale, they have no units. Stability constants can be used to compare the stability of any two ligands. The values quoted usually give the stability of the complex relative to the aqueous ion, where the ligand is water.

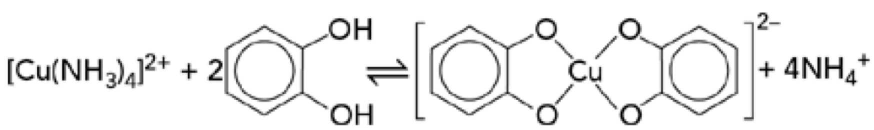
The following table gives some values of stability constants for various copper(II) complexes relative to their aqueous ions.

Ligand	**$\log_{10} K_{stab}$**
chloride, Cl^-	5.6
ammonia, NH_3	13.1
2-hydroxybenzoate 	16.9
1,2-dihydroxybenzene 	25.0

From Table, you can see that, in general, complexes with bidentate ions (2-hydroxybenzoate and 1,2-dihydroxybenzene) have higher stability constants than those with monodentate ligands.

We can use the values of the stability constants to predict the effect of adding different ligands to complex ions. For example, the addition of excess ammonia to the complex $[CuCl_4]^{2-}$(aq) should result in the formation of a dark blue solution of the ammonia complex. That is because the stability constant of the ammonia complex is higher than that of the chloride complex. The position of equilibrium is shifted to the right in the direction of the more stable complex.

Addition of excess 1,2-dihydroxybenzene, to the dark blue ammonia complex results in the formation of a green complex with 1,2-dihydroxybenzene. This is because the stability constant with the 1,2-dihydroxybenzene is much higher than that for ammonia. You can see from the Table that the log of *K*stab for 1,2-dihydroxybenzene is 25.0, so the actual value of *K*stab is 10 raised to the power 25, compared to the actual value for ammonia of 10 raised to the power 13.1. This shows that the copper complex with 1.2-ihydroxybenzene is much more stable than the complex ion formed with ammonia. These values of *K*stab predict that there will be hardly any of the copper complex with ammonia left in the equilibrium mixture shown in the equation below:



The stability constant, *K*stab, of a complex is the equilibrium constant for the formation of the complex ion in a solvent (from its constituent ions or molecules)

The higher the value of the stability constant, the more stable the complex.

2. Define stability constant, K_{stab}.

....

.... [1]

3. When a solution containing $EDTA^{4-}$ is added to a solution containing $[Cd(H_2O)_6]^{2+}$ a new complex is formed, $[CdEDTA]^{2-}$.
equilibrium 1 $[Cd(H_2O)_6]^{2+} + EDTA^{4-} \rightleftharpoons [CdEDTA]^{2-} + 6H_2O$
Write an expression for the stability constant, Kstab1, for equilibrium 1, and state its units.

Kstab1 =

units = ... [2]

4. A solution is made by dissolving $CuSO_4 \cdot 5H_2O$ in an excess of aqueous ammonia. This solution contains the copper complex $[Cu(NH_3)_4]^{2+}$.
Write an expression for the Kstab of $[Cu(NH_3)_4]^{2+}$.

Kstab = [1]

5. Nickel can form complexes with the ligands *en*, $H_2NCH_2CH_2NH_2$, and *tn*, $H_2NCH_2CH_2CH_2NH_2$, as shown.
equilibrium **1** $[Ni(H_2O)_6]^{2+} + 3en \rightleftharpoons [Ni(en)_3]^{2+} + 6H_2O$ Kstab = 6.76×10^{17}
equilibrium **2** $[Ni(H_2O)_6]^{2+} + 3tn \rightleftharpoons [Ni(tn)_3]^{2+} + 6H_2O$ Kstab = 1.86×10^{12}
Construct an expression for the stability constant, *K*stab, for equilibrium **1**.
State the units for *K*stab.

*K*stab =

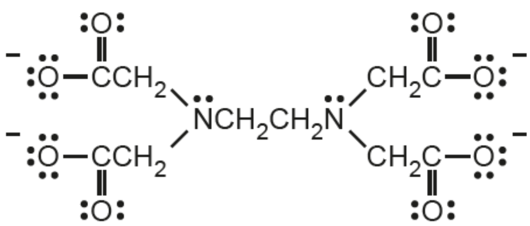
units = ...[2]

(iii) Describe what the *K*stab values indicate about the position of equilibrium for equilibrium **1** and **2**. Use the *K*stab values to deduce which complex, $[Ni(en)_3]^{2+}$ or $[Ni(tn)_3]^{2+}$, is more stable.

....

.... [1]

6. The structure of the polydentate ligand, $EDTA^{4-}$, is shown in Fig. 4.1.



The stability constants, at 298 K, of five octahedral complexes are given in Table 4.1.

Table 4.1

complex	K_{stab}
$[Cu(EDTA)]^{2-}$	6.31×10^{19}
$[Cr(EDTA)]^{2-}$	1.00×10^{13}
$[Cr(EDTA)]^{-}$	1.00×10^{24}
$[Fe(EDTA)]^{2-}$	2.00×10^{14}
$[Fe(EDTA)]^{-}$	1.26×10^{25}

(i) Deduce the number of lone pairs donated by each $EDTA^{4-}$ ligand in a single $[Fe(EDTA)]^{2-}$ complex ion.

.... [1]

(ii) Identify the most stable complex in Table 4.1. Explain your choice.

....

.... [1]

7. **(i)** Write an expression for the stability constant, K_{stab}, of the $[Co(NH_3)_6]^{2+}$ complex ion.

K_{stab} = [1]

(ii) Give the units of the stability constant, K_{stab}, of the $[Co(NH_3)_6]^{2+}$ complex ion.

units = ... [1]

(iii) The numerical value of the stability constant, K_{stab}, of the $[Co(NH_3)_6]^{2+}$ complex ion is 7.7×10^4.
In an aqueous solution, the concentration of the $[Co(NH_3)_6]^{2+}$ complex ion is 0.0740 mol dm^{-3} and the concentration of NH_3 is 0.480 mol dm^{-3} at equilibrium.
Calculate the concentration of $[Co(H_2O)_6]^{2+}$ in this solution.

concentration = ... mol dm^{-3} [1]

8. The stability constants, Kstab, of three complexes of mercury(II) are given in the table.

complex	Kstab
$[Hg(CN)_4]^{2-}$	2.5×10^{41}
$[HgCl_4]^{2-}$	1.7×10^{16}
$[HgI_4]^{2-}$	2.0×10^{30}

(i) Write an expression for the Kstab of $[Hg(CN)_4]^{2-}$.

Kstab = [1]

(ii) An aqueous solution containing Hg^{2+} is added to a solution containing equal concentrations of CN^{-}(aq), Cl^{-}(aq) and I^{-}(aq). The mixture is left to reach equilibrium.
Predict which of the complexes $[Hg(CN)_4]^{2-}$, $[HgCl_4]^{2-}$ and $[HgI_4]^{2-}$ is present in the resulting mixture in the highest concentration and which is present in the lowest concentration.
Explain your answer.

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....

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.... [2]

9. **(i)** Nickel can form complexes with the ligands *en*, $H_2NCH_2CH_2NH_2$, and *tn*, $H_2NCH_2CH_2CH_2NH_2$,as shown.
equilibrium **1** $[Ni(H_2O)_6]^{2+} + 3en \rightleftharpoons [Ni(en)_3]^{2+} + 6H_2O$ *K*stab = 6.76×10^{17}
equilibrium **2** $[Ni(H_2O)_6]^{2+} + 3tn \rightleftharpoons [Ni(tn)_3]^{2+} + 6H_2O$ *K*stab = 1.86×10^{12}
Construct an expression for the stability constant, *K*stab, for equilibrium **1**.
State the units for *K*stab.

*K*stab =

units = ...[2]

(ii) Describe what the *K*stab values indicate about the position of equilibrium for equilibrium **1** and **2**. Use the *K*stab values to deduce which complex, $[Ni(en)_3]^{2+}$ or $[Ni(tn)_3]^{2+}$, is more stable.

....

.... [1]

10. (a) Cd^{2+}(aq) ions form tetrahedral complexes with CH_3NH_2, OH^{-} and Cl^{-} ions, as shown in equilibria 1, 2 and 3.
equilibrium 1 $Cd^{2+}(aq) + 4CH_3NH_2(aq) \rightleftharpoons [Cd(CH_3NH_2)_4]^{2+}(aq)$ *K*stab = 3.3 × 106
equilibrium 2 $Cd^{2+}(aq) + 4OH^{-}(aq) \rightleftharpoons Cd(OH)_4^{2-}(aq)$ *K*stab = 5.0×10^{8}
equilibrium 3 $Cd^{2+}(aq) + 4Cl^{-}(aq) \rightleftharpoons CdCl_4^{2-}(aq)$ *K*stab = 6.3×10^{2}

(i) Give the units of *K*stab for equilibrium 1.

.... [1]

(ii) Write an expression for *K*stab for equilibrium 3.

*K*stab = [1]

(b) Cd^{2+}(aq) exists as a complex ion, $[Cd(H_2O)_6]^{2+}$(aq).
Identify the most stable and the least stable of the complexes in Table 4.1 by placing **one** tick (√) in each column.
Explain your answer.

Table 4.1

complex	most stable	least stable
$[Cd(H_2O)_6]^{2+}(aq)$		
$[Cd(OH)_4]^{2-}(aq)$		
$[Cd(CH_3NH_2)_4]^{2+}(aq)$		
$[CdCl_4]^{2-}(aq)$		

explanation ....

....

.... [2]

11. The complex $[CaEDTA]^{2-}$ can be used to remove toxic metals from the body.
Table 1.2 shows the numerical values for the stability constants, Kstab, for some metal ions with $EDTA^{4-}$.

Table 1.2

complex	K_{stab}
$[CaEDTA]^{2-}$	5.0×10^{10}
$[CrEDTA]^{-}$	2.5×10^{23}
$[FeEDTA]^{-}$	1.3×10^{25}
$[PbEDTA]^{2-}$	1.1×10^{18}

An aqueous solution containing $[CaEDTA]^{2-}$ is added to a solution containing equal concentrations of $Cr^{3+}(aq)$, $Fe^{3+}(aq)$ and $Pb^{2+}(aq)$. The resulting mixture is left to reach a state of equilibrium.

(i) State the type of reaction when $[CaEDTA]^{2-}$ reacts with $Cr^{3+}(aq)$, $Fe^{3+}(aq)$ and $Pb^{2+}(aq)$.

.... [1]

(ii) Deduce the relative concentrations of [CrEDTA]–, [FeEDTA]– and $[PbEDTA]^{2-}$ present in the resulting mixture.
Explain your answer.

.... > .... > ....
highest concentration lowest concentration

....

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.... [1]

12. The numerical value of the stability constant, *K*stab, of the copper(I) complex $[Cu(CN)_4]^{3-}$ is 2.0×10^{27}.

(i) Write an expression for the *K*stab of $[Cu(CN)_4]^{3-}$.

*K*stab =

[1]

(ii) In a solution the concentrations of CN^- and $[Cu(CN)_4]^{3-}$ are both 0.0010 mol dm^{-3}.
Use your expression from **(i)** and the value of *K*stab to calculate the concentration of $Cu^+(aq)$ in this solution.

concentration of Cu^{+}(aq) = ... mol dm^{-3} [1]

13. The mineral chromite contains a compound which has the formula $FeCrnO_4$. The oxidation state of iron in $FeCrnO_4$ is +2.
A sample of 4.18 g of $FeCrnO_4$ is dissolved in an excess of sulfuric acid. The resulting solution is made up to 250 cm3. This is solution F.
All the Fe^{2+} ions in 25.0 cm^3 of solution F are oxidised to Fe^{3+} ions by exactly 18.7 cm^3 of 0.0200 mol dm^{-3} $KMnO_4$.
One MnO_4^- ion reacts with five Fe^{2+} ions. Assume no other oxidation reaction occurs.
(i) Write an equation for the reaction of Fe^{2+} ions with MnO_4^- ions in acid solution.

.... [1]

(ii) Calculate the number of moles of Fe^{2+} ions in 25.0 cm^3 of solution F.

number of moles of Fe^{2+} ions = ... [2]

(iii) Calculate the Mr of $FeCr_nO_4$ and use your answer to deduce the value of n.

Mr of $FeCr_nO_4$ = ... value of n = ... [2]

14. Table 3.2 gives details of some complexes of Fe^{3+}.

Table 3.2

complex	colour	value of Kstab
$[Fe(H_2O)^{2+}]^{3+}$	violet	1
$[Fe(H_2O)_5SCN]^{2+}$	red	1.40×10^2
$[Fe(H_2O)_5F]^{2+}$	colourless	2.40×10^5

(i) Write an expression for Kstab of $[Fe(H_2O)_5SCN]^{2+}$.

Kstab = [1]

(iii) Use information in Table 3.2 to calculate the value of the equilibrium constant, Kc, for the following reaction.
$[Fe(H_2O)_5SCN]^{2+} + F^- \rightleftharpoons [Fe(H_2O)_5F]^{2+} + SCN^-$

Kc = ... [1]

(iv) A few drops of KF(aq) are added to a solution of $[Fe(H_2O)_6]^{3+}$(aq), followed by a few drops of KSCN(aq). Use information in Table 3.2 to describe any observations after each addition. Explain your answer.

....

....

....

.... [2]

15. Cadmium forms complex ion X, $[Cd(NH_3)_4]^{2+}$.
When a solution containing CN^- ions is added to an aqueous solution of X, a ligand exchange reaction takes place, forming complex ion Y. Y contains no NH_3 ligands and no H_2O ligands.
Y is in a much higher concentration in the mixture than X.
The oxidation state and coordination number of cadmium do not change in this reaction.
(i) Write an ionic equation for this reaction, using the formulae of the complex ions.

.... [2]

(ii) Cadmium forms complex ion Z in the same oxidation state and with the same coordination number as in X. All the ligands in Z are Cl – ions.
When NaCl (aq) is added to a solution of X, very little Z forms.
Write the three cadmium complexes, X, Y and Z, in order of increasing stability constant, Kstab.

....
smallest value of Kstab largest value of Kstab [1]

16. Ni^{2+} ions form a number of different complex ions, including $[Ni(H_2O)_6]^{2+}$, $[Ni(NH_3)_6]^{2+}$ and $[Ni(en)_3]^{2+}$.
The abbreviation en represents 1,2-diaminoethane. The numerical values of two stability constants, Kstab, are given in Table 5.1.
Table 5.1

Complex	Kstab
[Ni(NH3)6]$^{2+}$	4.8 × 107
[Ni(en)3]$^{2+}$	2.0 × 1018

(a) Complete the expression for the Kstab of [Ni(en)3]$^{2+}$.

Kstab =

[1]

(b) A solution of $[Ni(H_2O)_6]^{2+}$ is added to a solution that contains 0.10 mol dm–3 NH_3 and 0.10 mol dm–3 en.
(i) Predict which complex ion, $[Ni(NH_3)_6]^{2+}$ or $[Ni(en)_3]^{2+}$, is present in the resulting mixture in the highest concentration. Explain your answer.

complex ion present in largest concentration = ...

explanation .... [1]

(ii) Complete the equation for the ligand exchange reaction occurring in (i).

$[Ni(H_2O)_6]^{2+}$ + ... + ... [1]

17. Table 3.1 gives values for the stability constants, *K*stab, of different complexes of iron.
Table 3.1
complex stability constant, *K*stab
[Fe(H₂O)5(H2PO4)]$^{2+}$ 5.90 × 101
[Fe(H₂O)5SCN]$^{2+}$ 1.30 × 102
(i) [Fe(H₂O)5(H₂PO4)]$^{2+}$ can form when H3PO4 reacts with $[Fe(H_2O)_6]$3+.

Write an equation for this reaction.

.... [1]

(ii) Write an expression for *K*stab of $[Fe(H_2O)5SCN]^{2+}$ and give its units.

*K*stab = units = ... [2]

(iii) Use the stability constant data in Table 3.1 to calculate the value of the equilibrium constant, *K*c, for the following equilibrium.
$[Fe(H_2O)5(H_2PO4)]^{2+}$ + SCN– $[Fe(H_2O)5SCN]^{2+}$ + $H_2PO4–$

value of *K*c = ... [1]

18. The numerical value of the stability constant, *K*stab, of the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex ion is 1.40×10^{13}.
(ii) Compare the stabilities of the $[Cu(H_2O)_6]^{2+}$ and $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex ions. Explain your answer.

....

.... [1]

(iii) Write an expression for the stability constant, *K*stab, of the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex ion.
State the units of the stability constant.

*K*stab = units = ...[2]

(iv) In a particular solution the concentration of the $[Cu(NH_3)_4(H_2O)_2]^{2+}$ complex ion is
0.0074 mol dm^{-3} and the concentration of NH_3 is 0.57 mol dm^{-3}.
Use your expression in **(c)(iii)** and the *K*stab value of 1.40 × 1013 to calculate the concentration of the $[Cu(H_2O)_6]^{2+}$ complex ion in this solution.

concentration of $[Cu(H_2O)_6]^{2+}$ = ... mol dm–3 [1]

19. Write an expression for the stability constant, *K*stab, of [Fe(edds)]–(aq).

*K*stab = [1]

(v) The table shows the values for the stability constants, *K*stab, of both complexes.
complex *K*stab / mol–1 dm3
[Fe(edds)]– 3.98×10^{20}
[Fe(edta)]– 1.26×10^{25}
Predict which of the [Fe(edds)]– and [Fe(edta)]– complexes is more stable.
Explain your answer with reference to the *K*stab value for each complex.

....

....

.... [1]

(vi) When an excess of $edta^{4-}$(aq) is added to [Fe(edds)]–(aq), the following equilibrium is established.
[Fe(edds)]–(aq) + edta4–(aq) [Fe(edta)]–(aq) + edds4–(aq)
Calculate the equilibrium constant, *K*c, for this equilibrium, using the *K*stab values given in the table in **(c)(v)**.

*K*c = ... [1]

(c) Silver bromide, AgBr, dissolves in an aqueous solution of $S_2O_3^{2-}$ ions to form the complex ion $[Ag(S_2O_3)_2]$3–. The $S_2O_3^{2-}$ ions act as monodentate ligands.
equilibrium 1 AgBr(s) + $2S_2O_3^{2-}$–(aq) $[Ag(S_2O_3)_2]$3–(aq) + Br –(aq)
(ii) Write an expression for the equilibrium constant, *K*c, for equilibrium 1.

*K*c = [1]

(iii) Some additional data are given about the dissolution of AgBr in $S_2O_3^{2-}$(aq).
equilibrium constant numerical value
solubility product, *K*sp, of AgBr 5.4×10^{-13}
stability constant, *K*stab, of $[Ag(S_2O_3)_2]$3– 2.9×10^{13}
Use your answer to **(c)(ii)** and these data to calculate *K*c for equilibrium 1. Include the units for *K*c.

*K*c = ... units .... [2]

(d) The numerical values for the stability constants, *K*stab, of two other silver(I) complexes are given.
silver(I) complex numerical value of *K*stab
$[Ag(CN)_2]$– 5.3 × 1018
$[Ag(NH_3)_2]$+ 1.6 × 107

20. An aqueous solution containing Ag+ is added to a solution containing equal concentrations of CN–(aq), NH_3(aq) and $S_2O_3^{2-}$(aq). The mixture is left to reach equilibrium.
Deduce the relative concentrations of $[Ag(CN)_2]$–, $[Ag(NH_3)_2]$+ and $[Ag(S_2O_3)_2]$3– present in the resulting mixture. Explain your answer.

.... > .... > ....

highest concentration lowest concentration

....

.... [2]

21. An excess of aqueous ammonia is added to a solution containing the complex ion $[Co(H_2O)_6]^{2+}$.
$[Co(H_2O)_6]^{2+} + 6NH_3$ $[Co(NH_3)_6]^{2+} + 6H_2O$
(a) Complete the sentence to describe the colour change that will be seen during this reaction.

The colour changes from .... to .... [1]

(b) Write an expression for the stability constant, *K*stab, of $[Co(NH_3)_6]^{2+}$.

*K*stab = [1]

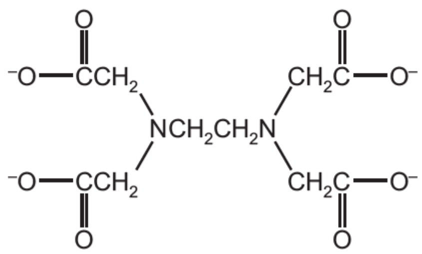
(c) The numerical value of *K*stab of $[Co(NH_3)_6]^{2+}$ is 7.7×10^4.
What deduction about the properties of $[Co(NH_3)_6]^{2+}$ and $[Co(H_2O)_6]^{2+}$ can be made from this *K*stab value?

.... [1]

22. **(ii)** When a solution containing $EDTA^{4-}$ is added to a solution containing $[Cd(H_2O)_6]^{2+}$ a new complex is formed, $[CdEDTA]^{2-}$.
equilibrium **1** $[Cd(H_2O)_6]^{2+} + EDTA^{4-} \rightleftharpoons [CdEDTA]^{2-} + 6H_2O$

Circle, on the structure of $EDTA^{4-}$, the **six** atoms that form bonds with the metal ion.

$EDTA^{4-}$



Write an expression for the stability constant, K_{stab1}, for equilibrium **1**, and state its units.

*K*stab1 =

units = ... [2]

(b) Cadmium ions form complexes with methylamine, CH_3NH_2, and with 1,2-diaminoethane, $H_2NCH_2CH_2NH_2$, as shown in equilibriums **2** and **3**. 1,2-diaminoethane is shown as en.
equilibrium **2** $[Cd(H_2O)_6]^{2+} + 4CH_3NH_2$ $[Cd(CH_3NH_2)_4(H_2O)_2]^{2+} + 4H_2O$ $K_{stab2} = 3.60 \times 10^6$
equilibrium **3** $[Cd(H_2O)_6]^{2+} + 2en$ $[Cd(en)_2(H_2O)_2]^{2+} + 4H_2O$ $K_{stab3} = 4.20 \times 10^{10}$

An equilibrium is set up between these two complexes as shown in equilibrium **4**.
equilibrium **4** $[Cd(CH_3NH_2)_4(H_2O)_2]^{2+} + 2en$ $[Cd(en)_2(H_2O)_2]^{2+} + 4CH_3NH_2$ $\Delta H^o = +0.840$ kJ mol^{-1}
$\Delta S^o = +80.9$ J K^{-1} mol^{-1}
(i) K_{eq4} is the equilibrium constant for equilibrium **4**.
Write an expression for K_{eq4} in terms of K_{stab2} and K_{stab3}.

*K*eq4 = [1]

(ii) Calculate the value of the standard Gibbs free energy change, ΔG^o, for equilibrium **4** at 298 K.

ΔG^o = ... kJ mol^{-1} [2]

(iii) State how the value of ΔG^o changes as the temperature increases. Explain your answer.

....

....

.... [1]

*****This is not a complete note. It is just to guide you. It is recommended to study the prescribed textbooks along with this material. *****