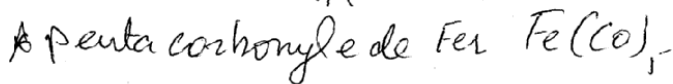
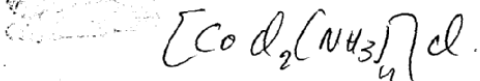
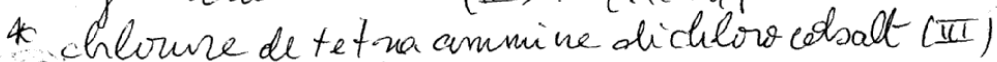
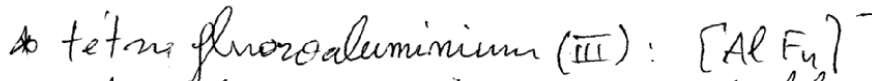
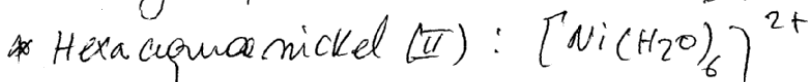
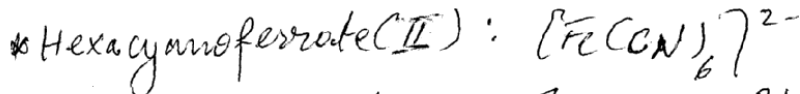
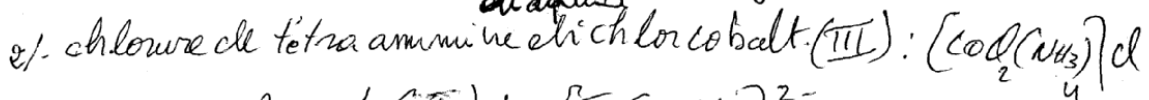
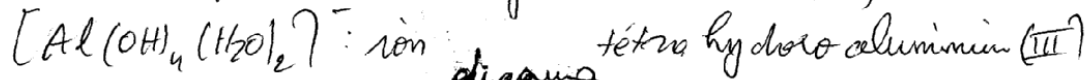
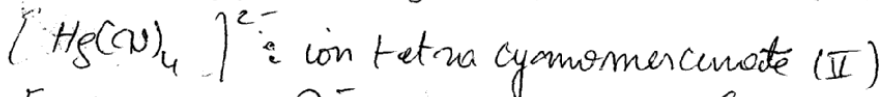
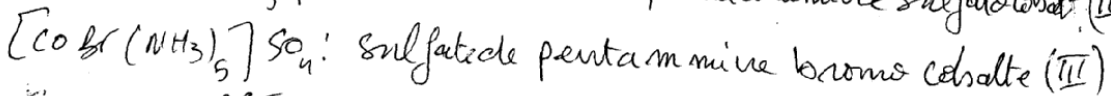
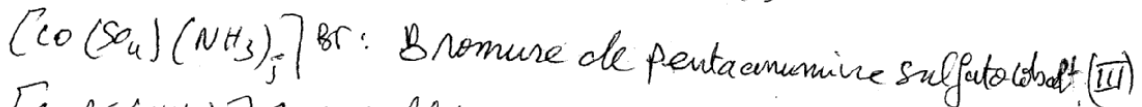
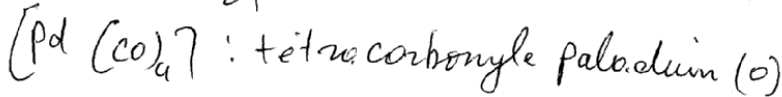
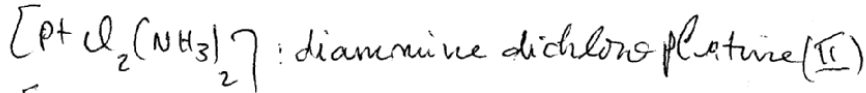
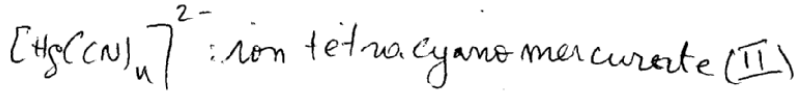


T.D N° 5 corrigé type
équilibre de complexation

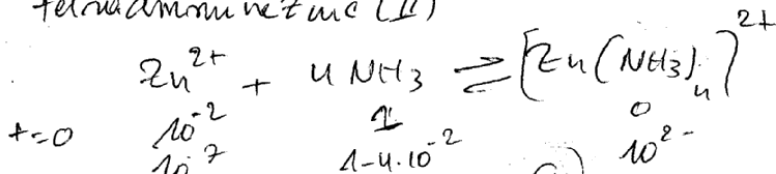
1/ Exercice 1:



EX N° 2

Calculer la constante de formation du complexe

tétrammine zinc (II)



- la dissolution du sel $\text{NaF} \rightarrow \text{Na}^+ + \text{F}^-$

$$[\text{Na}^+] = [\text{F}^-] = 0,2 \text{ M}$$

- après mélange les concentrations deviennent

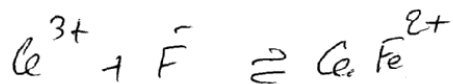
$$[\text{Na}^+] = [\text{F}^-] = \frac{0,2 \times 20}{40} = 0,1 \text{ M}$$

- la dissolution du sel $\text{Ce}_2(\text{SO}_4)_3$ donne :

$$[\text{Ce}^{3+}] = 2 \times 0,1 = 0,2 \text{ M}, \quad [\text{SO}_4^{2-}] = 3 \times 0,1 = 0,3 \text{ M}$$

- après mélange les concentrations deviennent

$$[\text{Ce}^{3+}] = 0,1 \text{ M} \quad \text{et} \quad [\text{SO}_4^{2-}] = 0,15 \text{ M}$$



$$t=0 \quad 0,1 \quad 0,1 \quad 0$$

$$t_{\text{eq}} \quad 0,1-x=\varepsilon \quad 0,1-x=\varepsilon \quad x=0,1$$

on calcul ε

$$K_f = \frac{[\text{CeF}^{2+}]}{[\text{Ce}^{3+}][\text{F}^-]} = \frac{[\text{CeF}^{2+}]}{\varepsilon^2}$$

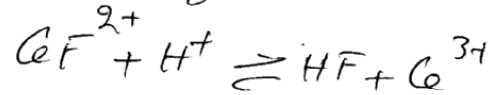
$$\varepsilon = \sqrt{\frac{[\text{CeF}^{2+}]}{K_f}} \Rightarrow \varepsilon = 2,81 \cdot 10^{-3} \text{ M}$$

$$K_f = 10^{4,1}$$

$$\text{donc: } [\text{F}^-] = 2,81 \cdot 10^{-3} \text{ M}, \quad [\text{Ce}^{3+}] = 2,81 \cdot 10^{-3} \text{ M}, \quad [\text{CeF}^{2+}] = 0,1 \text{ M}$$

② à la solution obtenue, on ajoute sans dilution une quantité d'acide HCl jusqu'à ce que 50% du cérium (Ce^{3+}) soit complexé. Déterminer les concentrations de Ce^{3+} , F^- , CeF^{2+} après cet ajout de HCl, en déduire le pH de la solution

Après l'ajout de HCl on aura une dissociation du complexe selon la réaction globale :



on a dissociation de 50% du complexe alors

$$[\text{CeF}^{2+}] = [\text{Ce}^{3+}] = 0,05 \text{ M}$$

$$K_f = \frac{[\text{CeF}^{2+}]}{[\text{Ce}^{3+}][\text{F}^-]} = \frac{1}{[\text{F}^-]} \Rightarrow [\text{F}^-] = \frac{1}{K_f} = 7,94 \cdot 10^{-5} \text{ M}$$

* en déduire le pH de la solution

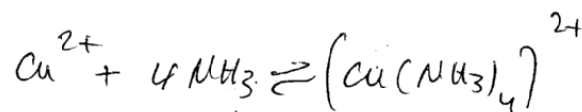
$$[\text{HF}] = 0,05 \text{ M} \text{ et } [\text{F}^-] = 7,94 \cdot 10^{-5} \text{ M}$$

et aussi : $K_a = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]} \Rightarrow [\text{H}_3\text{O}^+] = K_a \frac{[\text{HF}]}{[\text{F}^-]}$

$$[\text{H}_3\text{O}^+] = 0,4 \text{ M} \Rightarrow -\lg [\text{H}_3\text{O}^+] = \text{pH} = \underline{0,4}$$

Exercice 5:

1L donc $V = 1 \text{ litre}$



EI $n_0 = 0,1 \quad n_1 = 0,4 \quad 0$

EF $n_0 - x \quad n_1 - 4x \quad x = 0,1 \text{ mol.}$

$$= 0,1 - x = 0 \quad = 0,4 - 4x = 0$$

$$x = 0,1$$

$$K_f = \frac{[(\text{Cu}(\text{NH}_3)_4)^{2+}]_{\text{EF}}}{[\text{Cu}^{2+}]_{\text{EF}} \cdot [\text{NH}_3]_{\text{EF}}^4} = \frac{1}{K_D} = 4 \cdot 10^{12} \text{ donc on peut}$$

(4)

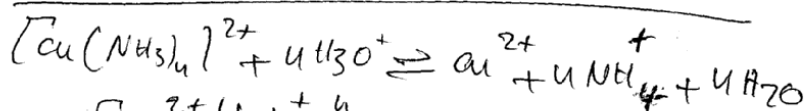
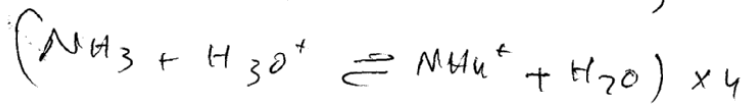
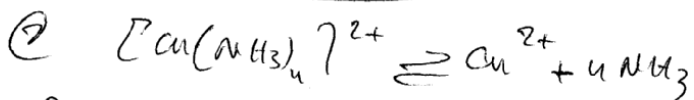
considérer la réaction comme totale

$$[\text{NH}_3] = 4[\text{Cu}^{2+}] \quad K_f = \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}}{[\text{Cu}^{2+}][\text{NH}_3]^4}, \quad K_d = \frac{1}{K_f}$$

$$K_d = \frac{[\text{Cu}^{2+}][\text{NH}_3]^4}{[\text{Cu}(\text{NH}_3)_4]^{2+}} \quad \text{et} \quad [\text{Cu}^{2+}] = \frac{[\text{NH}_3]}{4}$$

$$K_d = \frac{[\text{NH}_3]^5}{4 \times 0,1} \Rightarrow [\text{NH}_3] = \sqrt[5]{0,4 \times K_d} = 2,5 \cdot 10^{-3} \text{ M}$$

$$\boxed{[\text{Cu}^{2+}] = 6,25 \cdot 10^{-4} \text{ M}}$$



$$K = \frac{[\text{Cu}^{2+}][\text{NH}_4^+]^4}{[\text{Cu}(\text{NH}_3)_4]^{2+}[\text{H}_3\text{O}^+]^4} = \frac{1}{K_f \cdot K_a^4} = 1,6 \cdot 10^{24}$$

Non 50% du complexe est détruit à 50%, il doit rester 50% de la quantité initiale dans la forme de Cu^{2+}

$$[\text{Cu}^{2+}] = \frac{50}{100} C_0 = 0,05 \text{ M}, \quad [\text{Cu}(\text{NH}_3)_4]^{2+} = \frac{50}{100} C_0 = 0,05 \text{ M}$$

$$[\text{NH}_4^+] = 4[\text{Cu}^{2+}] = 0,2 \text{ M}$$

$$[\text{H}_3\text{O}^+] = \sqrt[4]{\frac{[\text{Cu}^{2+}][\text{NH}_4^+]^4}{K [\text{Cu}(\text{NH}_3)_4]^{2+}}} = 1,8 \cdot 10^{-7} \text{ M}$$

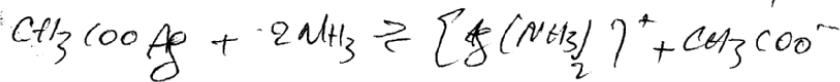
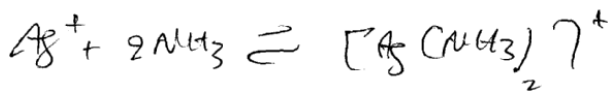
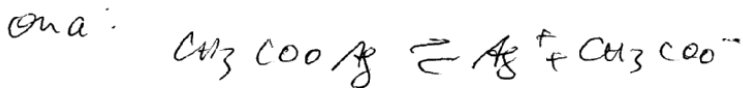
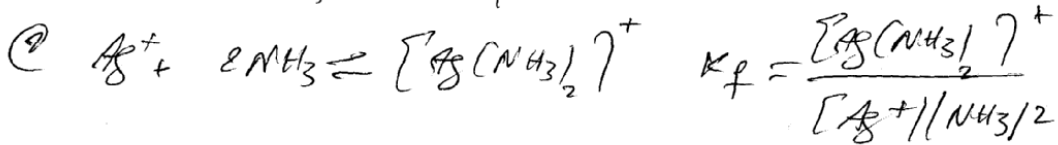
$$\Rightarrow \text{pH} = 6,7$$

Ex N° 6

Détermination de concentration de Ag^+ et CH_3COO^-

$$K_s = [Ag^+][CH_3COO^-] = S^2 \Rightarrow S = \sqrt{K_s} = 4,5 \cdot 10^{-2} \text{ M}$$

$$[Ag^+] = [CH_3COO^-] = 4,5 \cdot 10^{-2} \text{ M}$$



À 0 n_0 n_1 0 0

Apr $n_0 - x$ $n_1 - 2x$ x x

$$n_0 - x = 0 \Rightarrow n_0 = x = 0,1 \text{ mol}$$

$$n_1 [Ag(NH_3)_2]^+ = n_0, \quad CH_3COO^- = x = 0,1$$

$$K = \frac{[Ag(NH_3)_2]^+ [CH_3COO^-]}{[NH_3]^2} = K_s \cdot K_f = 3,2 \cdot 10^4$$

litre seque $V = 1 \text{ litre} \Rightarrow C = \frac{n}{V} \Rightarrow C = n$

$$K = \frac{(0,1)^2}{(n_1 - 0,2)^2} = 3,2 \cdot 10^4 \Rightarrow n_1 = n_{NH_3} = 0,2 \text{ mol}$$