

$$[A]_0 = 0,6 \text{ M}, \quad [A](5 \text{ min}) = 0,25 \text{ M}$$

$$a) \quad [B]_e = K \cdot [A]_e$$

$$[B]_e = [A]_0 - [A]_e$$

$$[A]_0 - [A]_e = K \cdot [A]_e$$

$$\frac{[A]_0}{K+1} = \underline{[A]_e} = 0,1905 \text{ M} \rightarrow \underline{[B]_e} = 0,4095 \text{ M}$$

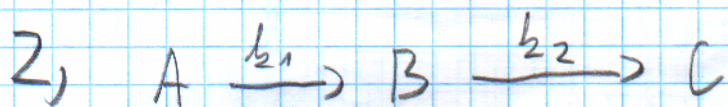
$$b) \quad k_1 = k_2 \cdot K$$

$$k_1 + k_2 = -\frac{1}{5 \text{ min}} \cdot \ln \left[\frac{[A](5 \text{ min}) - [A]_e}{[A]_0 - [A]_e} \right]$$

$$= 0,3858 \frac{1}{\text{min}}$$

$$k_2(K+1) = 0,3858 \frac{1}{\text{min}}$$

$$\underline{k_2} = 0,1225 \frac{1}{\text{min}} \rightarrow \underline{k_1} = 0,2633 \frac{1}{\text{min}}$$



$$t_{1/2}^A(296\text{K}) = 2500 \text{ min}$$

$$\text{Starting from pure B: } t_{1/2}^B(304\text{K}) = 150 \text{ min}$$

$$\frac{k_2(311\text{K})}{k_2(296\text{K})} = 3$$

$$k_2(296\text{K})$$

$$t_{B \max}(296\text{K}) = ? = \frac{\ln \frac{k_1(296\text{K})}{k_2(296\text{K})}}{k_1(296\text{K}) - k_2(296\text{K})}$$

$$k_1(296\text{K}) \text{ from the half-life of A: } \frac{\ln 2}{t_{1/2}^A(296\text{K})} = 2,77 \cdot 10^{-4} \frac{1}{\text{min}}$$

$$k_2(296\text{K}) \text{ from the Arrhenius Eq.: } k_2(296\text{K}) = A_2 \cdot e^{-\frac{E_{a,2}}{R \cdot 296\text{K}}}$$

↳ need $E_{a,2}$ and A_2

$$E_{a,2} = \frac{R \cdot \ln \frac{k_2(296\text{K})}{k_2(311\text{K})}}{\frac{1}{311\text{K}} - \frac{1}{296\text{K}}} = 56055 \frac{\text{J}}{\text{mol}}$$

$$\text{we can get } A_2 \text{ from the } 304\text{K} \text{ results: } k_2(304\text{K}) = \frac{\ln 2}{150 \text{ min}}$$

$$k_2(296\text{K}) = A_2 \cdot e^{-\frac{E_{a,2}}{R \cdot 296\text{K}}} \leftarrow A_2 = 1,98 \cdot 10^7 \frac{1}{\text{min}} \leftarrow A_2 \cdot e^{-\frac{56055 \text{ J/mol}}{R \cdot 304\text{K}}}$$

$$= 2,537 \cdot 10^{-3} \frac{1}{\text{min}}$$

$$t_{B \max} = \frac{\ln \frac{2,77 \cdot 10^{-4}}{2,537 \cdot 10^{-3}}}{2,77 \cdot 10^{-4} \frac{1}{\text{min}} - 2,537 \cdot 10^{-3} \frac{1}{\text{min}}} = 978 \text{ min} = 16,3 \text{ h}$$

$$\left(1 - \frac{[B](3h)}{[B]_{\max}}\right) \cdot 100\% = ?$$

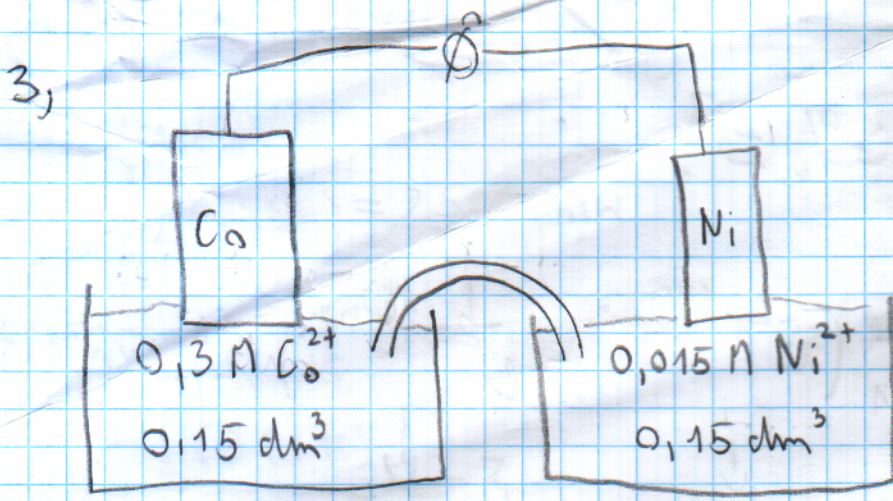
$$[B](3h) = [A]_0 \cdot \frac{k_1}{k_2 - k_1} \left(e^{-k_1 \cdot 180 \text{ min}} - e^{-k_2 \cdot 180 \text{ min}} \right) = [A]_0 \cdot 0,03897$$

\downarrow \downarrow \downarrow
 0,1226 0,9514 0,6334

$$[B]_{\max} = [A]_0 \cdot \left(\frac{k_1}{k_2} \right)^{\frac{k_2}{k_2 - k_1}} = 0,08324 \cdot [A]_0$$

\downarrow \downarrow
 0,1092 1,1226

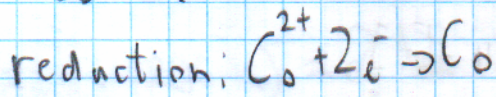
$$\left(1 - \frac{0,03897}{0,08324}\right) \cdot 100\% = \underline{\underline{53,18\%}}$$



$$T = 298 \text{ K}$$

$$\frac{RT}{F} = 0,02568 \text{ V}$$

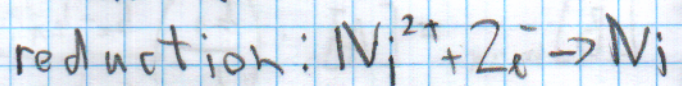
$$E_{\text{Co}} = ?$$



$$E_{\text{Co}} = E_{\text{Co}}^{\circ} - \frac{RT}{2F} \ln \frac{1}{0,13}$$

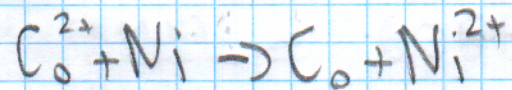
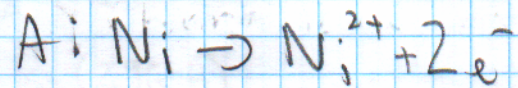
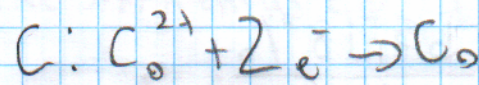
$$= -0,2925 \text{ V}$$

$$E_{\text{Ni}} = ?$$



$$E_{\text{Ni}} = E_{\text{Ni}}^{\circ} - \frac{RT}{2F} \ln \frac{1}{0,015}$$

$$= -0,3109 \text{ V}$$



Equilibrium: $E_{\text{Co}} = E_{\text{Ni}}$

Diluting the electrolytes lowers both E_s (see the

E_{Co} needs to be lowered

Nernst eqs.!)

$$E'_{\text{Co}} = E_{\text{Ni}} = -0,3109 \text{ V} = E_{\text{Co}}^{\circ} - \frac{RT}{2F} \ln \frac{1}{x} \Rightarrow x = 0,04135 \text{ M}$$

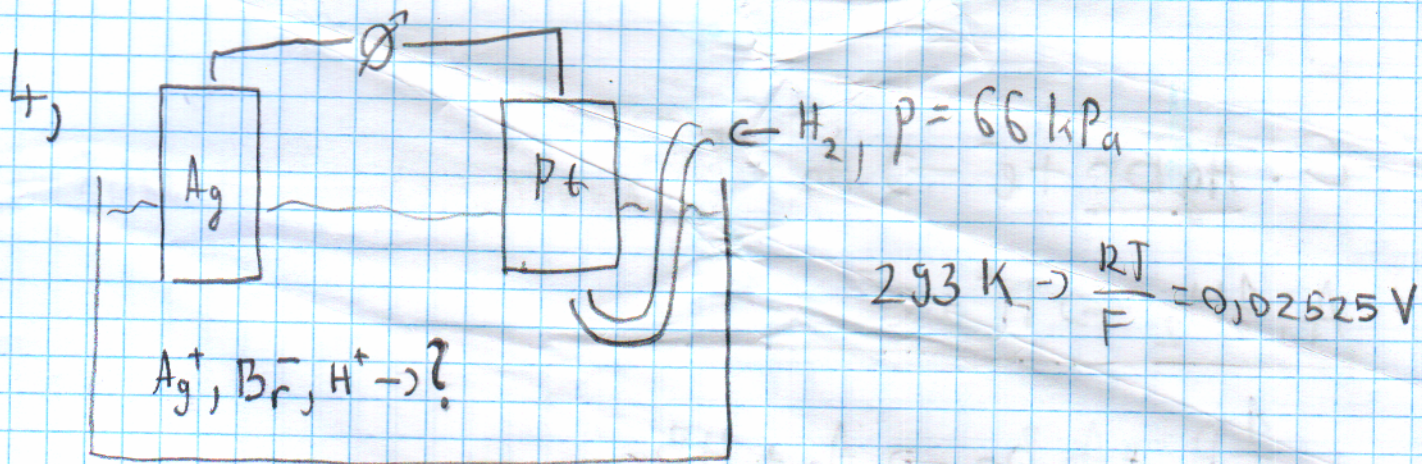
How much water is needed?

$$n = c_0 \cdot V_0 = 0,13 \text{ M} \cdot 0,15 \text{ dm}^3$$

$$= 0,045 \text{ mol}$$

$$\rightarrow V' = \frac{0,045 \text{ mol}}{0,04135 \text{ M}} = 0,631 \text{ dm}^3$$

+ 0,1481 dm³ water needed



What will be the activities of the soluted ions?

INITIALLY: Ag^+ : $0,025 \text{ dm}^3 \cdot 0,035 \text{ M} = \underline{0,000875 \text{ mol}}$

(NO_3^- is not interesting) H^+ and Br^- : $0,045 \text{ dm}^3 \cdot 2 \text{ M} = \underline{0,09 \text{ mol}}$

AFTER REACTION: Ag^+ : practically 0 mol in solution

Br^- : $0,09 \text{ mol} - 0,000875 \text{ mol} = \underline{0,089125 \text{ mol}}$

H^+ : 0,09 mol

activities: $a_{Br^-} = \frac{[Br^-] \cdot \gamma_{\pm}}{C_0} = \left(\frac{0,089125 \text{ mol}}{0,07 \text{ dm}^3} \cdot 0,95 \right) / 1 \text{ M}$

$= \underline{1,2096}$

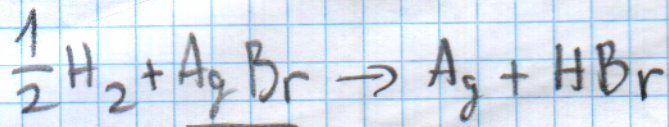
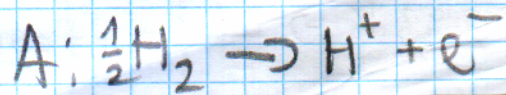
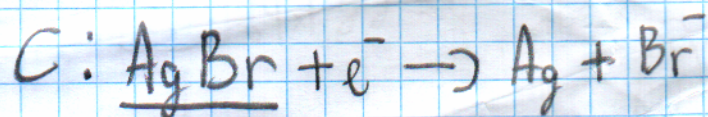
$a_{H^+} = \underline{1,2214}$

$E_{Ag/AgBr}?$ $AgBr + e^- \rightarrow Ag + Br^-$, $E_{Ag/AgBr} = E_{Ag/AgBr}^0 - \frac{RT}{F} \ln 1,2096$

$= \underline{0,0565 \text{ V}}$

$E_{H_2}?$ $H^+ + e^- \rightarrow \left(\frac{1}{2} \right) H_2$, $E_{H_2} = - \frac{RT}{F} \ln \frac{\sqrt{66 \text{ kPa} / p_0}}{1,2214} = \underline{0,0193 \text{ V}}$

!!!



$$\underline{E} = E_{Ag/AgBr} - E_{H_2} = \underline{0,0562 V}$$

$$\underline{\Delta_r G} = -z \cdot F \cdot E = \underline{5422 \frac{J}{mol}}$$

$$\underline{L} = e \left(E_{Ag/AgBr}^\circ - E_{Ag}^\circ \right) \frac{F}{RT} = \underline{2,966 \cdot 10^{-13}}$$

NOTES :- if there was ^(much) more Ag^+ than Br^- ,
it would have been easier to
treat the Ag electrode as a
metal electrode

- if we did not ignore the $AgBr$
that gets soluted back, we would
have had to first calculate
 L , then compute the true a_{Br^-}
from $L = (1,2096 + x) \cdot x = a_{Br^-} \cdot a_{Ag^+}$