Friday Worksheet **Name: ………………**

Electrolysis worksheet 4

1) An electrolytic cell attempts to produce chlorine gas via the electrolysis of aqueous 0.01M KCl solution using inert electrodes. A gas is produced at both electrodes.

a) Give the likely equation to the reaction occurring at the anode.

$2H_2O(l)\ ^{-}\rightarrow\ 4H^+ (aq) + O_2(g)\ \ 4e$

b) Give the likely equation to the reaction occurring at the cathode.

$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^- (aq)$

c) What happens to the pH of the solution surrounding the anode? Explain

Decreases as $[H^-]$ increases

2) What mass (in grams) of nickel could be electroplated from a solution of nickel(II) chloride by a current of 0.450 amperes flowing for 5.50 hours?

$Ni^{2+}(aq) + 2e \Rightarrow Ni(s)$

Step 1 find the total charge delivered

=> Q =It = 0.450 X 5.50 X 60 X 60 =8910

Step 2 find the mol of electrons

=> n_e = 8910/96500 = 0.0923

Step 3 find mol of Ni

=> n_{Ni} = 0.0923 / 2 = 0.0462

Step 4 find the mass of Ni

=> 0.0462 X 58.7 = 2.71 g

3) Pure aluminium is to be extracted from a large sample of $AlCl_3$. An electrolytic cell is set up run for 4.50 hours with a current of 0.490 amperes using inert electrodes.

a) One student suggested setting up an electrolytic cell using (1.00 M) $AlCl_3$. What are the products formed at the:

i. cathode ------ $2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^- (aq)$

Ii. anode ------ $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

b) Another student suggested molten $AlCl_3$ with the exclusion of water. What are the products formed at the:

i. cathode ------ The reaction is $Al^{3+} (l) + 3e^- \rightarrow Al(l)$ hence Al(l)

ii. anode ------ The reaction is $2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$ **hence Cl_2(g) and electrons**

iii. How many litres of the gas produced at the anode when measured at 0°C and 101.3 kPa pressure, are produced when the electrode efficiency is only 65%?

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

Step 1 Find the total charge delivered

=> Q =It = 0.490 X 4.50 X 60 X 60 X 0.65 = 5160 C

Step 2 find the mol of electrons

=> n_e = 5160 / 96500 = 0.0535

Step 3 find the mol of chlorine

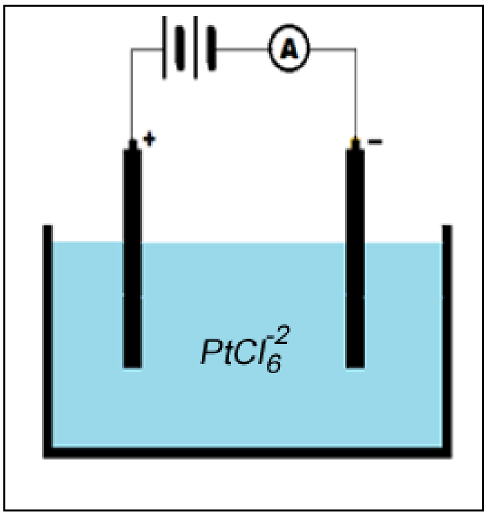
=> 0.0535 / 2 = 0.0267

Step 4 find the volume of chlorine (V = nRT/P)

=> 0.0267 X 8.31 X 273/101.3 = 0.599 Litres

4) A fine layer of platinum is to be plated onto an iron rod from a solution of $[PtCl_6]^{2-}$, using an average current of 10.0 amperes at an electrode efficiency of 70.0%?

a) The electrolytic cell shown on the right is used.



i. What material should the positive electrode be made from? platinum

ii. What is the reaction occurring at the cathode?

$Pt^{+4}(aq) + 4e \Rightarrow Pt(s)$

b) How long, in hours, would be required for the electroplating of 88.0 g of platinum

$Pt^{4+}(aq) + 4e \Rightarrow Pt(s)$

Step 1 find the mol of platinum..

=> n_{Pt} = 88.0 / 195 = 0.451

Step 2 find the mol of electrons needed

=> n_e = 0.451 X 4 = 1.804

Step 3 find the charge that this represents

=> Q = 1.804 X 96500 = 174086 C

Step 4 At 70.0% efficiency find the charge that must be delivered to achieve 174086 C

=> Charge needed = 174086/0.700 = 248694 C

Step 5 find the time necessary to deliver this amount of charge at a current of 10.0 A

=> 248694 / 10.0 = 24869.4 seconds

=> 6.91 hours