Friday Worksheet **Name: ……………….**

Electrolysis worksheet 1

1) A student prepares 1.0 M aqueous solutions of $AgNO_3$, $Fe(NO_3)_2$ $Cu(NO_3)_2$ and KNO_3.
Equal volumes of each solution are placed in separate beakers, identical platinum electrodes are placed in each beaker and each solution undergoes electrolysis with a current of 1.21Amps applied for 5.00 minutes at a voltage of 4.50 volts, under SLC.
Each cathode is then dried and weighed to determine mass change.

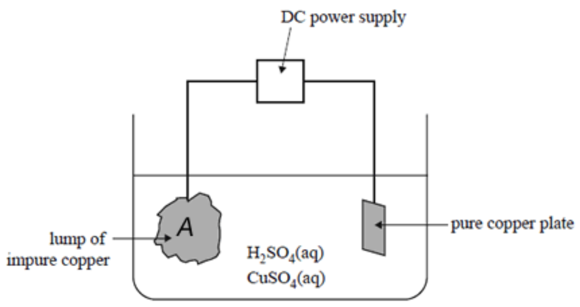
a) Write the equation for the half-reaction occurring at the anode and cathode of each cell.

$AgNO_3$
Anode (+)= $2H_2O(l) \Rightarrow O_2(g) + 4H^+ + 4e(aq)$
Cathode(-) = $Ag^+(aq) + e \Rightarrow Ag(s)$
$Cu(NO_3)_2$
Anode(+) = $2H_2O(l) \Rightarrow O_2(g) + 4H^+ + 4e(aq)$
Cathode(-) = $Cu^{2+}(aq) + 2e \Rightarrow Cu(s)$
KNO_3
Cathode(-)= $2H_2O(l) + 2e \Rightarrow H_2(g) + 2OH^-(aq)$
Anode (+)= $2H_2O(1) \Rightarrow O_2(g) + 4H^+(aq) + 4e$
Potassium will not be deposited as H_2O is a stronger oxidant than K^+, hence, H_2O will react rather than K^+ at the cathode.
$Fe(NO_3)_2$
Anode(+) = $Fe^{2+}(aq) + 2e \Rightarrow Fe(s)$
Cathode(-) = $Fe^{2+}(aq) \Rightarrow Fe^{3+}(aq) + e$

b) What is the mass of each metal deposited in grams?

Step 1 find the mol of electrons
=> n_e = It /96500 = 1.21 X 5.00 X 60.0 / 96500 = 0.00376
Step 2 find the mol of each metal that is deposited.
=> $Ag^+(aq) + e \Rightarrow Ag(s)$
n_{Ag} = 0.00376
=> $Cu^{2+}(aq) + 2e \Rightarrow Cu(s)$
=> n_{Cu} = 0.00188
=> $Fe^{2+}(aq) + 2e \Rightarrow Fe(s)$
=> n_{Fe} = 0.00188
Step 3 find the mass of each metal
=> $mass_{Ag}$ = 0.00376 X 107.9 = 0.406 g
=> $mass_{Fe}$ = 0.00188 X 55.8 = 0.105 g
=> $mass_{Cu}$ = 0.00188 X 63.6 = 0.120 g (3 sig fig)

2) An electrolytic process known as electro-refining is the final stage in producing highly purified copper. In a small-scale trial, a lump of impure copper is used as one electrode and a small plate of pure copper is used as the other electrode. The electrolyte is a mixture of aqueous sulfuric acid and copper sulfate.



	Mass of lump of impure copper	Mass of pure copper
before electrolysis	9.65 kg	1.62 kg
after electrolysis	0.963 kg	9.75 kg

a) What is the polarity of the terminal to which the lump of impure copper is to be attached.

Positive

b) In a trial experiment, the electrodes were weighed before and after electrolysis. Conditions in the electrolytic cell shown on the right are carefully controlled to ensure a high degree of copper purity and electrical efficiency. The results are provided in the table on the right.

On the basis of these results

i. calculate a percentage purity of the lump of impure copper

It is a common expectation of the examiners that a student be able to apply their chemical knowledge to unfamiliar contexts. Students must assume that the amount of impure copper that was removed from the lump labelled A of impure copper, included both copper and impurities. It is not correct to assume that the remaining 0.963 kg of impure copper contains just the impurities.

Mass of impure copper that was removed = 9.65 -0.963 = 8.69 kg

Mass of pure copper deposited = 9.75 – 1.62 = 8.13 kg

Percentage purity = (8.13 / 8.69) X 100 = 93.6%

ii. indicate one factor that may affect the accuracy of these results.

Not all the copper that was removed from the impure sample attached to the pure sample of copper.

Some of the copper deposited on the pure sample dislodged and fell off and could not be measured.

c) Use the mass of pure copper deposited, that is given in the table above, to determine the time, in hours, taken for this electrolysis reaction to be completed. Assume the current was a constant 28 A.

Step 1 find the mass of pure copper deposited.

Mass of pure copper deposited = 9.75 – 1.62 = 8.13 kg

Step 2 find the mol of pure copper deposited = 8,130 /63.5 = 128

Step 3 find the mol of electrons that will deposit this amount of copper.

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

$n_e = 2 \times 128 = 256$

step 4 find the charge of 256 mol of electrons

$Q = 256 \times 96500 = 2.47 \times 10^7 C$

Step 5 find the time required to deliver this amount of charge

$Q = It$

$\Rightarrow 2.47 \times 10^7 C / 28 = t(sec)$

$\Rightarrow 8.83 \times 10^5$ seconds $= 8.83 \times 10^5 / (60 \times 60) = 250$ hours (2 sig fig)