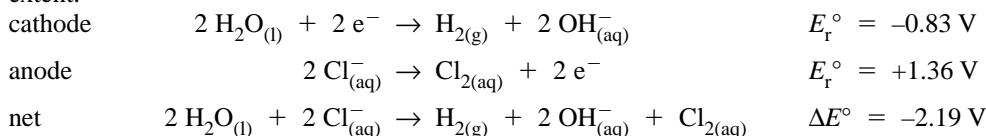**Evaluation**

(d) The experimental design is judged to be adequate with no obvious flaws. The problem was answered with reasonable certainty and the controls for the litmus test were adequate. There are two ways to improve the electrolysis of aqueous copper(II) sulfate. One is to use an apparatus so that the gas can be collected and identified. The other is to use pH paper or a pH meter instead of litmus paper. These changes would increase the certainty of the results.

(e) The overall quality of the evidence is good. There is some uncertainty about the results for the litmus test on the copper(II) sulfate solution. This could be solved by using pH paper or a pH meter. Overall, I am quite certain because the majority of the products were clearly identified with the diagnostic tests.

(f) The predictions for the cathode product of copper(II) sulfate and the cathode and anode products of the sodium sulfate solution were verified because the predicted products clearly agree with the experimental results. The prediction of the anode product for copper(II) sulfate could not be judged with sufficient certainty. The prediction for the cathode products of sodium chloride was also verified but the prediction for the anode products was falsified by the evidence obtained.

(g) The redox concepts and table remain acceptable until further results are obtained. Most of the predictions were verified except for the anode product of the sodium chloride electrolysis. It is necessary to determine how unique the result for sodium chloride is. It may be necessary to restrict or revise the concepts or procedure used to predict the products.

Synthesis

(h) The oxidation of water at the anode of the sodium chloride cell did not seem to occur, or occurred to a very limited extent.

cathode $\qquad$ $2\ H_2O_{(l)}\ +\ 2\ e^-\ \rightarrow\ H_{2(g)}\ +\ 2\ OH^-_{(aq)}$ $\qquad$ $E_r° = -0.83\ V$

anode $\qquad$ $2\ Cl^-_{(aq)}\ \rightarrow\ Cl_{2(aq)}\ +\ 2\ e^-$ $\qquad$ $E_r° = +1.36\ V$

net $\qquad$ $2\ H_2O_{(l)}\ +\ 2\ Cl^-_{(aq)}\ \rightarrow\ H_{2(g)}\ +\ 2\ OH^-_{(aq)}\ +\ Cl_{2(aq)}$ $\qquad$ $\Delta E° = -2.19\ V$

The minimum potential difference is +2.19 V.

(i) The minimum potential difference for oxygen as a product is 2.06 V and for chlorine as a product is 2.19 V. There is only 0.13 V difference between the two.

(j) The applied voltage in the electrolysis was 6 V, which is considerably higher than both minimum potential differences. Based on this fact and the fact that the minimum potential differences are so close, it is possible that the production of chlorine is much faster than that of oxygen. This could explain the larger quantity of chlorine produced.

(k) The rules should be retained until additional electrolyses of chloride solutions are conducted. If the same results appear, then the rules should be revised to account for this new information.

INVESTIGATION 10.3.1 INVESTIGATING AN ELECTROLYTIC CELL
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Prediction

(a) According to a modern reference, the value of the Faraday, F, is 9.65×10^4 C/mol.

Evidence

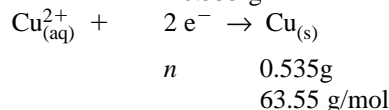|	Masses of copper electrodes	
Electrode	Initial mass (g)	Final mass (g)
labelled A	2.88	2.35
labelled C	2.74	3.28

• electric current decreased slightly near the start and was adjusted to a relatively constant 0.75 A

• elapsed time = 35 min 4 s

• cathode developed some dark brown crystals particularly at the edges

• some very thin, darker-coloured flakes fell off the anode when dipped in distilled water and acetone

• the final masses did not change when remeasured after a few minutes

Analysis

(b) for the cathode: $\qquad$ $\Delta m = 3.28\ g\ -\ 2.74\ g$
$\qquad\qquad\qquad\qquad = 0.54\ g$

for the anode: $\qquad$ $\Delta m = 2.88\ g\ -\ 2.35\ g$
$\qquad\qquad\qquad\qquad = 0.53\ g$

(c) average $\Delta m = (0.54 \text{ g} + 0.53 \text{ g})/2$
$\qquad\qquad\quad = 0.535 \text{ g}$

$$Cu^{2+}_{(aq)} + 2 e^- \rightarrow Cu_{(s)}$$
$$\qquad\qquad n \qquad\quad \dfrac{0.535 \text{g}}{63.55 \text{ g/mol}}$$

$n_{Cu} = 0.535 \text{ g} \times \dfrac{1 \text{ mol}}{63.55 \text{ g}}$

$n_{Cu} = 0.0084 \text{ mol}$

$n_{e^-} = 0.0084 \text{ mol} \times \dfrac{2}{1}$

$n_{e^-} = 0.017 \text{ mol}$

(d) $q = It$
$\qquad = 0.75 \dfrac{\text{C}}{\text{s}} \times 2104 \text{ s}$
$\quad q = 1.6 \times 10^3 \text{ C}$

(e) $F = \dfrac{1.6 \times 10^3 \text{ C}}{0.017 \text{ mol}}$

$F = 9.4 \times 10^4 \text{ C/mol}$

According to evidence collected and the copper(II) half-reaction, the value of the Faraday is 9.4×10^4 C/mol.

Evaluation

(f) The experimental design appears quite adequate to answer the question with no obvious flaws. It was important to control the electric current because it tended to change slightly during the electrolysis. The procedure was clear and produced sufficient evidence to answer the question. With the particular power supply used, it was not possible to get a current between one and two amperes. However, this is not a significant problem as long the current is known and controlled. A slight quantity of material was lost from the anode when removing the electrodes and rinsing. This could suggest that the copper foil used was not completely pure and some other substance was present but not oxidized. One improvement that could be made is to use a better quality of copper metal for the electrodes. Another improvement would be to run the cell for a longer period of time and/or at a higher current to obtain three significant digits for the change in mass. A more precise ammeter would also be required to improve the precision of the result.

(g) Considering the design, materials, and procedure used, I am fairly certain that the overall quality of the evidence is quite adequate. There do not appear to be any significant problems or sources of error other than the usual measurement uncertainties and the loss of a small amount of some material from the anode.

(h)
$\% \text{ difference} = \dfrac{|\ 9.4 \times 10^4 \text{ C/mol} - 9.65 \times 10^4 \text{ C/mol}\ |}{9.65 \times 10^4 \text{ C/mol}} \times 100\%$

$\% \text{ difference} = 2.9\%$

(i) The evidence suggests that copper is both oxidized and reduced in this cell.

at the cathode: $\qquad Cu^{2+}_{(aq)} + 2 e^- \rightarrow Cu_{(s)}$

at the anode: $\qquad Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2 e^-$
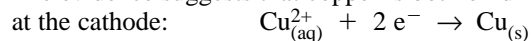
The change in masses of the electrodes is virtually identical and the small percent difference can easily be explained by experimental uncertainties. There is some evidence to suggest that the copper was not completely pure but this seems to be a minor problem.

Synthesis

(j) This type of cell is used in the electrorefining of copper to convert impure copper to copper of much higher purity.