Begin your response to **QUESTION 5** on this page.

Half-Reaction	$E°$ (V)
$Mg^{2+} + 2\,e^- \longrightarrow Mg$	-2.37
$Cl_2 + 2\,e^- \longrightarrow 2\,Cl^-$	$+1.36$

5. Molten $MgCl_2$ can be decomposed into its elements if a sufficient voltage is applied using inert electrodes. The products of the reaction are liquid Mg (at the cathode) and Cl_2 gas (at the anode). A simplified representation of the cell is shown above. The reduction half-reactions related to the overall reaction in the cell are given in the table.

(a) Draw an arrow on the diagram to show the direction of electron flow through the external circuit as the cell operates.

(b) Would an applied voltage of 2.0 V be sufficient for the reaction to occur? Support your claim with a calculation as part of your answer.
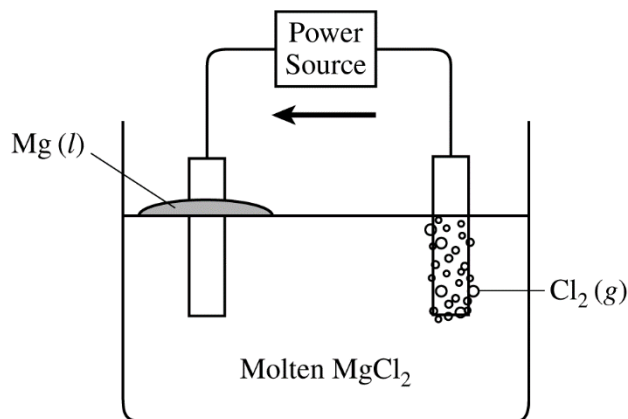
(c) If the current in the cell is kept at a constant 5.00 amps, how many seconds does it take to produce 2.00 g of $Mg(l)$ at the cathode?

GO ON TO THE NEXT PAGE.

Question 5: Short Answer   4 points

(a)	For the correct answer:	**1 point**

Electron flow should be indicated only in a counter-clockwise direction in the external circuit, from the Cl₂ anode to the Mg cathode.

(b)	For the correct answer and calculated value:	**1 point**

No, because 2.0 V is less than 3.73 V, which is the minimum voltage needed for electrolysis to occur.

$$E^{\circ}_{cell} = -2.37 \text{ V} + (-1.36 \text{ V}) = -3.73 \text{ V}$$

(c)	For the correct calculated value of moles of electrons (may be implicit):	**1 point**

$$2.00 \text{ g Mg} \times \frac{1 \text{ mol Mg}}{24.30 \text{ g Mg}} \times \frac{2 \text{ mol } e^-}{1 \text{ mol Mg}} = 0.165 \text{ mol } e^-$$

For the correct calculated number of seconds:	**1 point**

$$0.165 \text{ mol } e^- \times \frac{96,485 \text{ C}}{1 \text{ mol } e^-} \times \frac{1 \text{ s}}{5.00 \text{ C}} = 3180 \text{ s}$$

Total for part (c)	**2 points**

Total for question 5	**4 points**