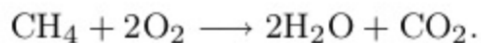
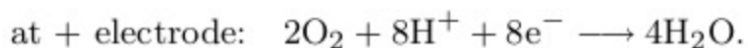
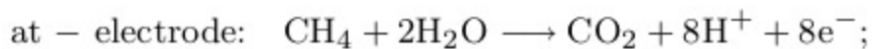


**Problem 5.5.** Consider a fuel cell that uses methane (“natural gas”) as fuel. The reaction is



- (a) Use the data at the back of this book to determine the values of  $\Delta H$  and  $\Delta G$  for this reaction, for one mole of methane. Assume that the reaction takes place at room temperature and atmospheric pressure, and that the water comes out in liquid form.
- (b) Assuming ideal performance, how much electrical work can you get out of the cell, for each mole of methane fuel?
- (c) How much waste heat is produced, for each mole of methane fuel?
- (d) The steps of this reaction are



What is the voltage of the cell?

## Methane fuel cell

$\text{CH}_4 + 2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{CO}_2$   
 assume 1 mole of methane, room temp, atmospheric pressure, all the final  $\text{H}_2\text{O}$  is liquid.

(a)  $\Delta H$  and  $\Delta G$  values

substance	$\Delta_f H$ (kJ)	$\Delta_f G$ (kJ)
$\text{CH}_4$	-74.81	-50.72
$\text{O}_2$	0	0
$\text{H}_2\text{O}(\text{l})$	-285.83	-237.13
$\text{CO}_2$	-393.51	-394.36

for the whole reaction

$$\begin{aligned}
 \Delta H &= \Delta H(\text{products}) - \Delta H(\text{reactants}) \\
 &= 2 \cdot (-285.83) + (-393.51) \\
 &\quad - (-74.81 + 2 \cdot 0) \quad \text{kJ} \\
 \Delta H &= -890.36 \text{ kJ}
 \end{aligned}$$

$$\begin{aligned}
 \Delta G &= 2 \cdot (-237.13) + (-394.36) \\
 &\quad - (-50.72 + 2 \cdot 0) \quad \text{kJ} \\
 \Delta G &= -817.90 \text{ kJ}
 \end{aligned}$$

(b) work? recall

$$\begin{aligned}
 \Delta G &= (Q - T\Delta S) + W_{\text{other}} \\
 \text{if } \Delta S &= Q/T, \text{ then (the best you can do)} \\
 W_{\text{other}} &= -817.90 \text{ kJ},
 \end{aligned}$$

i.e. you get 817.90 kJ of work out for each mole of methane used.

(c) waste heat? recall  $G = U - TS + pV$   
 $G = H - TS$

$\therefore \Delta G = \Delta H - T\Delta S$  at constant temperature

Assuming  $\Delta S$  is as small as possible so

$$Q = T\Delta S$$

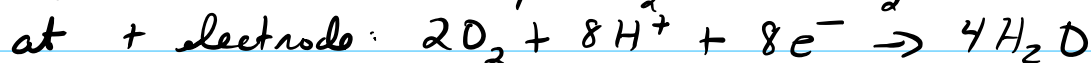
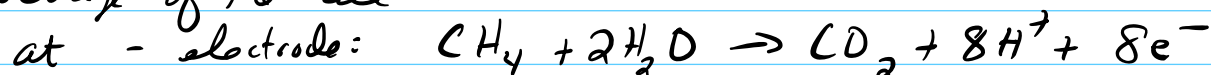
$$Q = \Delta H - \Delta G = -890.36 \text{ kJ} - (-817.90 \text{ kJ})$$

$$Q = -72.46 \text{ kJ}$$

$Q =$  heat added to the system

so  $-72.4 \text{ kJ} \Rightarrow$  that is heat leaving the system.

(d) Voltage of the cell?



(note net effect is



and 8 electrons get moved from - electrode to + electrode.

For one mole of methane

$$W_{\text{other}} = 817.90 \text{ kJ}$$

$$\Delta q = 8 \text{ moles} \times (1.602 \times 10^{-19} \text{ C})$$

$$\Delta V = \frac{W_{\text{other}}}{\Delta q} = \frac{817.90 \times 10^3 \text{ J}}{8 \cdot 6.02 \times 10^{23} \cdot 1.602 \times 10^{-19} \text{ C}} \times 1.02 \text{ J/C}$$

$$\Delta V = 1.02 \text{ Volts}$$