"Fuel Cell Efficiency" with the Solar Hydrogen Starter Kit

By Matt Kuhn. Adapted from "The Efficiency of a Fuel Cell" Heliocentris – Chemistry through Hydrogen

OBJECTIVES:

- Students will realize the efficiency benefits of fuel cells compared to internal combustion engines.
- Students will understand the concept of efficiency → the ratio of the output to the input of any system.
- Students will learn the basic mechanism of how a fuel cell produces electricity.
- Students will recognize how hydrogen can be produced by renewable means and thus a fuel cell can be used as a zero pollution energy source.
- Students will use data to calculate the energy efficiency of a fuel cell and apply this knowledge to societal energy issues.

MATERIALS:

1) 1 *Solar Hydrogen Starter Kit* with all components including:
   - 1 Solar module
   - 1 Electrolyzer
   - 1 Fuel cell
   - 1 Load measurement box or multimeter
   - 1 Set of connecting leads
   - 1 Stop watch
   - 2 Long tubes
   - 2 Short tubes
   - 2 Tubing stoppers
2) 1 Lamp (100-150 watt)
3) 1 Small bottle of distilled water

Available for $200-$300 at www.fuelcellstore.com or www.heliocentris.com

INTRODUCTORY DISCUSSION:

Review of Chemical Bonding

Hydrogen and Oxygen bond to form water (H_2O) because a single Hydrogen atom has 1 electron in its outer electron cloud, or energy level. It would prefer 2 electrons in order to be chemically stable. Oxygen has 6 outer electrons but prefers 8 in its outer energy level. Thus, 4 Hydrogen ($2H_2$) combine in an exothermic reaction with 2 Oxygen ($1O_2$) to form two covalently bonded water molecules. This reaction drives the fuel cell.

Review of Electricity

Electricity is the flow of electrons. Electrons flow when a build-up of electric charge on an anode causes electrons to be attracted to a positive electric charge at the cathode through a closed circuit. This electron flow can be converted into mechanical energy.

What is a fuel cell and how does it work?

- Electrochemical energy conversion device
- Fuel cells convert Hydrogen or Hydrogen-reformed fuels directly into electricity in a reaction with Oxygen
- They are quiet, compact, efficient, and clean
- The electrolyte membrane must stay hydrated at temperatures where water can exist as a liquid and vapor ($\approx 80^{o}$C)
- Reaction is too slow unless a Platinum catalyst is used
- A single fuel cell produces about 0.7 volts and 1 Amp/cm^2. Therefore, fuel cells are stacked together in series to increase the voltage by the number of cells.

Hydrogen is split into a proton and an electron with the aid of a Platinum catalyst. The protons pass through an electrolyte membrane while the electrons must go around the membrane through a circuit, doing work along the way. Finally, the protons and electrons recombine with Oxygen to form water and heat.

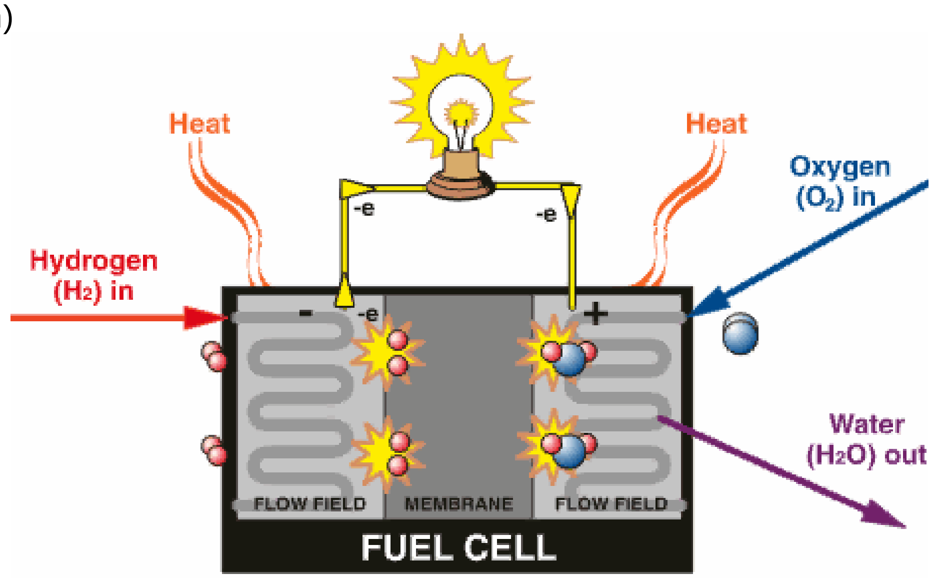
Fuel Cell Reactions

Anode (Hydrogen in):
$2H_2 \rightarrow 4H^+ + 4e^-$ (oxidation)

Cathode (Oxygen in):
$O_2 + 4e^- + 4H^+ \rightarrow 2H_2O_{(l)}$
(reduction)

Overall Reaction:
$2H_2 + 0_2 \rightarrow 2H_2O_{(l)}$ + Heat

?H = - 286 kJ/mole

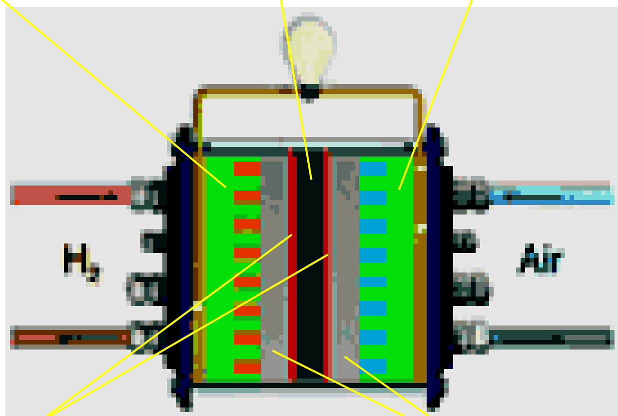


Parts of the Fuel Cell

Hydrated polymer electrolyte, also called the proton exchange membrane-PEM (wet plastic wrap)

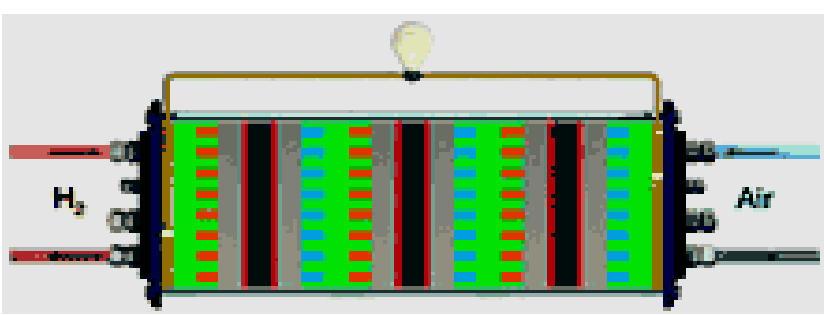
Anode (-)

Cathode (+)



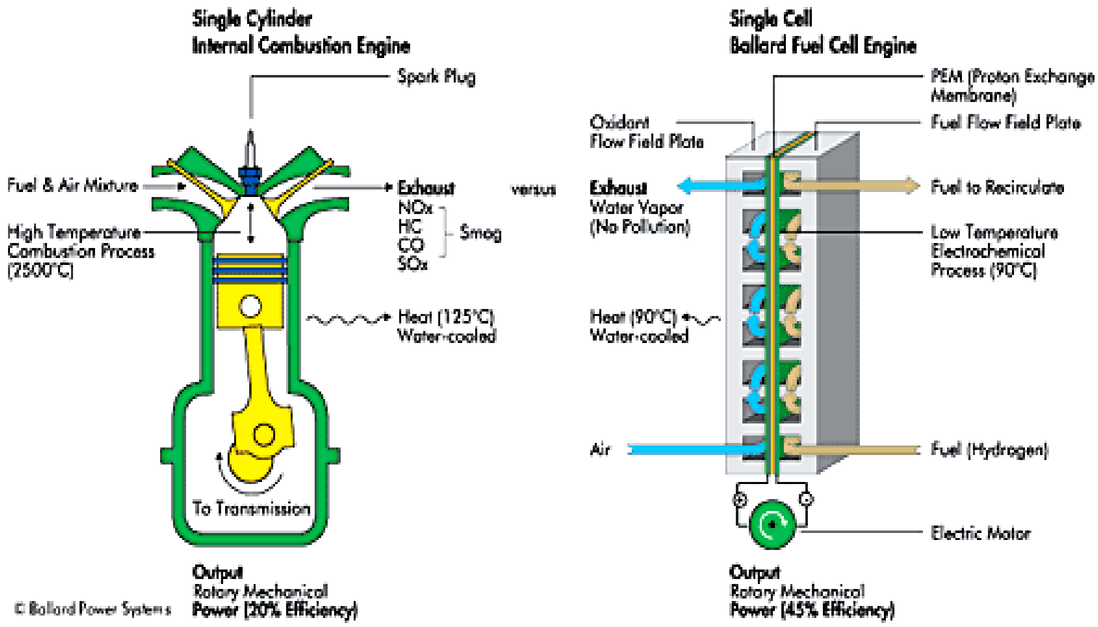
Porous catalyst layers with dispersed Platinum

Grooved Carbon-Teflon coated, gas diffusion layers



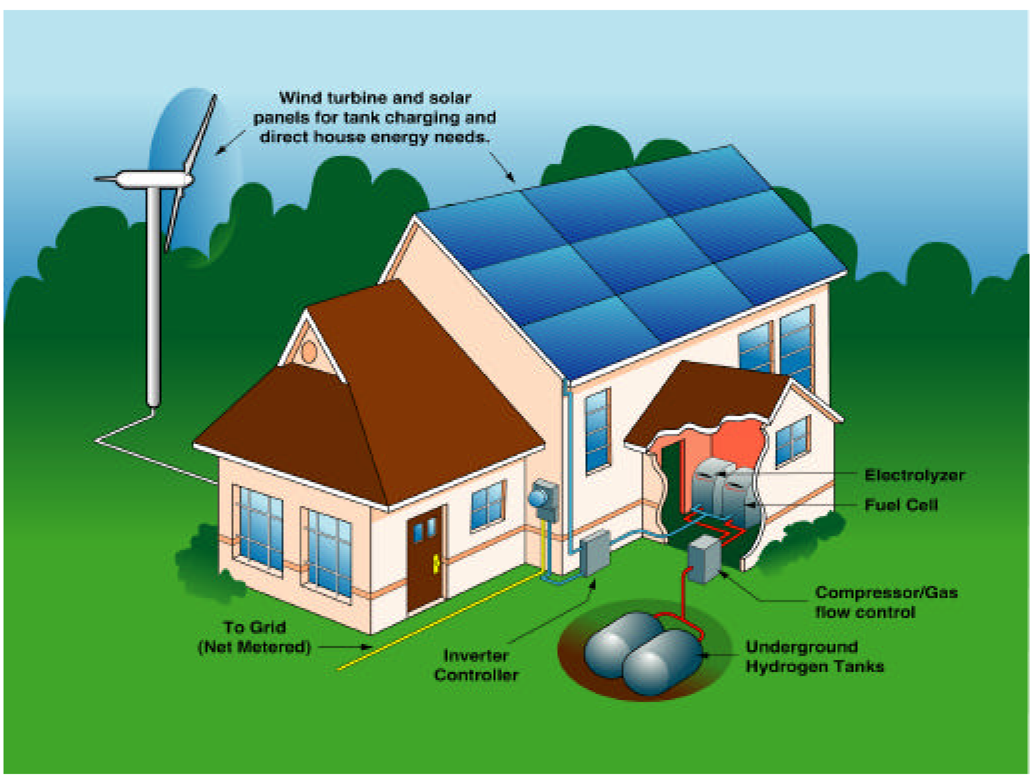
←Fuel Cell Stack

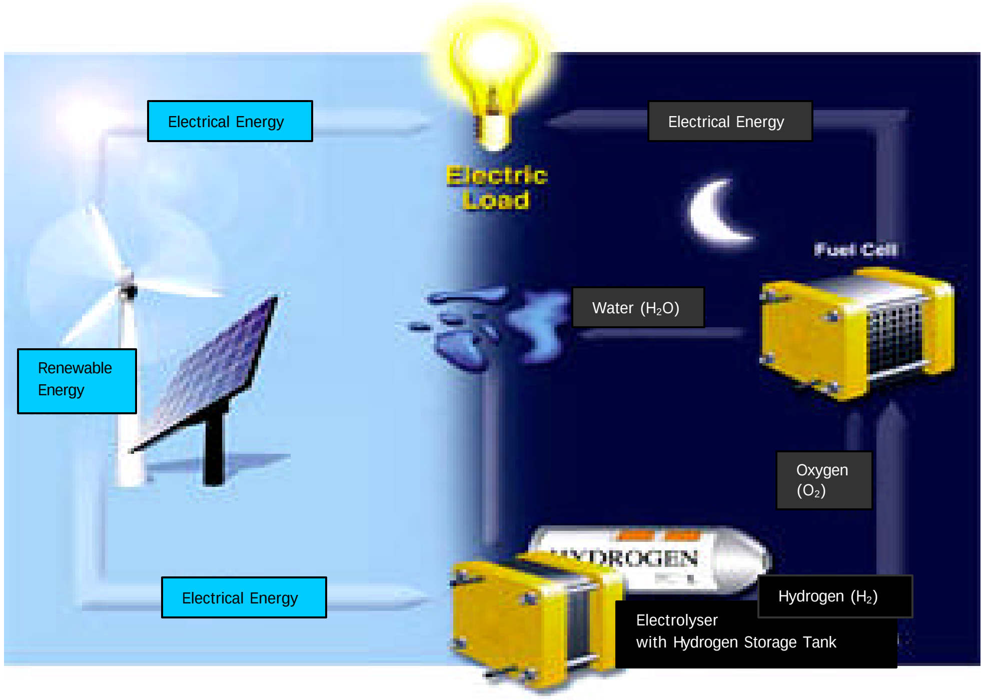
Efficiency comparison between fuel cells and internal combustion engines



Internal combustion = 20% compared to fuel cell = 45%

Fuel cells as a renewable energy storage device





From the Congressional Record 1875

"A new source of power... called gasoline has been produced by a Boston engineer. Instead of burning the fuel under a boiler, it is exploded inside the cylinder of an engine... The dangers are obvious. Stores of gasoline in the hands of people interested primarily in profit would constitute a fire and explosive hazard of the first rank. Horseless carriages propelled by gasoline might attain speeds of 14, or even 20 miles per hour. The menace to our people of this type hurtling through our streets and along our roads and poisoning the atmosphere would call for prompt legislative action even if the military and economic implications were not so overwhelming... the cost of producing (gasoline) is far beyond the financial capacity of private industry... In addition the development of this new power may displace the use of horses, which would wreck our agriculture."

The Experiment

Setting up the fuel cell and purging the system

1. Set up the apparatus as shown in *Diagram A*. As an alternative to solar power, you could also use a DC power supply if you want the electrolyser to fill the tanks more quickly. (**The DC voltage and current must not exceed 1.8 V and 3 A respectively. Check for correct polarity of the electrolyser)**

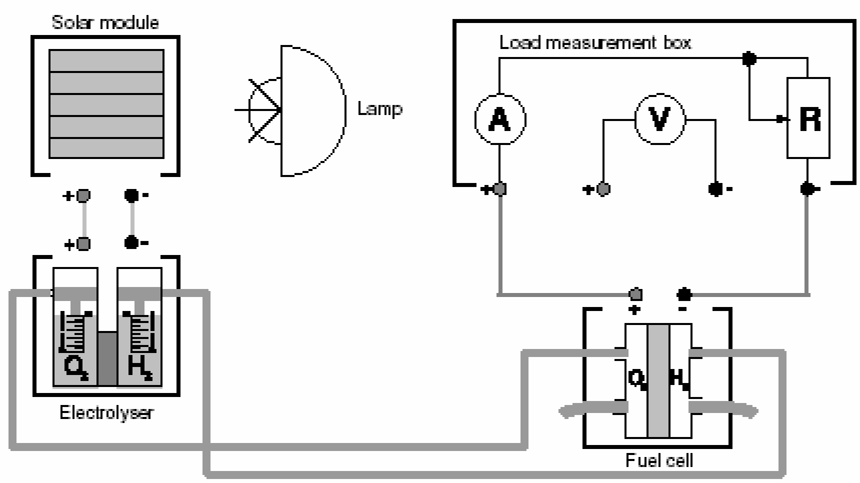


Diagram A

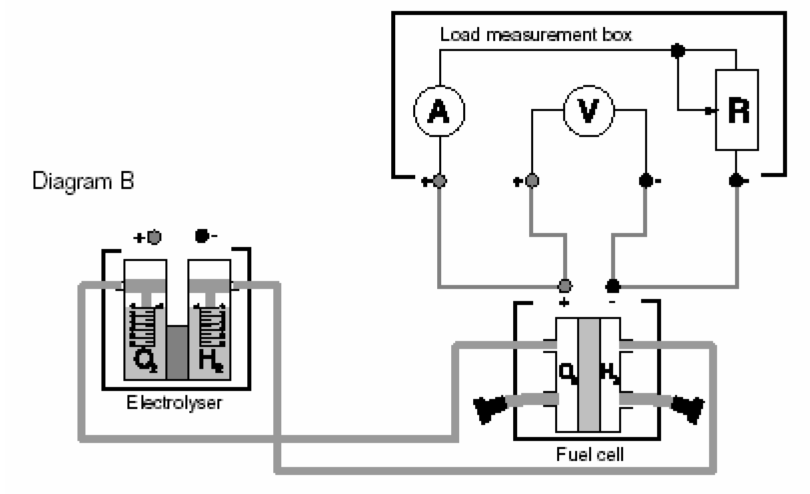
2. Check for correct tube connection between the electrolyser and fuel cell. Keep the outlet tubes running out of the fuel cell open. Adjust the load measurement box to "open" circuit.

3. Make sure that both of the electrolyser's gas storage tanks are filled with distilled water up to the 0 mL mark. Use the solar panel to provide a current of 700-900 mA to the electrolyser. You should be able to observe gas bubbles discharging oxygen and hydrogen into the tanks.

4. Run the complete system with the gases produced from the electrolyser for 10 minutes, thus purging the system of unwanted gases. After purging, set the load resistance to about 2 Ω and check the voltage and current. You should be getting about 0.6 – 0.9 V and 200 – 600 mA. If this is not the case, your fuel cell needs another purging.

5. Close both outlet tubes from the fuel cell with stoppers or clamps. Hydrogen and oxygen gases will now collect in the tanks. Disconnect the power supply to the electrolyser when 10 mL of hydrogen is stored.

6. Connect leads between the voltmeter and the fuel cell according to *Diagram B.*



7. With the load in the “open” circuit setting, and a disconnected electrolyser, calculate the hydrogen leakage rate by starting with a 10 mL full hydrogen tank and recording the level of hydrogen left after 180 seconds. Record this rate for later use.

8. Design a data table to record and measure the following during at least 2 trials:

- Initial Voltage in volts – V (${E_I}$)
- Final Voltage in volts – V (${E_F}$)
- Initial Current in amperes – A (${I_I}$)
- Final Current in amperes – A (${I_F}$)
- Initial Volume in milliliters – mL (Vol_I)
- Final Volume in milliliters – mL (Vol_F)
- Elapsed Time in seconds – s (set at 180 s)
- Load in ohms – Ω (set at 3 Ω)

9. Reconnect the electrolyser and refill the tanks to 10 mL. Then disconnect the electrolyser and adjust the load to 3 Ω. The fuel cell is now connected to a load and will consume the stored hydrogen. Record the time and immediately begin the first trial.

10. Run the fuel cell for at least two trials and record the measurements on your data table from step #8. After running the second trial, switch the load to “open” circuit and open the fuel cell outlet tubes.

Calculating the Faraday Efficiency and Energy Efficiency

You will need another data table with the following values calculated from the 1st table:

3 Ω for 180 s	**E** Average	**I** Average	**H_2** Consumed (Vol$_I$ – Vol$_F$ – Step 7 Leakage Rate)
1st Trial	V	A	mL
2nd Trial	V	A	mL
Trial Avg.	V	A	mL

The Faraday Efficiency

The Faraday Efficiency is a percentage that tell us how much of the hydrogen gas is being used for intended electrical energy production and how much is lost to other factors. The symbol for efficiency is the Greek letter, eta "h." Therefore, Faraday Efficiency's symbol is "h_F." It is the ratio between the theoretically calculated volume of hydrogen consumed by the load at a certain current *and* the experimentally calculated volume of hydrogen consumed. Hydrogen naturally exists in pairs, so we use "H_2" to represent it.

$$h_F = H_2\ Vol_{(th.)} \ / \ H_2\ Vol_{(exp.)}$$

If the fuel cell were perfect and all conditions were ideal, the Faraday Efficiency would be equal to "1" or 100%. Faraday's 2nd Law enables us to calculate the theoretical volume of hydrogen we would expect to be consumed.

$$H_2\ Vol_{(th.)} = \frac{\text{[Electrical charge in Coulombs (C)]}}{\text{[Electrical charge delivered by one mol } H_2\text{]}} \times \text{[}H_2\text{ Volume per mol]}$$

- [Electrical Charge in Coulombs (C)] = I × t → $I_{\text{Trials Average}} \times 180$ s = _____ C

- [H_2 Volume per mol] is a given value. At standard pressure and temperature, the volume of H_2 per mol of H_2 is <u>24,000 mL/mol</u>.

- $H_2 \rightarrow 2H^+ + 2e^-$, therefore, 1 mole of H_2 yields 2 moles of electrons (e^-). The electrical charge of one mole of electrons is given as 96,500 C. Since we have two moles of electrons,
  [Electrical charge delivered by one mol H_2] = 2 × 96,500 C = <u>193,000 C/mol</u>

The experimental volume of H_2 can be calculated from the experimental measurements:

$$H_2\ Vol_{(exp.)} = \text{Trials Average in milliliters}$$

Now we can calculate the Faraday Efficiency:

$$H_2\ Vol_{(th.)} = \frac{(I \times t)\ C}{193{,}000\ C/mol} \times 24{,}000\ mL/mol = \_\_\_\_\ mL$$

$$h_F = H_2\ Vol_{(th.)} \ /\ H_2\ Vol_{(exp.)} = \_\_\_\_\%$$

The Energy Efficiency

The energy efficiency of a fuel cell is a percentage that tells us how much energy from the combination of hydrogen and oxygen resulted in the intended electrical energy and how much energy was not converted to the intended electrical energy. It is a ratio between the electricity produced by the consumed hydrogen *and* the calculated theoretical energy contained in the consumed hydrogen.

$$h_E = \frac{[\text{Electrical energy in Joules } H_2]}{[\text{Energy content of consumed } H_2]_{\text{Theoretical}}}$$

- [Electrical energy in Joules H_2] = $E_{\text{Trail Average}} \times I_{\text{Trial Average}} \times$ time (s) = ____ J

- The energy content of hydrogen is calculated by using the amount of energy given off when hydrogen is combusted. This is a given value of 286,000 J/mol (exothermic) of hydrogen. This is called the *molar enthalpy* of hydrogen at standard temperature and pressure.

- We also have the given value for the amount of space H_2 takes up per mole. H_2 has a molar volume of 24,000 mL/mol. If we divide the average trial volume of hydrogen used by the molar volume, the millimeters will cancel out and we will have the number of moles of H_2 used.

$$\frac{\text{Trial Volume of } H_2 \text{ used in mL}}{24{,}000\ mL/mol} = \_\_\_\_\ mol\ H_2$$

- [Energy content of consumed H_2]$_{\text{Theoretical}}$ can now be calculated by multiplying the amount of energy in Joules/mol H_2 by the number of moles used.

____ mol $H_2 \times$ 286,000J/mol = ____ J = [Energy content of consumed H_2]$_{\text{Theoretical}}$

Now we can calculate the Energy Efficiency:

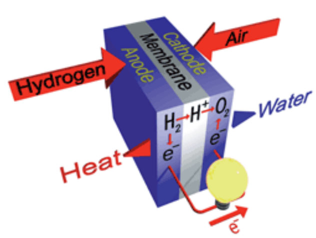
$$h_E = \frac{\text{[Electrical energy in Joules } H_2\text{]}}{\text{[Energy content of consumed } H_2\text{]}_{\text{Theoretical}}} = \_\_\_\_\%$$

Optional extended experimentation:

- You could vary the resistance to change the current and plot the results to find the characteristic curve of the fuel cell, and thus its optimal settings for peak efficiency.
- You could determine the maximum power output and its efficiency by running trials with graduated current, voltage, and resistance.

Questions to Ponder

1. How does the Faraday Efficiency compare to the Energy Efficiency?
2. How does the efficiency of fuel cells compare to the efficiency of the internal combustion engine?
3. Why doesn't the efficiency equal 100%? What happened to the uncaptured energy?
4. What are the two main emissions of fuel cells? What are the environmental benefits of fuel cell powered transportation?
5. What should come first, widespread use of fuels or widespread hydrogen fueling stations? What are the economic implications of building a hydrogen power infrastructure?
6. How is a fuel cell similar to a battery?
7. What are some current fuel cell applications in our society?
8. How does the volatility of hydrogen compare to common fossil fuels?
9. Compare the number of moving parts in an internal combustion engine to a fuel cell. How does this comparison affect their lifespan?
10. What are the price comparisons between automotive fuel cells and automotive internal combustions engines? What are the primary causes for the difference?
11. How does the storage of a liquid fuel compare to the storage of a gaseous fuel?
12. Why can a fuel cell be considered a renewable energy storage device?
13. Where does the United States get much of its oil? How would the widespread use of fuel cells affect this?
14. What are the political implications of imported oil compared to domestically produced hydrogen?



Teachers guide

The Faraday Efficiency of the fuel cell is less than one for the following reasons:

- Competing, simultaneous electrochemical reactions in the fuel cell, which supply fewer electrons for the same volume of hydrogen consumed.
- Chemical reactions between hydrogen and oxygen at the catalysts (catalytic oxidation/combustion).
- Hydrogen and oxygen recombination, or diffusion, by leakage through the electrolyte membrane. (Common in educational demonstration fuel cells.)

These websites will answer most questions you may have about fuel cells:
http://www.eere.energy.gov/hydrogenandfuelcells/
http://www.crest.org/hydrogen/index.html
http://www.fuelcells.org/

Sample Data and Calculations

Leakage rate calculated at 0.6 mL every 180 s

3 Ω for 180 s	**E**	**I**	**H_2**
1st Trial	E_I = 0.756 V	I_I = 0.210 A	Vol_I = 10 mL
	E_F = 0.754 V	I_F = 0.212 A	Vol_F = 3 mL
1st Trial Avg.	E_{AVG1} = 0.755 V	I_{AVG1} = 0.211 A	Vol_{I-F} – 0.6 = 6.4 mL
2nd Trial	E_I = 0.752 V	I_I = 0.210 A	Vol_I = 10 mL
	E_F = 0.746 V	I_F = 0.208 A	Vol_F = 3 mL
2nd Trial Avg.	E_{AVG2} = 0.749 V	I_{AVG2} = 0.209 A	Vol_{I-F} – 0.6 = 6.4 mL

3 Ω for 180 s	**E** Average	**I** Average	**H_2** Consumed (Vol$_I$ – Vol$_F$ – Step 7 Leakage Rate)
1st Trial	0.755 V	0.211 A	6.4 mL
2nd Trial	0.749 V	0.209 A	6.4 mL
Trial Avg.	0.752 V	0.210 A	6.4 mL

The Faraday Efficiency

$$h_F = H_2\ Vol_{(th.)} / H_2\ Vol_{(exp.)}$$

$$H_2\ Vol_{(th.)} = \frac{\text{[Electrical charge in Coulombs (C)]}}{\text{[Electrical charge delivered by one mol } H_2\text{]}} \times \text{[}H_2\text{ Volume per mol]}$$

- [Electrical Charge in Coulombs (C)] = I × t → 0.210 A × 180 s = 37.8 C

- [H_2 Volume per mol] is a given value. At standard pressure and temperature, the volume of H_2 per mol of H_2 is 24,000 mL/mol.

- $H_2 \rightarrow 2H^+ + 2e^-$, therefore, 1 mole of H_2 yields 2 moles of electrons (e^-). The electrical charge of one mole of electrons is given as 96,500 C. Since we have two moles of electrons,
  [Electrical charge delivered by one mol H_2] = 2 × 96,500 C = 193,000 C/mol

The experimental volume of H_2 can be calculated from the experimental measurements:

$$H_2\ Vol_{(exp.)} = \underline{6.4\text{ mL}}$$

Now we can calculate the Faraday Efficiency:

$$H_2\ Vol_{(th.)} = \frac{37.8\text{ C}}{193{,}000\text{ C/mol}} \times 24{,}000\text{ mL/mol} = \underline{4.7\text{ mL}}$$

$$h_F = H_2\ Vol_{(th.)} / H_2\ Vol_{(exp.)} = 4.7\text{ mL} / 6.4\text{ mL} = 0.73 \times 100\% = \underline{\textbf{73\% Efficient}}$$

The Energy Efficiency

$$h_E = \frac{[\text{Electrical energy in Joules } H_2]}{[\text{Energy content of consumed } H_2]_{\text{Theoretical}}}$$

- [Electrical energy in Joules H_2] = $E_{\text{Trails Average}} \times I_{\text{Trials Average}} \times$ elapsed time =
  0.752 V × 0.210 A × 180 s = <u>28.40 J</u>

- The energy content of hydrogen is calculated by using the amount of energy given off when hydrogen is combusted. This is a given value of <u>286,000 J/mol</u> (exothermic) of hydrogen. This is called the molar enthalpy of hydrogen at standard temperature and pressure. (Often given as –285.6kJ/mol, "–" means exothermic)

- We also have the given value for the amount of space H_2 takes up per mole. H_2 has a molar volume of <u>24,000 mL/mol</u> (rounded). If we divide the average trial volume of hydrogen used by the molar volume, the millimeters will cancel out and we will have the number of moles of H_2 used.

$$\frac{\text{Trial Volume of } H_2 \text{ used in mL}}{24{,}000 \text{ mL/mol}} = 6.4 \text{ mL} \div 24{,}000 \text{ mL/mol} \;\; \underline{0.00027} \text{ mol } H_2$$

- [Energy content of consumed H_2]$_{\text{Theoretical}}$ can now be calculated by multiplying the amount of energy in Joules/mol H_2 by the number of moles used.

0.00027 mol H_2 × 286,000J/mol = <u>76.27 J</u> = [Energy content of consumed H_2]$_{\text{Theoretical}}$

Now we can calculate the Energy Efficiency:

$$h_E = \frac{[\text{Electrical energy in Joules } H_2]}{[\text{Energy content of consumed } H_2]_{\text{Theoretical}}} = 28.4 \text{ J}/76.27 \text{ J} = 0.37 \times 100\% = \underline{\mathbf{37\%\ Efficient}}$$