Gasoline vs. Ethanol

Introduction

Controversy continues to swirl about the issue of changing from gasoline-fueled engines, to ethanol-powered ones (specifically, the current trend is toward E85- a mixture of 85% ethanol and 15% gasoline).

Some people contest the economics of the problem. The main economic issue is centered on whether or not it takes more energy to *make* a gallon of ethanol than the amount that ethanol produces in return. Some argue the relative merits of ethanol as a means of reducing carbon dioxide emissions. These arguments tend to become quite complex- especially when factors whose variables are quite unknown, are brought into the discussion. For example, regardless of the mechanisms for *producing* the carbon dioxide, the natural mechanisms for removal of this gas involve debatable concepts on *extraction* by plankton undergoing convective actions in the oceans. The natural cycles are just not known well enough for a deterministic evaluation of carbon dioxide extraction. Given more time though, research should eliminate the controversial areas.

This paper deals only with the relative production of energy and of carbon dioxide by these two fuels, under their most optimal conditions.

Production considerations alone, suggest that ethanol is not all that its proponents claim for it.

Summary of Findings

Assume two vehicles, each equipped with a **20 gallon** fuel tank, and each operating at **150 horsepower**. One uses iso-octane (a primary ingredient in gasoline) and the other uses ethanol.

1. *On a trip from point "A" to point "B" the gasoline engine used every drop of fuel and completed the trip in **6 hr. 16 min.***

   *On this same trip, the ethanol engine ran out of fuel in **4 hr. 25 min.** It would need a significant refueling to complete the trip.*

2. *For this trip from "A" to "B", the amount of carbon dioxide gas generated, was computed for the complete trip by both vehicles.*

   The amount of carbon dioxide produced, was ***almost identical*** *for the two vehicles.*

If you are concerned about a potential role of CO_2 in global warming, do not bother looking to ethanol as part of your solution. With a fully fueled ethanol-powered vehicle, you will not go as far as with an equivalent gasoline-powered vehicle... and when you do get to your destination, you will have spewed essentially the same amount of carbon dioxide into the air, regardless of your choice for the type of fuel.

Even if ethanol is considerably cheaper than gasoline, the larger amount needed, may offset any financial gain by using the ethanol fuel. As indicated, the economics of this is very fluid (pardon the pun) and so will not be covered in this paper.

Analysis - Energetics

Gasoline is a mixture of hydrocarbons, dominated by the presence of alkanes like hexane, heptane, octane, nonane, etc. It even has some cyclics like benzene and cyclohexane, and others. The preponderance of the alkanes with their very similar properties enables this calculation to proceed with the valid assumption that gasoline consists of only iso-octane.

I. Iso - Octane

The *complete* combustion of iso-octane is given by this reaction...

$$2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O \quad . \tag{1}$$

theoretical moles:	1	25/2	8	9
Actual moles:	(a)	(b)	(c)	(d)

where, (see Appendix - A) (a) = 459.45 moles

$$(b) = \frac{25}{2} \bullet 459.45 \rightarrow 5743.125 \text{ moles}$$

$$(c) = 8 \bullet 459.45 \rightarrow 3675.6 \text{ moles}$$

$$(d) = 9 \bullet 459.45 \rightarrow 4135.05 \text{ moles}$$

The Heat of Reaction in (1) is computed from the difference between the total Heats of Formation of the products, and the total Heats of Formation of the reactants...

$$\Delta H_{REACTION} = \sum_{PRODUCTS}^{i} (H_f) \bullet (\#moles)_i \ \text{-} \ \sum_{REACTANTS}^{j} (H_f) \bullet (\#moles)_j \qquad \textbf{(2)}$$

The following Heats of Formation are used for the combustion reaction in (1) above, and reaction (3), below (see Appendix A and Appendix - B),

iso-octane: $H_f = -224286 \dfrac{Joules}{mole}$

ethanol : $H_f = -277281 \dfrac{Joules}{mole}$

CO_2 gas : $H_f = -393601 \dfrac{Joules}{mole}$

H_2O gas : $H_f = -285720 \dfrac{Joules}{mole}$

Doing the arithmetic for (1) and (2), the Heat of Reaction for the combustion of our iso-octane is -2525.14 megaJoules (see Appendix - A). **The "minus sign" just means that the reaction is exothermic, as expected.**

II. Ethanol

The complete combustion of ethanol is given by this reaction...

$$2C_2H_5OH + 6O_2 \rightarrow 4CO_2 + 6H_2O \qquad \textbf{(3)}$$

theoretical moles:	1	3	2	3
Actual moles:	(a)	(b)	(c)	(d)

where, (see Appendix - B) (a) = 1298.75 moles

(b) = $3 \bullet 1298.75 \rightarrow 3896.25$ moles

(c) = $2 \bullet 1298.75 \rightarrow 2597.5$ moles

(d) = $3 \bullet 1298.75 \rightarrow 3896.25$ moles

Doing the arithmetic for (3) and (2), the Heat of Reaction for the combustion of our ethanol is -1775.50 megaJoules (see Appendix - B).

III. Time Until 20 Gallon Fuel Tank is Empty

A) Iso - octane

Assume a 150 HP engine and use the following definition and conversion of units...

$$Power\,(watts) = \frac{energy\,(Joules)}{Time\,(seconds)}\ ; \qquad \textbf{(4)}$$

1 HP = 745.70 watts

Substituting the power value (150 watts) and the energy production from the combustion of gasoline, the time until the 20 gallons of fuel is exhausted, is (see Appendix - C),

$$(\Delta T) = \text{6 hr. 16 min.} \quad \textbf{(gasoline)} \qquad \textbf{(5)}$$

B) Ethanol

Repeat the above calculation, but use the energy released from the Ethanol reaction. In this case, the time until the 20 gallons of that fuel is exhausted, is (see Appendix - C),

$$(\Delta T) = \text{4 hr. 25 min.} \quad \textbf{(ethanol)} \qquad \textbf{(6)}$$

Conclusion- Energetics

The ethanol-powered engine will run out of fuel before the gasoline-powered engine, even though both engines started with the same amount (i.e., 20 gallons) of their fuels.

This means that if you just made it to your destination from point "A" to point "B" using the gasoline, you will not make the destination on ethanol, unless you ***re-fuel*** the tank, en route to point "B."

Depending on the relative costs of the fuels at the time of the trip, you might not have much of a savings (if any at all) by using the ethanol.

Fuel me once... shame on me; Fuel me twice... shame on the ethanol proponents.

Analysis - Carbon Dioxide Production

Some people believe that carbon dioxide created by human activity plays a role in causing global warming. One of the claims for ethanol over gasoline is that the former fuel is supposed to produce less carbon dioxide than the latter, into the atmosphere.

Here, the results just obtained can be used to test that idea.

In going from point "A" to point "B" we used every drop of the 20 gallons of gasoline.

The "Actual" number of moles of carbon dioxide produced by combustion of the gasoline is 3675.6 moles.

Converting to pounds, the amount of carbon dioxide produced on the trip by *gasoline* combustion is...

\# moles = 3675.6 moles

$$\#Kg \;=\; (3675.6 moles)\bullet 0.044\frac{Kg}{mole} \;\rightarrow\; 161.73 \text{ kg of } CO_2$$

$$\#Lb \;=\; (161.73Kg)\bullet 2.205\frac{pounds}{Kg} \quad \text{--->} \quad \underline{\textbf{356.6 lb}} \text{ of } CO_2. \qquad \textbf{(7)}$$

Since the ethanol-fueled vehicle did *not* complete the trip from point "A" to point "B" it is necessary to compute the additional amount of ethanol needed to *complete* the same trip as the gasoline-fueled vehicle. Then, the pounds of CO_2 produced, can be computed on a trip-equivalent basis.

When the ethanol tank was depleted, 2597.5 moles of CO_2 are produced (see the lines below Eq.3). That number of moles of CO_2 is equivalent to 252.0 lb of CO_2...

$$\#\text{lb of } CO_2 \;=\; (2597.5 moles)\bullet(.044\frac{Kg}{mole})\bullet(2.205\frac{lb}{Kg}) \;\rightarrow\; 252.0 \text{ lb of } CO_2 \;.$$

The rate of production of CO_2 from the ethanol vehicle is, therefore...

$$\text{Rate of } CO_2 \text{ production} = \frac{252.0}{4.41}\frac{lb}{hr} \;\rightarrow\; 57.143\;\frac{lb}{hr}$$

For the ethanol vehicle to run as long as the Gasoline vehicle, the former would have to run an additional time, δT ...

δT = 6.27 - 4.41 hr

δT = 1.86 hr.

The additional CO_2 produced by the ethanol vehicle to complete the same trip, is then...

Additional CO_2 = $(57.143\frac{lb}{hr}) \bullet 1.86hr$

Additional CO_2 = 106.3 lb.

The *total* CO_2 produced on the trip by the ethanol engine is...

total CO_2 from Ethanol = 252.0 + 106.3 lb.,

total CO_2 from Ethanol = **358.3 lb** of CO_2 . **(8)**

Conclusion - Carbon dioxide Production

Comparing (7) and (8) see that *on a trip-equivalent basis*, the ethanol fuel actually exhausts slightly ***more*** CO_2 into the air than does the equivalent gasoline-powered engine.

That difference would all-but disappear, if one compared E85 to gasoline, since the major portion of the former consists of gasoline.

Appendix -A

The Actual moles of iso-octane (459.45) is determined from an assumed full (20 gallon) fuel tank:

20 gallons = 0.0757 m^3

#kg = volume * density

#kg = (0.0757 m^3)*(691.90 $\frac{kg}{m^3}$) → #kg = 52.3768 kg

Molecular weight of iso-octane:

MW = (8)*(12) + (18)*(1)

MW = 114 $\frac{gm}{mole}$ → MW = 0.114 $\frac{kg}{mole}$

Actual Moles = $\frac{52.3768\_kg}{0.114\_\frac{kg}{mole}}$ → Actual Moles = **<u>459.45</u>** moles

The Heat of Reaction for the gasoline reaction (1) is...

$\Delta H_{REACTION}$ =

$$(-393601\frac{J}{mole})\bullet(3675.6moles)+(-285720\frac{J}{mole})\bullet(4135.05moles)-(-224286\frac{J}{mole})\bullet(459.45moles)$$

$\Delta H_{REACTION}$ = $\underline{-2525.14}$_megajoules

Appendix - B

The actual moles of ethanol (1298.75) is determined from an assumed full (20 gallon) fuel tank:

20 gallons = 0.0757 m^3

#kg = volume * density

#kg = (0.0757 m^3)*(789.20 $\frac{kg}{m^3}$) → #kg = 59.7424 kg

Molecular weight of ethanol:

MW = (2)*(12) + (6)*(1) + (1)*(16)

MW = 46 $\frac{gm}{mole}$ → MW = 0.046 $\frac{kg}{mole}$

Actual Moles = $\frac{59.7424\_kg}{0.046\_\frac{kg}{mole}}$ → Actual Moles = **<u>1298.75</u>** moles

The Heat of Reaction for the ethanol reaction (3) is...

$$\Delta H_{REACTION} =$$

$$(-393601\frac{J}{mole})\bullet(2597.5moles)+(-285720\frac{J}{mole})\bullet(3896.25moles)-(-277281\frac{J}{mole})\bullet(1298.75moles)$$

$\Delta H_{REACTION}$ = $\underline{-1775.50}$ megajoules

Appendix - C

1 HP = 745.70 watts → 150 HP = 111855 watts

Using Eq.4 for the **gasoline**-powered engine,

$$Time(\sec onds)=\frac{energy(Joules)}{Power(watts)} \rightarrow \frac{2525.14\times10^6}{111855} \rightarrow$$ 22575.1 sec.

Time = 6.27 hr. i.e., 6 hr. 16 min.

Using Eq.4 for the **ethanol**-powered engine,

$$Time(\sec onds)=\frac{energy(Joules)}{Power(watts)} \rightarrow \frac{1775.50\times10^6}{111855} \rightarrow$$ 15873.2 sec.

Time = 4.41 hr. i.e., 4 hr. 25 min.