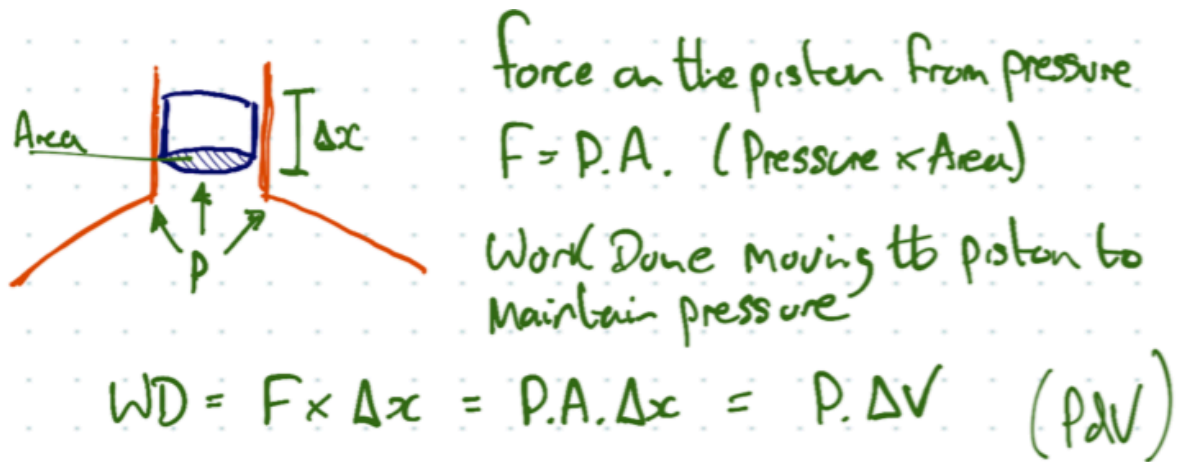


Application of the First Law

Dr Owen Williams – University of Wales TSD:

From last year's Powertrain and Control notes:



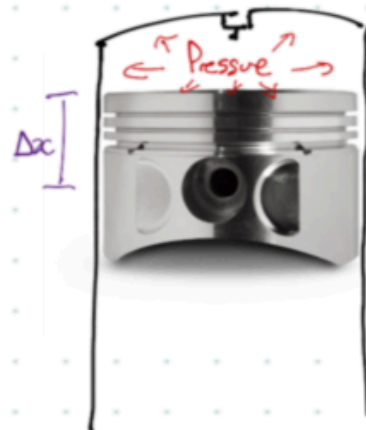
and

This is the same for a piston
in an engine

Force = Pressure $\times$ Area

$WD = \text{Pressure} \times \Delta V$

(important to
Remember!)



The first law is often formulated as:

$$dU = dQ - dW$$

It states that the change in the internal energy ΔU of a closed system is equal to the amount of heat Q supplied to the system, minus the amount of work W done by the system on its surroundings. We can of course write this as:

$$dQ = dU + dW$$

or

$$dQ = dU + PdV$$

In a piston engine, the flame is really thin, and the chemical burning is not fixed in a tight volume, so the flame has room to expand as it burns.

The heat energy of different things is measured at constant pressure in the laboratory. So... Enthalpy...

$$\begin{aligned}H &= U + PV \\dH &= d(U + PV) \\dH &= dU + PdV\end{aligned}$$

So at constant pressure, P

$$dH = dU + PdV = dQ$$

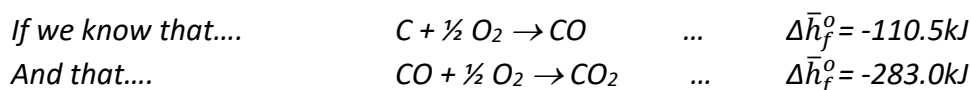
Meaning that at constant pressure, the heat in or out is equal to the change in enthalpy.

Enthalpy of Formation

Dr Owen Williams – University of Wales TSD:

One way to report the heat absorbed or released by chemical reactions would be to compile a massive set of reference tables that list the enthalpy changes for all possible chemical reactions, which would require an incredible amount of effort. Fortunately, Hess's law allows us to calculate the enthalpy change for virtually any conceivable chemical reaction using a relatively small set of tabulated data, starting from the elemental forms of each atom at 25 °C and 1 atm pressure.

Example:



Enthalpy of formation ($\Delta \bar{h}_f^{\circ}$) is the enthalpy change for the formation of 1 mol of a compound from its component elements, such as the formation of carbon dioxide from carbon and oxygen.

$$\begin{aligned}\Delta \bar{h}_{f, \text{Net}}^{\circ} &= \Delta \bar{h}_{f, 1}^{\circ} + \Delta \bar{h}_{f, 2}^{\circ} \\ \Delta \bar{h}_{f, \text{Net}}^{\circ} &= -110.5 \text{ kJ} + (-283.0 \text{ kJ}) \\ \Delta \bar{h}_{f, \text{Net}}^{\circ} &= -393.5 \text{ kJ}\end{aligned}$$

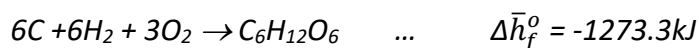
The magnitude of ΔH for a reaction depends on the physical states of the reactants and the products (gas, liquid, solid, or solution), the pressure of any gases present, and the temperature at which the reaction is carried out.

The standard conditions for which most thermochemical data are tabulated are a *pressure* of 1 atmosphere (atm) for all gases and a *concentration* of 1 M for all species in solution (1 mol/L). In addition, each pure substance must be in its standard state, which is usually its most stable form at a pressure of 1 atm at a specified temperature. We assume a temperature of 25°C (298 K) for all enthalpy changes given in this module, unless otherwise indicated.

Enthalpies of formation measured under these conditions are called **standard enthalpies of formation** ($\Delta \bar{h}_f^{\circ}$) which is the enthalpy change for the formation of 1 mol of a compound from its component elements when the component elements are each in their standard states. The standard enthalpy of formation of any element in its most stable form is zero by definition.

It is not possible to measure the value of $\Delta \bar{h}_f^{\circ}$ for glucose, -1273.3 kJ/mol , by simply mixing appropriate amounts of C, O₂, and H₂ and measuring the heat evolved as glucose is formed;

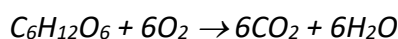
The reaction shown below does not occur at a measurable rate under any known conditions.



Glucose is not unique; most compounds cannot be prepared by the chemical equations that define their standard enthalpies of formation. Instead, values of $\bar{h}_f^\circ$ are obtained using Hess's law and standard enthalpy changes that have been measured for other reactions, such as combustion reactions.

Values of $\bar{h}_f^\circ$ for an extensive list of compounds are given in standard tables.

To demonstrate the use of tabulated $\bar{h}_f^\circ$ values, we will use them to calculate $\bar{h}_f^\circ$ for the combustion of glucose, the reaction that provides energy for your brain:



We can write:

$$\Delta H_f = \{6 \bar{h}_f^\circ [\text{CO}_2] + 6 \bar{h}_f^\circ [\text{H}_2\text{O}]\} - \{ \bar{h}_f^\circ [\text{C}_6\text{H}_{12}\text{O}_6] + 6 \bar{h}_f^\circ [\text{O}_2] \}$$

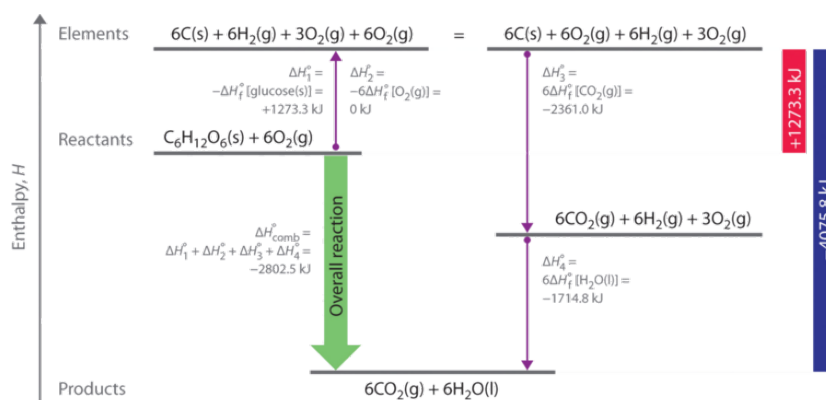
From tables, the relevant $\bar{h}_f^\circ$ values are $\bar{h}_f^\circ [\text{CO}_2] = -393.5 \text{ kJ/mol}$, $\bar{h}_f^\circ [\text{H}_2\text{O}] = -285.8 \text{ kJ/mol}$, and $\bar{h}_f^\circ [\text{C}_6\text{H}_{12}\text{O}_6] = -1273.3 \text{ kJ/mol}$.

Because O_2 is a pure element in its standard state, $\Delta \bar{h}_f^\circ [\text{O}_2] = 0 \text{ kJ/mol}$.

Inserting these values into the equation above and changing the subscript to indicate that this is a combustion reaction, we obtain

$$\begin{aligned} \Delta \bar{h}_{f, \text{comb}}^\circ &= [6(-393.5 \text{ kJ/mol}) + 6(-285.8 \text{ kJ/mol})] - [-1273.3 + 6(0 \text{ kJ/mol})] \\ \Delta \bar{h}_{f, \text{comb}}^\circ &= -2802.5 \text{ kJ/mol} \end{aligned}$$

The figure below shows two pathways from reactants (middle left) to products (bottom). The more direct pathway is the downward green arrow labelled $\Delta \bar{h}_{f, \text{comb}}^\circ$. The alternative hypothetical pathway consists of **four separate reactions** that convert the reactants to the elements in their standard states (upward purple arrow at left) and then convert the elements into the desired products (downward purple arrows at right).

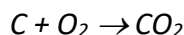


Our 4 Constituent Elements

Dr Owen Williams – University of Wales TSD:

For the combustion chemistry that we are interested in, the 4 main building blocks that we need to consider are O_2 , N_2 , C and H_2 . All of the other molecules are relative to these because the other enthalpies of other compounds are relative to them.

4 constituent elements



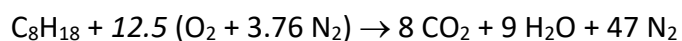
This is 1 kmol of C , 1 kmol of O_2 which becomes 1 kmol of CO_2

Enthalpy out from the reaction is the same as the heat out at constant pressure = 393.5MJ

Example values of 'Enthalpies of Combustion' can be seen in the table below.

Substance	Combustion Reaction	Enthalpy of Combustion $\Delta H_c^\circ \left(\frac{\text{kJ}}{\text{mol}} \text{ at } 25^\circ\text{C} \right)$
carbon	$C(s) + O_2(g) \rightarrow CO_2(g)$	-393.5
hydrogen	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	-285.8
magnesium	$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$	-601.6
sulfur	$S(s) + O_2(g) \rightarrow SO_2(g)$	-296.8
carbon monoxide	$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$	-283.0
methane	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	-890.8
acetylene	$C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l)$	-1301.1
ethanol	$C_2H_5OH(l) + 3O_2(g) \rightarrow CO_2(g) + 3H_2O(l)$	-1366.8
methanol	$CH_3OH(l) + \frac{3}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	-726.1
isooctane	$C_8H_{18}(l) + \frac{25}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l)$	-5461

Let's look at iso-octane...



Using standard tables and fuels data we should be able to do some calculations for some of the more complex fuels.

Shown below is a table detailing the Properties of Fuels for a range of fuels.

Formula	Fuel	M (kg/kmol)	T _b (°C)	c _{pg} ^a (kJ/kg-K)	T _{ig} (°C)	HHV (MJ/kg)	LHV (MJ/kg)	h _{fg} ^b (kJ/kg)	AFRs	T _f (K) ^c	Δh _f ⁰ (kJ/mol)	RON ^d	MON ^e
CH ₄	Methane	16.04	-161	2.21	537	55.536	50.048	510	17.2	2,226	-74.4	120	120
C ₂ H ₂	Acetylene	26.04	-84	1.60	305	49.923	48.225	-	13.2	2,540	8.7	50	50
C ₂ H ₄	Ethylene	28.05	-104	1.54	490	50.312	47.132	-	14.7	2,380	52.4	-	-
C ₂ H ₆	Ethane	30.07	-89	1.75	472	51.902	47.611	489	16.1	2,370	-83.8	115	99
C ₃ H ₈	Propane	44.10	-42	1.62	470	50.322	46.330	432	15.7	2,334	-104.7	112	97
C ₄ H ₁₀	n-Butane	58.12	-0.5	1.64	365	49.511	45.725	386	15.5	2,270	-146.6	94	90
C ₄ H ₁₀	iso-Butane	58.12	-12	1.62	460	49.363	45.577	366	15.5	2,310	-153.5	102	98
C ₅ H ₁₂	n-Pentane	72.15	36	1.62	284	49.003	45.343	357	15.3	2,270	-173.5	62	63
C ₅ H ₁₂	iso-Pentane	72.15	28	1.60	420	48.909	45.249	342	15.3	2,310	-178.5	93	90
C ₆ H ₁₄	n-Hexane	86.18	69	1.62	233	48.674	45.099	335	15.2	2,271	-198.7	25	26
C ₆ H ₁₄	iso-Hexane	86.18	50	1.58	421	48.454	44.879	305	15.2	2,300	-207.4	104	94
C ₇ H ₁₆	n-Heptane	100.20	99	1.61	215	48.438	44.925	317	15.2	2,273	-224.2	0	0
C ₈ H ₁₈	n-Octane	114.23	126	1.61	206	48.254	44.786	301	15.1	2,275	-250.1	20	17
C ₈ H ₁₈	iso-Octane	114.23	114	1.59	418	48.119	44.651	283	15.1	2,300	-259.2	100	100
C ₉ H ₂₀	n-Nonane	128.6	151	1.61	-	48.119	44.688	295	15.1	2,274	-274.7	-	-
C ₁₀ H ₂₂	n-Decane	142.28	174	1.61	210	48.002	44.599	277	15.1	2,278	-300.9	-41	-38
C ₁₀ H ₂₂	iso-Decane	142.28	171	1.61	-	48.565	44.413	-	15.1	2,340	-	113	92
C ₁₂ H ₂₆	n-Dodecane	170.33	216	1.61	204	47.838	44.574	256	15.0	2,276	-350.9	-88	-90
CH ₄ O	Methanol	32.04	65	1.37	385	22.663	19.915	1,099	6.5	2,183	-201.5	106	92
C ₂ H ₆ O	Ethanol	46.07	78	1.42	365	29.668	26.803	836	9.0	2,144	-235.1	107	89
H ₂	Hydrogen	2.02	-253	14.47	400	141.72	119.96	451	34.3	2,345	0	-	-

^aGas phase specific heat evaluated at 25°C

^bHeat of vaporization at 1 atm

^cEstimated equilibrium flame temperature

^dResearch octane number

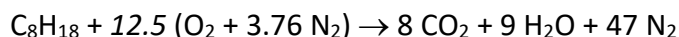
^eMotor octane number

From this table, we can see that Δh_f⁰ for iso-octane is -259.2 kJ/mol (or MJ/kmol)

We also know that Δh_f⁰ CO₂ is -393.5 kJ/mol (see table above)

We also need water, which is ΔH = -241.8 kJ/mol as a gas (g)

and Δh_f⁰=285.8 kJ/mol as a liquid (l)



Enthalpy needed for products

$$= 259.2 + 0 + 0 \text{ MJ}$$

$$= 259.2 \text{ MJ}$$

Enthalpy released from reaction

$$= (8 \times 393.5) + (9 \times 241.8) + (47 \times 0)$$

$$= 5324.2 \text{ MJ}$$

The net enthalpy is the difference between what you got out and what you had to put in...

$$5324.2 - 259.2 = 5065 \text{ MJ}$$

So burning 1 kmol of C₈H₁₈ at constant pressure releases 5065 MJ of heat energy (enthalpy)

The 'Lower Heating Value' is the term used when H₂O is a gas (LHV)

The mass of 1 kmol of C₈H₁₈ = (8 x 12) + (18 x 1) = 114 kg of iso-octane.

Which releases 5065 MJ of heat at constant pressure.

$$\text{So 1kg of iso-octane releases } \frac{5065}{114} = 44.4 \frac{\text{MJ}}{\text{kg}}$$

You can see that this gives you the calorific value you will be familiar with, 44x10⁶ J/kg

The 'Lower Heating Value' is the term used when H₂O is a gas (LHV)

When the H₂O in the products is a liquid, we get $= (8 \times 393.5) + (9 \times \mathbf{285.8}) + (47 \times 0)$
 $= 5720.2 \text{ MJ}$

And this gives us a new enthalpy of $5720.2 - 259.2 = 5461 \text{ MJ}$

Giving a 'Higher Heating Value' of $\text{HHV} = 5461/114 = 47.9 \text{ MJ/kg}$

So as a result, we can determine the higher and lower heating values for iso-octane and compare with the table above:

$$\mathbf{LHV = 44.4 \times 10^6 \text{ J/kg}}$$

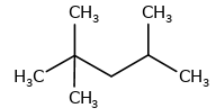
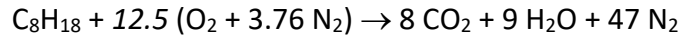
$$\mathbf{HHV = 47.9 \times 10^6 \text{ J/kg}}$$

More about Fuels

Dr Owen Williams – University of Wales TSD:



Iso-octane:



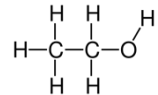
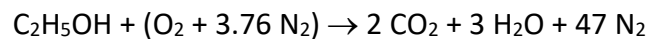
Enthalpy needed for products
= 259.2 + 0 + 0 MJ
= 259.2 MJ

Enthalpy released from reaction
= (8 x 393.5) + (9 x 241.8) + (47 x 0)
= 5324.2 MJ

Remember that this gives us the calorific value we are familiar with, 44×10^6 J/kg

Let's have a look at another fuel.

Ethanol:



Enthalpy needed for products
= 234.6 + 0 + 0 MJ
= 234.6 MJ

Enthalpy released from reaction
= (2 x 393.5) + (3 x 241.8) + (11.28 x 0)
= 1512.4 MJ

The mass of 1 kmol of $\text{C}_2\text{H}_5\text{OH}$ = (2 x 12) + (5 x 1) + (16) + (1) = 46 kg of ethanol.
Which releases 1277.8 MJ of heat at constant pressure.

$$\text{So 1kg of ethanol releases } \frac{1277.8}{46} = 27.78 \frac{\text{MJ}}{\text{kg}}$$

So, this gives you the calorific value of 27.8×10^6 J/kg

Can we put Ethanol in the BMW S1000RR in the workshop?

This engine has roughly 250cc displacement / cylinder
Assuming the volumetric efficiency to be 100%
So, 100% of the air/fuel mixture enters the engine.

If you are going to use ethanol, you'd need more
fuel per air (AFR).

Therefore, more rich fuel mixture compared to gasoline
(iso-octane)



More about Physics

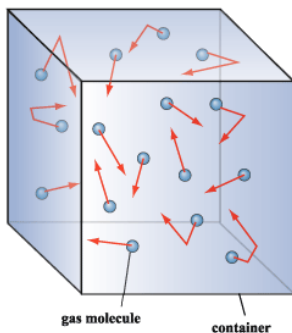
Dr Owen Williams – University of Wales TSD:

All gases follow the ideal gas law

$$P V = n R T$$

The average kinetic energy of **any** molecule in a gas = $\frac{1}{2} k_B T$

$$k_B = 1.38 \times 10^{-23} \text{ J/K}$$



Assuming the container has a 'hard wall' such that the molecules bounce back at the same speed. (no energy loss). Consider a molecule hitting the right hand (RH) wall.

Before the impact, the momentum to the right = mv_x is +ve
After the impact, the momentum to the left = mv_x is -ve
(same magnitude, different direction)

$$\text{Change in momentum per impact on RH wall} = mv_x - (-mv_x) = 2mv_x$$

$$\text{The number of times that the molecule hits the RH wall per second} = \frac{v_x}{2L_x}$$

The change in momentum per second per molecule on the RH wall

$$\text{change in momentum per second per molecule} = 2mv_x \times \frac{v_x}{2L_x} = \frac{mv_x^2}{L_x}$$

Supposing that the number of molecules in the box is η .

$$\text{change in momentum per second} = \eta \times \frac{mv_x^2}{L_x}$$

Now, let's go back and have a look at Newton's law a moment...

Force = Mass x Acceleration = mass x velocity per second = **momentum per second**

$$\text{Force} = \eta \times \frac{mv_x^2}{L_x}$$

$$\text{Pressure} = \frac{F}{A}$$

$$Pressure = \frac{F}{L_y L_z}$$

$$Pressure = \frac{\eta \times mv_x^2}{L_x L_y L_z}$$

$$Pressure = \frac{\eta \times mv_x^2}{Volume}$$

$$PV = \eta \times mv_x^2$$

Now remember, Boltzmann said all molecules have the same KE

$$\frac{1}{2}mv_x^2 = \frac{1}{2}k_B T$$

$$mv_x^2 = k_B T$$

$$\boxed{PV = \eta k_B T}$$

There are 6.02214×10^{26} molecules in a kmol so

$$\eta \times 6.02214^{26} = n \text{ kmols}$$

$$PV = n \times 6.02214 \times 10^{26} \times k_B T$$

$$PV = n \times 6.02214 \times 10^{26} \times 1.3806 \times 10^{-23} T$$

$$PV = n \times 8314 \times T$$

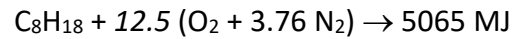
$$\boxed{PV = n R_U T}$$

(since $R_U = 8314 \text{ J/kmol}$)

Therefore, ALL gases obey the ideal gas law

$$PV = nRT \text{ (n in kmoles)}$$

So once again...



For our 250cc cylinder at 100% volumetric efficiency
Assuming $P = 1 \text{ Bar}$, $t = 20^\circ\text{C}$, $R_U = 8314$ are fixed, then...

$$V \propto n$$

Thus, if the volume is simply the displacement at a fixed pressure and temperature, then the number of kmoles is fixed.

The number of kmoles of mixture in port injection is the same for all fuels

$$\text{Energy released per kmol of gasoline} = \frac{5065}{1+12.5(1+3.76)} = \frac{5065}{60.5} = 83.72 \text{ MJ/kmol}$$

$$\text{Energy released per kmol of ethanol} = \frac{1277.8}{1+3(1+3.76)} = \frac{1277.8}{15.28} = 83.6 \text{ MJ/kmol}$$

$$\text{Energy released per kmol of methanol} = \frac{675.9}{1+1.5(1+3.76)} = \frac{675.9}{8.14} = 83.03 \text{ MJ/kmol}$$

FUEL	Port Injection	Energy released
	Energy Released / kmol mix	Per kmol of Air
Gasoline	83.72 MJ	85.12 MJ
Ethanol	83.6 MJ	89.48 MJ
Methanol	83.03 MJ	94.66 MJ

Since the volume is proportional to the number of moles, we can say...

$$\frac{\text{volume of methanol}}{\text{total volume}} = \frac{1}{1 + 1.5(1 + 3.76)} = 0.122 = 12\%$$

$$\frac{\text{volume of ethanol}}{\text{total volume}} = \frac{1}{1 + 3(1 + 3.76)} = 0.065 = 6.5\%$$

$$\frac{\text{volume of gasoline}}{\text{total volume}} = \frac{1}{1 + 12.5(1 + 3.76)} = 0.016 = 1.6\%$$

This basically means that you can't get more air in since methanol takes up 12% of the volume, while gasoline only takes up 1.6%.

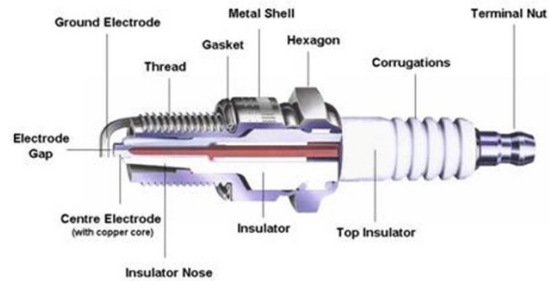
Total mass = Air mass + Fuel Mass

$$\begin{aligned} &= x (\text{O}_2 + 3.76 \text{ N}_2) + \text{C}_8\text{H}_{18} + \text{CH}_3\text{OH} \\ &= 14 (2 \times 16 + 3.76 \times 2 \times 14) + (8 \times 12 + 18) + (12 + 3 + 16 + 1) \\ &= 1921.92 + 114 + 32 \end{aligned}$$

$$\begin{aligned} \text{So } \text{C}_8\text{H}_{18} &= (114/1921.9) = 5.9\% \\ \text{and } \text{CH}_3\text{OH} &= (32/1921.9) = 1.6\% \end{aligned}$$

Ideal Fuel and Combustion

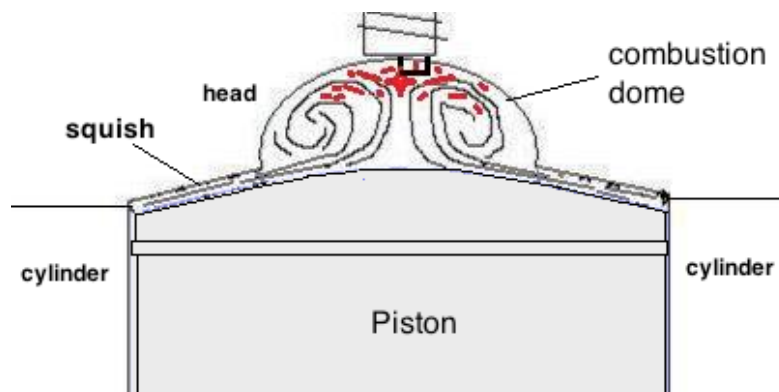
Dr Owen Williams – University of Wales TSD:



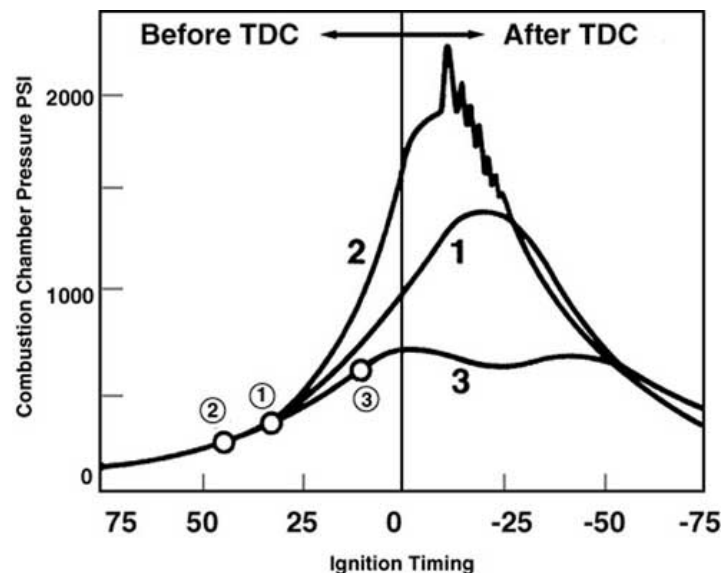
The ignition system creates a voltage of about 30kV which results in a spark being generated across the spark gap. There is an ignition delay between the spark occurring and the visible rise in pressure as the flame kernel builds up and a flame leaves the spark plug electrodes.



The flame front propagates through the combustion chamber volume creating a 'burned' and an 'un-burned' zone (we will model this later as a 'two zone' model). As the flame front propagates and the piston reaches the TDC position. The squish area creates turbulence as the gap closes and this resulting turbulence encourages combustion.



The best ignition point for maximum brake torque is called MBT timing (MBT = maximum brake torque). Generally, this is found by trial and error on the dynamometer. Advancing the ignition builds pressure earlier and therefore the combustion starts earlier.



Advancing before MBT timing increases the negative work occurring BTDC and therefore the Brake Torque drops. Retarding the ignition beyond MBT timing reduces the negative work before BTDC, but it reduces the positive pressure after TDC because it is burning in a bigger volume and again the Brake Torque drops.

Slow burning fuel takes a longer crank angle duration to burn in and thus it needs to be ignited earlier to have its MBT earlier, otherwise it burns in too big a volume and then lacks pressure on the downstroke.

The typical burn duration at full load is approximately 50° - 60°

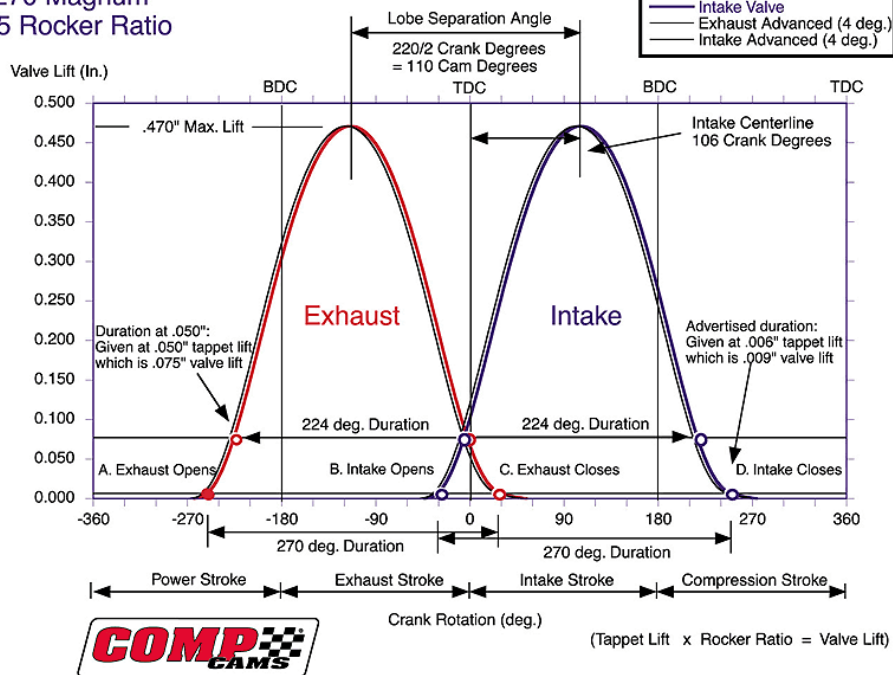
The downsides of slow burning include:

1. More negative work before TDC
2. There is more time at high combustion temperatures (of over 2000K) for heat loss into the oil and water.
3. Even at MBT timing the end of combustion is later. You get less work out during expansion and then hotter exhaust gasses and higher temperatures in the exhaust/head.

You need an expansion ratio after the end of combustion and before the exhaust valve opens so that the pressure can do positive work on the piston and while it is doing this the temperature of the gas is dropping, before the exhaust valve opens.

For the first 20° after the exhaust valve begins to open, the gas comes out at the speed of sound and results in choked flow.

270 Magnum 1.5 Rocker Ratio



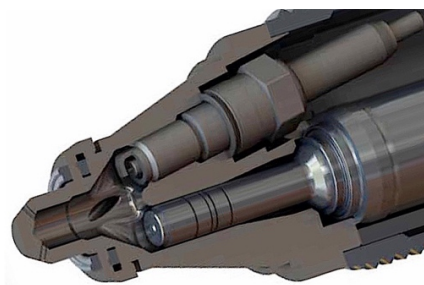
F1 Combustion

Dr Owen Williams – University of Wales TSD:

During the compression stroke, around 97 percent of the fuel charge is injected directly into the cylinder and the remaining three percent of fuel is directed into the chamber where a traditional spark plug ignites that small, overly rich charge within the tiny confines of the chamber. The resulting high-pressure jet stream of hot gasses are forced through a number of tiny holes in the chamber and into the main cylinder to ignite the rest of the air-fuel charge, which is by now thoroughly mixed and lean.



At first, observers surmised that F1 engine builders had solved all the mysteries and problems of homogeneous charge compression ignition (HCCI), or at least were using a form of HCCI along with spark ignition. HCCI is somewhat the holy grail of combustion for an ICE, since it theoretically leverages the best benefits of both spark and compression engines. In its simplest form, an HCCI engine draws an air-fuel mixture into the cylinder, similar to a gas spark engine, then then ignites the mixture using compression, similar to a diesel engine. There is no spark or fuel-delivery timing, but there's a much lower temperature combustion. That means lower NOx emissions and higher heat efficiency, which cuts down on CO2 emissions. The only problem is that there are no timing controls. Ignition is solely at the mercy of the laws of chemistry and physics, so high-speed and high-load conditions where cylinder temperatures rise can be a quite unpredictable.



<https://www.mahle-powertrain.com/media/mahle-powertrain/experience/mahle-jet-ignition/mahle-jet-ignition-product-information.pdf>

<https://www.mahle-powertrain.com/en/experience/mahle-jet-ignition/>

Generally, the process uses a separate chamber housing both the injector and the spark plug. The small chamber uses a rich (or possibly stoichiometric) mixture, where as the main chamber is lean but hot at TDC from an 18:1 compression ratio. After the spark occurs in the small chamber you get jets of flame or radicals (In chemistry, a **radical** is an atom, molecule, or ion that has an unpaired valence electron. With some exceptions, these unpaired electrons make radicals highly chemically reactive).

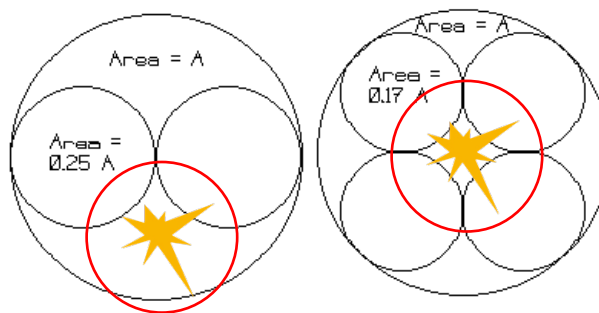
Combustion Processes

Dr Owen Williams – University of Wales TSD:

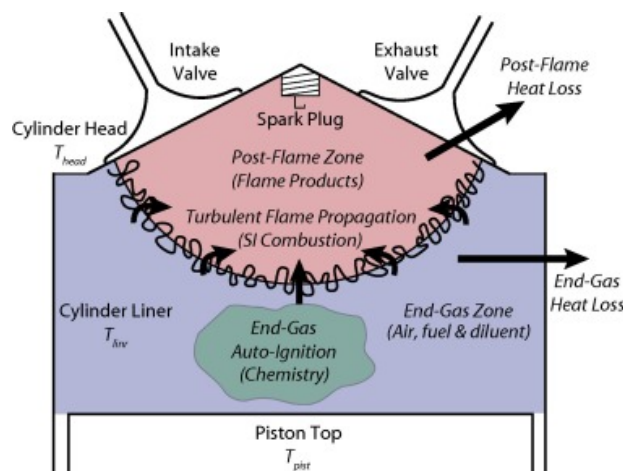
Faster burning fuel needs a later ignition (nearer TDC) for MBT timing and therefore less negative work. The pressure rises faster, and the engine combustion process sounds harsher (more like an explosion and less of a controlled burn) (NVH – Noise Vibration Harshness).

Turbulent air velocity before flame initiation feeds the flame with air and fuel and makes the flame front propagate faster resulting in a faster burn.

In terms of piston and combustion chamber design, the 4-valve head creates a geometry with less distance for the flame front to travel. Shown below is a schematic representation comparing geometries where you can see the flame front position after a time t .



In terms of NVH, you generally want to limit the pressure rise in the cylinder to $5 \text{ Bar}/^\circ$ to sound acceptable.



Generally, the burned zone temperatures are in the region of $\sim >2000\text{K}$ and the unburned zone is generally $\sim <1000\text{K}$. The unburned zone gets hotter since it is being compressed due to the moving flame front.

FORCED INDUCTION

As you pressure charge an engine, you burn more air and fuel in the same volume and the peak pressures obviously go up. Turbo charged engines thus have higher cylinder pressure to compress the unburned zone more and this then raises its temperature and the danger from this is that it can become:

1. Too hot
2. And/or waits too long for the flame and the mixture auto-ignites.

Typical road going pressure charged engines cannot run at MBT timing but have to be retarded (over 10° retard) to keep pressures down by burning in a bigger volume to protect unburned zone auto-ignition.

$$\text{The work done per degree by the gas} = \frac{P_1 + P_2}{2} \times (V_2 - V_1)$$

IMEP = indicated mean effective pressure

IMEP x Cylinder Displacement = Work Done in 2 Revolutions

The general idea is to maximise the IMEP at MBT timing.

Turbo charged road engines are generally knock limited (KL) IMEP ~ 20 Bar

Octane Rating

Dr Owen Williams – University of Wales TSD:

In the UK we measure the:

RON (Research Octane Number) @ 600 RPM
&
MON (Motored Octane Number) @ 900 RPM

This is achieved using a test engine with variable compression ratio.

The octane rating of gasoline is measured in a test engine and is defined by comparison with the mixture of 2,2,4-trimethylpentane (iso-octane) and normal heptane that would have the same anti-knocking capacity as the fuel under test: the percentage, by volume, of 2,2,4-trimethylpentane in that mixture is the octane number of the fuel. For example, gasoline with the same knocking characteristics as a mixture of 90% iso-octane and 10% heptane would have an octane rating of 90.



A rating of 90 does not mean that the gasoline contains just iso-octane and heptane in these proportions, but that it has the same detonation resistance properties (generally, gasoline sold for common use never consists solely of iso-octane and heptane; it is a mixture of many hydrocarbons and often other additives). Because some fuels are more knock-resistant than pure iso-octane, the definition has been extended to allow for octane numbers greater than 100.

Example:

Our 2 reference fuels are:

Iso-octane (C_8H_{18}) ~ RON 100
n-Heptane (C_7H_{16}) ~ RON 0

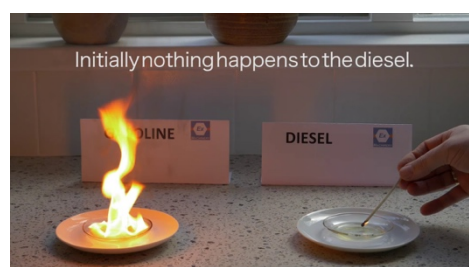
If during the test, your fuel takes a CR that is the same as a mixture of 95% iso-octane and 5% normal n-heptane it will be given a RON of 95

MON is measured with the same method as RON but at 900RPM with a heated intake.

$$\text{The 'Anti Knock Index' AKI} = \frac{RON + MON}{2}$$

In a diesel, the compression ratio is around 15:1 and the air is heated, and the fuel injected into it. As such, it needs to have a low octane rating in order for it to ignite without a spark. Diesel does not vaporise because of the higher flash point.

<https://www.youtube.com/watch?v=7soVqyGq4i4&t=4s>



Race Fuels

Dr Owen Williams – University of Wales TSD:

The 2019 FIM MotoGP guidelines for fuel stipulate the following:

Property	Units	Min.	Max.	Test Method
RON		95.0	102.0	ISO 5164
MON		85.0	90.0	ISO 5163
Oxygen	% (m/m)		2.7	ISO 22854
Benzene	% (v/v)		1.0	ISO 22854
Vapour Pressure (DVPE)	kPa		90	EN 13016-1
Lead	mg/L		5.0	EN 237
Density at 15 °C	kg/m ³	720.0	775.0	ASTM D 4052
Oxidation Stability	minutes	360		ASTM D 525
Existent gum	mg/100 mL		5.0	EN ISO 6246
Sulphur	mg/kg		10	ASTM D 5453
Nitrogen	% (m/m)		0.2	ASTM D 4629
Copper Corrosion	Rating		Class 1	ISO 2160
Distillation:				
At 70 °C	% (v/v)	22.0	50.0	ISO 3405
At 100 °C	% (v/v)	46.0	71.0	ISO 3405
At 150 °C	% (v/v)	75.0		ISO 3405
Final Boiling Point	°C		210.0	ISO 3405
Residue	% (v/v)		2.0	ISO 3405
Olefins(*)	% (v/v)		18.0	ISO 22854
Aromatics(*)	% (v/v)		35.0	ISO 22854
Total di-olefins	% m/m		1.0	GCMS / HPLC GCMS / HPLC
Appearance		clear and bright		visual inspection

For the MotoGP race, no fuel on the motorcycle may be more than fifteen degrees C (15°C) below ambient temperature. The use of any device on the motorcycle to artificially decrease the temperature of the fuel below ambient temperature is forbidden.

(The old FIA F1 regs limited fuels to a RON of 102)