

Applied Thermodynamics



Notes

UNIT-1 COMBUSTION THERMODYNAMICS

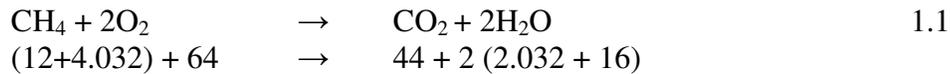
1.1 Introduction

All conventional fossil fuels, whether, solid, liquid or gaseous, contain basically carbon and hydrogen which invariably react with the oxygen in the air forming carbon dioxide, carbon monoxide or water vapour. The heat energy released as a result of combustion can be utilized for heating purposes or for generation of high pressure steam in a boiler or as power from an engine or a gas turbine.

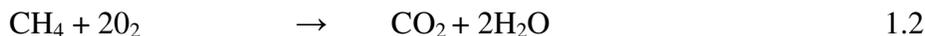
The solid fuels are burned in beds or in pulverised form suspended in the air stream. The liquid fuels are burned either by vaporising and mixing with air before ignition, when they behave like a gaseous fuel. The gaseous fuels are either burned in burners when the fuel and air are premixed or the fuel and air flow separately into a burner or a furnace and simultaneously mix together as combustion proceeds.

The Kg-mole or gram-mole is widely used in combustion calculations as a unit of weight. The molecular weight of any substance in kg represents one kilogram mole or 1K mole. 1Kmol of hydrogen has a mass of 2.016Kg and 1Kmol of carbon has a mass of 12Kg.

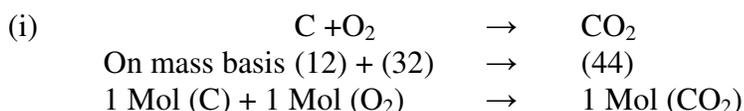
Consider a reaction

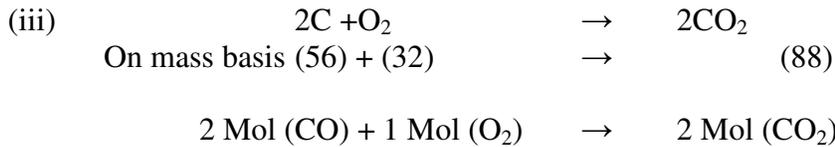
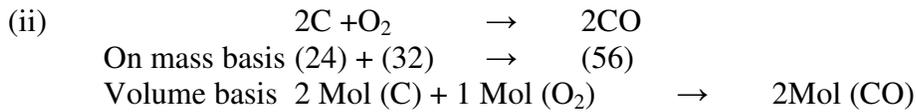


16.032kg of methane reacts with 64Kg of oxygen to form 44kg of carbon dioxide and 36.032kg of water. We can also simply state that 1Kmol of methane reacts with 2Kmol of oxygen to form 1Kmol of carbon dioxide and 2K mol of water, this has advantage of permitting easy conversion between the mass and volumetric quantities for the gaseous fuel and the product of combustion. If the gases are considered ideal then according to Avogadro hypothesis, all gases contain the same number of molecules per unit volume. It implies that 1K mole of any gaseous substance occupies the volume of 22.4m^3 at NTP i.e., 1.013bar and 273K.



1 volume of methane reacts with 2 volume of oxygen to form one volume of CO_2 and two volumes of H_2O . Therefore in any reactions, the mass is confirmed but the no. of mol or volumes may not be considered.





1.2 Combustion Stoichiometry

A balanced chemical equation for complete Combustion of the reactions with no excess air in the product is known as a stiochiometric equation. A stiochiometric mixture of the reactants is one in which the molar proportions of the reactants are exactly as given by the stiochiometric coefficients, so that no excess of any constituent is present. In general a chemical reaction may be written as



Where the reactants A and B react to form the products C and D. The small letters a, b, c and d are known as the stiochiometric coefficients.

For the combustion of any fuel the most common oxidizer is air which is a mixture of 21% O₂ and 79% N₂ (on volume basis). One mol of oxygen is accompanied by 79/21 (3.76) mol of Nitrogen. The Chemical equation for the stiochiometric combustion of carbon with air is written as

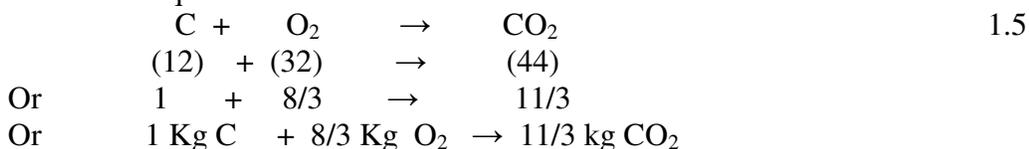


The minimum amount of air required for the complete combustion of a fuel is known as theoretical air. However in practice it is difficult to achieve complete combustion with theoretical air. Therefore fuel requires some excess air for different application and may vary from 5% ~ 20% and in gas turbine it may go up to 400% of theoretical quantity.

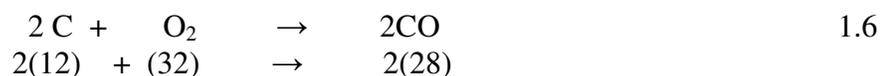
1.3 Theoretical air required for complete combustion.

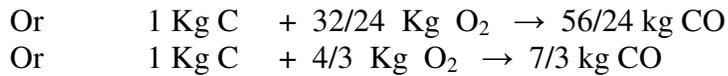
If the fuel composition is known, the requirement of oxygen or air can be calculated either by mass balance or by mole method.

Consider a equation



Similarly

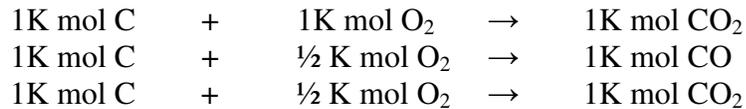




Similarly



On molal basis



1.4 Conversion of Gravimetric analysis to volumetric basis and vice versa

If the composition of fuel is given on gravimetric (or weight) basis it can be converted to volumetric (or mole) basis as follows. Divide the weight of each constituents of the mixture by its molecular weight. This will give the relative volume (or mole) of each constituents. Add all the relative volumes of the constituents then,

$$\frac{\text{Individual (relative) volume of the constituents}}{\text{Total volume of all the constituents}} \times 100$$

will give the %age by volume of each constituents in the fuel.

If the volumetric composition of a fuel is given, it can be converted to gravimetric (or weight) basis as follows. Multiply the individual volume of each constituent by its molecular weight. This will give relative weight of each constituent. Add all the relative weights of the constituents then

$$\frac{\text{Individual weight of the constituents}}{\text{Total (relative) weights of the constituents}} \times 100$$

will give the %age by weight of each constituent in the fuel.

1.4.1 Calculation of the minimum amount of air for a fuel of known composition.

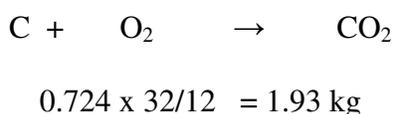
Example 1

Calculate the minimum volume of air required to burn one Kg of coal having the following composition by weight

C = 72.4%, H₂ 5.3%, N₂ = 1.81, O₂ = 8.5%, moisture 7.2%
 S = 0.9% and ash 3.9%

On weight basis:

Taking 1kg coal as basis weight of oxygen required to burn 1kg of coal



$$0.53 \times 16/2 = 0.424 \text{ kg}$$

$$0.009 \times 32/32 = 0.009 \text{ kg}$$

$$\text{Total O}_2 = 2.363 \text{ kg per kg of coal}$$

But 0.085kg O₂ is available in coal, therefore O₂ required

$$= 2.363 - 0.085 = 2.278 \text{ kg per Kg of coal.}$$

Air contains 23% of oxygen by weight. Therefore the weight of the air supplied is

$$2.278 \times 100/23 = 9.9 \text{ kg per kg of coal}$$

Density of air required at NTP

$$P v = mRT$$

$$P = m/v RT = \rho RT,$$

$$\rho = \frac{\text{Molecular weight}}{\text{Volume}}$$

$$= P/RT = 1.013 \times 10^5 / 287 \times 273 = 1.29 \text{ kg/m}^3$$

$$\text{Therefore volume of air required} = 9.9(\text{kg})/1.29(\text{kg}) = 7.67 \text{ m}^3$$

On mole basis

Consider 100kg of coal

C	=	72.4/12	=	6.03K mol,	O ₂	=	8.5/32	=	0.265K mol
H ₂	=	5.3/2	=	2.65K mol,	H ₂ O	=	7.2/18	=	0.4K mol
N ₂	=	1.8/28	=	0.064K mol,	S	=	0.9/32	=	0.028K mol



Therefore 6.03 K mol of carbon requires
6.03 K mol of oxygen



$$\text{H}_2 - 2.65 \times \frac{1}{2} = 1.325\text{K mol}$$

$$\text{S} - 0.028 \times 1 = 0.028$$

$$\text{Total O}_2 \text{ required } 6.03 + 1.325 + 0.028 = 7.383$$

The oxygen present in coal 0.265K mol

$$\text{Net O}_2 \text{ required} = 7.383 - 0.265 = 7.118\text{K mol}$$

Air required

$$7.118 \times 100/21 = 33.89\text{K mol} / 100\text{kg of coal} = 0.3389\text{K mol} / 1\text{kg coal}$$

Volume of air supplied
 $0.3389\text{K mol/kg} \times 22.4\text{m}^3 = 7.59\text{m}^3/\text{kg of coal}$

Example 2

Calculate the volumetric analysis of the flue gases when coal burns with 20% excess air from the previous calculation the actual air required 33.89K mol/100kg coal. Therefore the actual air is

$$33.89 \times 120/100 = 40.67\text{K mol/ 100 kg coal}$$

The amount of N₂ associated with this

$$40.67 \times 79/100 = 32.13\text{K mol}$$

The amount of O₂ present $40.67 \times 21/100 = 8.54\text{K mol}$

The actual amount of O₂ required was 7.118K mol excess O₂ will appear in exhaust gas = 8.54 – 7.118 = 1.422K mol.

Therefore:

$$\begin{aligned} \text{CO}_2 &= 6.03\text{K mol} \\ \text{SO}_2 &= 0.028\text{K mol} \\ \text{N}_2 &= 32.13\text{K mol (air)} + 0.064 \text{ (fuel)} \\ &= 32.194\text{K mol} \\ \text{O}_2 &= 1.422\text{K mol os excess oxygen.} \end{aligned}$$

Therefore the Total volume = (6.03 + 0.028 + 32.194 + 1.422)
= 39.674K mol

The volumetric composition of the gas

$$\begin{aligned} \text{CO}_2 &= (6.03/39.674) \times 100 = 15.12\% \\ \text{SO}_2 &= (0.028/39.674) \times 100 = 0.07\% \\ \text{N}_2 &= (32.13/39.674) \times 100 = 81.15\% \\ \text{O}_2 &= (1.422/39.674) \times 100 = 3.58\% \end{aligned}$$

1.5 Calculation of the composition of fuel and excess air supplied from the exhaust gas analysis:

Some times the composition of fuel is unknown and it becomes necessary to judge whether the amount of air supplied is sufficient or not, or excess. This can be obtained by analyzing the sample of exhaust gases.

Example 3

The composition of dry flue gases obtained by burning a liquid fuel containing only hydrogen and carbon is CO₂ 10.7%, O₂ 5.1%, N₂ 84.2%. Calculate the composition of fuel by weight and excess air used.

Solution: consider 100K mol of dry flue gases. They will contain 10.7K mol of O₂ (from CO₂) + 5.1K mole of (as max. oxygen) = 15.8K mol

Using nitrogen balance the actual air used $84.2 \times 100/79 = 106.58$ K mol of dry flue gases and oxygen in the air supplied $106.58 \times 21/100 = 22.38$ K mol. Therefore the amount of O₂ present in the water produced by the combustion of H₂ is $22.38 - 15.8 = 6.58$ K mol O₂. We know that 1 K mole of H₂ combines with $\frac{1}{2}$ K mol O₂ to produce water. Therefore the amount of hydrogen present is $6.58 \times 2 = 13.16$ K mol/100K mol of dry flue gases, and the carbon present is $12 \times 10.7 = 128.4$ kg/100K mol of dry flue gas. Therefore the composition of fuel (by weight) is 128.4kgC and 26.32Kg H₂ on the %age basis.

$$C = (128.4/(128.4+26.32)) \times 100 = 82.99\%$$

$$H = (26.32/(128.4+26.32)) \times 100 = 17.01\%$$

Excess air supplied

The amount of O₂ required to burn 10.7K mol C is 10.7K mol and to burn 13.16K mol H₂ is 13.16 X $\frac{1}{2} = 6.58$

Total O₂ required = $10.7 + 6.58 = 17.28$ K mol/100K mol of dry flue gases

$$\%age \text{ of excess air} = (22.38 - 17.28)/(17.28) \times 100 = 29.5\%$$

1.6 Dew point of products:

The product of combustion containing water vapour are known as wet products. The water vapour present in combustion product is cooled down to a point of condensation the vapour turn in to liquid and volume will be reduced. Knowing the partial pressure exerted by the water before condensing, it is possible to find the saturation temp. corresponding to partial pressure from the steam tables.

1.7 Flue gas analysis

A fuel has the following %age volumetric analysis

$$H_2 = 48, CH_4 : 26, CO_2 : 11, CO : 5, N_2 = 10$$

The %age volumetric analysis of the dry exhaust gases in CO₂:8.8, O₂: 5.5, N₂: 85.7

Determine the air/fuel ratio by volume if air contains 21% O₂ by volume

Solution: the chemical equation for the reaction of 100 moles of fuel gas with air may be written as



$$\begin{aligned}
 \text{Carbon balance} \quad (C) &\rightarrow 26 + 11 + 5 = a = H_2 \\
 H_2 &\rightarrow 48 + 52 = b = 100 \\
 O_2 &\rightarrow 11 + 2.5 + x = a + c \quad (i) \\
 N_2 &\rightarrow 10 + 3.76x = d \quad (ii)
 \end{aligned}$$

Solving (i) and (ii) we, have

$$\text{From (i)} \quad 11 + 2.5 + x = 100/2 + a + c$$

$$13.5 + x = 50 + a + c$$

$$\text{Adding} \quad \underline{10 + 3.76x = d}$$

$$23.5 + 4.76x = 50 + a + c + d$$

$$\text{Or} \quad a + c + d = 4.76x - 26.5$$

% CO₂ by volume in dry gas

$$(a/a+c+d) \times 100 = 8.8$$

$$\text{or} \quad (42/4.76x - 26.5) = 0.088$$

$$4.76x = 503.77$$

$$AF = \frac{\text{Total mol air}}{\text{Total mol fuel}} = \frac{503.77}{100} = 5.038\%$$

Example 4

A blast furnace gas has the following volumetric analysis H₂ CO-24%, CH₄ – 2%, CO₂-6%, O₂-3% and N₂-56%

Determine the Ultimate gravimetric analysis

Given volumetric analysis, H₂ – 9%, CO-24%, CH₄ – 2%, CO₂-6%, O₂-3% and N₂-56%

Solution: The volumetric analysis may be converted into mass or granite metric analysis by completing the table as follows:

Constituent	Volume in 1m ³ of flue gas (a)	Molecular mass (b)	Proportional mass (c)=(a)x(b)	Mass in kg per kg of the gas (d)=(c)/ Σ ©	% by mass = (d)x100
CO	0.24	28	6.72	6.72/18.48 = 0.36	36%
CH ₄	0.02	16	0.32	0.32/18.48 = 0.0173	1.73%
CO ₂	0.06	44	2.64	264/18.48 = 0.142	14.2%
O ₂	0.03	32	0.96	0.96/18.48 = 0.0519	5.19%
N ₂	0.56	14	7.84	7.84/18.48 = 0.42	42%
			Σc = 18.48	Σ (d) = 1	100

The volumetric analysis of flue gas components becomes

CO-0.36, CH₄ – 0.0173, CO₂- 0.142, O₂-0.0519 and N₂-0.42

Example 5

Determine the fuel gas analysis and air fuel ratio by weight when fuel oil with 84.9% carbon, 11.4% hydrogen, 3.2% sulphur, 0.4% oxygen and 0.1% ash by weight is burnt with 20% excess air, assume complete combustion.

Solution: Consider 1kg of fuel

Oxygen required / Kg of fuel

For burning of 1kg C - $0.849 \times 32/12$

For burning of 1kg H - $0.114 \times 16/2$

For burning of 1kg S - $0.032 \times 32/32$

Total O₂ required is 3.208 kg.

Amount of O₂ contained in the fuel = 0.004Kg

Net O₂ supplied / kg of fuel = 3.208 – 0.004

$$= 3.204 \text{ kg O}_2$$

Net air supplied = $3.204 \times 100/23 = 13.93$ kg/kg of fuel

When 20% excess air supplied

Total air supplied = $13.93 \times 1.2 = 16.716$ kg/kg of fuel.

N₂ actually supplied = $16.716 \times 77/100 = 12.871$ kg/kg of fuel

O₂ actually supplied = $16.716 \times 23/100 = 3.845$ kg/kg of fuel

Total free O₂ in fuel gas = $3.845 - 6.204$

$$= 0.641 \text{ kg/kg of fuel}$$

Total free N₂ in fuel gas = 12.87 kg/kg of fuel

Flue gas analysis:

C converted to CO₂ = $0.849 \times 44/12 = 3.113$ kg CO₂

H converted to H₂O = $0.114 \times 18/2 = 1.026$ kg H₂O

S converted to SO₂ = $0.032 \times 64/32 = 0.064$ kg SO₂

Flue gas / kg of fuel:

$$= 3.113 + 1.26 + 0.064 + 0.641 + 12.871$$

CO₂ H₂O SO₂ O₂ N₂

$$= 17.715 \text{ kg.}$$

Therefore:

$$\text{CO}_2 = (3.113/17.715) \times 100 = 17.573\%$$

$$\text{SO}_2 = (0.064/17.715) \times 100 = 0.36\%$$

$$\text{O}_2 = (0.641/17.715) \times 100 = 3.618\%$$

$$\text{H}_2\text{O} = (1.026/17.715) \times 100 = 5.79\%$$

$$\text{N}_2 = (12.871/17.715) \times 100 = 72.656\%$$

Air fuel mixture ratio is = 16.716 : 1

Example 6

A blast furnace gas has the following volumetric analysis.

$\text{H}_2 = 9\%$, $\text{CO} = 24\%$, $\text{CH}_4 = 2\%$, $\text{CO}_2 = 6\%$, $\text{O}_2 = 3\%$ and $\text{N}_2 = 56\%$

Determine the ultimate gravimetric analysis.

Solution:

Total H₂ in the blast furnace gas.

$$\% \text{ volumetric analysis} = 9\text{H}_2 + 2\text{H}_4$$

$$\begin{aligned} \text{Proportional mass} &= \% \text{ volumetric analysis} \times \text{mol. Mass of element} \\ &= (9 \times 2) + (2 \times 4) = 18 + 8 = 26 \text{ kg.} \end{aligned}$$

Total 'C' in the blast furnace gas.

$$\% \text{ of volumetric analysis} = 24\text{C} + 2\text{C} + 6\text{C}$$

$$\begin{aligned} \text{Proportional mass} &= (24+2+6) \times 12 \\ &= 384 \text{ kg} \end{aligned}$$

Total O₂ in the blast furnace gas

$$\% \text{ of volumetric analysis} = 24 \times \text{O} + 6\text{O}_2 + 3\text{O}_2$$

$$\begin{aligned} \text{Proportional mass} &= (24+16) \times 9 \quad (32) \\ &= 672 \text{ kg} \end{aligned}$$

Total N₂ in the blast furnace gas

$$\% \text{ of volumetric analysis} = 56 \text{ N}_2$$

$$\text{Proportional mass of N}_2 = 56 \times 28 = 1568 \text{ Kg.}$$

Total weight of blast furnace gas:

$$\begin{aligned} &= 384\text{kg C} + 26\text{kg H}_2 + 672\text{kg O}_2 + 1568\text{kg N}_2 \\ &= 2650\text{kgs} \end{aligned}$$

Gravimetric %age composition:

$$\text{C} = (384/2650) \times 100 = 14.49\%$$

$$\text{H}_2 = (26/2650) \times 100 = 0.98\%$$

$$\text{O}_2 = (672/2650) \times 100 = 25.36\%$$

$$\text{N}_2 = (1568/2650) \times 100 = 59.17\%$$

Example 7

The analysis of coal used in a boiler trial is as follows. 82% carbon, 6% hydrogen, 4% oxygen, 2% moisture and 8% ash. Determine the theoretical air required for complete combustion of 1kg of coal. If the actual air supplied is 18kg per kg of coal the hydrogen is completely burned & 80% carbon burned to CO₂, the remainder is CO, Determine the volumetric analysis of the dry products of combustion.

Solution: For complete combustion.

O₂ required is

$$\text{For carbon} - 0.82 = 2.186 \text{ kg of O}_2$$

$$\text{For hydrogen} - 0.006 = 0.48 \text{ kg of O}_2$$

$$\text{Total O}_2 \text{ required} = 2.666\text{kg.}$$

$$\text{Net O}_2 \text{ supplied} = \text{Total O}_2 \text{ required} - \text{O}_2 \text{ present in the fuel}$$

$$= 2.66 - 0.004$$

$$= 2.662 \text{ kg/kg of coal}$$

Theoretical minimum air required for complete combustion [C burns to CO₂ totally]

$$\text{Air supplied} = 2.626 \times 100/23 = 11.417 \text{ kg/kg of coal}$$

Flue gas analysis:

But actually only 80% carbon is burnt to CO₂

$$\text{CO}_2 = 0.8 \times 0.82 \times 44/12 = 2.405 \text{ kg of CO}_2$$

20% carbon is burnt to CO

$$\text{CO} = 0.2 \times 0.82 \times 28/12 = 0.383 \text{ kg of CO}$$

O₂ actually required for 80% carbon burnt to CO₂

$$= 0.8 \times 0.82 \times 32/12 = 1.749 \text{ kg of O}_2$$

O₂ actually required for 20% carbon burnt to CO

$$= 0.2 \times 0.82 \times 16/12 = 0.219 \text{ kg of O}_2$$

O₂ required by Hydrogen:

$$= 0.06 \times 8 = 0.48 \text{ kg of O}_2.$$

H₂O produced = 0.06 x 9 = 0.54 kg of H₂O

But actual air supplied = 18kg

$$\text{Actually O}_2 \text{ supplied} = 18 \times 23/100 = 4.14 \text{ kg of O}_2$$

$$\begin{aligned} \text{Free O}_2 \text{ in the flue gas} &= 4.14 + 0.04 - 1.749 - 0.219 - 0.48 \\ &= 1.732 \text{ kg of O}_2/\text{kg of coal} \end{aligned}$$

$$\text{N}_2 \text{ in the flue gas} = 18 \times 77/100 = 13.86 \text{ kg/kg of coal}$$

Volumetric analysis of the dry products of combustion.

$$\text{CO}_2 = (2.405/44) \times 100 = 0.0546 \text{ m}^3/\text{K.mol}$$

$$\text{CO} = (0.383/28) \times 100 = 0.0137 \text{ m}^3/\text{K.mol}$$

$$\text{O}_2 = (1.732/32) \times 100 = 0.0541 \text{ m}^3/\text{K.mol}$$

$$\text{N}_2 = (13.86/28) \times 100 = 0.495 \text{ m}^3/\text{K.mol}$$

In % of volume:

$$\text{CO}_2 = (0.0546/0.6174) \times 100 = 8.84\%$$

$$\text{CO} = (0.0137/0.6174) \times 100 = 2.22\%$$

$$\text{O}_2 = (0.0541/0.6174) \times 100 = 8.76\%$$

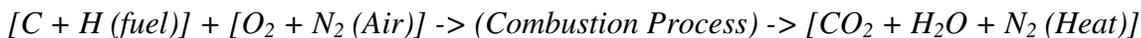
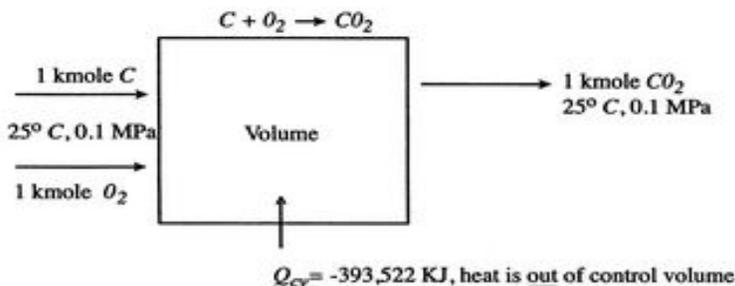
$$\text{N}_2 = (0.495/0.6174) \times 100 = 80.70\%$$

1.8 Enthalpy of reaction

Enthalpy of a reaction is defined as the difference between the enthalpy of the products at a specified state and the enthalpy of the reactants at the same state for a complete reaction. For combustion process, the enthalpy of a reaction is usually referred to as the “enthalpy of combustion” it is obviously a very useful property for analyzing the combustion processes of fuels. However there are so many different fuels and fuel mixtures that is not practical to list enthalpy of combustion values for all possible cases. Besides, the enthalpy of combustion is not of much use when the combustion is incomplete. Therefore a more practical approach would be have a more fundamentally property to represent the chemical energy of an element or compound at some reference state. This property is the “enthalpy of formation” which can be viewed as the enthalpy of a substance at a specified state due to its chemical composition. To establish a starting point it is assigned the enthalpy of formation for all stable elements such as O₂, N₂, H₂ and C a value of zero at standard reference state of 25°C and 1 atm. For all stable compounds.

In a chemical reaction bonds are broken in the reactants and new bonds formed in the products. Energy is required to break bonds and energy is released when bonds are formed. The energy associated with a chemical reaction depends on the number and type of bonds broken and/or formed.

Every chemical species has a certain amount of "heat content," or enthalpy, H, which cannot be measured. However, differences in enthalpy can be measured. The net energy change for a reaction performed at constant pressure is the enthalpy change for the reaction. This enthalpy change, ΔH , has units kJ/mol and is defined:



where

C = Carbon, H = Hydrogen, O = Oxygen, N = Nitrogen

$$(1) \quad \Delta H = H(\text{products}) - H(\text{reactants})$$

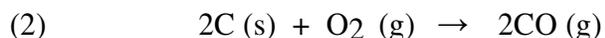
If energy is given off during a reaction, such as in the burning of a fuel, the products have less heat content than the reactants and ΔH will have a negative value; the reaction is said to be exothermic. If energy is consumed during a reaction, ΔH will have a positive value; the reaction is said to be endothermic.

The enthalpy change for a chemical change is independent of the method or path by which the

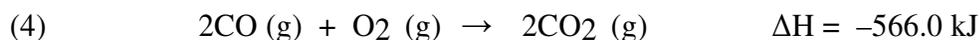
change is carried out as long as the initial and final substances are brought to the same temperature. This observation, known as HESS'S LAW, has important practical utility.

Thermochemical equations may be treated as algebraic equations: they may be written in the reverse direction with a change in the sign of ΔH – even though the reverse reaction may not actually occur; they may be added and subtracted algebraically; the equation and associated ΔH value may be multiplied or divided by factors. Hess's Law allows the calculation of enthalpy changes that would be difficult or impossible to determine directly, i.e. by experiment.

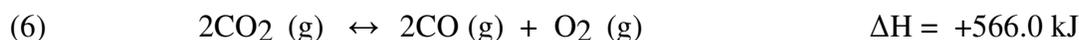
The enthalpy change for the reaction:



cannot be determined directly because carbon dioxide will also form. However, ΔH can be measured for:



Multiplying equation (3) by 2 gives equation (5), and reversing equation (4) gives equation (6):



Adding equations (5) and (6) gives the desired information:



For a reaction in which a compound is formed from the elements, the enthalpy change is called the heat of formation, ΔH_f° , for the compound. The superscript "o" indicates standard conditions of one atmosphere pressure. Equation (2) and (3) are such reactions. Some others:



In reactions (2), (3), (7), and (8) ΔH for the reaction is ΔH_f° for the compound. For the reaction:



the heat of reaction is associated with the formation of two moles of SO_3 . But heat of formation is per mole of compound, so ΔH_f° for SO_3 is half of -790.4 , or -395.2 kJ .

Extensive listings of heats of formation are available in handbooks. With these values of ΔH_f° , you can calculate virtually any heat of reaction. The heat of a reaction is the sum of ΔH_f° values for the products minus the sum of ΔH_f° values for the reactants. Expressed as a formula:

$$(10) \quad \Delta H_{\text{rxn}} = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

Heats of formation for several compounds are given below. Note that the phase of the compound

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