Step 1: Establish the Fuel Properties

The fuel given is a hydrocarbon represented by the chemical formula C_5H_8 (alkyne/diene series). We calculate the molecular weight of the fuel (M_F) using the baseline values from the manual:

$$m_F = 5\,(molar\ mass\ of\ C) + 8(molar\ mass\ of\ H) = 5(12) + 8(1) = 68\ \frac{kg}{kmol_{fuel}}$$

Step 2: Formulate and Balance the Stoichiometric Equation

To determine how much oxygen is chemically required for a perfect burn, we write the ideal stoichiometric reaction where all Carbon is oxidized to CO_2 and all Hydrogen is oxidized to H_2O:

$$C_5H_8 + a\left(O_2 + 3.76N_2\right) \rightarrow 5CO_2 + 4H_2O + 3.76aN_2$$

To find the theoretical oxygen coefficient (a), balance the oxygen atoms across both sides of the equation: $2a = 5(2) + 4(1) = 14$

$$a = 7\ moles\ of\ CO_2$$

Step 3: Compute the Actual Air Supply (with 35% Excess Air)

Because real-world burners cannot achieve perfect molecular mixing at exactly stoichiometric ratios, excess air (e = 35% = 0.35) is introduced to ensure complete combustion. The actual moles of oxygen supplied (a) scale up by a factor of (1 + e):

$$a_{act} = a(1 + e) = 7(1 + 0.35) = 9.45\ moles\ of\ O_2$$

Since every mole of atmospheric O_2 brings along 3.76 moles of inert N_2, the total mass of actual air (m_a') required to burn 1 kmol (68 kg) of this fuel is:

$$m_a' = n_{O_2}M_{O_2} + n_{N_2}M_{N_2}$$

$$m_a' = 9.45(32) + 9.45(3.76)(28) = 1,297.296\ \frac{kg_{air}}{kmol_{fuel}}$$

Step 4: Calculate the Actual Air-Fuel Ratio (A:F')

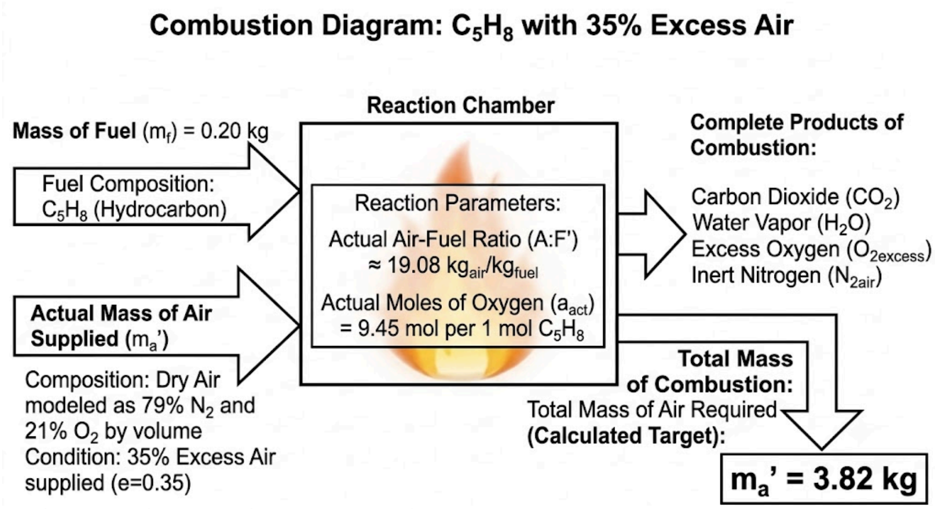
The actual air-fuel ratio represents the mass of air supplied per unit mass of fuel consumed:

$$A{:}F' = \frac{Mass\ of\ actual\ air}{Mass\ of\ fuel} = \frac{1{,}297.296\ \frac{kg_{air}}{kmol_{fuel}}}{68\ \frac{kg}{kmol_{fuel}}} = 19.0779\ \frac{kg_{air}}{kg_{fuel}}$$

<u>Step 5:</u> Determine Mass of Air Required for 0.20 kg of Fuel

Given that the physical mass of the fuel sample is exactly m_f = 0.20 kg, the mass of air needed is computed by scaling the ratio:

$$\text{Mass of Air Required} = (A{:}F')(m_f) = \left(19.0779\ \frac{kg_{air}}{kg_{fuel}}\right)\left(0.20\ kg_{fuel}\right) = 3.8156\ kg$$



This problem demonstrates how thermal plants manage combustion stability. Supplying exactly the theoretical amount of air (e=0) would lead to incomplete combustion, producing unburned carbon soot and hazardous carbon monoxide (CO), which wastes chemical fuel energy. By supplying 35% excess air, mechanical engineers ensure there is plenty of oxygen to achieve complete combustion.

While excess air prevents fuel waste, it introduces a practical engineering trade-off: **Nitrogen Ballast Effect**. Atmospheric air consists of 79% inert Nitrogen (N_2).

This means that for this 0.20 kg fuel sample, nearly 2.9 kg of the required 3.8 kg of air is inert nitrogen that does not contribute to heat generation. Instead, it passes through the system, absorbing combustion energy and carrying it out the exhaust stack as waste heat.

In real-world power plants, engineers use combustion control systems to keep excess air as low as possible to maximize boiler efficiency, while still providing enough oxygen to ensure a complete, clean burn.