Biological oils are different from petroleum oils ("regular" diesel fuel) in molecular structure and properties.

Some cities are converting their bus lines to biodiesel. 

Because of their large size and consequent large intermolecular attractions, the viscosity of biological oils is generally too high for use in conventional diesel engines. Biological oils also burn a little less readily, and with a sootier flame than petroleum diesel. Biological oils can be used in conventional diesel engines if they are preheated to reduce their viscosity, but this requires an auxiliary electrical heater until the engine warms up. For these reasons, biological oils require processing for use as biodiesel.

A biological oil is an ester, which is a type of organic compound having the atom linkage shown below.
hydrogen atoms. R = -CH₃ and 
R’ = -C_{18}H_{35}O₂ for the methyl 
steare in Example 1

The ester linkage in biological oils is created when a glycerol molecule reacts with organic acids. The glycerol molecule has a chain of 3 carbon atoms, each with an -OH (alcohol) group on it. The figure below shows how an organic alcohol reacts with a organic acid. Organic chemists abbreviate molecular structures—the "zig-zag" lines in the figure represent carbon chains with a C atom at each "zig" or "zag". Each carbon has 4 bonds, and if fewer than 4 are shown, it's assumed that they go to H atoms. So the alcohol is C₂H₅OH (ethanol), and the acid is acetic acid (or ethanoic acid, CH₃COOH) in the Figure:

\[ \text{Organic Acid} + \text{Alcohol} \rightarrow (\text{with sulfuric acid catalyst}) \text{ an ester} \text{ and water} \]

Since glycerol has 3 -OH groups, 3 long chain organic "fatty acids" attach to make the bulky "triglyceride".

Stearin, or glyceryl tristearate

But just as easily as esters can be made from alcohols and acids, they can switch alcohols or acids. In the presence of a strong base catalyst, like NaOH, a triglyceride can react with 3 small alcohol molecules, like methanol (CH₃OH), which replace the glycerol "backbone", making 3 separate esters of lower molecular weight

Quite often a mixture of two or more products is formed. For example, when a vegetable oil reacts with methanol, only one or two of the acids may be displaced from the glycerine, producing only 1 or 2 FAMEs.

\[ C₃H₅(C₁₈H₃₅O₂)₃ + \text{NaOH} + 2 \text{CH₃OH} \rightarrow C₃H₅(C₁₈H₃₅O₂)₂(\text{OH}) + 2 \text{C}_{1₇}H₃₅\text{COOCH₃} \]

\[ C₃H₅(C₁₈H₃₅O₂)₃ + \text{NaOH} + 1 \text{CH₃OH} \rightarrow C₃H₅(C₁₈H₃₅O₂)(\text{OH})₂ + 1 \text{C}_{1₇}H₃₅\text{COOCH₃} \]

\[ C₃H₅(C₁₈H₃₅O₂)₃ + \text{NaOH} + 3 \text{CH₃OH} \rightarrow C₃H₅(\text{OH})₃ + 3 \text{C}_{1₇}H₃₅\text{COOCH₃} \]
Usually, a large excess of methanol and sodium hydroxide are added, so that the reaction produces the maximum amount of FAME.

But in the case of a transesterification, even though none of the reactants is completely consumed, no further increase in the amounts of the products occurs. We say that such a reaction does not go to completion. When a mixture of products is produced or a reaction does not go to completion, the effectiveness of the reaction is usually evaluated in terms of percent yield of the desired product. A theoretical yield is calculated by assuming that all the limiting reagent is converted to product. The experimentally determined mass of product is then compared to the theoretical yield and expressed as a percentage:

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EXAMPLE 1 When 100.0 g C$_3$H$_5$(C$_{18}$H$_{35}$O$_2$)$_3$ gas and 15.0 g CH$_3$OH are mixed at 55°C with NaOH catalyst, they react to form 90.96 g C$_{17}$H$_{35}$COOCH$_3$ methyl stearate biodiesel. Calculate the percent yield.

Solution We must calculate the theoretical yield of NH$_3$, and to do this, we must first discover whether N$_2$ or H$_2$ is the limiting reagent. For the balanced equation

$$C_3H_5(C_{18}H_{35}O_2)_3 + NaOH + 3 CH_3OH \rightarrow C_3H_5(OH)_3 + 3 C_{17}H_{35}COOCH_3$$

stearin + sodium hydroxide + 3 CH$_3$OH $\rightarrow$ glycerol + 3 methyl stearate

The stoichiometric ratio of the reactants is

$$\begin{array}{cccc}
\text{C}_3\text{H}_5\text{(C}_{18}\text{H}_{35}\text{O}_2)_3(s)} & + & 3 \text{CH}_3\text{OH (l)} & \rightarrow 1 \text{C}_3\text{H}_5\text{(OH)_3 (l)} & + & 3 \text{C}_{18}\text{H}_{35}\text{O}_2)3\text{CH}_3(s) \\
\text{m, g} & 100.0 \text{g} & 15.00 \text{g} & \rightarrow & 1 \text{C}_3\text{H}_5\text{(OH)_3 (l)} & + & 3 \text{C}_{18}\text{H}_{35}\text{O}_2)3\text{CH}_3(s) \\
\text{M, g/mol} & 891.5 & 32.04 & 298.5 & 92.1 \\
\text{n present, mol} & 0.1122 \text{ mol} & 0.4682 \text{ mol} & \text{} & \text{} \\
\text{n actual, mol} & 0.1122 & 0.3366 & 0.1122 & 0.3366 \\
\text{m actual, mass} & 100.0 & 10.78 & 10.33 & 100.47
\end{array}$$

Transesterification is a classic example of a reaction which does not go to completion.

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References

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