3. The molar enthalpy of combustion of octane is reported to be -1.3 MJ/mol.

a) Write the balanced equation using whole number coefficients, and determine the enthalpy change for the reaction.

$2\,C_8H_{18(l)} + 25\,O_{2(g)} \rightarrow 16\,CO_{2(g)} + 18\,H_2O_{(g)}$

$\Delta H = n\,\Delta_r H \quad 2\,mol \times \frac{-1.3\,MJ}{mol} = -2.6\,MJ = -2.6 \times 10^3\,kJ$

b) If 500 grams of octane is burned, calculate the quantity of energy released in MJ.

$\Delta H = n\,\Delta_r H = \frac{500\,g}{114.26\,\frac{g}{mol}} \times \frac{-1.3\,MJ}{mol} = \boxed{-5.7\,MJ}$

c) How many grams of octane must be burned to produce 500 kJ of energy?

$n = \frac{\Delta H}{\Delta_r H} = \frac{-500\,kJ}{-1.3 \times 10^3\,\frac{kJ}{mol}} = 0.38\,mol$

$m = nM = 0.38\,mol \cdot 114.26\,g/mol = \boxed{44\,g}$

4. Glycerol ($C_3H_8O_3(l)$) combusts according to the following chemical equation:

$2\,C_3H_8O_3(l) + 7\,O_2(g) \rightarrow 6\,CO_2(g) + 4\,H_2O(l) \quad \Delta_f H\,(C_3H_8O_3(l)) = -1654\,kJ/mol$

a) Calculate the enthalpy change when 25.0 g of glycerol combusts.

$\Delta H = n\,\Delta_r H = \frac{25.0\,g}{92.11\,g/mol} \times \frac{-1654\,kJ}{mol} = \boxed{-449\,kJ}$

b) What is the enthalpy change if 350 g of carbon dioxide is released from the reaction?

$\Delta H = n\,\Delta_r H$

$= \frac{350\,g}{44.01\,g/mol} \times \frac{-551.1\,kJ}{mol}$

$= \boxed{-4.38 \times 10^3\,kJ}$

Find $\Delta_r H$ per CO_2

① $\Delta H = n\,\Delta_r H = 2\,mol\,C_3H_8O_3 \cdot \frac{-1654\,kJ}{mol} = -3308\,kJ$

② $\Delta_r H = \frac{\Delta H}{n} = \frac{-3308\,kJ}{6\,mol\,CO_2} = \boxed{-551.3\,\frac{kJ}{mol}}$

c) If a sample of glycerol is burned and 4000 kJ of energy is released, calculate the mass of glycerol that burned.

$n = \frac{\Delta H}{\Delta_r H} = \frac{-4000\,kJ}{-1654\,\frac{kJ}{mol}} = 2.41837\,mol$

$m = nM = 2.41837\,mol \times 92.11\,g/mol$

$= \boxed{222.8\,g}$