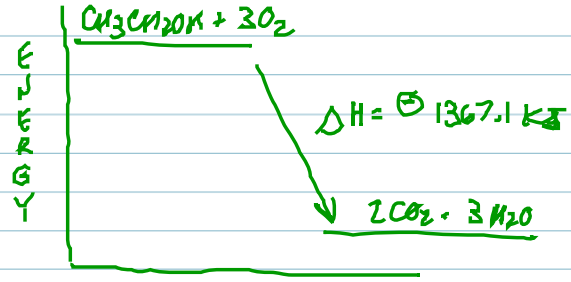
When 1.00 moles of ethanol, CH_3CH_2OH, is burned in air, 1367.1 kJ of heat is released.

$$CH_3CH_2OH + 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O + 1367.1\ kJ$$



How much heat is released when 11.39 g of ethanol is burned?

$$11.39\ g\ EtOH \times \frac{1\ mol}{46.08\ g} \times \frac{\ominus 1367.1\ kJ}{1\ mol\ EtOH} = \ominus 337.92\ kJ$$

Hess's Law – For any reaction that can be written in steps, the sum of the "standard heats of reaction" ΔH°_{rxn} of the steps is the same as the standard heat of reaction for the entire process

standard $T = 25^\circ C$ (298K)

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} \quad \Delta H^\circ = -393.5\ kJ$$

$$C_{(s)} + \tfrac{1}{2}O_{2(g)} \rightarrow CO_{(g)} \quad \Delta H^\circ = -110.5\ kJ$$

$$CO_{(g)} + \tfrac{1}{2}O_{2(g)} \rightarrow CO_{2(g)} \quad \Delta H^\circ = -283.0\ kJ$$

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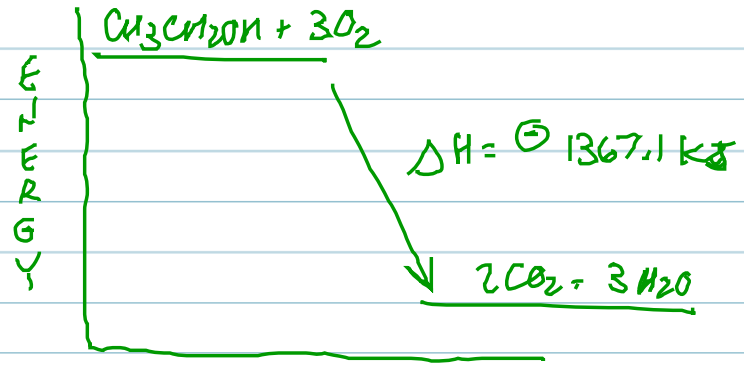

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TARGET: $3A + B \rightarrow E$

$2A + 3B \rightarrow C + D$ $\quad \Delta H = -64$ kJ ✓

$E + A \rightarrow C$ $\quad \Delta H = 100$ kJ ✓

$A + \frac{1}{2}D \rightarrow B$ $\quad \Delta H = 74$ kJ

$C \rightarrow E + A$ $\quad \Delta H = \ominus 100$ kJ

$2B + 2A \rightarrow C + D$ $\quad \Delta H = \ominus 64$ kJ

$2A + D \rightarrow 2B$ $\quad \Delta H = \oplus 148$ kJ

$3A + B \rightarrow E$ $\quad \Delta H = \ominus 16$ kJ

TARGET: $A + 2B \rightarrow C$

steps:

$E \rightarrow D + 2A$ $\quad \Delta H = 43.22$ kJ ✓

$B \rightarrow F + \frac{1}{4}D$ $\quad \Delta H = 342$ kJ ✓

$4F + E \rightarrow 2C$ $\quad \Delta H = 2.2$ kJ

$A + \frac{1}{2}D \rightarrow \frac{1}{2}E$ $\quad \Delta H = \ominus 21.61$ kJ

$2B \rightarrow 2F + \frac{1}{2}D$ $\quad \Delta H = \oplus 684$ kJ

$2F + \frac{1}{2}E \rightarrow C$ $\quad \Delta H = \oplus 1.1$ kJ

$A + 2B \rightarrow C$ $\quad \Delta H = \oplus 663.49$ kJ