6. Given the equation: $C_{12}H_{22}O_{11} + 12O_2 \rightarrow 12CO_2 + 11H_2O + 5638\ kJ$

a. How much heat is released during the formation of 9.6 moles of CO_2?

$9.6\ mol\ CO_2 \times \frac{5638\ kJ}{12\ mol\ CO_2} =$ Answer 4510.4 kJ

b. How much heat is released during the formation of 0.036 moles of H_2O?

$0.036\ mol\ H_2O \times \frac{5638\ kJ}{11\ mol\ H_2O} = 18.45\ kJ$ Answer 18.45 kJ

c. If 1026 grams of $C_{12}H_{22}O_{11}$ are consumed, how much heat is released?

$1026\ g\ C_{12}H_{22}O_{11} \times \frac{1\ mol}{342.0\ g} = 3\ mol\ C_{12}H_{22}O_{11} \times \frac{5638\ kJ}{1\ mol\ C_{12}H_{22}O_{11}} =$ 16914 kJ

Answer

d. If 23.76 grams of CO_2 are produced, how much heat is released?

$23.76\ g\ CO_2 \times \frac{1\ mol\ CO_2}{44.0\ g\ CO_2} = 0.54\ mol\ CO_2 \times \frac{5638\ kJ}{12\ mol\ CO_2} = 253.71\ kJ$

Answer

7. Calculate the amount of heat (in Joules) required to warm 350.0 g of water from 30°C to 35°C. (Heat Capacity (C) for H_2O is 4180 J/kg·°C)

$Heat = m \cdot C \cdot \Delta t$

$= 0.350\ kg \times 4180\ \frac{J}{kg \cdot °C} \times 5°C =$ 7315 J

Answer

8. 35.112 kJ of heat are added to a 500.0 gram sample of water initially at 7°C. Calculate the final temperature of the water sample. Be careful with units!

$Heat = m \cdot C \cdot \Delta t$

$35{,}112\ J = 0.500\ kg \times 4180\ \frac{J}{kg \cdot °C} \times \Delta t\ °C$

$\Delta t = \frac{35112}{(0.500)(4180)} = 16.8°C$

$t_{final} = t_{init} + \Delta t$

Answer $t_{final} = 7 + 16.8 =$ 23.8°C