



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

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

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

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

Answer 18.45 kJ

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

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

Answer _____

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

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

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 $4180 \text{ J/kg} \cdot ^\circ\text{C}$)

$$\text{Heat} = m \cdot C \cdot \Delta t$$

$$= 0.350 \text{ kg} \times 4180 \frac{\text{J}}{\text{kg} \cdot ^\circ\text{C}} \times 5^\circ\text{C} = \underline{7315 \text{ J}}$$

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!

$$\text{Heat} = m \cdot C \cdot \Delta t$$

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

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

Answer $t_{\text{final}} = 7 + 16.8 = \underline{23.8^\circ\text{C}}$

$$\Delta t = \frac{35112}{(0.500 \times 4180)} = \underline{16.8^\circ\text{C}}$$

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Brendan G. Carr

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