

Thermochemistry Worksheet Key

1. How much heat, in calories, is released or absorbed if a 15.0 g sample of copper is converted from
(a) the solid to liquid state.

$$q = m \times \Delta H_{fus} = 15.0 \text{ g} \times 206 \frac{\text{J}}{\text{g}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 738.5 = 739 \text{ cal}$$

So 739 cal is absorbed

- (b) the gas to the liquid state.

$$\begin{aligned} q &= m \times \Delta H_{con} = 15.0 \text{ g} \times -4.7 \frac{\text{kJ}}{\text{g}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = -16,849 \\ &= -1.68 \times 10^4 \text{ cal or } -16.8 \text{ kcal} \end{aligned}$$

So 16.8 kcal is released.

For copper m.p. = 1084 °C ; b.p. = 2927 °C; ΔH_{fus} = 206 J/g, ΔH_{vap} = 4.7 kJ/g

2. How much heat is absorbed or released if a 12.5 g sample of aluminum changes temperature from 25.0 to 87.5 °C? The specific heat capacity of aluminum is 0.900 J/g K

$$T_{final} = 87.5 + 273.15 = 360.65 \text{ K} \quad T_{initial} = 25.0 + 273.15 = 298.15 \text{ K}$$

$$\begin{aligned} q &= m \times c \times \Delta T = 12.5 \text{ g} \times 0.900 \frac{\text{J}}{\text{g} \cdot \text{K}} \times (360.65 - 298.15) \text{ K} \\ &= 12.5 \text{ g} \times 0.900 \frac{\text{J}}{\text{g} \cdot \text{K}} \times 62.5 \text{ K} = 703.1 = 703 \text{ J} \end{aligned}$$

3. What was the mass of a sample of water if the addition of 4.28×10^3 joules raised the temperature of the sample from 22.0 °C to 34.5 °C?

$$T_{final} = 34.5 + 273.15 = 307.65 \text{ K} \quad T_{initial} = 22.0 + 273.15 = 295.15 \text{ K}$$

T increases so endothermic so $q > 0$

$$q = m \times c \times \Delta T$$

$$\text{so } m = \frac{q}{c \times \Delta T} = \frac{4.28 \times 10^3 \text{ J}}{4.184 \frac{\text{J}}{\text{g} \cdot \text{K}} \times (307.65 - 295.15) \text{ K}} = \frac{4.28 \times 10^3 \text{ J}}{4.184 \frac{\text{J}}{\text{g} \cdot \text{K}} \times 12.5 \text{ K}} = 81.83 = 81.8 \text{ g}$$

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