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Calorimetry Calculations

1. If you wanted to cool your bottle of water, 500 g, from room temperature (22 °C) to its freezing point, → 0°C how much heat would be removed? The specific heat of water is 4.184 J/g °C.

$m = 500g$ $\Delta T = 0 - 22 = -22$ $c = 4.18$ $Q = (500)(4.18)(-22)$ $-45980J$ 46024J

2. You have a gold ring with a mass of 30 g and a copper ring with the same mass. How much heat would you have to add to each ring to raise their temperatures from 20 °C to 50 °C? The specific heat of gold is 0.129 J/g °C, and the specific heat of copper is 0.385 J/g °C.

Gold
$m = 30g$ $c = 0.129$ $\Delta T = 50 - 20 = 30$ $Q = (30)(.129)(30)$ 116.1J

Copper
$Q = (30)(.385)(30)$ 346.5J

Which metal is easier to heat up? How do you know?

Gold = requires less energy (specific heat is smaller also)

3. A blacksmith has heated a 500 g bar of iron to 200 °C and molded it into a horseshoe. He then puts it in a tub of water at 20 °C to cool it. How much heat is released by cooling the horseshoe? The specific heat of iron is .444 J/g °C.

$\Delta T = 20 - 200 = -180$ $c = .444$ $m = 500g$ $Q = mc\Delta T$ $Q = (500)(.444)(-180)$ $-39960J$

Where did that energy go? Is this process endothermic or exothermic? How can you tell by your calculation?

- Energy is transferred to the surroundings
- Value is negative = exothermic

4. A diamond has a specific heat of 0.502 J/g °C. How much energy is needed to raise the temperature of an 85 g diamond from 0 °C to 92 °C?

$Q = mc\Delta T$
$\Delta T = 92 - 0 = 92$
$m = 85g$
$c = 0.502$

$Q = (85)(.502)(92)$
$Q = 3925.64J$

Remember:
Endothermic ↑ Pos = adding Energy
Neg = releasing Energy ↓ Exothermic