

Specific Heat Capacity (c)

- the quantity of energy, in Joules (J), needed to change the temperature of one gram (g) of a substance by one degree Celsius ($^{\circ}\text{C}$).
- c** values are on p. 632 (will be provided on the public exam)

eg. **c** for water is $4.184 \text{ J/g} \cdot ^{\circ}\text{C}$

Formula: $q = mc\Delta T$ $q =$ heat lost or gained $\Delta T =$ change in temperature
 $m =$ mass of object heated/cooled $\Delta T = T_2 - T_1$
 $c =$ specific heat capacity

- You will be given all but one of the above variables and asked to find the missing variable.

Examples:

1. A student must use 225 mL of hot water in a lab procedure. Calculate the amount of heat required to raise the temperature of 225 mL of water from 20.0°C to 100.0°C .

Solution:

- since the density of water is 1.00 g/mL , the mass of 225 mL of water is 225 g
 - **c** for water is $4.184 \text{ J/g} \cdot ^{\circ}\text{C}$

$$q = mc\Delta T$$

$$q = (225 \text{ g})(4.184 \text{ J/g} \cdot ^{\circ}\text{C})(100.0^{\circ}\text{C} - 20.0^{\circ}\text{C})$$

$$= 755312 \text{ J}$$

$$= 75.5 \text{ kJ}$$

2. Calculate the specific heat capacity of a new alloy if a 15.4 g sample absorbs 393 J when it is heated from 0.0°C to 37.6°C .

Solution:

$$m = 15.4 \text{ g} \qquad T_2 = 37.6^{\circ}\text{C}$$

$$q = 393 \text{ J} \qquad T_1 = 0.0^{\circ}\text{C}$$

$$q = mc\Delta T$$

$$393 \text{ J} = (15.4 \text{ g})(c)(37.6^{\circ}\text{C} - 0.0^{\circ}\text{C})$$

$$393 \text{ J} = (579.04 \text{ g} \cdot ^{\circ}\text{C}) \times (c)$$

$$c = 0.679 \text{ J/g} \cdot ^{\circ}\text{C}$$

3. A 40.0 g sample of ethanol releases 2952 J as it cools from 50.0°C . Calculate the final temperature of the ethanol.

Solution:

Use 2 steps:

- First solve for ΔT

$$m = 40.0 \text{ g} \qquad T_1 = 50.0^{\circ}\text{C}$$

$$q = -2952 \text{ J} \qquad c = 2.46 \text{ J/g} \cdot ^{\circ}\text{C}$$

$$q = mc\Delta T$$

$$-2952 \text{ J} = (40.0 \text{ g})(2.46 \text{ J/g} \cdot ^{\circ}\text{C})(\Delta T)$$

$$-2952 \text{ J} = (98.4 \text{ J}^{\circ}\text{C})(\Delta T)$$

$$\Delta T = -30^{\circ}\text{C}$$

- Next find the final temperature

$$\Delta T = T_2 - T_1$$

$$-30^{\circ}\text{C} = T_2 - 50^{\circ}\text{C}$$

$$T_2 = 20^{\circ}\text{C}$$

Exercises: (Be careful with positive and negative signs!!)

1. Calculate the heat change involved when 2.00 L of water is heated from 20.0°C to 99.7°C in an electric kettle. (**667 kJ**)

2. Calculate the heat change associated with cooling a 350.0 g aluminum bar from 70.0°C to 25.0°C. Is the change endothermic or exothermic? Why? **(-14.2 kJ)**
3. A 175 g piece of iron and a 175 g piece of aluminum are placed in a hot water bath so that they are warmed to 99.7°C. The metal samples are removed and cooled to 21.5°C. Which sample undergoes the greater heat change? **(Al; -12.3 kJ Fe; - 6.08 kJ)**
4. A 63.5 g sample of an unidentified metal absorbs 355 J of heat when its temperature changes by 4.56°C. Calculate the specific heat capacity of the metal. **(1.23 J/g °C)**
5. Calculate the specific heat capacity of titanium if a 43.56 g sample absorbs 0.476 kJ as its temperature changes from 20.13°C to 41.06°C. **(0.522 J/g °C)**
6. The burning of a sample of propane generated 104.6 kJ of heat. All of this heat was used to heat 500.0 g of water that had an initial temperature of 20.0°C. What was the final temperature of the water? **(70.0 °C)**
7. 750.0 g of water that was just boiled (heated to 100.0 °C) loses 78.45 kJ of heat as it cools. What is the final temperature of the water? **(75.0 °C)**