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How many joules of heat energy are required to heat 1.0 gallons of water from 20°C to 85°C?

$$1.0 \text{ gallons} \times \frac{3.7 \text{ L}}{1 \text{ gal}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ g H}_2\text{O}}{1 \text{ mL H}_2\text{O}} = 3700 \text{ g H}_2\text{O}$$

$$q = (3700 \text{ g})(4.184 \frac{\text{J}}{\text{g K}})(85^\circ\text{C} - 20^\circ\text{C}) = 1.01 \times 10^6 \text{ J}$$

You have 125.0 grams of gold, that has a specific heat capacity of 0.13 J/g•K, and it is at a temperature of 37°C (i.e. a gold necklace that you just took off). You use a hair dryer to blow hot air on the gold, eventually adding a total of 3.2 kJ of heat to the gold. What is the temperature of the gold after this amount of heat energy has been added?

$$3200 \text{ J} = (125 \text{ g})(0.13 \frac{\text{J}}{\text{g K}}) \Delta T$$

$$\Delta T = 196.9 \text{ K}$$

$$37^\circ\text{C} = \overset{310}{\cancel{1000}} \text{ K} \rightarrow \Delta T = T_f - T_i$$

$$196.9 = T_f - \overset{310}{\cancel{1000}} \text{ K}$$

$$T_f = 506.9 \text{ K or } 233.6^\circ\text{C}$$

If the addition of 735 J of heat to 200.0 grams of a liquid causes its temperature to go from 31.5°C to 34.2°C, what is the specific heat capacity of that liquid?

$$\Delta T = 34.2^\circ\text{C} - 31.5^\circ\text{C} = 2.7^\circ\text{C}$$
$$735 \text{ J} = (200.0 \text{ g})(C_p)(2.7^\circ\text{C})$$

$$C_p = 1.36 \frac{\text{J}}{\text{g K}}$$

C or °C

The temperature inside your freezer is -20°C . You remove a chunk of ice from the freezer that weighs 750 grams. How much energy is required to turn all this ice into gaseous water?

$$750\text{g} \times 4.184 \frac{\text{J}}{\text{g}\cdot\text{K}} (0^{\circ}\text{C} - -20^{\circ}\text{C}) = 62760 \text{ J to warm ice to } 0^{\circ}\text{C}$$

$$750\text{g} \times 333 \frac{\text{J}}{\text{g}} = 249750 \text{ J to melt ice}$$

$$750\text{g} \times 4.184 \frac{\text{J}}{\text{g}\cdot\text{K}} (100^{\circ}\text{C} - 0^{\circ}\text{C}) = 313800 \text{ J to warm water to } 100^{\circ}\text{C}$$

$$750\text{g} \times 2256 \frac{\text{J}}{\text{g}} = 1692000 \text{ J to evaporate} \quad \text{Total} = 2,318,310 \text{ J}$$

You have a cube of iron (Fe) that measures 1.35 inches on a side. The density of iron is $7.86\text{g}/\text{cm}^3$ and the specific heat capacity of iron is $0.45 \text{ J}/\text{g}\cdot\text{K}$. Using a furnace, you heat the cube of iron to a temperature of 750°C and then drop it into a bucket containing 2.5 L of water originally at room temperature (20°C). What will the temperature of the water be when everything has equilibrated?

$$1.35 \text{ inches} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 3.43 \text{ cm} \leftarrow \text{length of side of cube}$$

$$\text{Volume of cube} = (3.43 \text{ cm})^3 = 40.35 \text{ cm}^3$$

$$\text{Mass of cube} = 40.35 \text{ cm}^3 \times \frac{7.86 \text{ g}}{\text{cm}^3} = 317.2 \text{ g Fe}$$

$$q_1 + q_2 = 0$$

$$\left[(317.2 \text{ g}) \left(0.45 \frac{\text{J}}{\text{g}\cdot\text{K}} \right) (T_f - 750^{\circ}\text{C}) \right] + \left[(2500 \text{ g}) \left(4.184 \frac{\text{J}}{\text{g}\cdot\text{K}} \right) (T_f - 20^{\circ}\text{C}) \right] = 0$$

$$142.745 T_f - 107055 \text{ J} + 10460 T_f - 209200 \text{ J} = 0$$

$$10603.5 T_f = 316255 \text{ J}$$

$$T_f = 29.8^{\circ}\text{C}$$