Unit 2 Matter and Energy **Name ______________________ Block ____**
Specific Heat Practice Problems

1) Calculate the **amount of energy** required ***to raise*** the temperature of **73.2 grams** of water from **23.1˚C** to **98.6˚C**. (Specific Heat of Water = **4.184 J/ g·°C**)

Raising the temperature means heat is absorbed, so Q is positive.

$$Q = m \cdot C_p \cdot (T_{final} - T_{initial})$$

$$Q = (73.2\ g) \cdot (4.184\ \frac{J}{g\ °C}) \cdot (98.6\ ℃ - 23.1\ ℃)$$

$$Q = +23123.2944\ J \qquad \text{sig figs} \rightarrow\ +23,100\ J\ \ or\ \ +23.1\ kJ$$

2) The **specific heat capacity of copper is 0.3845 J/ g·°C**. If **245 Joules of energy** is required to ***raise the temperature*** of a sample of copper from **25.0˚C** to **50.0˚C**, what is the **mass** of the sample?

$$m = \frac{Q}{C_p \cdot (T_{final} - T_{initial})}$$

$$m = \frac{+245\ J}{(0.3845\ \frac{J}{g\ °C}) \cdot (50.0\ ℃ - 25.0\ ℃)}$$

$$m = 25.48764629 \quad sig\ figs \rightarrow\ \ 25.5\ g$$

3) A **30.0 gram** sample requires **930 Joules of energy** to ***increase the temperature*** **by 25˚C**. What is the **specific heat capacit**y of the sample? Identify the sample using the table to the right.

$$C_p = \frac{Q}{m \cdot \Delta T}$$

$$C_p = \frac{+930\ J}{(30.0\ g) \cdot (25\ ℃)}$$

$$C_p = 1.24\ \frac{J}{g\ °C} \quad sig\ figs \rightarrow\ 1.2\ \frac{J}{g\ °C}$$

Element	Specific Heat (J/g·°C)
Beryllium	1.824
Sodium	1.224
Magnesium	1.024
Aluminum	0.9025
Calcium	0.6315

Note: ΔT means the CHANGE in Temperature. When using a Δ symbol, always measure the change by calculating the FINAL minus the INITIAL.

4) A **10.0 g brass** sample is **heated to 82.3°C** and dropped into a calorimeter containing **25.0 g of liquid water** at **22.0°C**. The accepted **specific heat of brass is known to be 0.377 J/g·°C.** Calculate the final temperature of the mixture? (The **specific heat of water is 4.184 J/g·°C.**)

> Note: Based on the ***Law of Energy Conservation***, the amount of HEAT absorbed **(+Q)** by the water is equal to the amount of HEAT transferred **(-Q)** by the metal.

$$+Q_{water} = -Q_{brass}$$

w = water
b = brass

$$+[m_w \cdot C_{p_w} \cdot (T_{final_w} - T_{initial_w})] = -[m_b \cdot C_{p_b} \cdot (T_{final_b} - T_{initial_b})]$$

$$+\left[(25.0\ g) \cdot (4.184\ \frac{J}{g\ °C}) \cdot (T_{final_w} - 22.0℃)\right] = -\left[(10.0\ g) \cdot (0.377\ \frac{J}{g\ °C}) \cdot (T_{final_b} - 82.3℃)\right]$$

> **The Key to solving this problem is to recognize that** $T_{final_w} = T_{final_b}$.

$$104.6\ \frac{J}{°C} \cdot (T_{final} - 22.0℃) = -3.77\ \frac{J}{°C} \cdot (T_{final} - 82.3℃)$$

$$104.6\ T_{final} - 2301.2 = -3.77\ T_{final} + 310.271$$

$$+3.77\ T_{final} \qquad = +3.77\ T_{final}$$

$$108.37\ T_{final} - 2301.2 = +310.271$$

$$+2301.2 = +2301.2$$

$$\frac{108.37\ T_{final}}{108.37} = \frac{2611.471}{108.37}$$

$$T_{final} = 24.09773\ ℃$$

$$with\ proper\ sig\ figs\ = 24.1℃$$