Heat and Temperature:

- Heat is **Thermal Energy** transferred from a **HOT** system to a **COLD** system that are in contact (e.g., convection currents & conduction).
- Temperature is a measure of the **average kinetic energy** of the atoms or molecules in the system.

Specific Heat Capacity:

- In the formula for calculating heat energy, there is a variable called the **specific heat capacity** symbol **c**.
- Every substance has its own **specific heat capacity** value.
- Specific heat capacity is a **physical property** of the material.
- Different substances have **different** specific heat capacities.
- By definition, it is the amount of **ENERGY** needed to raise the temperature of **1 g** of a substance by **1 °C = specific heat capacity (c)**.
- Water has a **very high** specific heat capacity.
  - It takes **4.18 J/g** to raise the temperature of 1 g of 1 ml of water by 1 °C. (c = 4.18 J/g)
- Most metals have much **lower** specific heats. The c value of Fe is **0.385 J/g °C**.

**Example:** When I cook a hard-boiled egg in the morning, sometimes I place water in a pot but forget the egg. If I remember within 30 seconds or so, it is still safe to place my hand in the water, but it would be a bad idea to touch the pot itself.

- Metals warm up **FASTER** than water does—so that’s what a high specific heat implies.
- The **higher** the c value, the **more** energy it is to warm up that substance (water has a higher c value than the metal).
- By the same idea, high specific heat substances also **lose** their heat slowly, while metals cool off **quickly** because they have low specific heat capacities.

You will **always** be given the specific heat capacity of any substance you need (unless you are asked to calculate it with other information).
### Calculating Heat Energy

We can calculate the heat released or absorbed using the specific heat capacity, the mass of the substance, and the change in temperature.

\[ Q = mc \Delta T \]

- **Q** is the **quantity** of **heat** in **J**.
- **m** is the **mass** of the substance in **kg**.
- **c** is the **specific heat capacity** in **J/(kg°C)**.
- **\Delta T** is the **temperature change** in **°C**.

**Tips for Using the Formula**

- After you find \( Q \), there is only one operation in this formula (multiply). So there is no order of operations to worry about.

#### Practice

**How much energy is needed to warm up 300.0 kg of water from 10.0°C to a comfortable 37.0°C?**

- **m** = 300.0 kg
- \( \Delta T = 37.0°C - 10.0°C = 27.0°C \)
- **\( c_{\text{water}} = 4.182 \text{ J/(g°C)} \)**

\[ Q = m \cdot c \cdot \Delta T \]

\[ Q = (300.000 \text{ kg}) \cdot (4.182 \text{ J/(g°C)}) \cdot (27.0°C) \]

\[ Q = 3.39 \times 10^7 \text{ J} \]

**Practice**: A water tank contains 200.0 kg of water. The water is heated. How much energy is required to raise the temperature of the water from 15.0°C to 60.0°C?

- **m** = 200.0 kg
- \( \Delta T = 60.0°C - 15.0°C = 45.0°C \)
- **\( c_{\text{water}} = 4.182 \text{ J/(g°C)} \)**

\[ Q = m \cdot c \cdot \Delta T \]

\[ Q = (200.000 \text{ kg}) \cdot (4.182 \text{ J/(g°C)}) \cdot (45.0°C) \]

\[ Q = 3.76 \times 10^6 \text{ J} \]
Other Versions of the Same Formula

- The formula can be changed around to solve for mass or specific heat or \( \Delta T \). 

\[
\begin{align*}
\text{mass:} & \quad m = \frac{Q}{c \Delta T} \\
\text{specific heat:} & \quad c = \frac{Q}{m \Delta T} \\
\text{temperature change:} & \quad \Delta T = \frac{Q}{mc}
\end{align*}
\]

- For all of these formulas, you will find it easiest to solve for the denominator (bottom of the fraction) first, and then divide the \( Q \) by that result.

Practice: What final temperature will be attained by a 250.0 g sample of Cu at 10.0°C if it absorbs 1000.0 J of heat? 

\[
c_{\text{Cu}} = 0.385 \text{ J/(g°C)}
\]

\[
\begin{align*}
\Delta T &= \frac{Q}{mc} \\
\Delta T &= \frac{1000.0 \text{ J}}{250.0 \text{ g} \times 0.385 \text{ J/(g°C)}} \\
\Delta T &= 10.4 \text{ °C}
\end{align*}
\]

\[
\Delta T = T_f - T_i = 10.4 \text{ °C} + 10.0 \text{ °C} = 20.4 \text{ °C}
\]

Practice: 800.0 kJ were absorbed by a pond, sending its temperature rising from 20.0°C to 25.0°C. How much water was in the pond? 

\[
c_{\text{H}_{2}\text{O}} = 4.182 \text{ J/(g°C)}
\]

\[
\begin{align*}
\Delta T &= T_f - T_i = 25.0 \text{ °C} - 20.0 \text{ °C} = 5.0 \text{ °C} \\
m &= \frac{Q}{c \Delta T} \\
m &= \frac{800.0 \times 10^3 \text{ J}}{4.182 \times 20.91 \text{ J/(g°C)}} \\
m &= 38,259 \text{ g} \\
&= 38 \text{ kg}
\end{align*}
\]
Mixing Problems

- Assuming a well-insulated system, the amount of heat a hot object ______ can be assumed to be ____________ by a colder object (any heat lost in the system must be gained by the ____________________).
- Mathematically, these quantities can only become equal if an _________________ is inserted.

Practice: What final temperature will be attained if 300.0 g of 30.0°C water is mixed with an equal mass of 66.0°C alcohol? (c\text{H}_{2}O = 4.182 \text{ J} / (\text{g} ^\circ \text{C}) \text{ and } c_{\text{alcohol}} = 2.440 \text{ J} / (\text{g} \circ \text{C})). Comment on why the final temperature of the mixture is NOT average of the two temperatures.
1. 800.0 g of water is warmed from 10.0°C to 80.0°C. How much energy was absorbed by the water? \( (c_{\text{water}} = 4.182 \text{ J/(g°C)}) \)

2. 700.0 g of water is allowed to cool from its boiling point to 20.0°C. How much energy in was released into the room? \( (c_{\text{water}} = 4.182 \text{ J/(g°C)}) \)

3. How much heat energy is required to warm 3.5 kg of water from 16°C to 66°C? 
\( (c_{\text{water}} = 4.182 \text{ J/(g°C)}) \)
4. How much energy must a heater supply in order for 200 kg of bathwater to warm up from 10.0°C to our body temperature of 37.0°C? (c_{water} = 4.182 J/(g°C))

5. If 9000.0 J of heat is absorbed by 800.0 g of water at 5.0°C, what will be its final temperature? (c_{water} = 4.182 J/(g°C))

6. Find the specific heat of a material that lost 41,900 J of energy when 200.0 g of the material went down 50.0°C in temperature.
7. If a pile of snow with a mass of 525 kg loses 740,000 J of thermal energy during the night, how much will the temperature of the snow drop? \( c_{snow} = 2.090 \text{ J/(g C)} \)

8. What is the final temperature of 0.63 kg of water that releases 2290 joules of thermal energy? The water had an initial temperature of 48.2°C. \( c_{water} = 4.182 \text{ J/(g C)} \)

9. If one bar of copper and one bar of gold have equal masses of 250 g, and have equal starting temperatures, how much will their final temperatures differ by if 65,000 J of thermal energy is added to each of the bars? \( c_{Cu} = 0.385 \text{ J/(g C)} \) & \( c_{Au} = 0.129 \text{ J/(g C)} \)
10. What is the mass of a lead block, if it takes 67,000 J to raise the temperature by 100.0°C?  
\(c_p = 0.129 \text{ J/(g \cdot C)}\)

11. a) In his part time job as a blacksmith, Nathan was quenching (cooling) a 350 g iron horseshoe from 470.0°C to 30.0°C by plunging the horseshoe into a bucket of cold water. How much thermal energy was lost by the horseshoe?  
\(c_p = 0.449 \text{ J/(g \cdot C)}\)

b) If the bucket of water that Nathan used contained 14.0 kg of water that had an initial temperature of 25.0°C, how much would the water temperature increase when cooling the horseshoe?  
\(c_{\text{H2O}} = 4.182 \text{ J/(g \cdot C)}\)
12. 250.0 g of water at 24.0 °C was heated with 74,700 J of energy. What is the final temperature of the water? \( (c_{\text{H}_2\text{O}} = 4.182 \text{ J/(g °C)}) \)

13. What mass of copper, originally at 50.0°C, must be added to 1.0 kg of 10.0°C water to raise its temperature to 20.0°C? \( (c_{\text{Cu}} = 0.385 \text{ J/(g °C)} \) & \( c_{\text{H}_2\text{O}} = 4.182 \text{ J/(g °C)} \)

14. A 450 g sample of water is originally at 25.0°C. How cold will it get if we add 300.0 g of 0.5°C water to that sample? \( (c_{\text{H}_2\text{O}} = 4.182 \text{ J/(g °C)}) \)