

Specific Heat Capacity, c

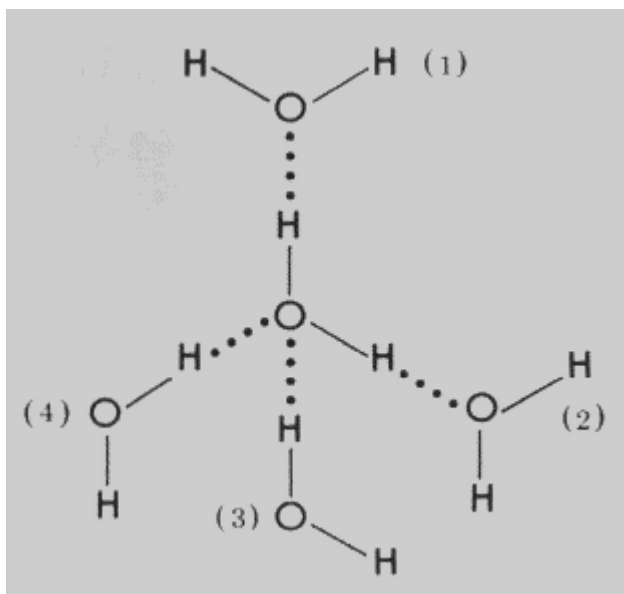
The specific heat capacity of a substance, c , is the amount of energy needed to raise the temperature of 1 g of a substance by 1 °C.

Specific heat capacity is a physical characteristic property. Different substances have different specific heats. Water has a very high specific heat: it takes 4.19 J to raise the temperature of 1 g or 1 ml of water by 1 °C. Most metals, on the other hand, have much lower specific heats. Cu's c value is only 0.39 J/(g C).

When I get hard boiled eggs ready for my kids in the morning, I sometimes 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 difficult it is to warm up that substance. By the same token, high specific heat substances also lose their heat slowly, while metals cool off quickly.

The high specific heat capacity of water helps temper the rate at which air changes temperature, which is why temperature changes between seasons is gradual, especially near large lakes or the ocean. Water is the reason why Toronto is milder than Montreal.

If water's specific heat was lower than it actually is, life would not be possible. The rate of evaporation would be too high, and it would be too difficult for the evolutionary precursors of cells to maintain homeostasis.



Why is water special? The reason it is difficult to raise water's temperature is that *hydrogen bonds* exist between molecules of water. The hydrogens of one molecule are attracted to the oxygen atom of another molecule. (see the dots = '••••' in the adjacent diagram. Each set represent a hydrogen bond.) To overcome this attraction, energy is needed. It's as if a stiff spring existed between different molecules and a lot of strength is needed to make that spring move or break. As water cools, the bonds reform and release energy, so it is also difficult for water to lose its heat.

A very useful formula allows us to calculate the amount of heat either absorbed or lost by a substance during a physical change.

$$Q = m c \Delta T$$

Q = quantity of heat in joules (J)

m = mass of the substance acting as the environment in grams (g)

c = specific heat capacity (4.19 for H₂O) in J/(g °C)

ΔT = change in temperature = T_{final} - T_{initial} in °C

Example 1

How much energy is needed to warm up 300 kg of water from 10 °C to a comfortable 37 °C?

Note 300 kg = 300 000 g

$$Q = m c \Delta T$$

$$= 300\,000(4.19)(37 - 10)$$

$$= 33\,939\,000 \text{ J}$$

$$= 33\,939 \text{ kJ}$$

Example 2

a. What final temperature will be attained by a 250 mL cup of 10 °C water if it absorbs 1000 J of heat?

Note, because water's density = 1 g/mL, 250 mL = 250 g

$$Q = m c \Delta T$$

$$1000 \text{ J} = 250(4.19)(x - 10)$$

$$x - 10 = 1000 / 250 / 4.19$$

$$x - 10 = 0.95$$

$$x = 10.95 \text{ °C}$$

b. What final temperature will be attained by a 250 g sample of Cu at 10 °C if it absorbs 1000 J of heat? Cu's c = 0.39 J/(g °C)

$$Q = m c \Delta T$$

$$1000 \text{ J} = 250(0.39)(x - 10)$$

$$x - 10 = 1000 / 250 / 0.39$$

$$x - 10 = 10.25$$

$x = 20.25 \text{ °C}$ Notice how the same amount of energy makes the same mass of Cu go up by 10 degrees, not just 0.95 degree as in water's case.

