University of Science and Technology

Faculty of Engineering

First Year - Semester One - Academic Year:2025-2026

Physics (1)

Lec (8)

Specific Heat Capacity

SOME SUBSTANCES CHANGE TEMPERATURE MORE EASILY THAN OTHERS.

- **SPECIFIC HEAT:** the amount of energy required to raise the temperature of 1 gram of a substance by 1°C

. Definition of Specific Heat Capacity

Specific heat capacity is **the amount of heat required to raise the temperature of a unit mass of a substance by one degree**.

- In other words, it measures **how much energy a material can absorb before its temperature rises**.
- Each material has a different specific heat capacity.

Mathematical Formula

$$Q = m \cdot c \cdot \Delta T$$

Where:

- Q = heat energy absorbed or released (Joules or calories)
- m = mass of the substance (kg or g)
- c = specific heat capacity (J/kg·°C or cal/g·°C)
- ΔT = change in temperature (°C or K)

This formula shows that the **heat required** depends on **the mass of the object, the material type, and the temperature change**.

Units of Measurement

1. **SI Unit:**
   - Joule per kilogram per degree Celsius: **J/(kg·°C)**
2. **Traditional Unit:**
   - Calorie per gram per degree Celsius: **cal/(g·°C)**

Note: **1 cal ≈ 4.186 J**

Typical Specific Heat Capacities of Common Materials

Material	Specific Heat Capacity (J/kg·°C)	Specific Heat Capacity (cal/g·°C)
Water	4186	1
Copper	385	0.092
Iron	450	0.108
Aluminum	900	0.215
Wood	1700	0.406

Observation: Water has the **highest specific heat capacity** among common materials, which is why it **heats and cools slowly** compared to metals.

Practical Example : Suppose we have **2 kg of water at 25°C** and we want to raise its temperature to **75°C**.

$Q=m\cdot c\cdot\Delta T$ $\quad Q=2\cdot 4186\cdot(75-25)$ $\quad Q=2\cdot 4186x50$ $\quad Q=418{,}600J$

- We need **418.6 kJ of energy** to heat the water.

Insight: If we used copper instead, it would require much less heat due to its lower specific heat capacity.

Why Specific Heat Capacity Matters

1. **Material Selection:**
   - Materials with high specific heat (like water) are used in cooling systems, e.g., car radiators, because they **absorb and store heat efficiently**.
2. **Daily Applications:**
   - Cooking: Water takes longer to heat than oil because of its higher specific heat.
   - Climate moderation: Oceans regulate temperature because water absorbs heat without large temperature changes.
3. **Industrial Applications:**
   - Heat exchangers, steam systems, and energy storage all rely on knowledge of specific heat capacity.

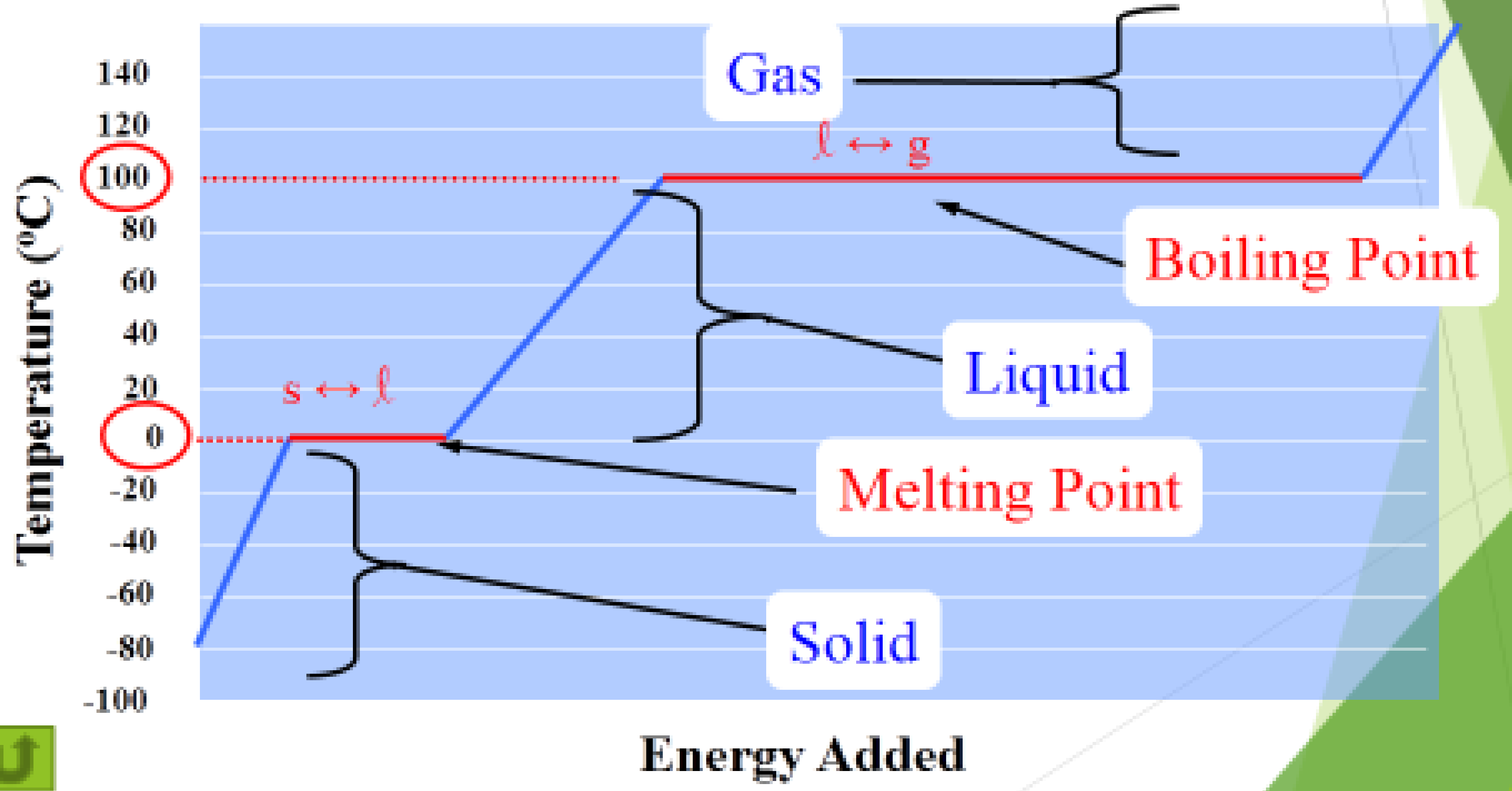
Extended Example: Comparing Materials

Suppose we have 1 kg of **water** and 1 kg of **copper**, both starting at 20°C. We supply **100 kJ of heat** to each.

- Water: $\Delta T = Q/m \cdot c = 100{,}000/[1 \times 4186] \approx 23.9°C$
- Copper: $\Delta T = 100{,}000/[1 \times 385] \approx 259.7°C$

Conclusion: Copper heats up much faster than water under the same heat input because it has a much lower specific heat capacity.

Heating Curve of water



Definition of Latent Heat : Latent heat is **the amount of heat energy required to change the state of a substance without changing its temperature**.

- This energy is absorbed or released **during a phase change**.
- Common phase changes include:
    - Melting (Solid → Liquid)
    - Freezing (Liquid → Solid)
    - Boiling/Evaporation (Liquid → Gas)
    - Condensation (Gas → Liquid)

Formula for Latent Heat

$$Q = m \cdot L$$

Where:

- Q = heat energy absorbed or released (Joules or calories)
- m = mass of the substance (kg or g)
- L = latent heat of the substance (J/kg or cal/g)

Note:

- L_f = Latent Heat of Fusion (for melting/freezing)
- L_v = Latent Heat of Vaporization (for boiling/condensation)

Types of Latent Heat

1. Latent Heat of Fusion (L_f)

- Heat required to convert **1 kg of a solid to a liquid** at the same temperature.
- Example: Melting 1 kg of ice at 0°C requires approximately **334 kJ**.

2. Latent Heat of Vaporization (L_v)

- Heat required to convert **1 kg of liquid to gas** at its boiling point without temperature change.
- Example: Boiling 1 kg of water at 100°C requires approximately **2260 kJ**.

. Practical Examples

Example 1: Melting Ice

- Mass of ice: 0.5 kg
- Latent heat of fusion of ice: 334 kJ/kg

$Q=m\cdot L_f=0.5x334=167$ kJ

- **Interpretation:** 0.5 kg of ice absorbs 167 kJ to melt completely at 0°C, temperature remains constant.

Example 2: Boiling Water

- Mass of water: 0.2 kg
- Latent heat of vaporization: 2260 kJ/kg

$$Q = m \cdot L_v = 0.2 \times 2260 = 452\ \text{kJ}$$

- **Interpretation:** 0.2 kg of water absorbs 452 kJ to turn completely into steam at 100°C, temperature stays 100°C.

Example 3: Condensing Steam

- Mass of steam: 0.3 kg
- Latent heat of vaporization: 2260 kJ/kg

$$Q = m \cdot L_v = 0.3 \cdot 2260 = 678\ \text{kJ}$$

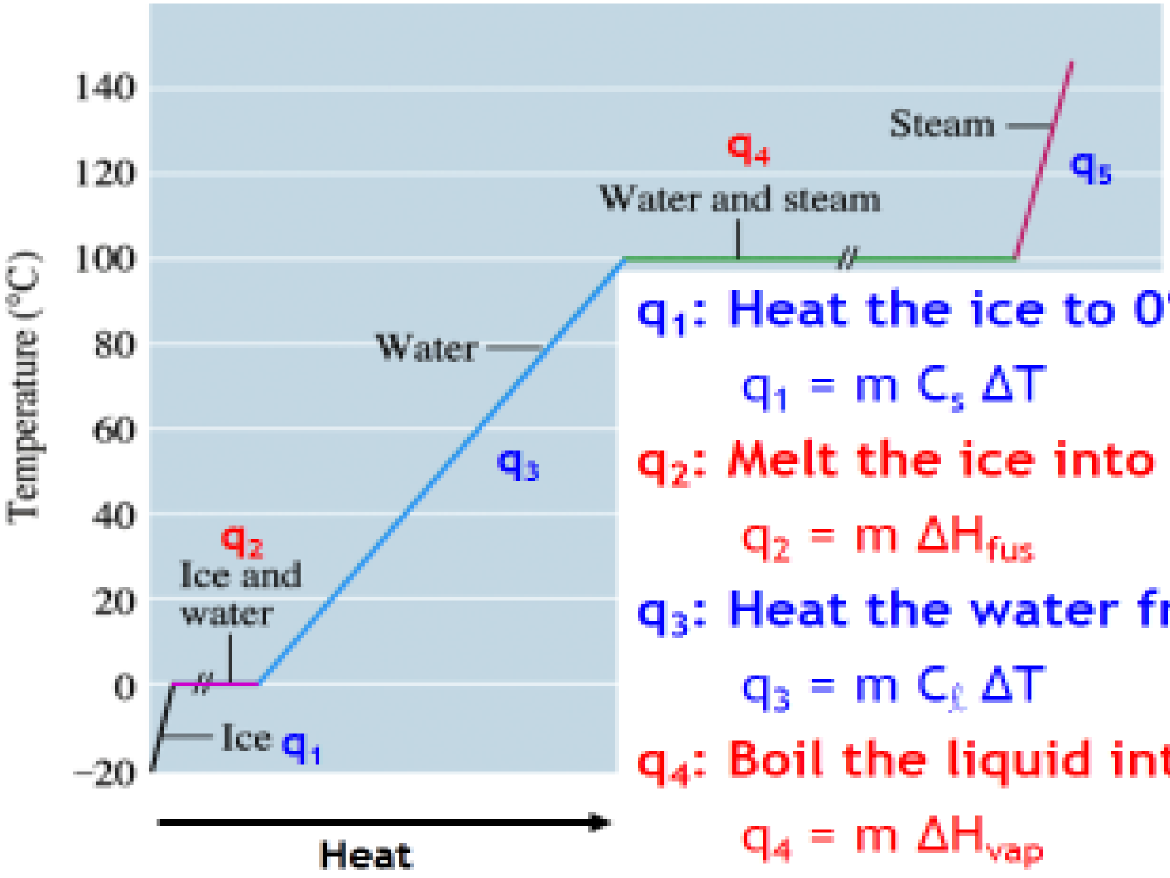
- **Interpretation:** Steam releases 678 kJ to condense into liquid water at 100°C.

Important Notes

- During phase change, **temperature remains constant**, even though heat is absorbed or released.
- Different substances have different latent heats:
    - Water has a **high latent heat**, so it **heats and cools slowly**.
    - Metals have low latent heat, so they **heat up and cool down quickly**.

Heating Curve of Water

From Ice to Steam in Five Easy Steps



q_1: Heat the ice to 0°C

$q_1 = m\,C_s\,\Delta T$

q_2: Melt the ice into a liquid at 0°C

$q_2 = m\,\Delta H_{fus}$

q_3: Heat the water from 0°C to 100°C

$q_3 = m\,C_\ell\,\Delta T$

q_4: Boil the liquid into a gas at 100°C

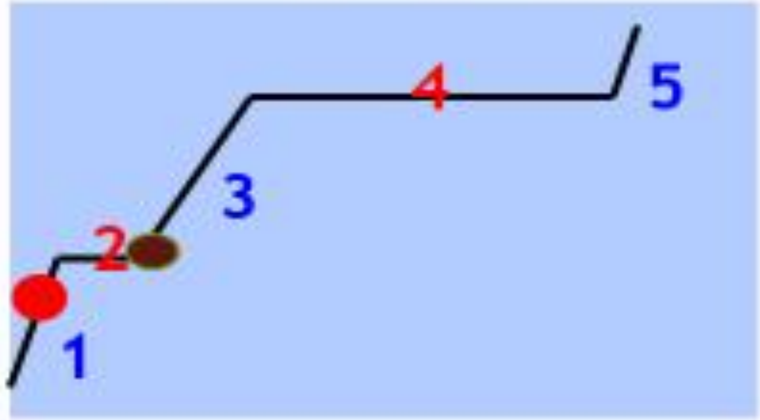
$q_4 = m\,\Delta H_{vap}$

q_5: Heat the gas above 100°C

$q_5 = m\,C_g\,\Delta T$

$\boxed{q_{tot} = q_1 + q_2 + q_3 + q_4 + q_5}$

1. How much energy (J) is required to heat 12.5 g of ice at -10.0 °C to water at 0.0 °C?



Notice that your q values are positive because heat is added…

q_1: Heat the ice from -10 to 0°C

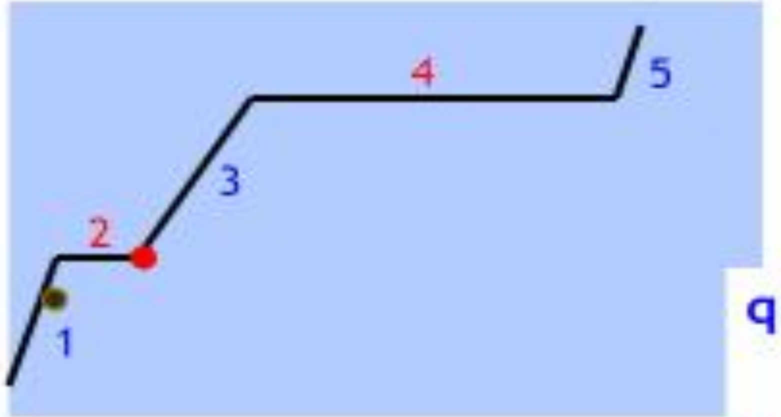
$$q_1 = 12.5\text{ g }(2.077\text{ J/g °C})(0.0 - -10.0\text{ °C}) = 259.63\text{ J}$$

q_2: Melt the ice at 0°C to liquid at 0 °C

$$q_2 = 12.5\text{ g }(333\text{ J/g}) = 4162.5\text{ J}$$

$$q_{tot} = q_1 + q_2 = 259.63\text{ J} + 4{,}162.5\text{ J} = \boxed{4{,}420\text{ J}}$$

2. How much energy (J) is required to heat 25.0 g of ice at -25.0 °C to water at 95.0 °C?



Notice that your q values are positive because heat is added...

q_1: Heat the ice from -25 to 0°C

$q_1 = 25.0\ g\ (2.077\ J/g\ °C)(0.0 - -25.0\ °C) = 1298.1\ J$

q_2: Melt the ice at 0°C to liquid at 0 °C

$q_2 = 25.0\ g\ (333\ J/g) = 8325\ J$

q_3: Heat the water from 0°C to 95 °C

$q_3 = 25.0\ g\ (4.184\ J/g\ °C)(95.0 - 0.0°C) = 9937\ J$

$q_{tot} = q_1 + q_2 + q_3 = 1298.1\ J + 8,325\ J + 9937\ J = 19,560\ J$

A 75.0 g piece of lead (specific heat = 0.130 J/g°C),initially at 435°C, is set into 125.0 g of water, initially at 23.0°C. What is the final temperature of the mixture?

q = m x C x ΔT for both cases, although specific values differ

Plug in known information for each side

$$q_{water} = -q_{Pb}$$

$$m_{water}\, C_{water}\, \Delta T_{water} = -m_{Pb}\, C_{Pb}\, \Delta T_{Pb}$$

$$125\ (4.18)\ (T_f - 23) = -75\ (0.13)\ (T_f - 435)$$

$$522.5\ T_f - \cancel{12017.5} = \cancel{-9.75\ T_f} + 4241.25$$

$$+9.75\ T_f \quad \cancel{+12017.5} \qquad \cancel{+9.75\ T_f} \quad +12017.5$$

$$532.25\ T_f = 16258.75$$

$$\boxed{T_f = 30.5^{o}C}$$