Thermal Equilibrium

Observation:

When two objects at different T are brought "together", they will eventually reach the same temperature and the system reaches an *equilibrium* state when no further physical changes occur in the system.

e.g. warm soda cans in a cooler filled with ice

Thermal Equilibrium

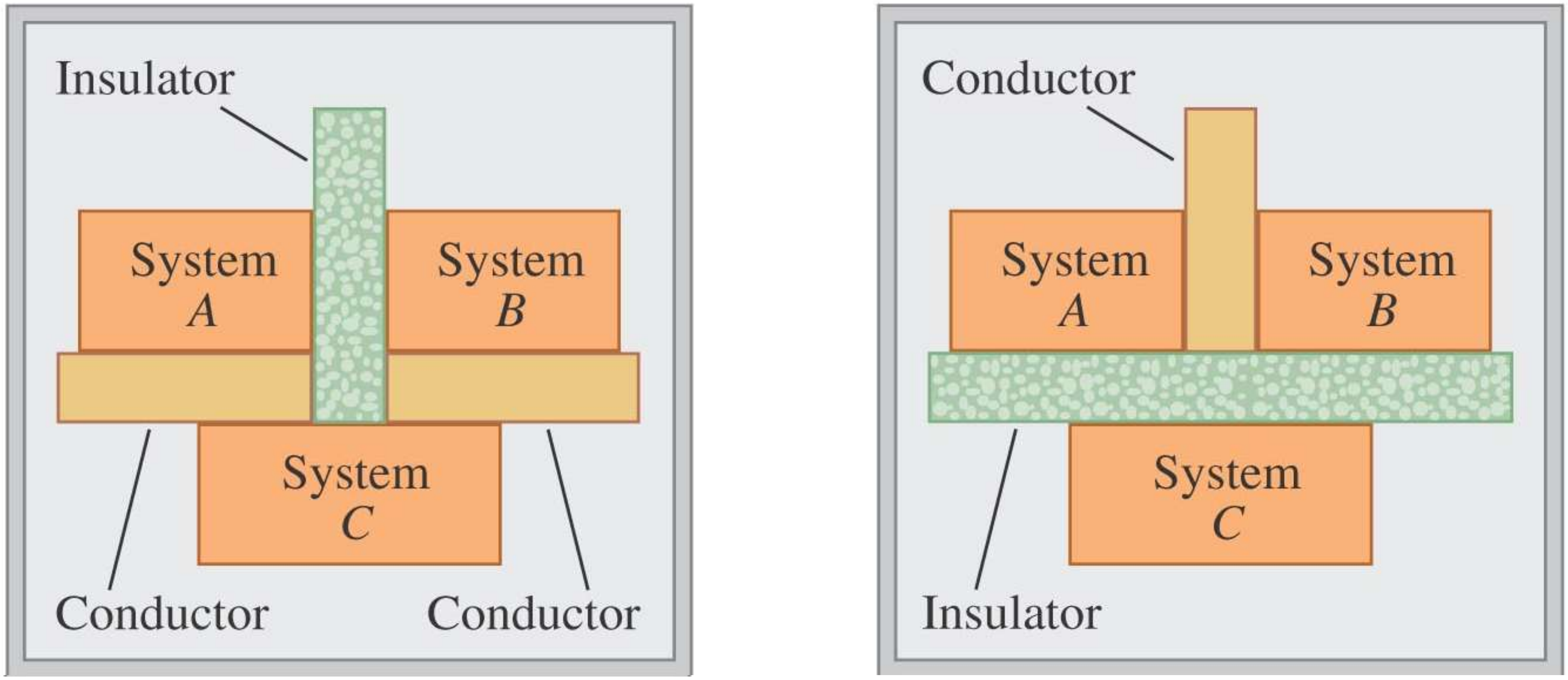
Additional concepts…

→ **Heat** (more on this later): the *transfer of energy* between objects with different *T*.

→ **Thermal Contact**: two objects are in thermal contact if heat can transfer between them (not necessary in physical contact).

→ **Thermal equilibrium**: the situation in which two objects in *thermal contact* cease to exchange energy by the process of *heat*.

The 0th Law of Thermodynamics



(a) If systems *A* and *B* are each in thermal equilibrium with system *C* …

(b) … then systems *A* and *B* are in thermal equilibrium with each other.

No heat flow between A-C and B-C.

No heat flow between A-B.

Note: this *transitive* property is not true for all physical process, e.g. two pieces of iron with a magnet but the 0th Law have been shown to be true experimentally.

Heat

Caution: in everyday usage, **Temp & Heat** are usually interchangeable.

But, in physics, they are **not** the same!

Temperature: a macroscopic state variable ~ avg. KE of molecules in the system (later).

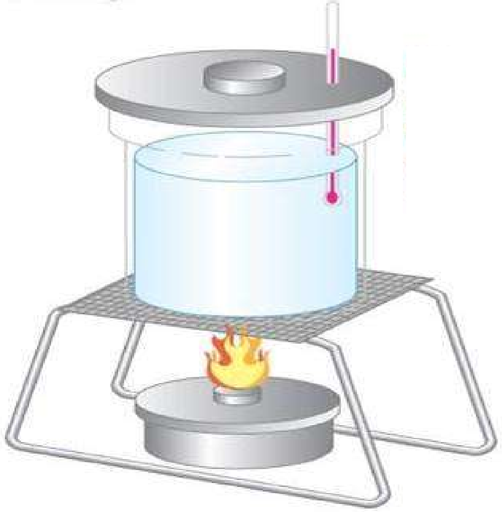
Heat: the *transfer of energy* between bodies due to a temperature difference.

Mechanical Equivalent of Heat (Sir James Joules, 1818-1889)

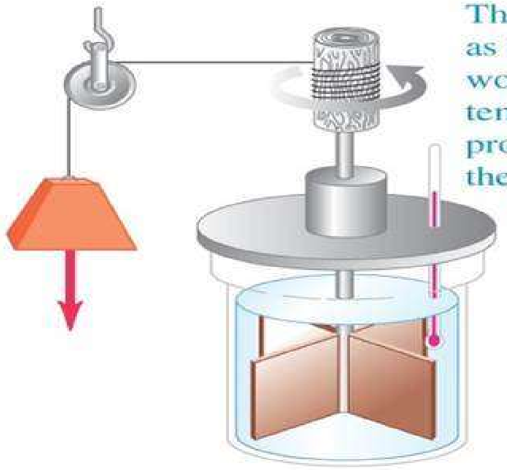
- Joule found that he can raise the water temperature by the same amount with either providing heat or by doing mechanical work !

- His conclusion: *mechanical work* and *heat* are equivalent in raising the temperature of the water.

Raising the temperature of water by direct heating



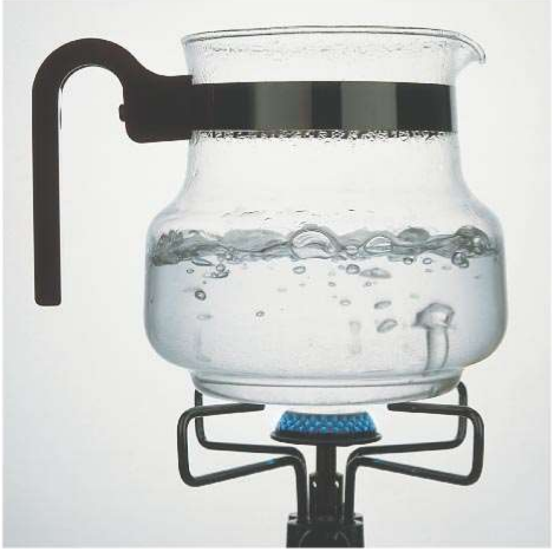
Raising the temperature of water by doing work on it



The water warms as the paddle does work on it; the temperature rise is proportional to the amount of work done.

D colliding balls

Quantity of Heat



Energy transfer due to temp diff is called **heat**.

Historically, the unit of heat is defined in terms of temp changes of *water*:

1 calorie (cal) = amount of energy transfer (heat) needed to raise the temp of 1g of water from 14.5^{o}C to 15.5^{o}C.

From Joules experiment, we now know that this amount of heat is equivalent to 4.186 J amount of mechanical energy, i.e.,

$$1\text{cal} = 4.186\text{ J}$$

Specific Heat/Heat Capacity

Different type of materials will need different amount of heat to raise its temp by $1^{o}C$ (or 1K).

We can quantify this using *specific heat c*:

$Q = mc\,\Delta T$ *or* $dQ = mc\,dT$

Q → quantity of heat needed to raise T from T_1 to T_2 $(\Delta T = T_2 - T_1)$

m → mass of the material

c → "specific heat" is characteristic of the type of material [$J/kg\cdot K$]

Specific Heat Values

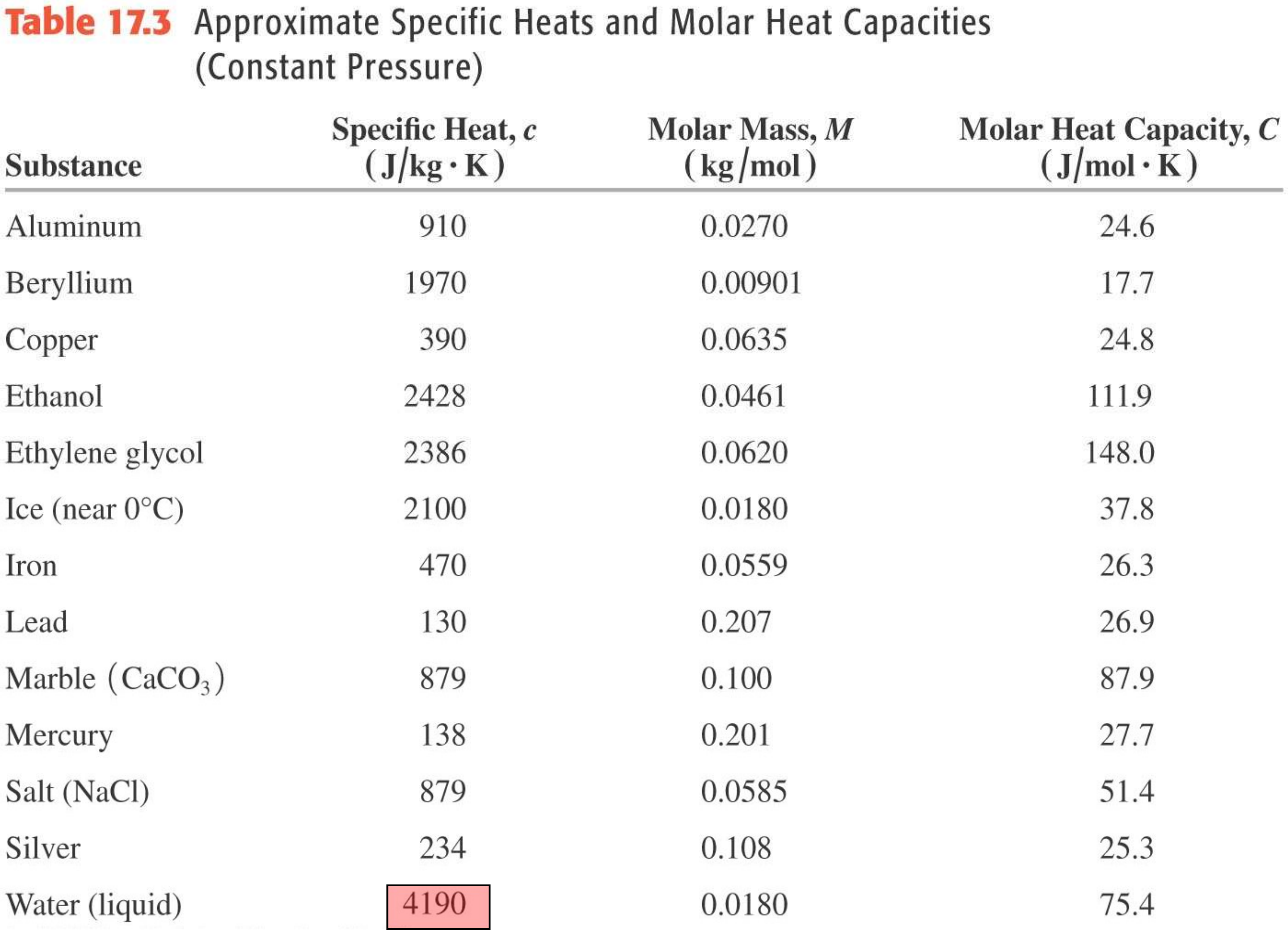


Table 17.3 Approximate Specific Heats and Molar Heat Capacities (Constant Pressure)

Substance	Specific Heat, c $(\mathrm{J/kg \cdot K})$	Molar Mass, M $(\mathrm{kg/mol})$	Molar Heat Capacity, C $(\mathrm{J/mol \cdot K})$
Aluminum	910	0.0270	24.6
Beryllium	1970	0.00901	17.7
Copper	390	0.0635	24.8
Ethanol	2428	0.0461	111.9
Ethylene glycol	2386	0.0620	148.0
Ice (near 0°C)	2100	0.0180	37.8
Iron	470	0.0559	26.3
Lead	130	0.207	26.9
Marble $(CaCO_3)$	879	0.100	87.9
Mercury	138	0.201	27.7
Salt (NaCl)	879	0.0585	51.4
Silver	234	0.108	25.3
Water (liquid)	4190	0.0180	75.4

Molar Specific Heat/Heat Capacity

- One can also specify the amount of materials by the number of molecules (or mole n) instead of its mass (m in kg). With $m=nM$,

$$Q = mc\,\Delta T = (nM)c\,\Delta T = nC\Delta T \qquad (note:\ cM = C)$$

n → number of mole

M → molar mass (mass per mole)

C → molar specific heat (note upper case)

$(1\ \text{mole} = 6.022\times10^{23}\ particles)$

Specific Heat Values

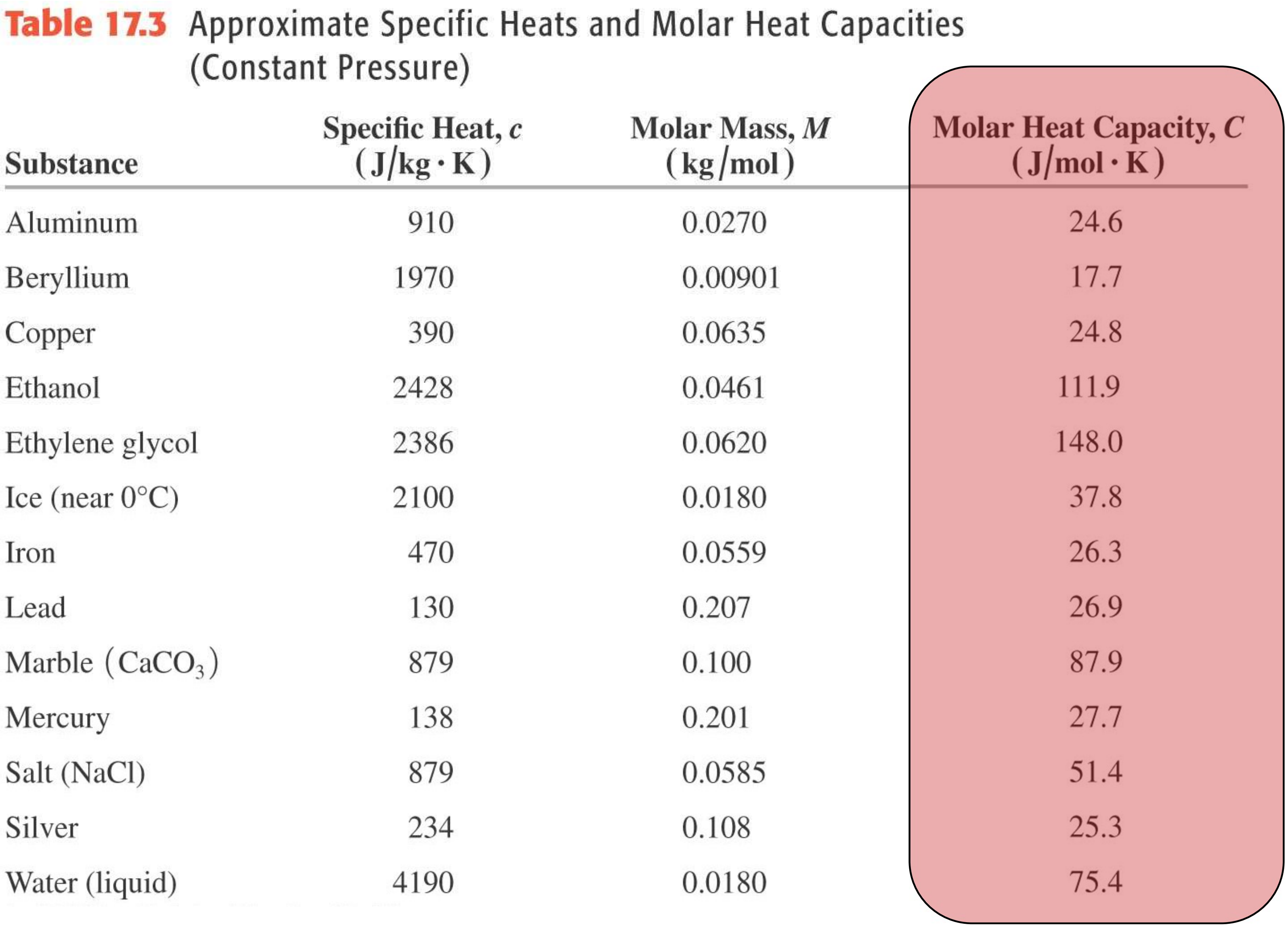


Table 17.3 Approximate Specific Heats and Molar Heat Capacities (Constant Pressure)

Substance	Specific Heat, *c* (J/kg · K)	Molar Mass, *M* (kg/mol)	Molar Heat Capacity, *C* (J/mol · K)
Aluminum	910	0.0270	24.6
Beryllium	1970	0.00901	17.7
Copper	390	0.0635	24.8
Ethanol	2428	0.0461	111.9
Ethylene glycol	2386	0.0620	148.0
Ice (near 0°C)	2100	0.0180	37.8
Iron	470	0.0559	26.3
Lead	130	0.207	26.9
Marble ($CaCO_3$)	879	0.100	87.9
Mercury	138	0.201	27.7
Salt (NaCl)	879	0.0585	51.4
Silver	234	0.108	25.3
Water (liquid)	4190	0.0180	75.4

Molar Specific Heat

- Specific heat also depends on the *process* by which heat is being transferred into the system
- Two often used molar specific heats:

  C_p : molar specific heat at *constant pressure*

  (heating a liquid in an open container)

  C_v : molar specific heat at *constant volume*

  (heating a gas in a closed container)

- For most materials, $C_p > C_v$.

Specific Heat (examples)

Physical intuition for *c*:

"For a given amount of heat flow *Q*, *specific heat c* is a measure of the *thermal sensitivity* of the material !"

- Hot food on metal/wood plate
- Colliding metal balls

D **water balloon**

Example (Iron vs. Wood Plate)

Metal Plate

Hot food

Q=2 kJ

Wood Plate

m=0.25kg
c_{iron}=470 J/kg K

$$\Delta T = Q/mc$$

m=0.25kg
c_{wood}=2500 J/kg K

Putting the same amount of heat (hot food) on the plate, what is ΔT?

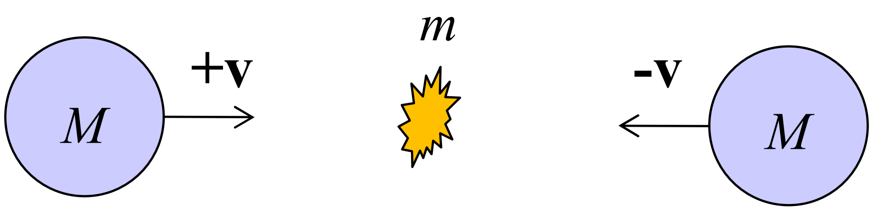
$$\Delta T = \frac{2000J}{0.25kg \cdot 470J / kg \cdot K} = 17K$$

$$\Delta T = \frac{2000J}{0.25kg \cdot 2500J / kg \cdot K} = 3.2K$$

hot to the touch!

much less temp increase !

Example (Colliding Balls)



$$M = 0.5kg$$

$$m = 10mg = 1\times10^{-5}\,kg$$

$$v = 5\ m/s$$

$$c_{wood} = 2500J/kg\cdot K$$

KE delivered to the small piece of paper by the two balls:

$$KE = 2\left(\frac{1}{2}Mv^2\right) = Mv^2 = 0.5kg\left(5m/s\right)^2 = 12.5J$$

KE → Q, what is the temperature increase for the small piece of paper?

$$Q = KE = 12.5J$$

$$\Delta T = Q\big/mc_{wood} = 12.5J/\left(1\times10^{-5}kg\cdot 2500J/kg\cdot K\right) = 500K \rightarrow$$

(starting @ room T)

$$793.15K$$

$$520^{o}C$$

paper will burn!

Extensive and Intensive Quantities

Extensive quantities: depend on the amount of substance:

→ double the amount ⇨ double the quantity

e.g., volume V is an extensive quantity

Intensive quantities: **not** depend on amount of substance

→ units are typically per kg $[\cdot/kg]$ or per mol $[\cdot/mol]$

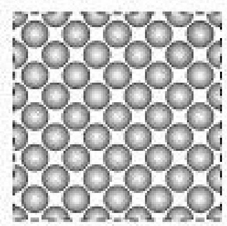
e.g. specific heat c and molar heat capacity C:

$$[J/kg\cdot K] \qquad\qquad [J/mol\cdot K]$$

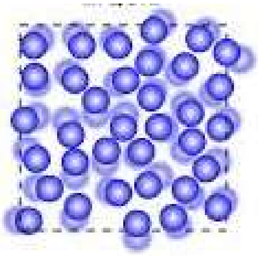
(Note: Typically, heat capacity is an extensive quality and specific heat is intensive but your book does not make this distinction.)

Phases of Matters

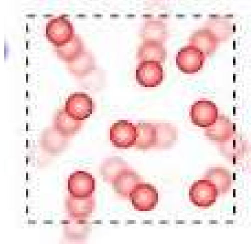
Solid



Liquid



Gas



H_2O

ice

$T < T_{freezing}$

water

$T_{freezing} < T < T_{boiling}$

steam

$T > T_{boiling}$

Heat Exchanges during Phase Changes

During **Phase Changes**, energy exchanged is used for *internal* structural changes (e.g., pulling molecules further apart) :

e.g. ice → water or water → steam

$$Q = m\,L$$

heat of fusion (water)

$L_f = 3.34\times10^5\, J/kg$

$<<$

heat of vaporization (water)

$L_v = 2.26\times10^6\, J/kg$

Since all energy is used for internal structural change, heat exchanged by substances during **Phase Changes** does NOT produce ΔT.

D heat pad

Heats of Fusion & Heats of Vaporization

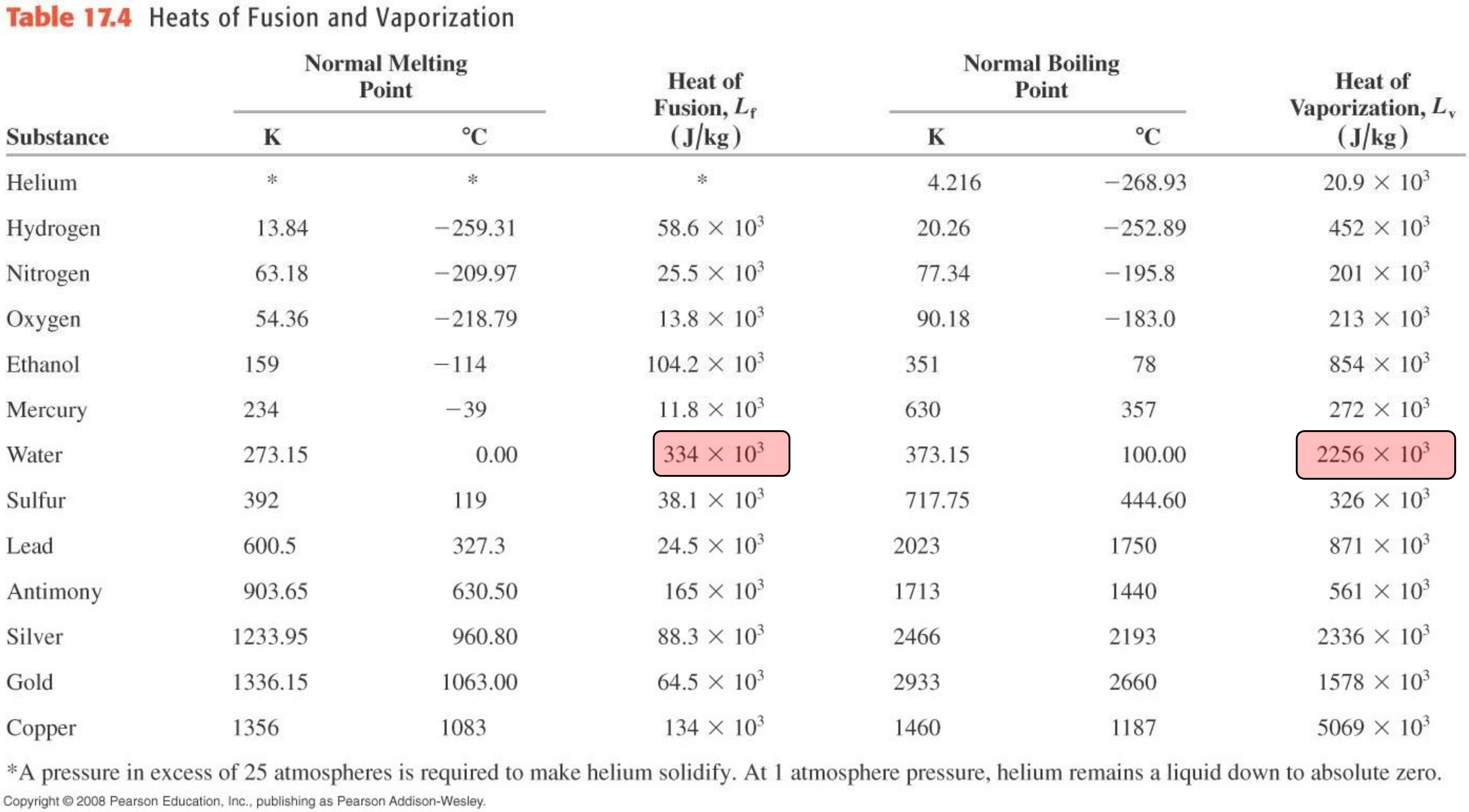


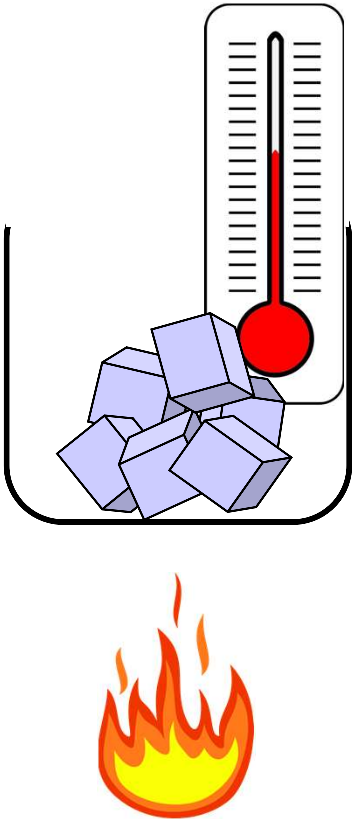
Table 17.4 Heats of Fusion and Vaporization

Substance	Normal Melting Point K	Normal Melting Point °C	Heat of Fusion, L_f (J/kg)	Normal Boiling Point K	Normal Boiling Point °C	Heat of Vaporization, L_v (J/kg)
Helium	*	*	*	4.216	−268.93	20.9×10^3
Hydrogen	13.84	−259.31	58.6×10^3	20.26	−252.89	452×10^3
Nitrogen	63.18	−209.97	25.5×10^3	77.34	−195.8	201×10^3
Oxygen	54.36	−218.79	13.8×10^3	90.18	−183.0	213×10^3
Ethanol	159	−114	104.2×10^3	351	78	854×10^3
Mercury	234	−39	11.8×10^3	630	357	272×10^3
Water	273.15	0.00	334×10^3	373.15	100.00	2256×10^3
Sulfur	392	119	38.1×10^3	717.75	444.60	326×10^3
Lead	600.5	327.3	24.5×10^3	2023	1750	871×10^3
Antimony	903.65	630.50	165×10^3	1713	1440	561×10^3
Silver	1233.95	960.80	88.3×10^3	2466	2193	2336×10^3
Gold	1336.15	1063.00	64.5×10^3	2933	2660	1578×10^3
Copper	1356	1083	134×10^3	1460	1187	5069×10^3

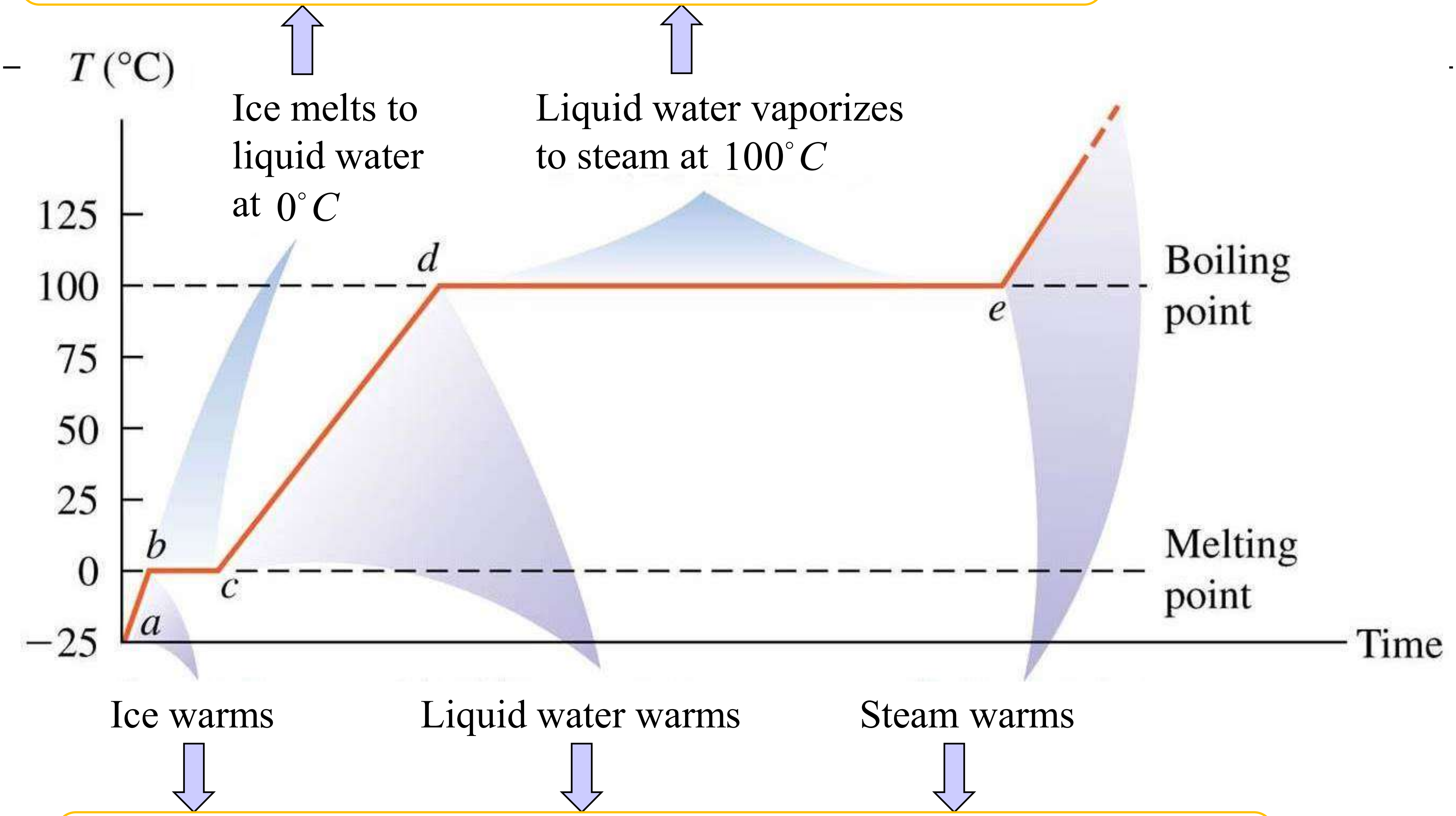
*A pressure in excess of 25 atmospheres is required to make helium solidify. At 1 atmosphere pressure, helium remains a liquid down to absolute zero.

Calorimetric Processes

Now we are ready to describe thermodynamic process such as the following:



Phase of water changes: During these periods, temperature stays constant as heat is added: $Q = +mL$



Temperature of water changes: During these periods, temperature rises as heat is added: $Q = mc\Delta T$