Physics 262/266

Recitation

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Chapter 17: Temperature & Heat

Topics in Chapter

- thermometers and temperature scales
- absolute zero and the Kelvin scale
- meaning of thermal equilibrium
- thermal expansion
- meaning of heat
- calorimetry calculations
- mechanisms of heat transfer

Temperature (T)

Physics definition: average KE of molecules (more on this next chapter)

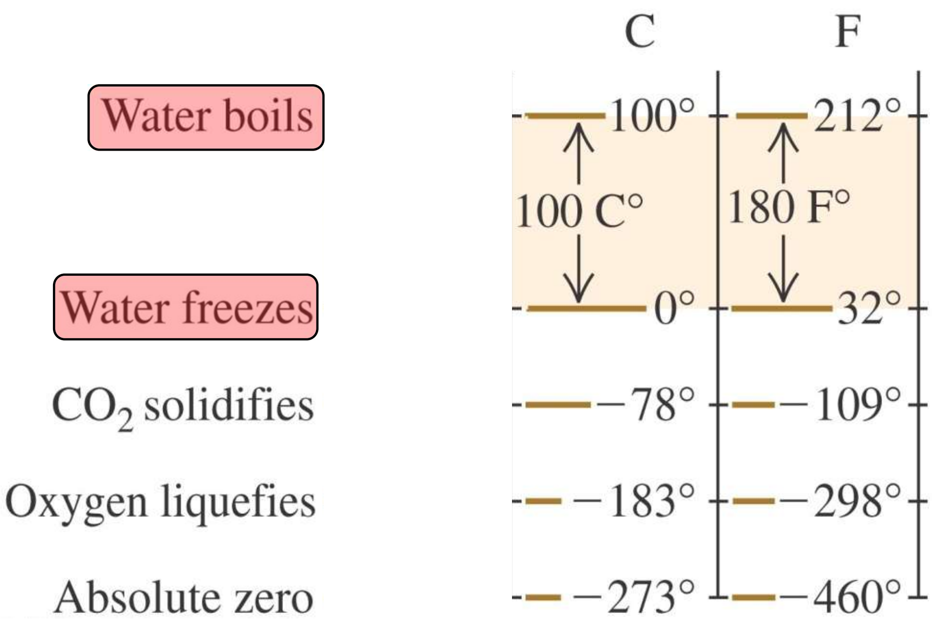
Common usage: a measure of “hot” & “cold”

Physical changes associated with ΔT:

- Most materials expand when heated
- Pressure of gas (in a closed container) ⇧ with ⇧ T
- Electrical resistance changes with T
- Materials radiate at different λ at different T
- State of matter change with T

  Ice ←→ water ←→ steam

Temperature Scales

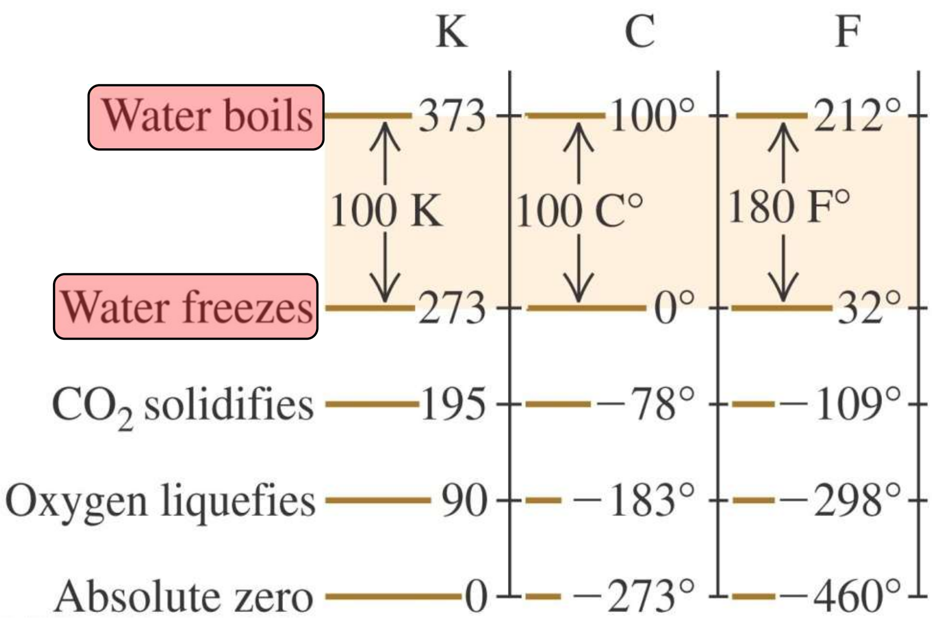


Celsius (°C) & Fahrenheit (°F) are traditionally defined with two readily reproducible reference states:

Freezing point & *Boiling* point of water at 1 atm.

$$\begin{cases} T_C = \frac{5}{9}\left[T_F - 32^o\right] \\ T_F = \frac{9}{5}T_C + 32^o \end{cases}$$

Temperature Scales



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(new)

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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.

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.,

$$1 cal = 4.186 J$$

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*:

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

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

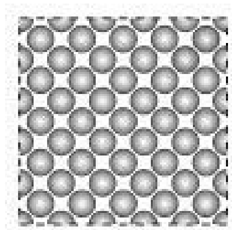
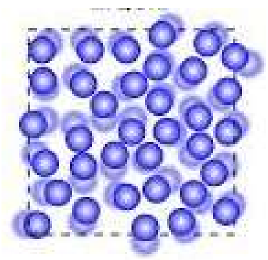
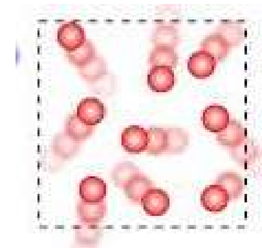
c → "specific heat" is characteristic of the type of material [$J/kg \cdot K$]

Specific Heat Values

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

Substance	Specific Heat, *c* (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

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

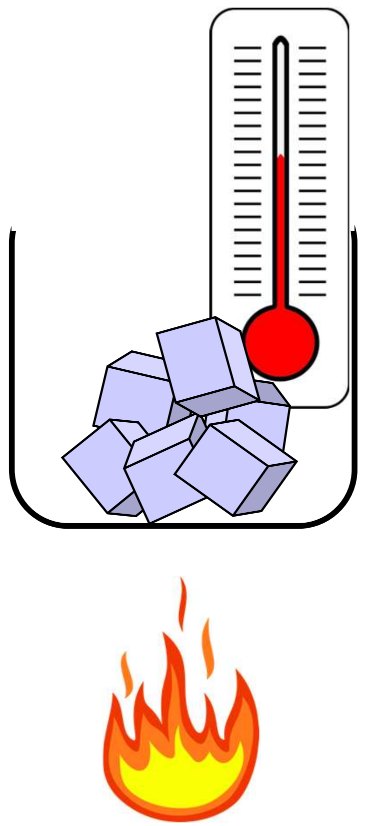
Table 17.4 Heats of Fusion and Vaporization

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

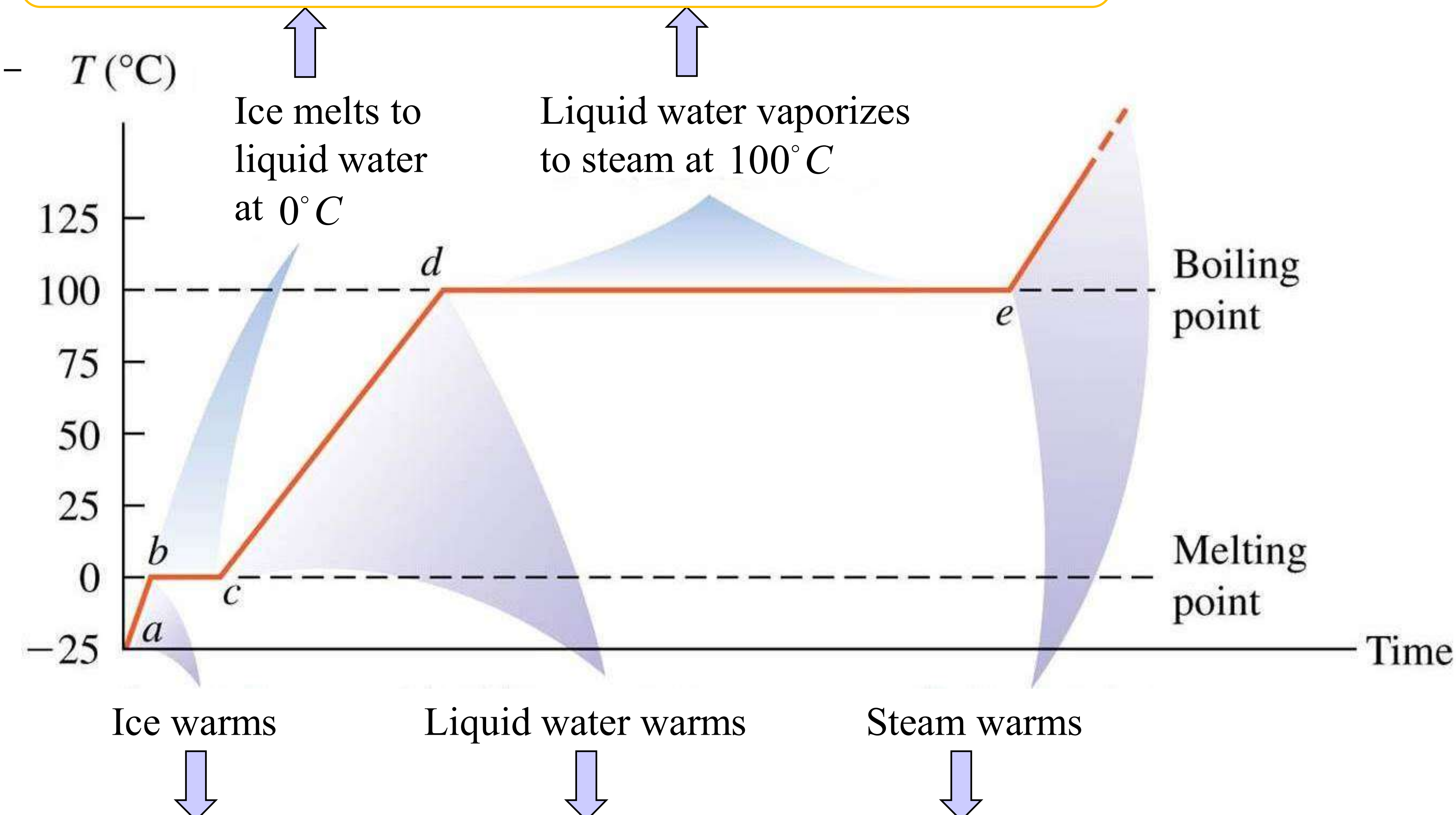
*A pressure in excess of 25 atmospheres is required to make helium solidify. At 1 atmosphere pressure, helium remains a liquid down to absolute zero.

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: $Q = mc\Delta T$

Calorimetry: Problem Solving with Heat Exchanges (method 1)

- Main Concept: Conservation of Energy

  ⇒ $\Sigma Q = 0$ (sum of all heat flows into and out of system =0)

- Sign Convention: heat enters a system is +

  heat leaves a system is –

- $\Delta T = T_f - T_i$

Calorimetry: Problem Solving with Heat Exchanges (method 2)

- Main Concept: Conservation of Energy

OR $$\sum |Q_{gain}| = \sum |Q_{loss}|$$

- Keep all heats as positive quantities

Calorimetry: Problem Solving with Heat Exchanges

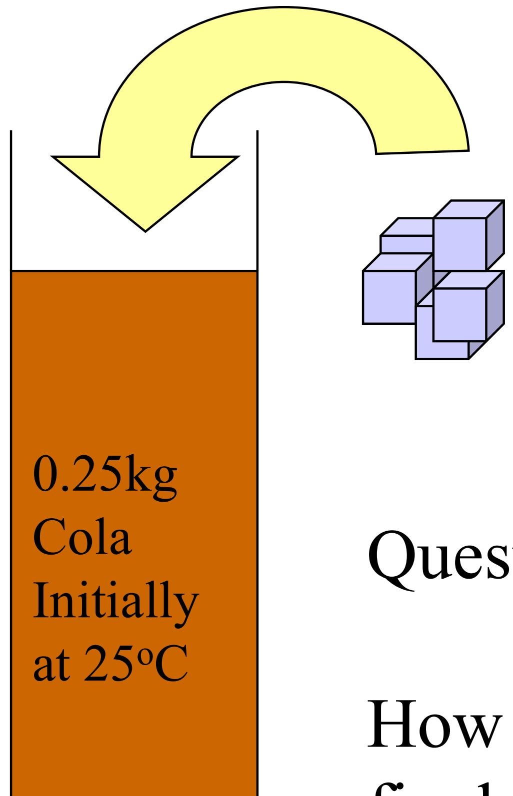
Steps:

1. **Identify all phase change pts**
2. Apply (either $Q=mc\Delta T$ or $Q=mL$) for each processes separately. (don't apply $Q=mc\Delta T$ across ph. changes!)
3. Use

$$\sum_{ALL} Q = 0 \quad \text{and follow sign convention}$$

or just do $$\sum \left|Q_{gain}\right| = \sum \left|Q_{loss}\right|$$

Calorimetry (example 17.8)



Ice initially at -20°C

note

Question:

How much ice needed so that the final mixture is all liquid water with a temperature of 0°C?

http://complex.gmu.edu/www-phys/phys262/soln/ex17.8.pdf