

## Lecture 3

# Phase change and heat capacity

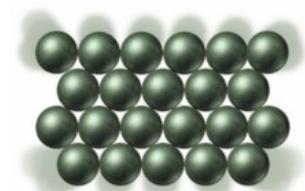
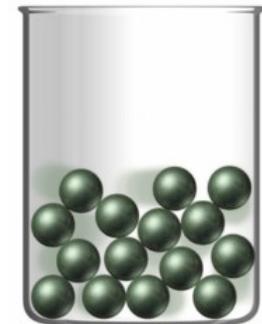
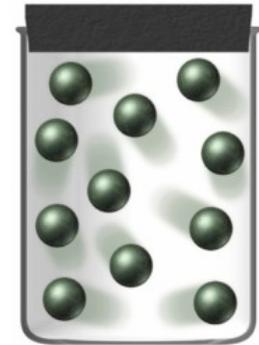
*Pre-reading: §17.6*

# Phases of matter

Matter can exist in different *phases*:

- **gas**: very weak intermolecular forces, rapid random motion
- **liquid**: intermolecular forces bind closest neighbors
- **solid**: strong intermolecular forces

A transition from one phase to another is called **phase change**.



# Phase changes

Phase changes involve absorption or emission of heat.

So the transfer of heat into a system can change the *temperature* of the system, or the *phase*, or both.

# Temperature change

*(no phase change)*

The amount of heat  $Q$  needed to increase the temperature of a mass  $m$  from  $T_1$  to  $T_2$  is proportional to the temperature change  $\Delta T = T_2 - T_1$ :

$$Q = m c \Delta T$$

$c$  is the *specific heat* of the material, and depends on the material.

Units:  $\text{J} \cdot \text{kg}^{-1} \cdot \text{K}^{-1}$

# Specific heat

Water has a remarkably large specific heat.

<u>Substance</u>	<u><math>c</math> (J.kg<sup>-1</sup>.K<sup>-1</sup>)</u>
Aluminum	910
Copper	390
Ice	2100
<b>Water</b>	<b>4190</b>
Steam	2010
Air	1000
Soils / sand	~500

# Molar heat capacity

Sometimes it's more convenient to describe the amount of a substance in terms of *moles*.

1 mole of any substance contains the same number of molecules: this number is called *Avogadro's number*  $N_A$ , where  $N_A = 6.02 \times 10^{23}$  *molecules / mol*

The mass  $m$  can be written as the mass per mole times the number of moles:  $m = nM$ .

# Molar heat capacity

Write the specific heat equation using  $M$ :

$$Q = mc\Delta T$$

$$Q = nMc\Delta T$$

so we define

$$Q = nC\Delta T$$

where  $C$  is the *molar heat capacity*.

Units:  $\text{J}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$

# Phase change

Heat transfer is involved in phase changes:



At the melting point

$$Q = \pm mL_f$$

$L_f$  is the *heat of fusion*.

At the boiling point

$$Q = \pm mL_v$$

$L_v$  is the *heat of vaporisation*.

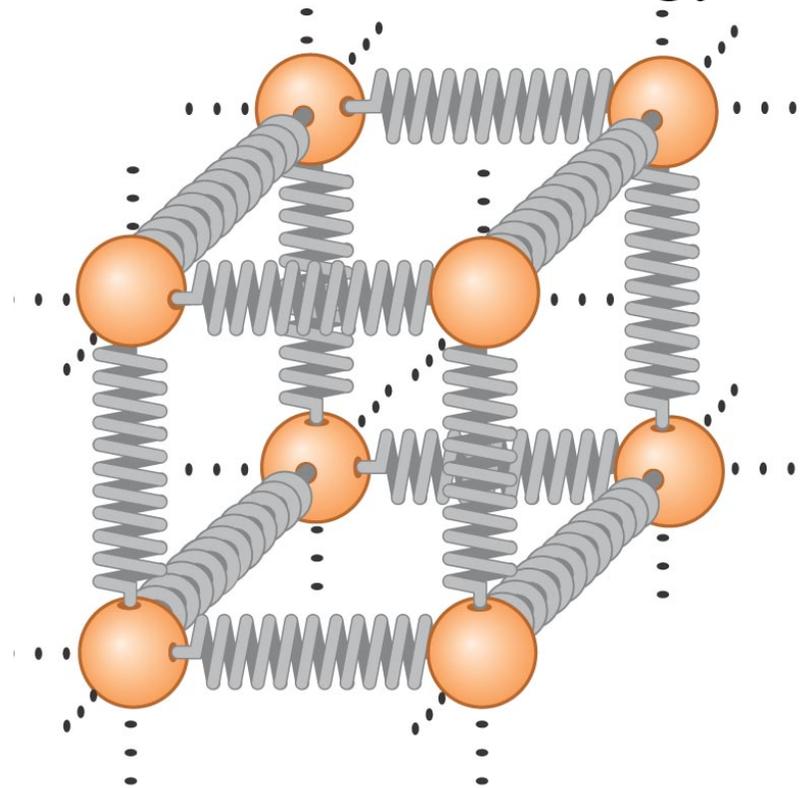
The process is reversible, and there is no change in  $T$ .  
 $Q > 0$ : energy *absorbed*,  $Q < 0$ : energy *released* during the phase change.

# Heats of fusion and vaporisation

<b>Substance</b>	<b>Melting Point (°C)</b>	<b>Latent Heat of Fusion (J/kg)</b>	<b>Boiling Point (°C)</b>	<b>Latent Heat of Vaporization (J/kg)</b>
Helium	- 269.65	$5.23 \times 10^3$	- 268.93	$2.09 \times 10^4$
Nitrogen	- 209.97	$2.55 \times 10^4$	- 195.81	$2.01 \times 10^5$
Oxygen	- 218.79	$1.38 \times 10^4$	- 182.97	$2.13 \times 10^5$
Ethyl alcohol	- 114	$1.04 \times 10^5$	78	$8.54 \times 10^5$
Water	0.00	$3.33 \times 10^5$	100.00	$2.26 \times 10^6$
Sulfur	119	$3.81 \times 10^4$	444.60	$3.26 \times 10^5$
Lead	327.3	$2.45 \times 10^4$	1 750	$8.70 \times 10^5$
Aluminum	660	$3.97 \times 10^5$	2 450	$1.14 \times 10^7$
Silver	960.80	$8.82 \times 10^4$	2 193	$2.33 \times 10^6$
Gold	1 063.00	$6.44 \times 10^4$	2 660	$1.58 \times 10^6$
Copper	1 083	$1.34 \times 10^5$	1 187	$5.06 \times 10^6$

# Where does the energy go?

The latent heat of vaporization is the energy required to give molecules sufficient kinetic energy to break the inter-molecular bonds that keep molecules relatively fixed in the solid state, so that the molecules can move around and enter a gaseous state.



# Evaporation and cooling

As a liquid evaporates, it extracts heat from its surroundings and hence the surroundings are cooled.



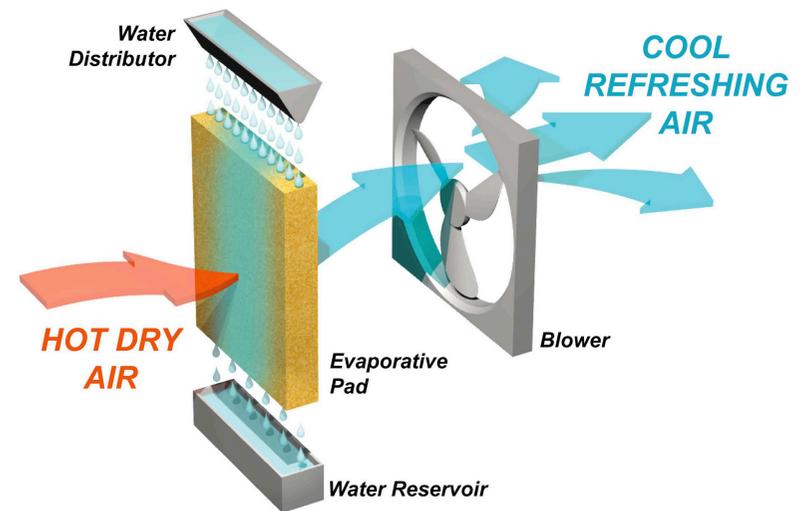
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KJF §17.5

# Evaporation and cooling

*Examples:*

- why do you feel uncomfortable on hot humid days?
- why do fans make you feel cooler?
- evaporative cooling can be used to cool buildings



# Problem

If you have 1.0 kg of water at a temperature of 10°C, how much heat energy needs to be added for it all to become steam?

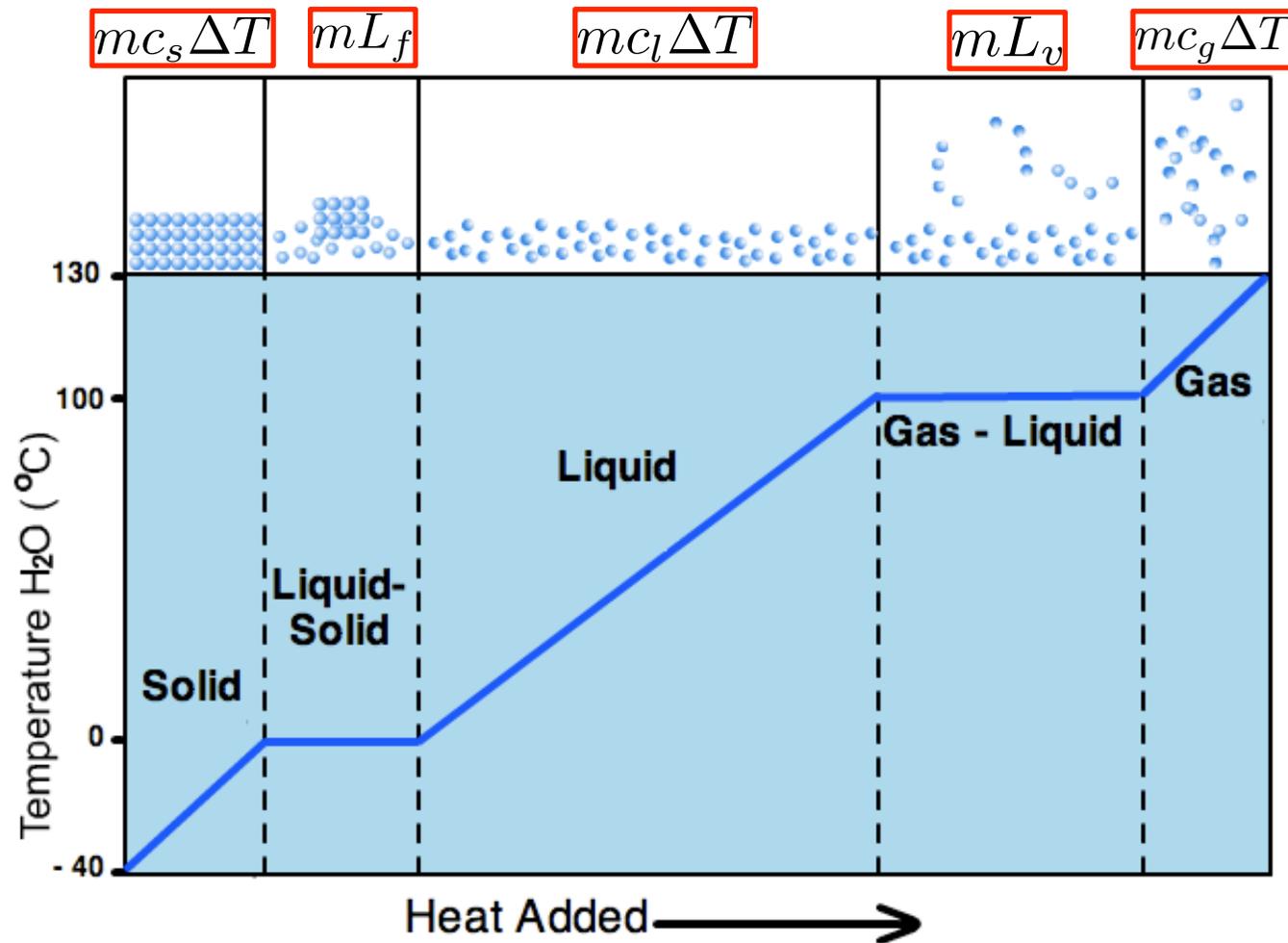
$$c_{\text{water}} = 4200 \text{ J} \cdot \text{kg}^{-1} \cdot \text{K}^{-1}$$

$$L_v = 2.3 \times 10^6 \text{ J} \cdot \text{kg}^{-1}$$



# Heating curve

Simple model for heating water at a constant rate:



## Specific heat

<u>Substance</u>	<u><math>c</math> (J.kg<sup>-1</sup>.K<sup>-1</sup>)</u>
Aluminum	910
Copper	390
Ice	2100
<b>Water</b>	<b>4190</b>
Steam	2010
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## Latent heats

Latent heat – phase change

**Water** - large values of latent heats at atmospheric pressure

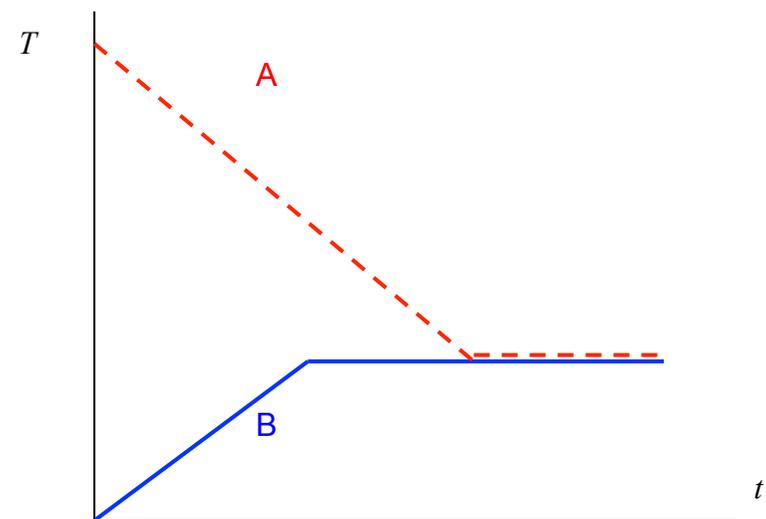
$$L_f = 3.34 \times 10^5 \text{ J.kg}^{-1} (273 \text{ K})$$

$$L_v = 2.26 \times 10^6 \text{ J.kg}^{-1} (373 \text{ K})$$

# Problem

A sample of liquid water A and a sample of ice B of identical masses, are placed in a thermally isolated container and allowed to come to thermal equilibrium. The diagram below is a sketch of the temperature  $T$  of the samples versus time  $t$ . Answer each of the following questions and justify your answer in each case.

1. Is the equilibrium temperature above, below or at the freezing point of water?
2. Does the ice partly melt, or does it undergo no melting?



## Problem *(2004 exam)*

What mass of steam at 130 °C must be condensed onto a 0.100 kg glass cup to warm the cup and the 0.200 kg of water it contains from 20.0 °C to 50.0 °C?

Data:

specific heat capacity of steam  $c_s = 2.01 \times 10^3 \text{ J.kg}^{-1}.\text{K}^{-1}$

specific heat capacity of water  $c_w = 4.19 \times 10^3 \text{ J.kg}^{-1}.\text{K}^{-1}$

specific heat capacity of glass  $c_g = 8.37 \times 10^2 \text{ J.kg}^{-1}.\text{K}^{-1}$

latent heat of vaporisation of water  $L = 2.26 \times 10^6 \text{ J.kg}^{-1}$

# Next lecture

## Mechanisms of heat transfer

*Read: KJF §17.7*