Lecture 45

(Ch13:3)

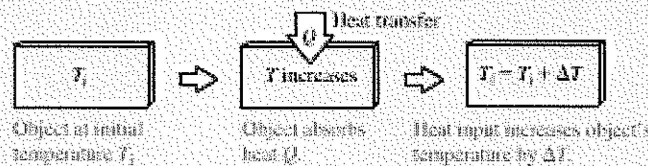
CHAPTER 13 SUMMARY

Heat and Thermal Energy

(Section 13.1) **Heat** is energy in transit as a result of a temperature difference. It's measured in joules, although commonly used alternatives are calories and food calories.

Calories and joules: 1 cal = 4.186 J

Food calories: 1 food calorie = 1 Cal = 1 kcal = 1000 cal = 4186 J



Heat Capacity and Specific Heat

(Section 13.2) Heat flows between two objects in thermal contact until their temperatures are equal; then they're in **thermal equilibrium**.

Heat capacity relates heat and temperature change. **Specific heat** is heat capacity per unit mass. Specific heats of gases are measured at constant volume or constant pressure.

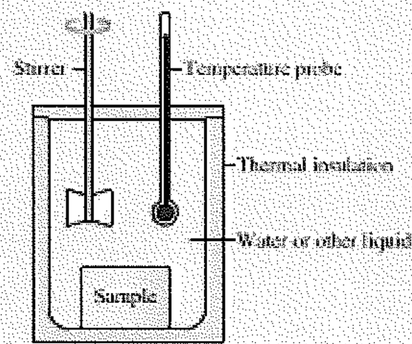
The **equipartition theorem** predicts specific heats of some gases and solids.

Heat capacity: $Q = C\Delta T$

Specific heat: $Q = mc\Delta T$

Molar specific heat (gases, constant volume): $Q = nc_V\Delta T$

Molar specific heat (gases, constant pressure): $Q = nc_P\Delta T$



A calorimeter, used to measure specific heats

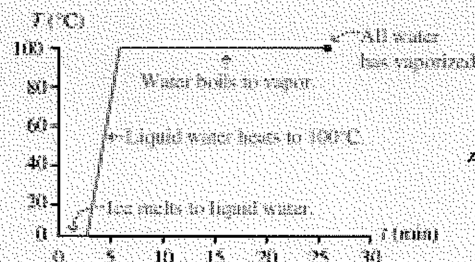
So far

Phase Changes

(Section 13.3) **Heats of transformation** describe the energy per unit mass needed for phase changes: the **heat of fusion** for melting and the **heat of vaporization** for vaporizing.

Heat of fusion: $Q = mL_f$

Heat of vaporization: $Q = mL_v$



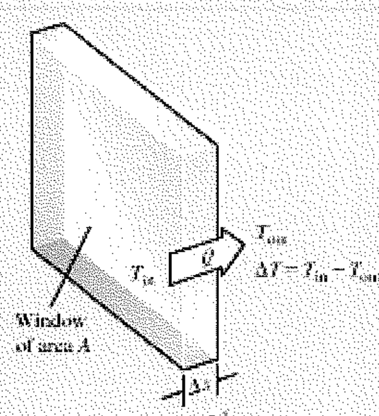
Today

Conduction, Convection, and Radiation

(Section 13.4) **Conduction** is transfer of heat by direct contact and involves collisions among molecules, atoms, and electrons. **Thermal conductivity** quantifies a material's heat conduction capability. **Convection** is the bulk motion of fluid, carrying thermal energy. **Radiation** is energy transfer by electromagnetic waves. Any object above absolute zero radiates power given by the **Stefan-Boltzmann law**.

Thermal conductivity: $\frac{Q}{t} = kA\frac{\Delta T}{\Delta x}$

Stefan-Boltzmann law: $P = e\sigma AT^4$



Review: Absorption of Heat

- Heat Capacity $Q = C\,(T_f - T_i)$
- Specific Heat $Q = c\,m\,(T_f - T_i)$
- Molar Specific Heat $Q = C\,n\,(T_f - T_i)$

  C_v C_p

- Heats of Transformation: $Q = L\,m$

  L_V, L_F

45. You mix 18 kg of water at 25 °C with 6 kg of water at 2 °C, what is the final temperature ?

All we can say is that the hotter water changes temperature by: $\Delta T_{hot} = T_f - 25°C$, while the colder water changes temperature by: $\Delta T_{cold} = T_f - 2.0°C$. We will be able to solve for T_f using Equation 13.2,

i.e. $Q = c.m.\ \Delta T$

and the fact that the heat lost by the hot water is gained by the cold water: $Q_{hot} = -Q_{cold}$, assuming of course that no heat is lost to the surroundings.

Known: $m_{hot} = 18$ kg, $m_{cold} = 6$ kg. and c_{water}= 4186 J/(kg.°C) (Table 13.1)

SOLVE The equal but opposite heat exchange implies:

$$Q_{hot} = -Q_{cold} \Rightarrow m_{hot} c \Delta T_{hot} = -m_{cold} c \Delta T_{cold}$$

Solving for the final temperature:

$$T_f - 25°C = -\frac{6\text{ kg}}{18\text{ kg}}(T_f - 2.0°C) \Rightarrow T_f = 19°C$$

56 g N_2 at 25°C + 12 g He at 45°C

$T_{eq} = ?$

57. **ORGANIZE AND PLAN** The nitrogen starts off colder, so it will gain heat from the helium: ($Q_N = -Q_{He}$). We'll assume that the gases are combined under fixed pressure, so that the heat gained or lost will come from Equation 8.4: $Q = nc_P\Delta T$. The molar specific heats can be taken from Table 13.2 for nitrogen ($c_{N_2} = 29.1$ J/mol·°C) and for helium ($c_{He} = 20.8$ J/mol·°C). We'll need to convert the given masses into moles, and write the temperature change for the nitrogen as: $\Delta T_N = T_f - 25°\text{C}$, and the helium as: $\Delta T_{He} = T_f - 45°\text{C}$.
Known: $m_N = 56$ g, $m_{He} = 12$ g.
SOLVE The equal but opposite heat exchange implies:

56g N_2 = n · 28 g/mol

12g He = n · 4 g/mol

$$-Q_N = -Q_{He} \Rightarrow n_N c_N \Delta T_N = -n_{He} c_{He} \Delta T_{He}$$

The molar masses are 28 g/mol for nitrogen gas and 4 g/mol for helium gas, so the number of moles are 2 mol of nitrogen and 3 mol of helium. Solving for the final equilibrium temperature:

$n_{N_2} = 2$

$n_{He} = 3$

$$\overbrace{T_f - 25°\text{C}}^{\Delta T\ N_2} = -\frac{(3\text{ mol})(20.8\text{ J/mol}\cdot°\text{C})}{(2\text{ mol})(29.1\text{ J/mol}\cdot°\text{C})}\overbrace{(T_f - 45°\text{C})}^{\Delta T\ He} \Rightarrow \underline{T_f = 35°\text{C}}$$

REFLECT The answer makes sense, since the final temperature is halfway between the initial temperatures of the nitrogen and the helium. If you assumed that the gases were mixed with constant volume, the result would be practically the same: $T_f = 34°\text{C}$. This is because the ratio of the molar specific heats (c_{He}/c_N) is practically the same for constant volume and constant pressure.

Table 13.2 | He → $c_p = 20.8$ mol·°C

N_2 → $c_p = 29.1$ mol·°C

- **Heats of transformation**

phase changes

  - *phases* or *states*: solid, liquid, gas
  - When a *phase* transform into another (*phase transformation*) the temperature is constant
  - Needs ***heat of transformation***

$$Q = L\,m \qquad L = \frac{Q}{m}$$

  - Units for L: J/Kg
  - Examples (Table 18-4); Heat of vaporization $L_{V;}$ Heat of fusion L_F

TABLE 19-4 **Some Heats of Transformation**

	Melting		Boiling	
Substance	Melting Point (K)	Heat of Fusion L_F (kJ/kg)	Boiling Point (K)	Heat of Vaporization L_V (kJ/kg)
Hydrogen	14.0	58.0	20.3	455
Oxygen	54.8	13.9	90.2	213
Mercury	234	11.4	630	296
Water	273	333	373	2256
Lead	601	23.2	2017	858
Silver	1235	105	2323	2336
Copper	1356	207	2868	4730

Table 13-3

TABLE 13.3 Heats of Transformation at $P = 1$ atm

Substance	Melting point (°C)	Heat of fusion L_f (J/kg)	Boiling point (°C)	Heat of vaporization L_v (J/kg)
Copper	1084	2.05×10^5	2560	3.92×10^5
Ethanol	−114	1.04×10^5	78	8.52×10^5
Gold	1064	6.45×10^4	2650	1.57×10^6
Helium	N/A	No solid phase at $P = 1$ atm	−269	2.09×10^4
Lead	328	2.50×10^4	1740	8.66×10^5
Mercury	−39	1.22×10^4	358	2.67×10^5
Nitrogen	−210	2.57×10^4	−196	1.96×10^5
Oxygen	−218	1.38×10^4	−183	2.12×10^5
Tungsten	3400	1.82×10^5	5880	4.81×10^6
Uranium	1133	8.28×10^4	3818	1.88×10^6
Water	0	3.33×10^5	100	2.26×10^6

How much heat, Q, is required to melt 500 g of ice at 0 °C ?

By definition $Q = mL_f$ where the heat of fusion of ice is $L_{ice} = 3.33 \times 10^5$ J/kg

Therefore, $Q = (0.5 \text{ kg}).(3.33 \times 10^5) = 166.5$ kJ

Sample Problem 18-3

(a) How much heat must be absorbed by ice of mass $m = 720$ g at $T_1 = -10^oC$ to take it to liquid state at $T_3 = 15^oC$?

Let $T_2 = 0^oC$. Then

$Q_{12} = c_{ice} m(T_2 - T_1) = (2{,}220 \text{ J/kg K})(0.72 \text{ kg})[0^oC - (-10^oC)]$

$= 15{,}984 \text{ J} = 15.98 \text{ kJ}$

Table 13.3

$Q_F = L_F\, m = (333 \text{ kJ/kg})(0.720 \text{ kg}) = 239.8 \text{ kJ}$

Table 13.1

$Q_{23} = c_w\, m\, (T_3 - T_2) = (4{,}190 \text{ J/kg K})(0.720 \text{ kg})(15^oC - 0^oC)$

$= 45{,}252 \text{ J} = 45.25 \text{ kJ}$

$Q = Q_{12} + Q_F + Q_{23} = 15.98 \text{ kJ} + 239.8 \text{ kJ} + 45.25 \text{ kJ} = 300 \text{ kJ}$

Sample Problem 18-3 (cont)

(b) If we supply the ice with a total energy of only 210 kJ (as heat), what then are the final state and the temperature of the water?

$Q_{12} = 15.98$ kJ, $Q_F = 239.8$ kJ, $Q_{23} = 45.25$ kJ

Final state: ICE and WATER, $T_f = 0^\circ$ C

$Q_{rem} = 210 \text{ kJ} - 15.98 \text{ kJ} = 194 \text{ kJ}$

$Q_{rem} = m_w L_F$

$m_w = Q_{rem} / L_F = 194 \text{ kJ}/333 \text{ kJ/kg} = 0.583 \text{ kg}$

$m_{ice} = 0.720 \text{ kg} - 0.583 \text{ kg} = 0.137 \text{ kg}$

15g Cu → at 600°C ; $T_m^{Cu} = 1084\,°C$

$L_{Cu} = 2.05 * 10^3$ J/kg

63. **ORGANIZE AND PLAN** The copper is in liquid form, so energy must be removed from it to cause it to solidify. This loss of heat is from Equation 13.5: $Q = -mL_f$, where we have included a negative sign to signify that this is heat taken away from the copper. The latent heat of fusion for copper from Table 13.3 is: $L_f = 2.05 \times 10^5$ J/kg. Once it turns completely solid, the copper temperature will be at its melting point: $T_i = 1084°C$ from Table 13.3. As it cools to 600°C, the heat removed from the copper will be $Q = mc\Delta T$, where the specific heat of copper from Table 13.1 is: $c = 385$ J/kg·°C. In the end we will sum these two ~~energies.~~ Table 13.3

Known: $m = 15$ g, $\boxed{\Delta T = 600°C - 1084°C = -484°C.}$

SOLVE The energy removed ~~while the copper~~ is solidifying is:

$$Q = -mL_f = -(0.015 \text{ kg})(2.05 \times 10^5 \text{ J/kg}) = -3100 \text{ J}$$

The energy removed while it is cooling is: ✱ Table 13.3

$$Q = mc\Delta T = (0.015 \text{ kg})(385 \text{ J/kg°C})(-484°C) = -2800 \text{ J}$$

The total energy removed is: ✱ Table 13.1

$$Q_{tot} = -3100 \text{ J} - 2800 \text{ J} = -5900 \text{ J}$$

REFLECT The energies are all negative, as they should be, because they represent heat loss from the copper.

ice + 9.53 kJ → H_2O at 0°C | Q = ? $m_{ice} = m_{H_2O}$ boil ↑

$m_{H_2O} = ?$

70. ORGANIZE AND PLAN We can do part (a) without knowing the mass. We just take the ratio of the heat of vaporization to the heat of fusion for water. For part (b), we can invert Equation 13.5: $m = Q_I / L_f$ to find the mass.

Known: $Q_I = 9.53$ kJ.

SOLVE (a) Because the mass is the same in both the melting and the boiling, we can relate the heats of transformation and thereby solve for the energy needed to boil the water:

→ H_2O vaporization / boiling

$Q = m.L$ $\quad m = \dfrac{Q_I}{L_f} = \dfrac{Q_W}{L_v} \Rightarrow Q_W = \dfrac{L_v}{L_f} Q_I = \dfrac{2.26 \times 10^6 \text{ J/kg}}{3.33 \times 10^5 \text{ J/kg}}(9.53 \text{ kJ}) = 64.7 \text{ kJ}$

→ H_2O fusion / melting

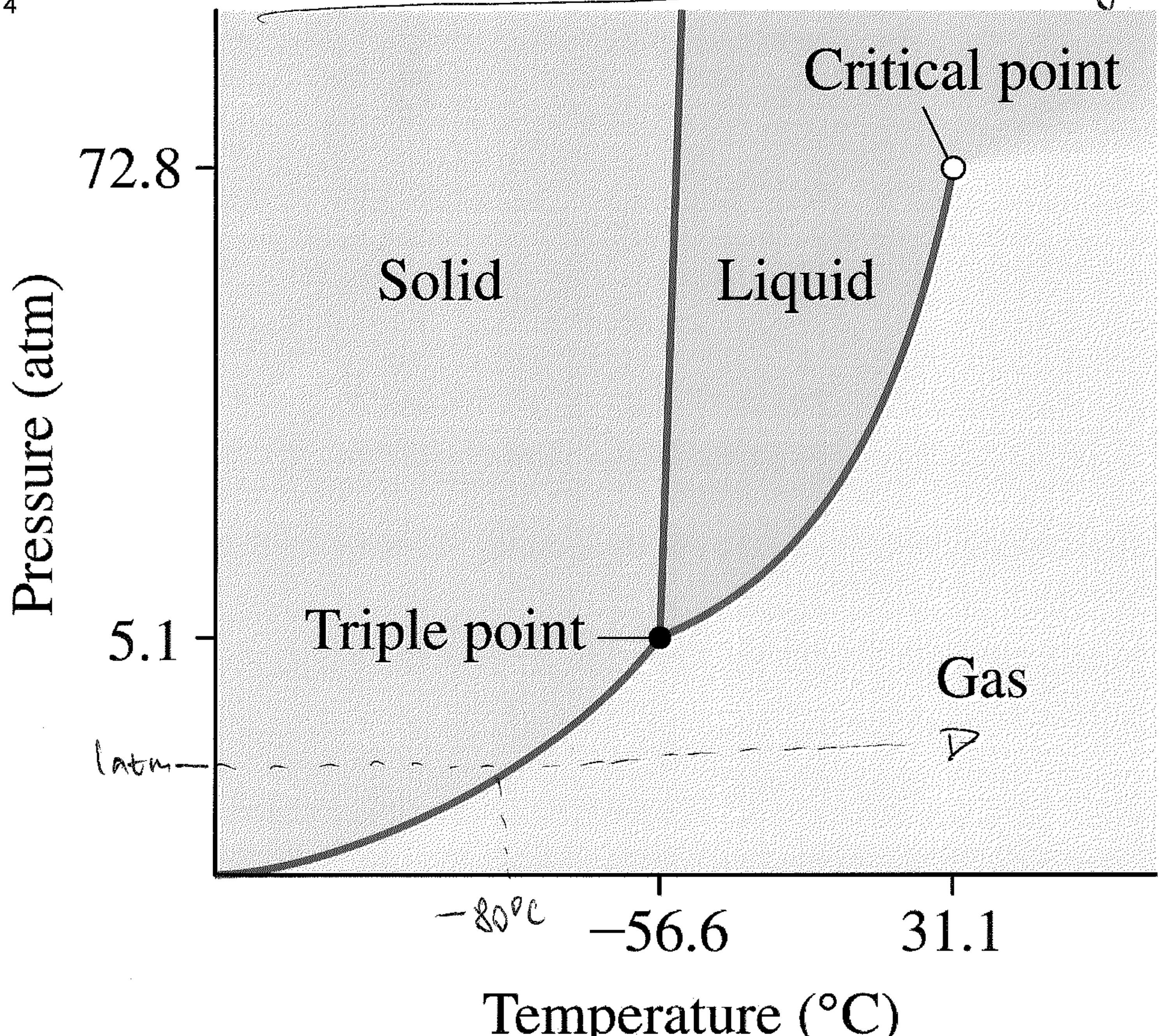
m = const (b) Finding the mass:

$$m = \frac{Q_I}{L_f} = \frac{9.53 \text{ kJ}}{3.33 \times 10^5 \text{ J/kg}} = 29 \text{ g}$$

REFLECT Boiling the water take more energy than melting because the latent heat of vaporization is almost 7 times larger than the latent heat of fusion.

Phase diagram of CO_2 / dry ice

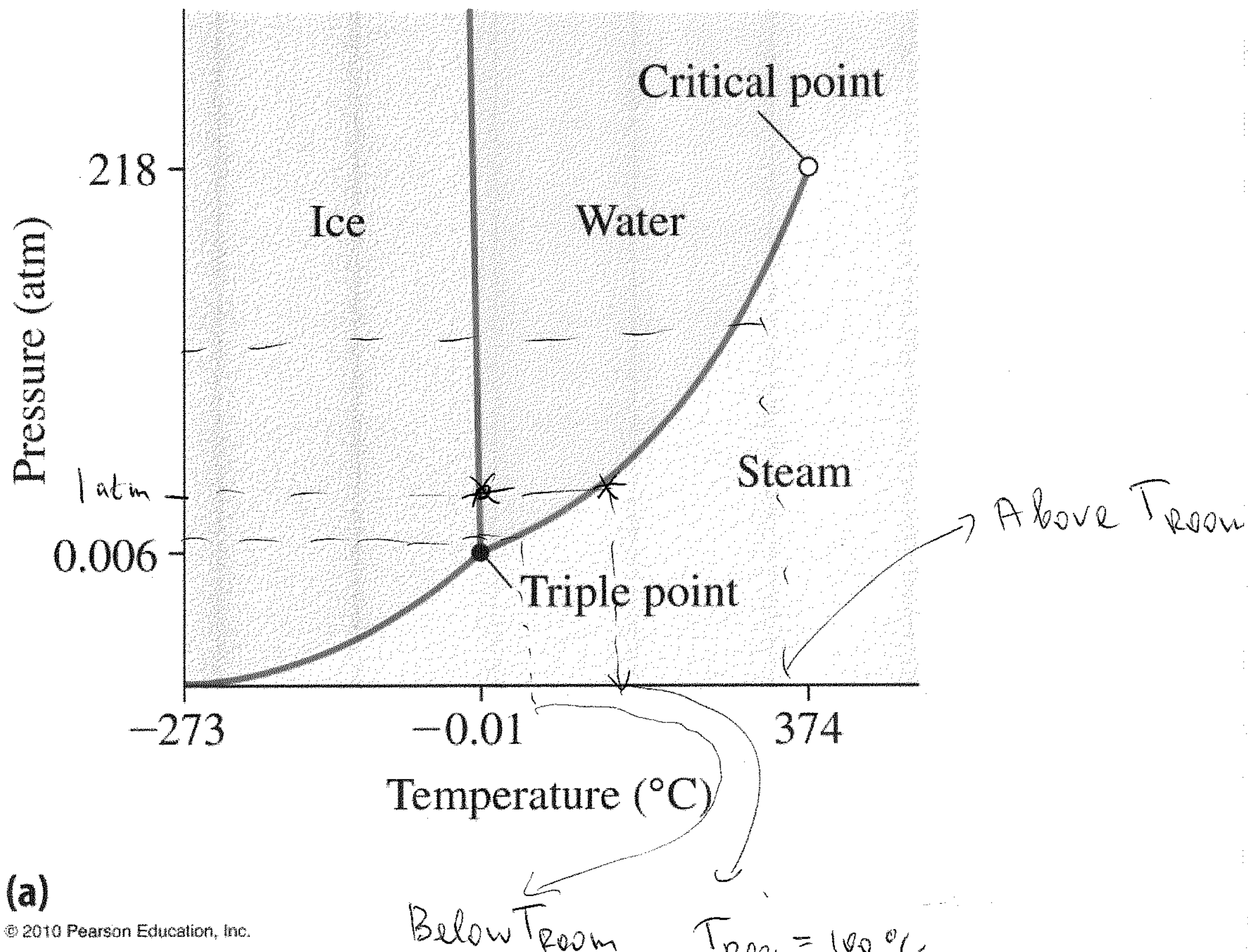
Figure 13.14



with T → ~~at~~ P = 1atm
solid CO_2 → gas CO_2

Figure 13.15A

Phase diagram of H_2O



(a)