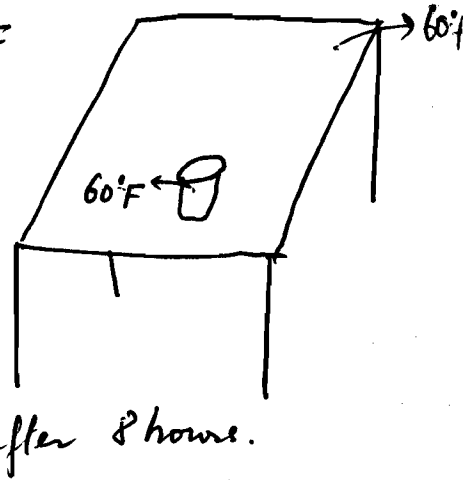
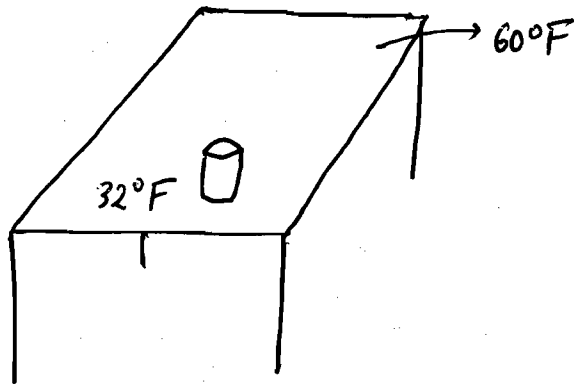


Ch 16 Temperature & Heat

①

Temperature: a macroscopic property, a consequence of the microscopic behavior

32°F
can of soda



Two objects in thermal contact can exchange heat to arrive at a thermodynamic equilibrium (same temperature)

Temperature can be measured with a thermometer (mercury; gas at constant volume; electronic, etc.)

Fahrenheit °F → $T_F = \frac{9}{5}T_C + 32$

Celsius °C → $T_C = T_K - 273.16^\circ\text{K}$

Kelvin °K

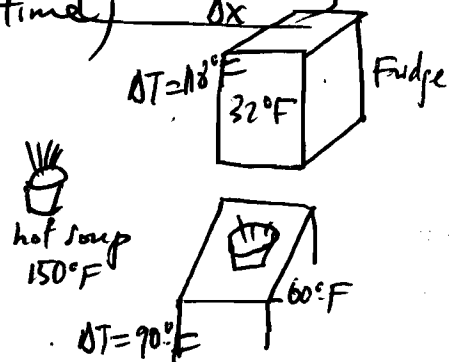
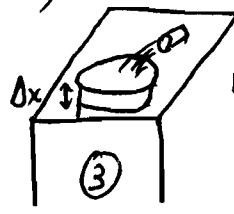
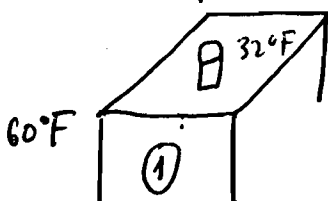
Triple point (water)

↳ "Absolute zero" is 0 °K.

Heat transfer (or exchange) rate:

(How fast heat can be transferred) (per unit time)

$$H = -kA \frac{\Delta T}{\Delta x}$$



Units: (S.I.)

$$[H] = [k][A] \frac{[\Delta T]}{[\Delta x]}$$

$\frac{[H]}{[A]} = \frac{[k][\Delta T]}{[\Delta x]}$
 Energy per unit time: $\frac{J}{s}$
 $[k] = \frac{m^2 \cdot \text{°K}}{m}$

length: m
 mass: kg
 time: s
 temperature: °K

k: thermal constant → depends on type of materials
 (metal conducts heat faster, etc.)

$$[k] = \frac{\frac{J}{s}}{m^2 \cdot \text{°K}} = \frac{J}{m^2 \cdot \text{°K} \cdot s} = \frac{W}{m \cdot \text{°K}} \quad (\text{SI})$$

$$[k] = \frac{\text{Btu} \cdot \text{in}}{\text{h} \cdot \text{ft}^2 \cdot \text{°F}} \quad (\text{British system})$$

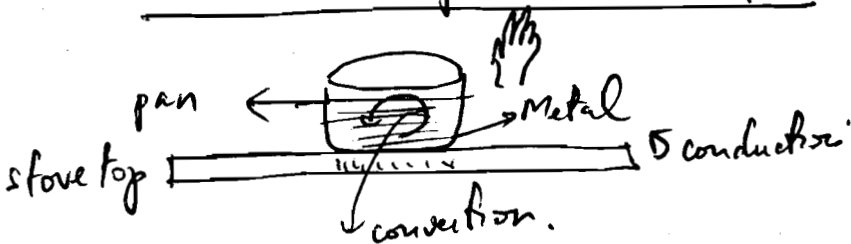
$$H = -kA \frac{\Delta T}{\Delta x} \longrightarrow H = -\frac{A \cdot \Delta T}{R}$$

R-factor: $R = \frac{\Delta x}{k}$

$\left\{ \begin{array}{l} \text{SI: } \frac{m^2 \cdot \text{°K}}{W} \\ \text{British: } \frac{\text{ft}^2 \cdot \text{°F}}{\frac{\text{Btu}}{h}} \end{array} \right.$

Mechanisms of heat transfer:

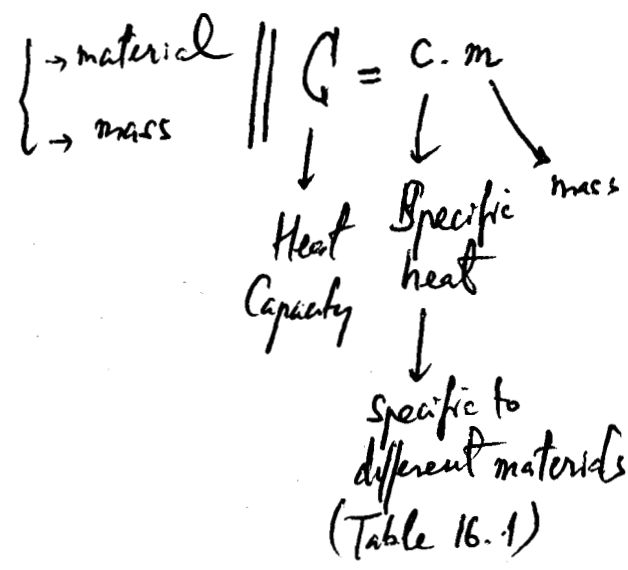
- conduction: (electrons, atoms, molecules)
- convection: (circulation)
- radiation



Heat exchange: $\Delta Q_{Table}^- + \Delta Q_{soda}^+ = 0$

(Soda $32^\circ F = T_{soda}$; Table: $T_{Table} = 60^\circ F$; T.D. equilibrium = $59.9^\circ F$)

$\Delta Q = C \Delta T$
Heat Capacity ($\frac{J}{^\circ K}$)
 $= cm \Delta T$

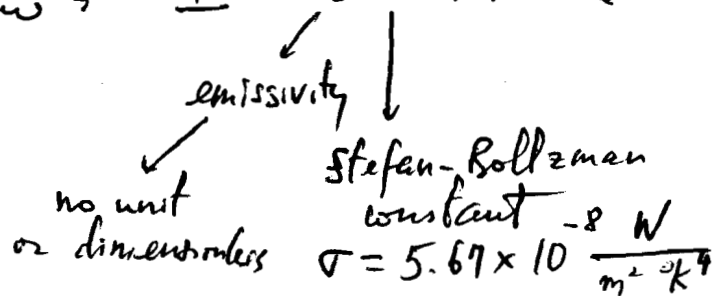


$C_{water} = 4184 \frac{J}{kg \cdot ^\circ K}$
 $C_{copper} = 386 \frac{J}{kg \cdot ^\circ K}$

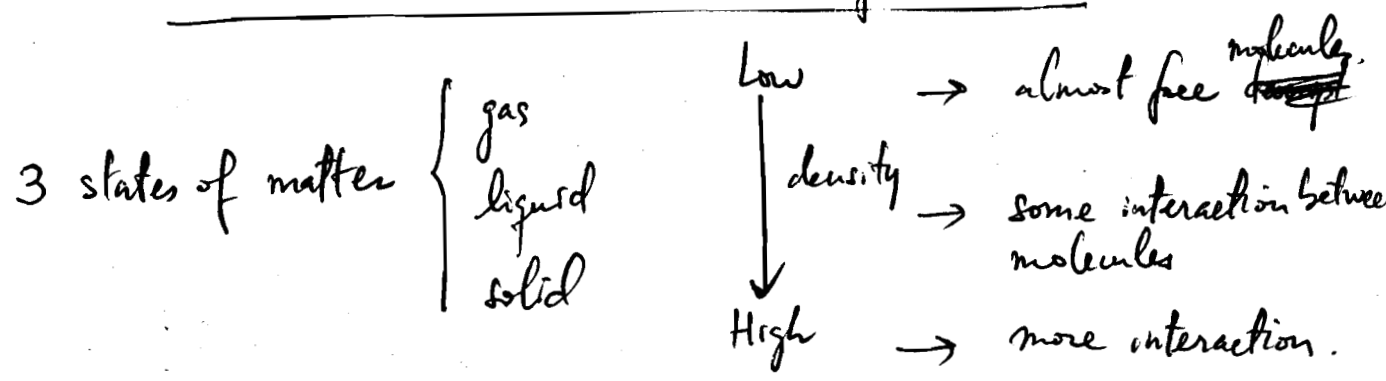
$C_{Table} m_{Table} (T_{TD} - T_{Table})^- + C_{soda} m_{soda} (T_{TD} - T_{soda})^+ = 0$

Heat loss rate (per unit time) by radiation of an object of area A & temperature T:

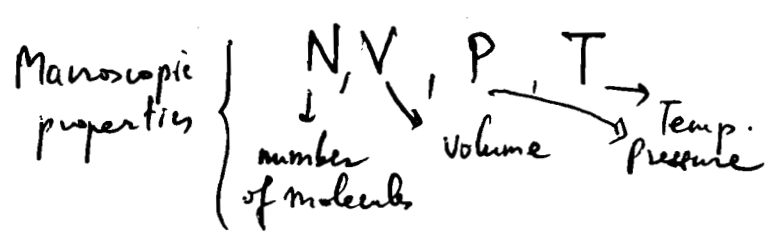
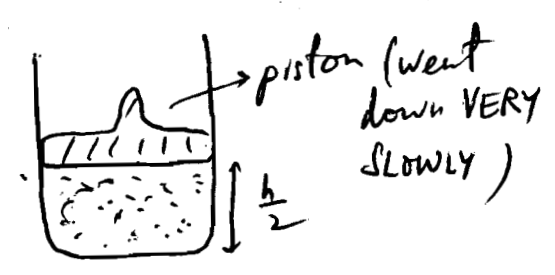
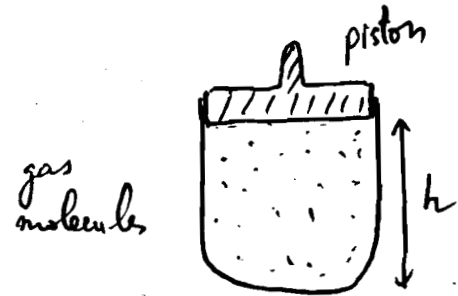
Stefan-Boltzmann Law $P = e \sigma A T^4$



Ch 17 Thermal Behavior of Matter:



Ideal gases : (completely free molecules)



$$N, \frac{V}{2}, 2P, T$$

①

$$\Rightarrow \boxed{PV = kNT}$$

②

$$2P \frac{V}{2} = kNT$$

Ideal Gas Law

k = Boltzman constant = $1.38 \times 10^{-23} \frac{J}{^{\circ}K}$

$N = n N_A$; $N_A = 6.022 \times 10^{23}$
of molecules

$$PV = k n N_A T = n \underbrace{k N_A}_{R} T \rightarrow \boxed{PV = nRT}$$

ideal gas

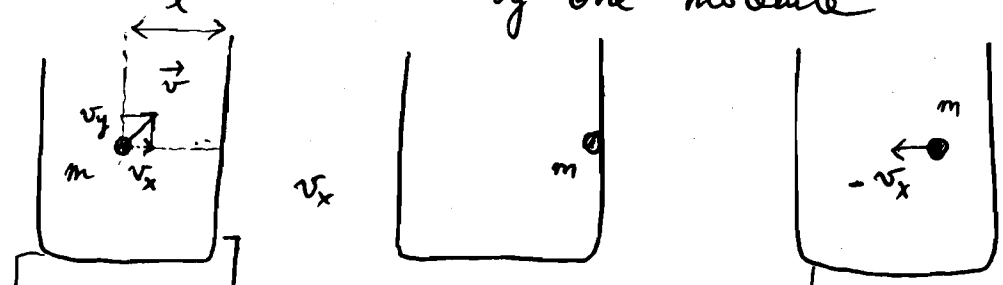
$R = 8.314 \frac{J}{^{\circ}K \text{ mol}}$

To take into consideration some interaction b/w gas molecules:

$$\left(P + \frac{n^2 a}{V^2} \right) (V - nb) = nRT$$

Real gas law or van der Waals law
 a & b are constants.

Ideal gas: mechanics of gas molecules : pressure exerted on the container wall by one molecule



focus on this wall
 Conservation of momentum for the molecule & wall system:
 $m v_x = -m v_x + \underbrace{2m v_x}_{\substack{\text{momentum} \\ \text{received by} \\ \text{the wall}}}$

$$P_i = \frac{F_i}{A} = \frac{\frac{\Delta p_i}{\Delta t}}{A} = \frac{\frac{\sum m v_{xi}}{\frac{\Delta l}{v_{xi}}}}{A} = \frac{m v_{xi}^2}{\underbrace{Al}_{\substack{\text{volume} = \text{vol}}}}$$

By N molecules: $P = \sum_{i=1}^N P_i = \frac{mN}{\text{vol}} \left(\frac{\sum_{i=1}^N v_{xi}^2}{N} \right)$

$$P = \frac{mN}{\text{vol}} \overline{v_x^2} \quad \left(\text{average of } v_x^2 \right)$$

$$\overline{v_x^2} = \frac{1}{3} \overline{v^2}$$

\downarrow x-component of total vel. \downarrow Total velocity

$$\rightarrow P = \frac{mN}{3Vol} \overline{v^2} \rightarrow \boxed{PV = mN \frac{\overline{v^2}}{3}}$$

Derived from microscopic mechanics of gas molecules.

Previously: $PV = kNT$

$$\rightarrow \text{Conclusion: } kT = m \frac{\overline{v^2}}{3} = \frac{2}{3} \frac{1}{2} m \overline{v^2}$$

Average K.E. of a gas molecule

→ The average kinetic energy of a gas molecule is $\frac{3}{2} kT$

$$\boxed{\underbrace{\frac{1}{2} m \overline{v^2}}_{\text{microscopic}} = \underbrace{\frac{3}{2} kT}_{\text{macroscopic}}}$$

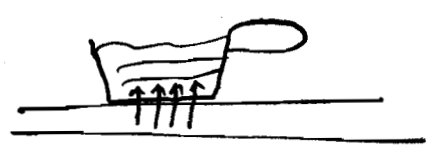
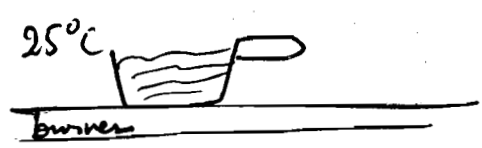
More on this in Statistical Mechanics

In 1D what is the average K.E. of a gas molecule?
 $\frac{1}{2} m \overline{v_x^2} = \frac{1}{2} kT.$

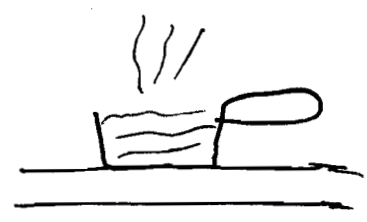
Thermal behavior of matter :

Heat is supplied, ΔQ ,
consequences are :

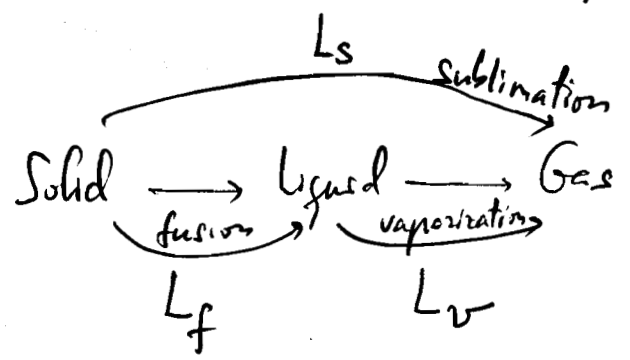
- 1) ΔT ($\Delta Q = cm\Delta T$)
- 2) Change of phase ($\Delta Q = mL$)
 - Latent heat Transformation
 - L_v (liq \rightarrow gas)
 - L_f (solid \rightarrow liq)
 - L_s (solid \rightarrow gas)
- 3) Expansion ΔL , or ΔV
 - linear $\alpha = \frac{\Delta L}{L \cdot \Delta T}$
 - volumic $\beta = \frac{\Delta V}{V \cdot \Delta T}$



$\Delta Q = cm\Delta T$
Temperature increase



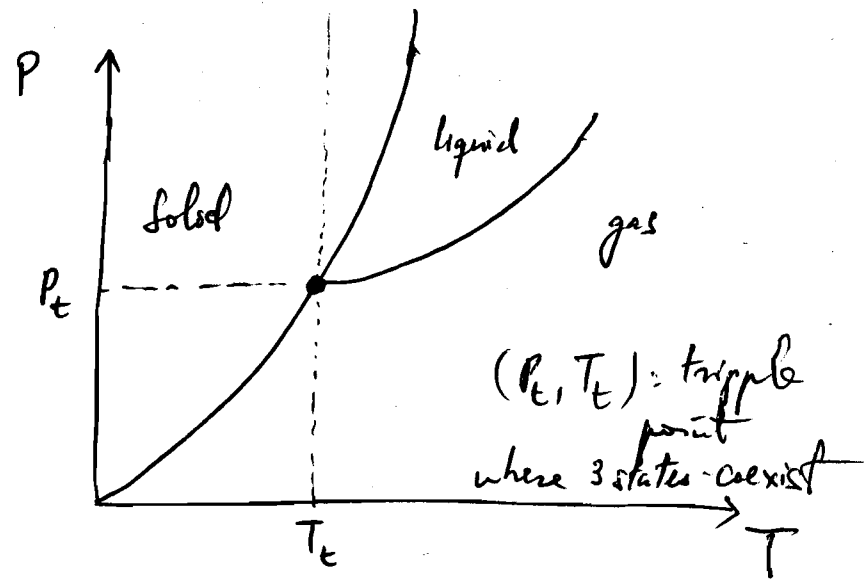
water starts to boil at 100°C.
After 2 min
 $T = 100^\circ\text{C}$
Change of phase.
($\Delta T = 0!$)



Phase Diagram:

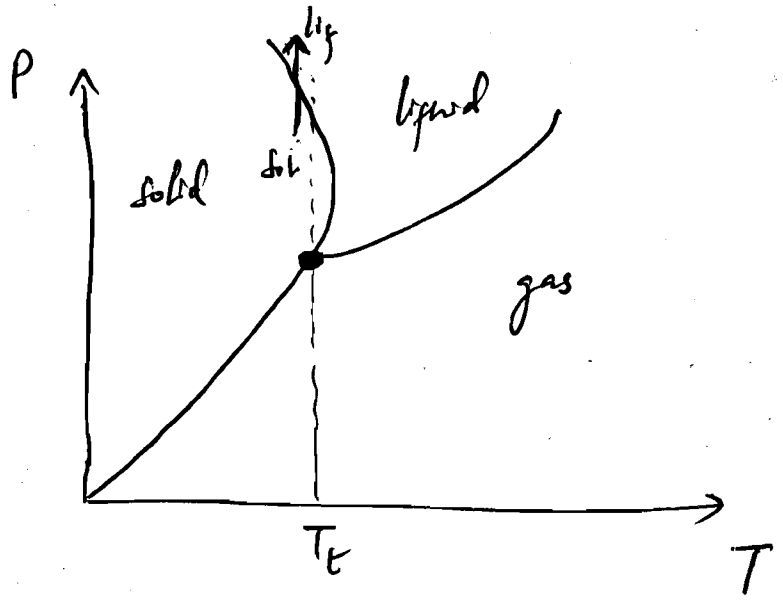
Normal
↓

Solid → liquid:
we need to
increase T



Water
↓

Solid → liquid
} → increase T
} → increase P at
same T



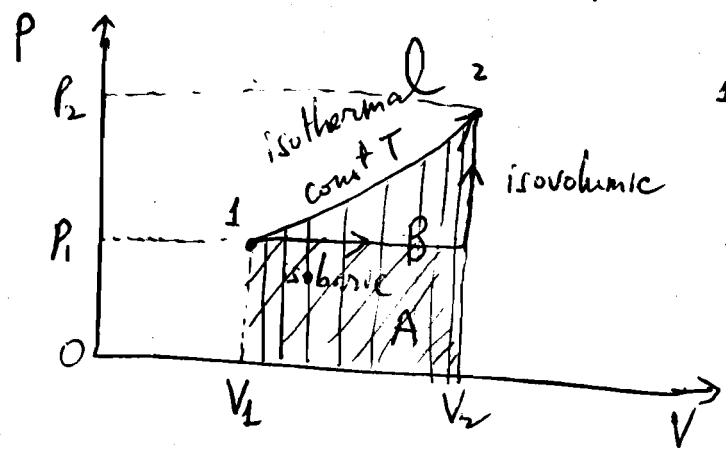
By applying pressure on an ice cube, it melts.
density $\rho_{ice} < \rho_{water}$

Ch 18 Heat, Work, and 1st Law of Thermodynamics

1st Law of Thermodynamics:

Heat absorbed by a system	Q	} $\Delta U = Q - W$ (conservation of energy)
Work done by system	W	
Change of internal energy:	ΔU	

$\Delta U, T, P, V$ are state variables (they do not depend on a particular process)
 Q, W are NOT state variables (they depend on a particular process)



$1, 2$: states
 State 1: P_1, V_1, T_1
 State 2: P_2, V_2, T_2

In a PV diagram: the work done is the area under the curve. Work done from 1 to 2 via a isobaric + isochoric process: $W = A$ is less than the work done if an isothermal process is followed ($W = B > A$)

Why the work done is the area under the curve in a PV diagram?

Work done is $W = F \cdot \Delta x = P \cdot \underbrace{A \cdot \Delta x}_{\Delta V}$

$\rightarrow dW = P dV \rightarrow [W = \int dW = \int P dV]$

\hookrightarrow area under the curve $P(V)$!

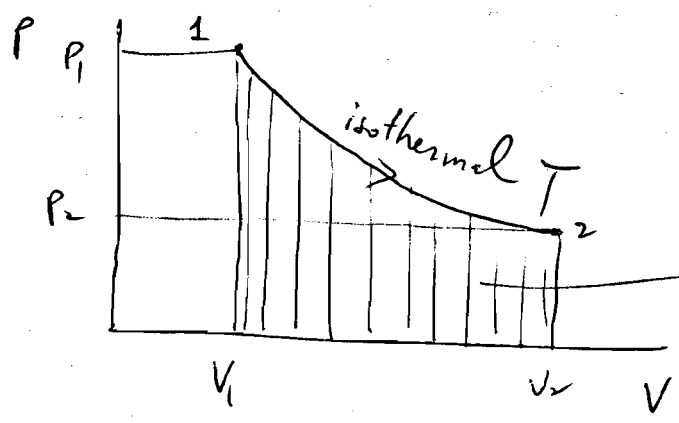
Isothermal process : (const. T) b/w state 1 and state 2

Work done in this process : $W = \int P dV$

\hookrightarrow Need $P(V)$

\rightarrow Ideal gas : $PV = nRT \rightarrow P = \frac{nRT}{V}$

Work done by an ideal gas in an isothermal process $\hookrightarrow W = \int_1^2 \frac{nRT}{V} dV = nRT \int_1^2 \frac{dV}{V} = nRT [\ln V_2 - \ln V_1] = nRT \ln \left(\frac{V_2}{V_1} \right) \checkmark$



For an ideal gas, this area is $nRT \ln \left(\frac{V_2}{V_1} \right)$

Ideal gas:

1) All internal energy comes from kinetic energy
(since we assume for an ideal gas, there is no interaction)

$$\Delta U = \Delta k.E$$

2) Ave. k.E per molecule: $\frac{1}{2} m \overline{v^2} = \frac{3}{2} kT$

3) Isothermal: $T_1 = T_2 \rightarrow$

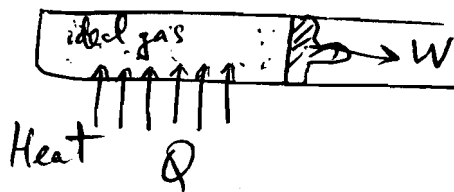
$$\Delta U = 0$$

isothermal, ideal gas.

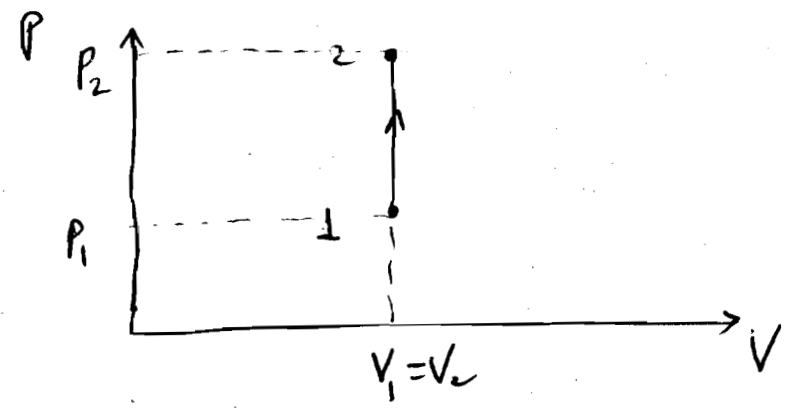
So then what is the heat absorbed by an ideal gas in an isothermal process?

$$\Delta U = Q - W = 0 \rightarrow Q = W = nRT \ln \frac{V_2}{V_1}$$

suppose: $V_1 < V_2 \rightarrow \begin{cases} W > 0 \\ Q > 0 \end{cases}$



Isochoric Process (const. volume) b/w state 1 & state 2.



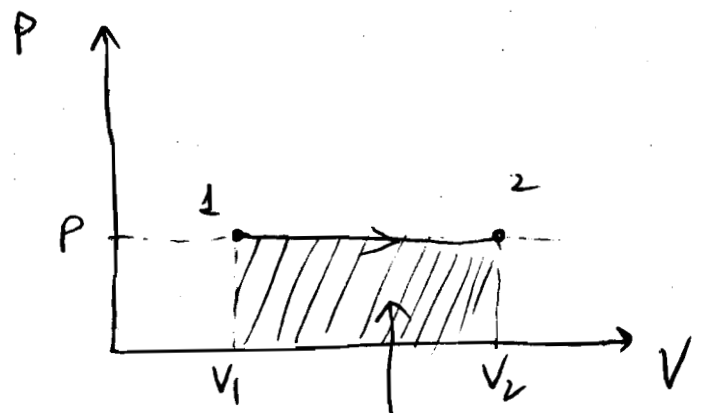
$$W = \int_1^2 P dV = 0$$

$\Delta U = Q$

Specific heat at constant V : $c_v = \frac{1}{n} \frac{Q}{\Delta T} = \frac{1}{n} \frac{\Delta U}{\Delta T}$

$$Q = n c_v \Delta T$$

Isobaric Process (const. pressure) b/w state 1 & state 2



$$W = \int_1^2 P dV = P \int_1^2 dV = P(V_2 - V_1)$$

$\underbrace{\hspace{10em}}_{P \Delta V}$

Specific heat at constant P : $c_p = \frac{1}{n} \frac{Q}{\Delta T}$

$$Q = n c_p \Delta T$$

For an ideal gas: $\frac{P \Delta V}{W} = n R \Delta T =$

1st Law of T.D: $\Delta U = Q - W = Q - P \Delta V$
 $\downarrow \qquad \downarrow \qquad \qquad \downarrow$ ideal gas.
 $n C_v \Delta T \quad n C_p \Delta T \qquad n R \Delta T$

$n C_v \Delta T = n C_p \Delta T - n R \Delta T$

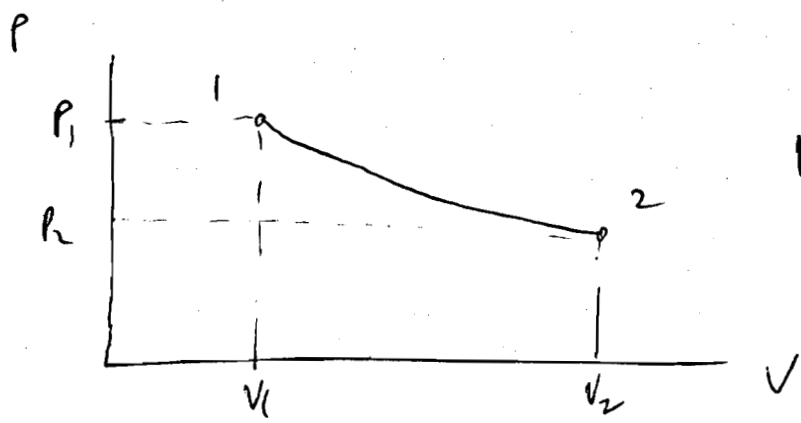
$C_v = C_p - R \quad \text{or} \quad \boxed{C_p = C_v + R}$

$R = 8.314 \frac{J}{K \cdot mol}$

Ideal gas: $C_p > C_v$

Does this make sense? Yes $\left\{ \begin{array}{l} \text{const. } P : Q = \Delta U + W \rightarrow \text{can absorb more heat} \\ \text{const. } V : Q = \Delta U + 0 \end{array} \right.$

Adiabatic Process ($Q=0$) b/w state 1 & state 2



$W = \frac{P_1 V_1 - P_2 V_2}{\gamma - 1}$

1st Law of T.D: $\Delta U = -W$

ideal gas $\left\{ \begin{array}{l} P(V^\gamma) \text{ is constant} : P_1 V_1^\gamma = P_2 V_2^\gamma \\ \gamma = \frac{C_p}{C_v} > 1 \text{ no unit (or dimensionless)} \\ TV^{\gamma-1} \text{ is constant} : T_1 V_1^{\gamma-1} = T_2 V_2^{\gamma-1} \end{array} \right.$

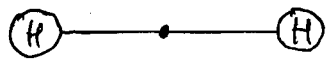
→ Ideal gas: what are possible values for γ ?

Ave. k.E per molecule: $\frac{1}{2} m \bar{v}^2 = \frac{3}{2} kT$ (3D)

↳ Ave. KE per molecule per degree of freedom: $\frac{1}{2} kT$

→ Monoatomic gases (Ne, Ar, etc.): molecules are single atoms → Ave. KE. per molecule is $\frac{3}{2} kT$

→ Diatomic gases (H₂, O₂, etc...): molecules are groups of 2 atoms



degrees of freedom = 5



→ total 6 (lose 1 if they are connected: along x: they can't move away from each other.)

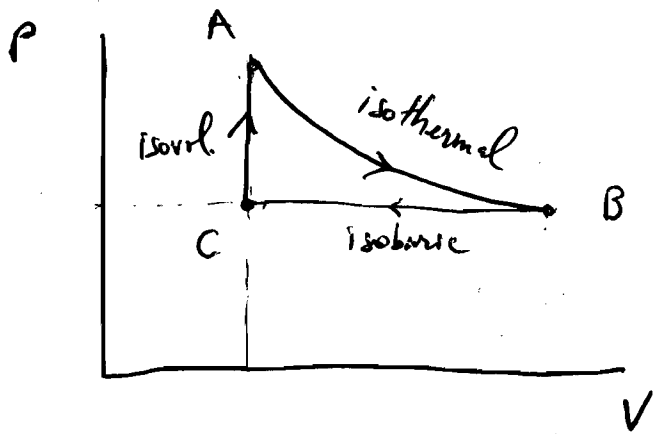
Ave kin. energy per molecule is: $\frac{5}{2} kT$

$C_v = \frac{1}{n} \frac{\Delta U}{\Delta T}$ { Monoatomic = $\Delta U = N \frac{3}{2} kT$ → $C_v = \frac{3R}{2}$
Diatomic = $\Delta U = N \frac{5}{2} kT$ → $C_v = \frac{5R}{2}$

$\gamma = \frac{C_p}{C_v} = \frac{C_v + R}{C_v}$ } Monoatomic: $\gamma = 1.67$
Diatomic: $\gamma = 1.4$

ideal gas.

18.42



$Q_{AB} = 35 \text{ J}$
 $W_{BC} = -22 \text{ J}$

a)

b) $W_{ABC} = W_{AB} + W_{BC} = 35 - 22 = 13 \text{ J}$

AB: isothermal (ideal gas: $\Delta U = 0 \rightarrow Q = W$)

$W_{AB} = Q_{AB} = 35 \text{ J}$

c)

$Q_{BCA} = Q_{BC} + Q_{CA} = n C_p \Delta T_{BC} + n C_v \Delta T_{CA}$

$C_p = \frac{1}{n} \frac{Q_{BC}}{\Delta T_{BC}}$

$C_v = \frac{1}{n} \frac{Q_{CA}}{\Delta T_{CA}}$

$\rightarrow \Delta U = Q - W \rightarrow \Delta U_{ABCA} = 0$

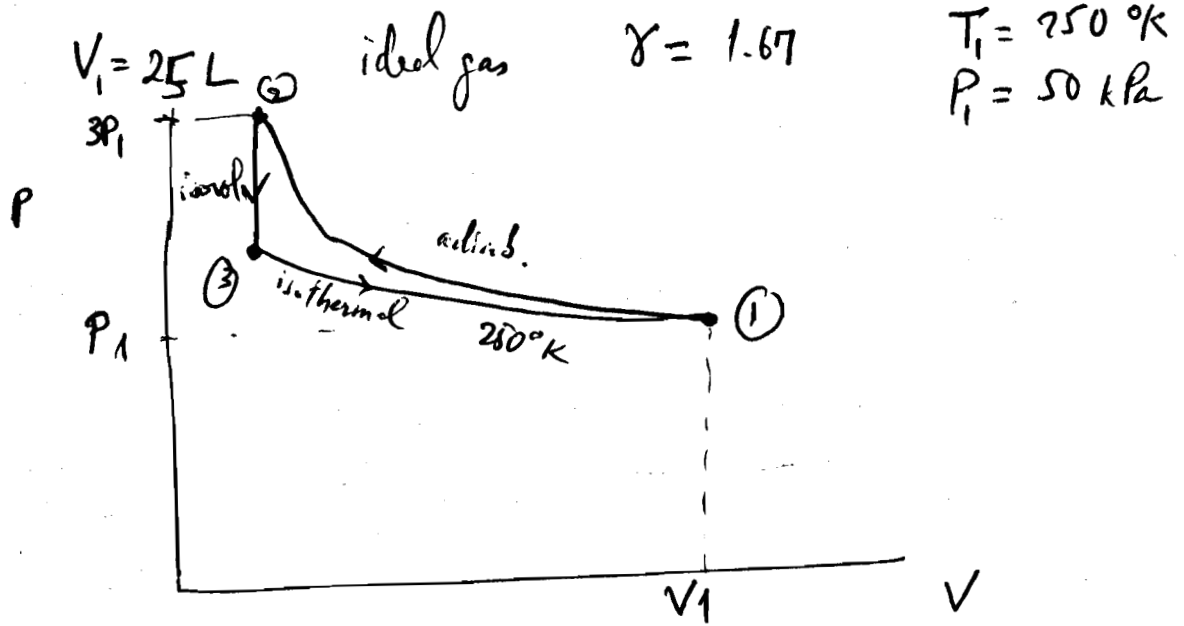
$\rightarrow Q_{ABCA} = W_{ABCA} = 13 \text{ J}$

$Q_{AB} + Q_{BCA} = 13 \text{ J}$

$Q_{BCA} = 13 \text{ J} - Q_{AB} = 13 - 35 = -22 \text{ J}$

(From B to C to A gas delivers 22 J)

18.49



a) Work done on gas?

We can find work done by the gas W_{1231}

$$W_{1231} = W_{12} + 0 + W_{31}$$

$$W_{31} = nRT \ln \frac{V_1}{V_3} = nRT \ln \frac{V_1}{V_2} \rightarrow \text{need}$$

$$\textcircled{2} \text{ from } \textcircled{1} \text{ via adiabatic: } P_1 V_1^\gamma = P_2 V_2^\gamma$$

$$V_1^\gamma = 3 V_2^\gamma$$

$$\left(\frac{V_1}{V_2}\right)^\gamma = 3 \rightarrow \frac{V_1}{V_2} = 3^{1/\gamma}$$

$$\text{Also: } P_1 V_1 = nRT$$

$$\rightarrow W_{31} = P_1 V_1 \ln \frac{V_1}{V_2} = P_1 V_1 \ln 3^{1/\gamma} = \frac{P_1 V_1}{\gamma} \ln 3$$

$$= \frac{50 \times 10^3 \times 25 \times 10^{-3}}{1.67} \ln 3 \text{ J} = 822 \text{ J}$$

$$W_{12} = \frac{P_1 V_1 - P_2 V_2}{\gamma - 1} = \frac{P_1 V_1 - 3P_1 V_1 \frac{1}{3^{3/8}}}{\gamma - 1}$$

① → ② adiabatic.

$$= P_1 V_1 \frac{1 - 3^{(1 - \frac{1}{\gamma})}}{\gamma - 1}$$

$$= 50 \times 10^3 \times 25 \times 10^{-3} \frac{1 - 3^{\frac{0.67}{1.67}}}{0.67}$$

$$= -1033 \text{ J}$$

$$W_{1231} = -1033 + 822 = -211 \text{ J}$$

→ Work done on gas is 211 J

(17.46)

Refrigerator extracts energy from its inside at $95 \frac{\text{J}}{\text{s}}$
 How long to freeze $m_w = 0.75 \text{ kg}$ at $T_w = 0^\circ\text{C}$?

↓
 Q = how much heat we need to extract from the water to freeze it? → Heat of transformation:

$$Q = m_w L_f = 0.75 \text{ kg} \times 334 \frac{\text{kJ}}{\text{kg}} = 250.5 \text{ kJ}$$

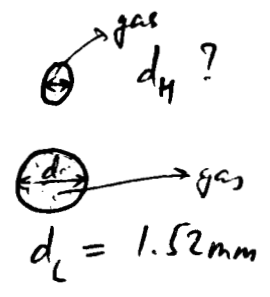
↑
 Table 17.1 = $334 \frac{\text{kJ}}{\text{kg}}$

$$t = \frac{Q}{95 \frac{\text{J}}{\text{s}}} = \frac{250500 \text{ J}}{95 \frac{\text{J}}{\text{s}}} = \frac{1 \text{ hr}}{60 \text{ s}} = 43.95 \text{ min.}$$

18.33; 18.57;

18.33

a) gauge pressure $\left\{ \begin{array}{l} P_H = 125 \text{ mmHg} \\ P_L = 80 \text{ mmHg} \end{array} \right. \rightarrow$
 (not including atmospheric pressure)



$PV = \text{constant}$, Why? b/c $PV = nRT$
 \hookrightarrow body temp. is const.

$P_L V_L = P_H V_H$

$P_L \cdot \frac{4}{3} \pi \left(\frac{d_L}{2}\right)^3 = P_H \cdot \frac{4}{3} \pi \left(\frac{d_H}{2}\right)^3 \rightarrow P_L d_L^3 = P_H d_H^3$

$\rightarrow d_H = \left(\frac{P_L}{P_H} d_L^3\right)^{1/3}$
 $= \left(\frac{P_L}{P_H}\right)^{1/3} d_L$
 $= \left(\frac{80 + 760}{125 + 760}\right)^{1/3} 1.52 \text{ mm}$
 $= 1.49 \text{ mm.}$

1 atm = 1.013×10^5 Pa
 1 mmHg = 133.3 Pa
 \hookrightarrow We can use mmHg instead of Pa in this problem b/c conversion factor will cancel up & down.
 What is 1 atm in mmHg?

$$\frac{1 \text{ atm}}{101300 \text{ Pa}} \cdot \frac{133.3 \text{ Pa}}{1 \text{ mmHg}} = \frac{133.3 \text{ atm}}{101300 \text{ mmHg}} = \frac{1 \text{ atm}}{\frac{101300}{133.3} \text{ mmHg}}$$

$$= \frac{1 \text{ atm}}{760 \text{ mmHg}}$$

b) Work done by heart?

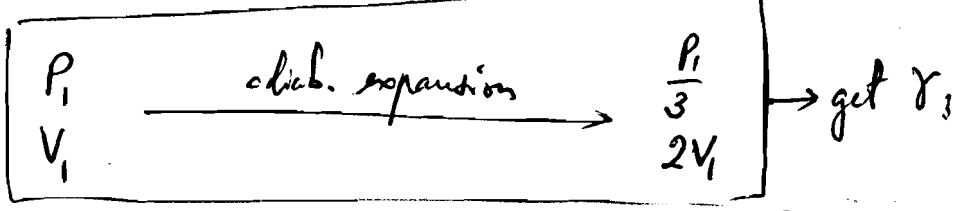
Work done by gas = isothermal process (const. temp at the body temp.)

$$\begin{aligned}
 W_{LH} &= nRT \ln\left(\frac{V_H}{V_L}\right) = P_L V_L \ln\left(\frac{d_H^3}{d_L^3}\right) = 3P_L V_L \ln\left(\frac{d_H}{d_L}\right) \\
 &= 3(80+760) \cdot \frac{133.3}{1} \cdot \frac{4}{3}\pi(0.76 \times 10^{-3})^3 \ln\left(\frac{1.49}{1.52}\right) \\
 &= 617675 \times 10^{-9} \ln\left(\frac{1.49}{1.52}\right) = -12312 \times 10^{-9} \text{ J} \\
 &= -12 \times 10^{-6} \text{ J} = -12 \mu\text{J}
 \end{aligned}$$

→ 12 μJ

18.57

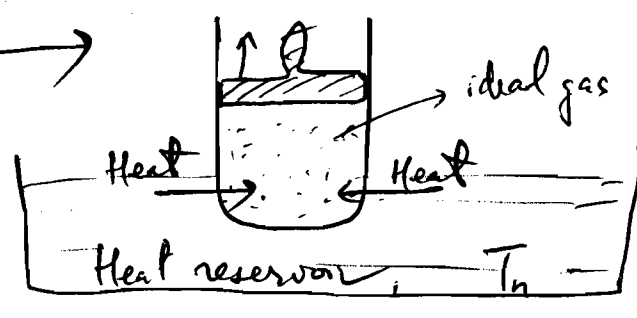
Ar (monatomic) & O₂ (diatomic)



What fraction of molecules are Ar?

Ch 19 2nd Law of Thermodynamics

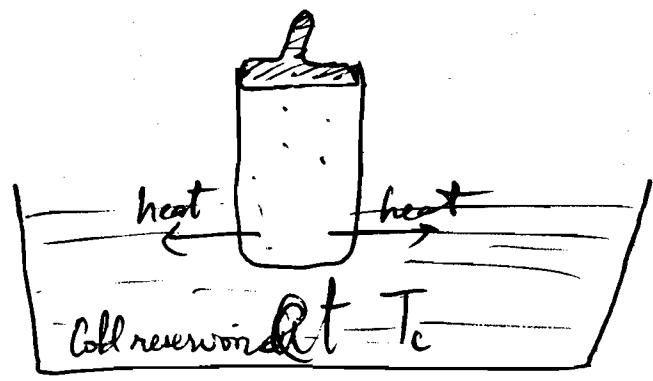
Heat reservoir: source of heat, at constant temperature



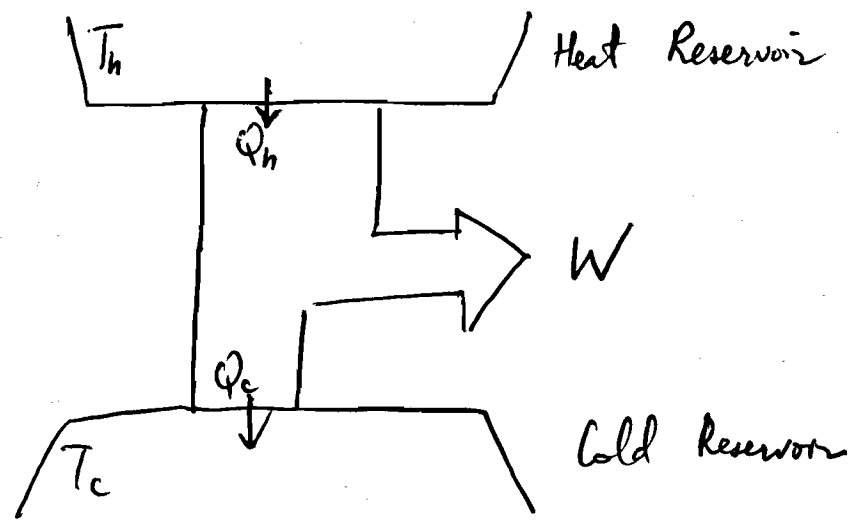
Putting an ideal gas in contact with a heat reservoir: it absorbs heat \rightarrow expands (doing work)

Repeating in cycles:
 \downarrow
 Heat Engine

How to continue to get work from the gas? : Repeat the process: bring the piston back to its original position by putting the heated gas in contact with cold reservoir



Next, putting gas in contact with a cold reservoir at T_c : it loses heat \rightarrow compresses: (receiving work)



$$\Delta U_{\text{engine}} = Q_{\text{net}} - W = Q_h - Q_c - W$$

↓
1st Law of T.D.

→ Ideal, isothermal process = 0
(const. T.)

$$Q_h - Q_c = W$$

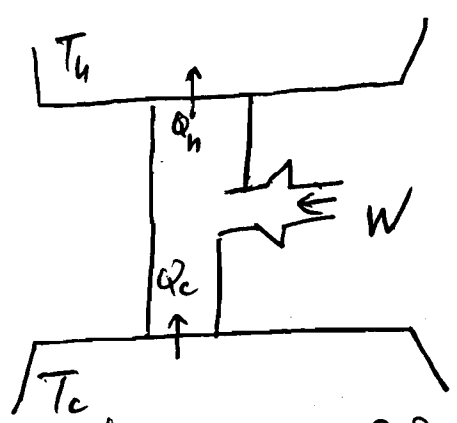
efficiency of an ideal gas heat engine: $e = \frac{W}{Q_h} = \frac{Q_h - Q_c}{Q_h}$

$$= 1 - \frac{|Q_c|}{|Q_h|} < 1$$

(Qc < Qh)

2nd Law of T.D.: it is impossible to build a heat engine operating in cycles that extracts heat from a hot reservoir (and returning some of it to a cold reservoir) that can deliver a 100% of work.

Refrigerators: reversed heat engines



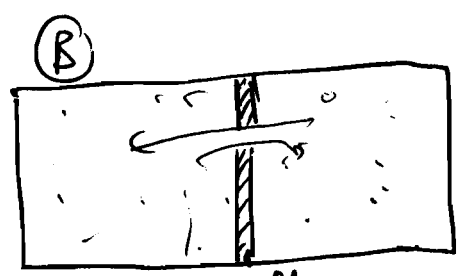
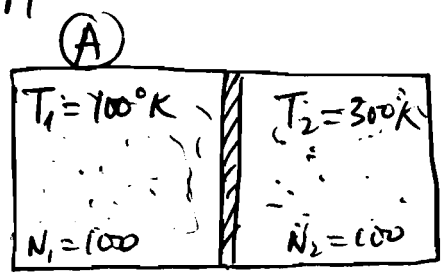
2nd of T.D.: it is impossible to transfer heat from a cold reservoir to a hot reservoir without requiring any work.

Coefficient of Performance: C.O.P. = $\frac{Q_c}{W}$

Entropy: $\Delta S \equiv \int_1^2 \frac{dq}{T}$

3rd Law of T.O: entropy of a closed system can never decrease $\Delta S \geq 0$

(entropy is related to order).



mixed together.

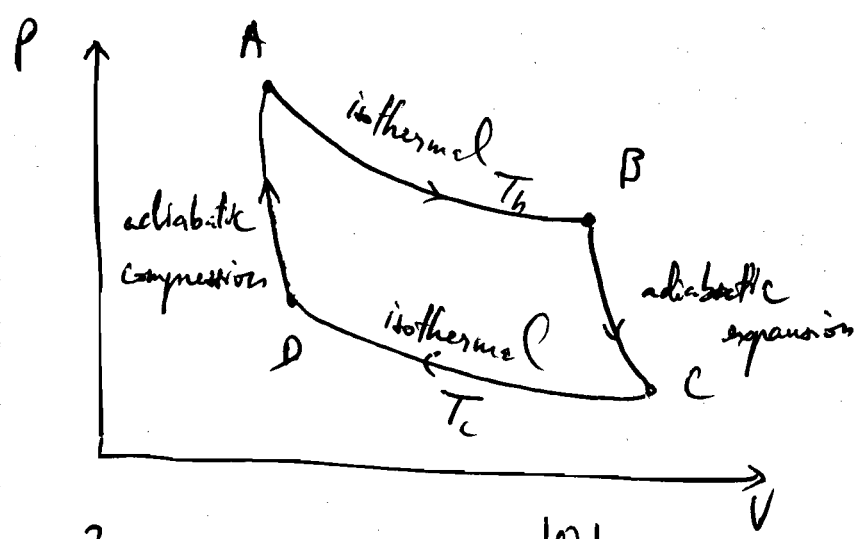
$V_1 = V_2$
 $P_2 > P_1$ why?

$PV = nRT$ (macroscopic)
 $v_{th} = \sqrt{\frac{3kT}{m}}$ (microscopic) : higher thermal speed, higher pressure on the wall

(A) \rightarrow (B) spontaneously mixed together : order has decreased. (ΔS increased).

\rightarrow Natural process is always increasing entropy.

Carnot Engine: 4 reversible processes $\left\{ \begin{array}{l} 2 \text{ isothermal} \\ 2 \text{ adiabatic} \end{array} \right.$
 ↳ ideal gas, max. achievable efficiency = e_{max}



$e_{max} = ?$; $e = 1 - \frac{|Q_c|}{|Q_h|}$

$Q_h = \text{heat absorbed from hot reservoir during isothermal exp. } A \rightarrow B = nRT_h \ln \frac{V_B}{V_A}$
 $Q_c = (\text{isothermal } C \rightarrow D) = nRT_c \ln \frac{V_D}{V_C}$

There is a connection b/w the volumes =

$B \rightarrow C$ adiabatic: $T_B V_B^{\gamma-1} = T_C V_C^{\gamma-1}$
 $\rightarrow \left(\frac{V_B}{V_C} \right)^{\gamma-1} = \frac{T_C}{T_B} = \frac{T_{cold}}{T_h} = \frac{T_c}{T_h}$

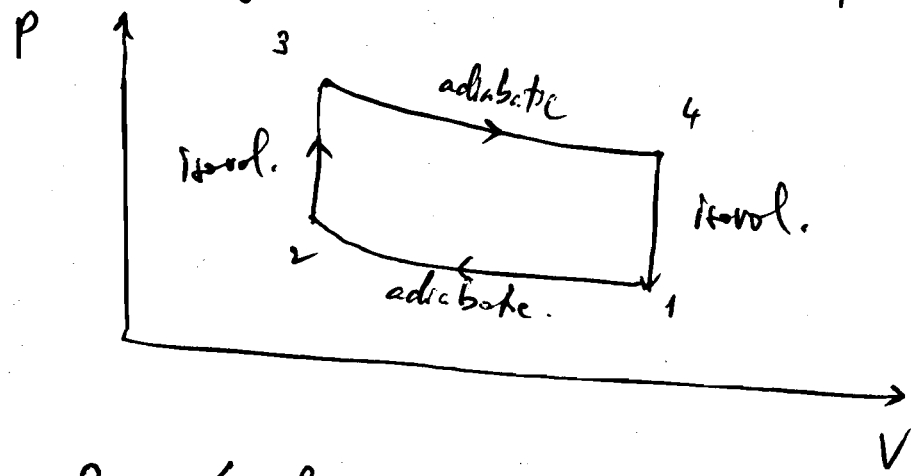
$D \rightarrow A$ adiabatic: $T_D V_D^{\gamma-1} = T_A V_A^{\gamma-1}$
 $\left(\frac{V_D}{V_A} \right)^{\gamma-1} = \frac{T_A}{T_D} = \frac{T_h}{T_c}$

$\left. \begin{array}{l} \frac{V_B}{V_C} \cdot \frac{V_D}{V_A} = 1 \\ \frac{V_D}{V_C} = \frac{V_A}{V_B} \end{array} \right\}$

$$e = 1 - \frac{|Q_c|}{|Q_h|} = 1 - \frac{|nRT_c \ln \frac{V_D}{V_C}|}{|nRT_h \ln \frac{V_B}{V_A}|} = 1 - \frac{nRT_c \ln \frac{V_C}{V_D}}{nRT_h \ln \frac{V_B}{V_A}}$$

$$e_{max} = 1 - \frac{T_c}{T_h} < 1$$

Otto Cycle Engines : 4 reversible processes $\left\{ \begin{array}{l} 2 \text{ adiabatic} \\ 2 \text{ isovolumic} \end{array} \right.$



$$e_{otto} < e_{max}$$

$$\Delta S = \int_1^2 \frac{dQ}{T} \left\{ \begin{array}{l} \text{isovol: } dQ = n c_v dT \quad (c_v = \frac{1}{n} \frac{\Delta Q}{\Delta T} = \frac{1}{n} \frac{dQ}{dT}) \\ \Delta S = \int_1^2 \frac{n c_v dT}{T} = n c_v \int_1^2 \frac{dT}{T} = n c_v [\ln T]_1^2 \\ \text{isothermal: } \Delta S = \frac{1}{T} \int_1^2 dQ = \frac{Q_2 - Q_1}{T} = \frac{\Delta Q}{T} \end{array} \right.$$