

Lecture Notes

→ "Conservation of energy" before the stipulation of the first law meant conservation of kinetic and potential energy.

→ i.e. $\frac{1}{2} m v^2 + mgh = \text{constant}$

→ There is no natural point of zero potential energy (but also for kinetic energy); everything is mostly expressed in terms of differences.

→ The above is now known as conservation of mechanical energy.

→ Macroscopic vs. Microscopic approaches

↓
treat average properties that are a result of the motions of large sums of molecules
- Thermodynamics
- Classical physics (Newton).

↓
treat each molecule individually and then calculate average properties over large samples
- Statistical Mechanics

Q: Did we do anything in class to connect the two 'worlds'?

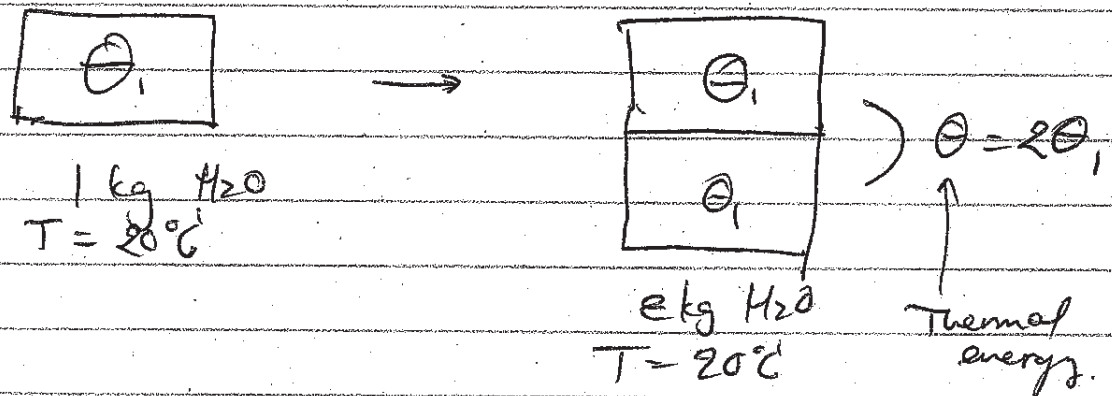
A: Yes, The pressure calculation from kin-theory

$$\frac{1}{2} m \langle v^2 \rangle = \frac{3}{2} kT \Rightarrow T = \frac{1}{3} \frac{m \langle v^2 \rangle}{k} \quad T \propto \langle v^2 \rangle \quad \text{abs } T$$

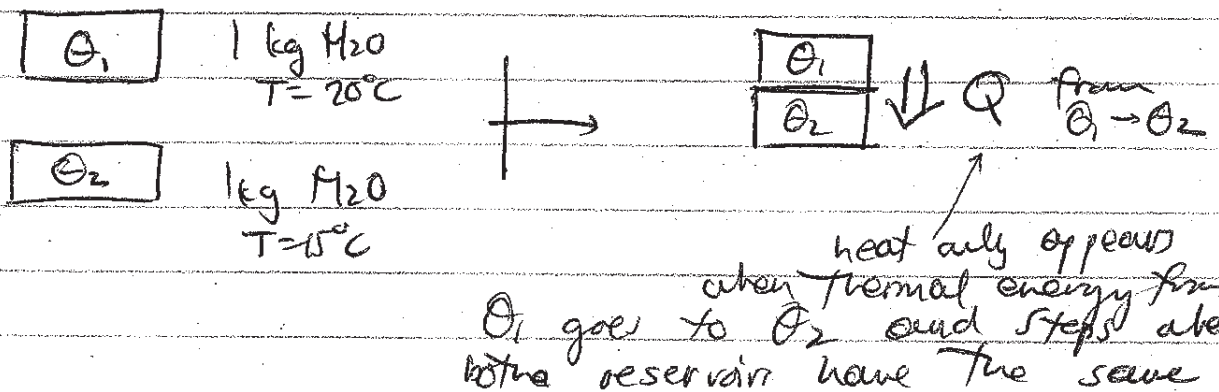
also we saw that $P = \frac{1}{3} N m \langle v^2 \rangle$ abs. P.

Heat vs. Temperature

Temperature is an intensive property and a measure of the energy level of the masses we are looking at. Thermal energy is an extensive property.



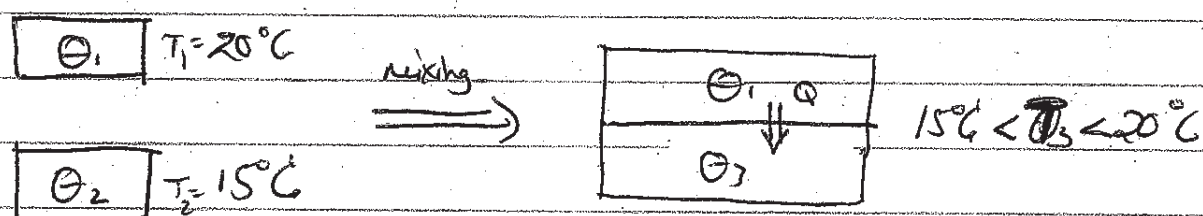
Heat is not a part of any system. It appears only when Thermal energy is exchanged between two systems.



The differentiation between heat and energy level of a solid took a very long time to understand.

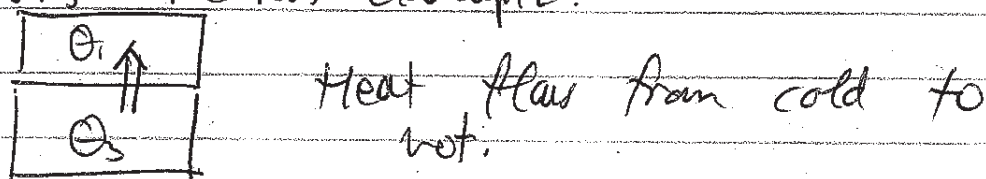
Processes

- Thermodynamic: When the state of a system changes. "State" means that temperature, pressure, molecular structure changed.
- Natural (or spontaneous): A process that would occur without any intervention. Example:



Heat always flows from the hot to cold.
You never see that heat flows from cold to hot (in nature).

- Unnatural: The reverse of a natural process.
e.g.; using previous example:

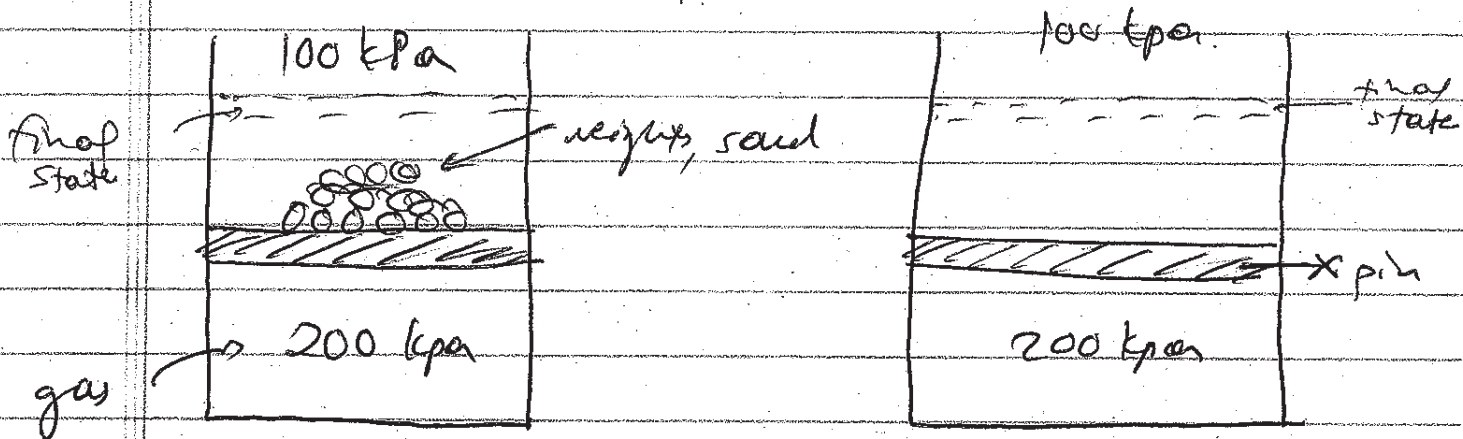


Q: Can unnatural processes happen?

A: Yes, but you have to pay a price (usually by adding work into a system).

Example: refrigerators pump heat from a compartment of low T into the environment of high T .

The rest of concepts can be introduced with the "piston" example



→ in this case we take one grain of sand at a time.

→ slow progression

→ at every time,

→ P inside cylinder

is always in equilibrium with gas above + weights

↑
infinitesimal process or reversible process

→ in this case we release pin and so $P \neq$ equilibrium with gas and (lack of) weights above.

↑
finite or irreversible process.

Isothermal process: If the walls are perfectly permeable to heat and the temperature of the cylinder and system is maintained constant (by, e.g. putting inside a water bath).

Adiabatic process: No heat is allowed to flow through the walls (insulated). In this case, T may change.

Equilibrium states: All the variables of the system are constant and uniform. No tendency for change. A reversible process is for which all states between initial and final are in thermodynamic equilibrium.

Work:

Definition: The work done by a force F moving a distance ds in the direction of the applied force is:

$$dw = F ds$$

So, if you move something a finite distance, the total work is:

$$W = \int_{s_1}^{s_2} F ds$$

work has units $F \times (\text{distance})$

↑ ↑
force length

or $(\text{kg m s}^{-2}) \times (\text{m}) = \text{kg m}^2 \text{s}^{-2} \text{ (energy)}$

so, work is a form of energy.

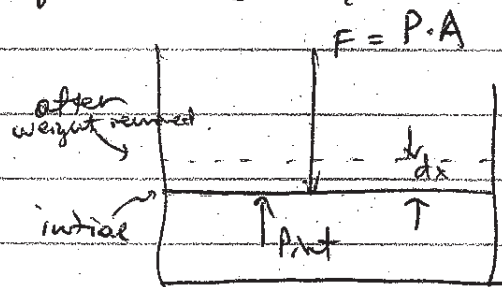
Convention: work done on the system is positive
work done by the system is negative

→ Let's go back to the cylinder example

The work done by the expanding cylinder is:

$$dw = -F dx$$

↑
because of our convention.



but, since P is $\left(\frac{\text{force}}{\text{area}}\right)$, and because A is constant:

$$dw = -P \cdot A dx$$

↪
differential change in volume of our system $\therefore = dV$

or, $dw = -P dV$

for a finite expansion we just use the integral:

$$W_{\text{expansion}} = - \int_{V_{\text{initial}}}^{V_{\text{final}}} P dV \quad (1)$$

If we have a reversible process, then we know P from an equation of state

Take ideal gas: $P = \frac{nRT}{V}$ and

Substitute in equation (1):

$$W_{\text{expansion}} = - \int_{V_{\text{initial}}}^{V_{\text{final}}} \frac{nRT}{V} dV$$

If we assume $n = \text{const}$ and $T = \text{const}$ (i.e. isothermal process) we can take out of integral:

$$W_{\text{expansion}} = - nRT \int_{V_{\text{int}}}^{V_{\text{final}}} \frac{dV}{V} = - nRT \ln V \Big|_{V_{\text{int}}}^{V_{\text{final}}} =$$

$$= - nRT \ln \frac{V_{\text{final}}}{V_{\text{int}}} = - nRT \ln \frac{P_{\text{int}}}{P_{\text{final}}}$$

$$\text{from } \frac{P_1}{P_2} = \frac{V_2}{V_1}$$

Compare with reversible expansion work of ideal gases:

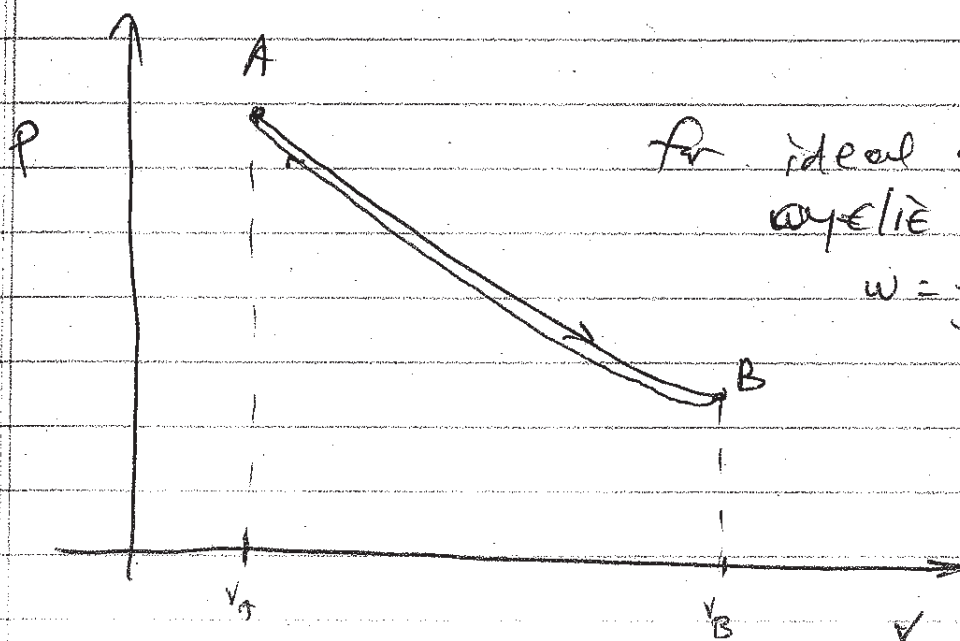
$$w = -nRT \left(1 - \frac{P_{fin}}{P_{init}} \right)$$

Q: Calculate work for both processes at $P_{fin} = 200 \text{ atm}$, $P_{init} = 100 \text{ atm}$ and see which one has more work.

What does this imply?

P-V diagrams

The total work done in a process can be shown on a P-V plot.

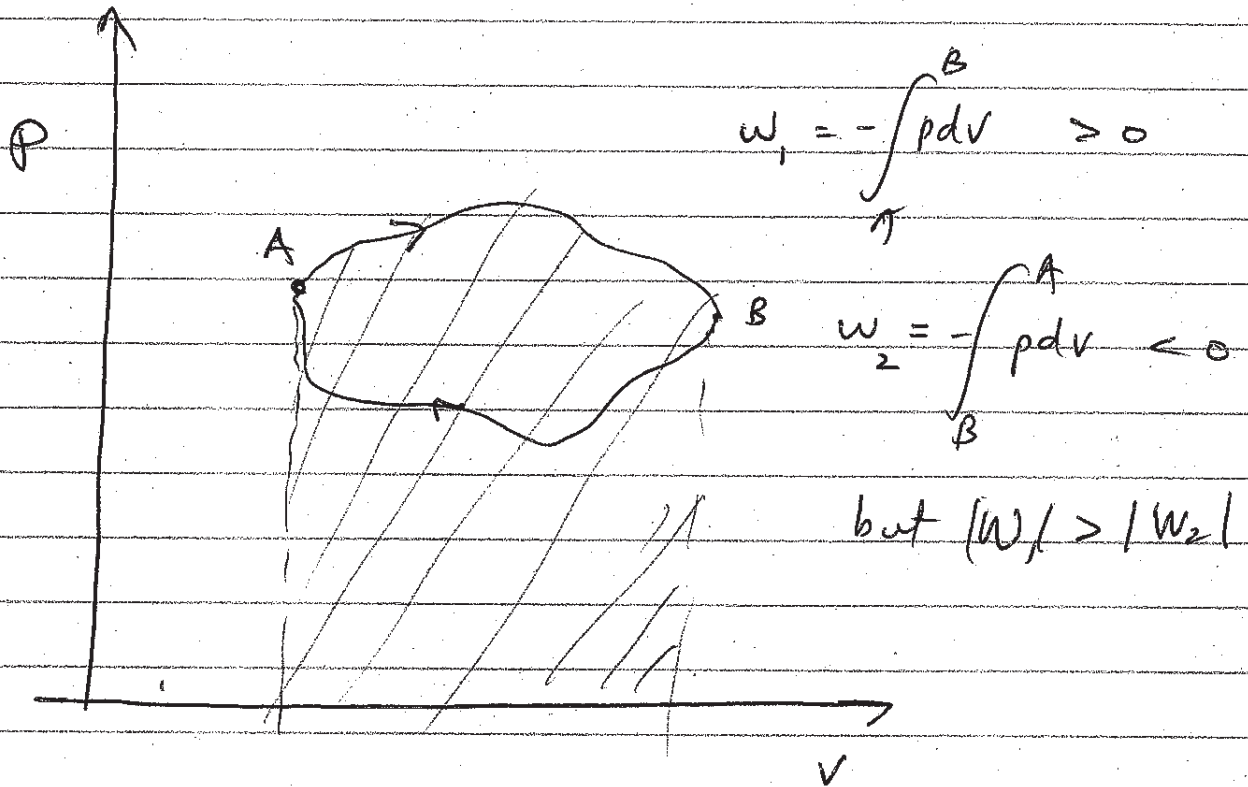


For ideal gas, and a cyclic process,

$$w = - \int_A^B p dv \neq \int_B^A p dv = 0$$

Not necessarily the same for any process. Eg:

what happens in the following process?



So total work for cyclic process is
 $W_1 - W_2 > 0$.

→ This is the principle on which steam engines work.