Atmospheric pressure is around 100kPa. The force of air on a 10cm by 10cm square is closest to

A) 1 N
B) 10 N
C) 100 N
D) 1000 N
E) 10000 N

Atmospheric pressure is around 100kPa, where Pa=Pascal is the SI unit of pressure. The force of air on a 10cm by 10cm square is closest to:

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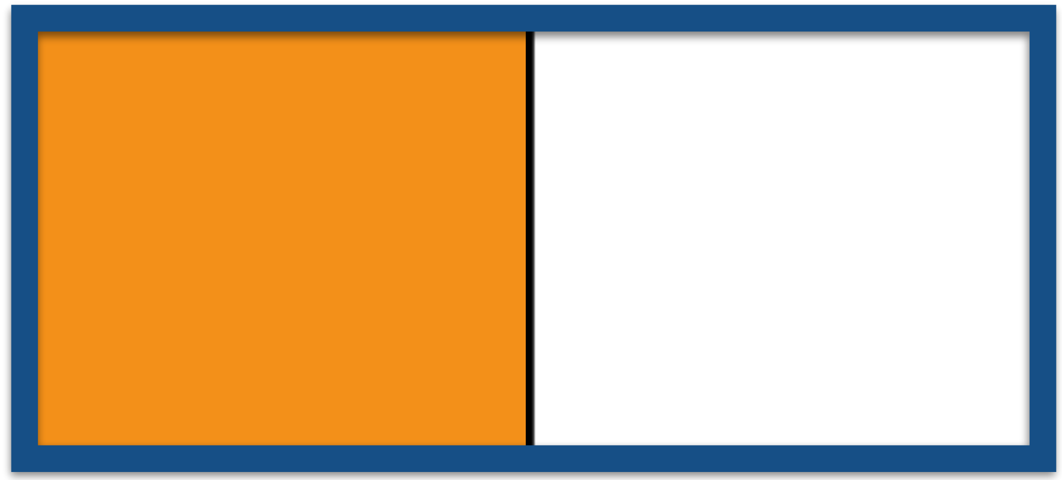
Force is equal to pressure times area. The area is $(0.1m)^2 = 0.01m^2$ and the pressure is 100,000 Pa, so the force is 1000N.

Using the ideal gas law, explain what happened to the gas in the can in each stage of the demo, and explain the final outcome.

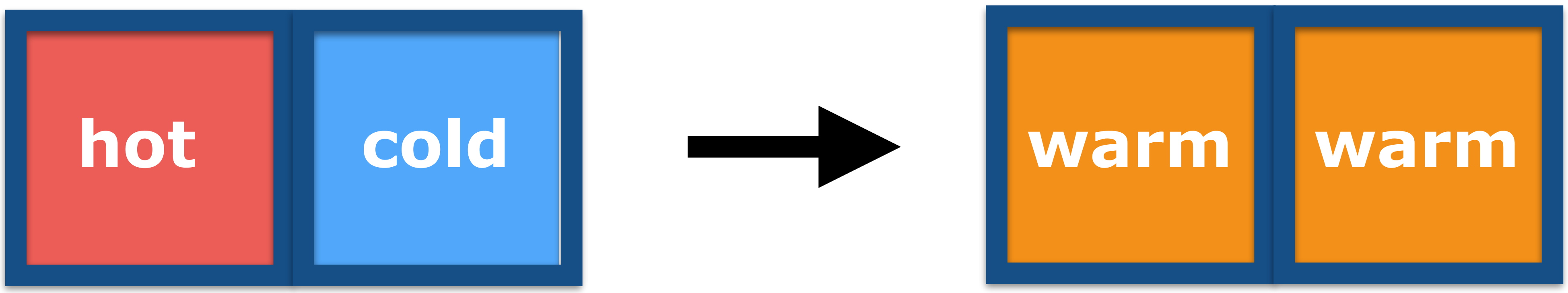


A container with a partition in the middle is filled halfway with gas.
The partition is removed instantaneously so that the gas is allowed to fill the box.

For an ideal gas, we argued last time that the temperature must remain constant (by energy conservation).

For a real gas, the temperature is observed to decrease. How is this possible?

Macroscopic Definition of Temperature



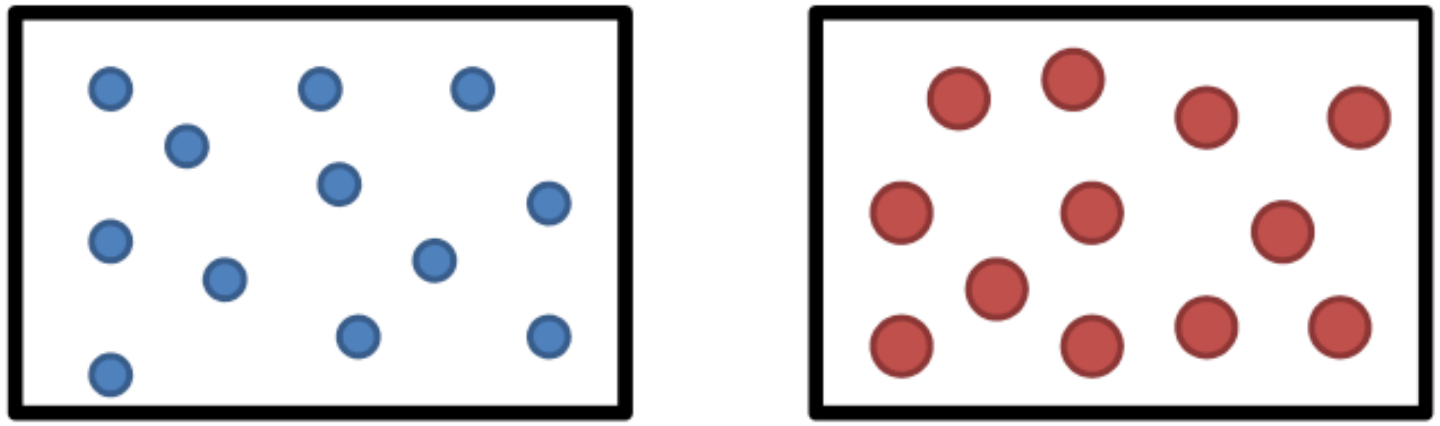
cools heats in thermal equilibrium (nothing changes)

Consider three boxes A, B, and C.

If A and B are in thermal equilibrium and B and C are in thermal equilibrium, then A and C are in thermal equilibrium.

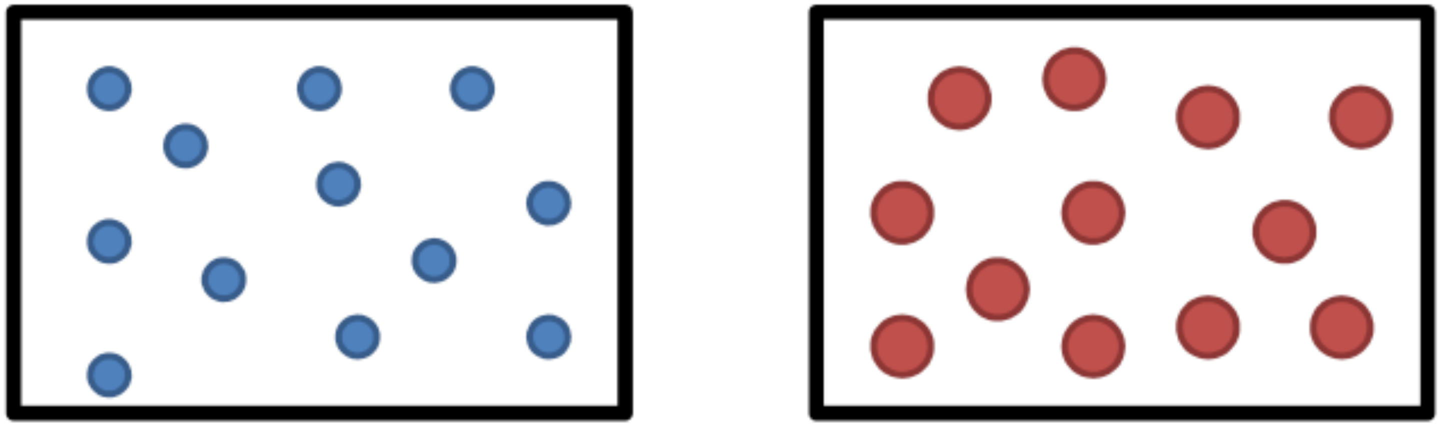
- Zeroth Law of Thermodynamics

We can define temperature to be the same for any two systems in thermal equilibrium.
The zeroth law allows us to build thermometers in a consistent way.



1 mole of helium and 1 mole of neon in identical containers are each heated from 300K to 400K. Compared to the amount of energy required to heat the helium, the amount of energy required to heat the neon is

A) significantly larger
B) significantly smaller
C) approximately the same



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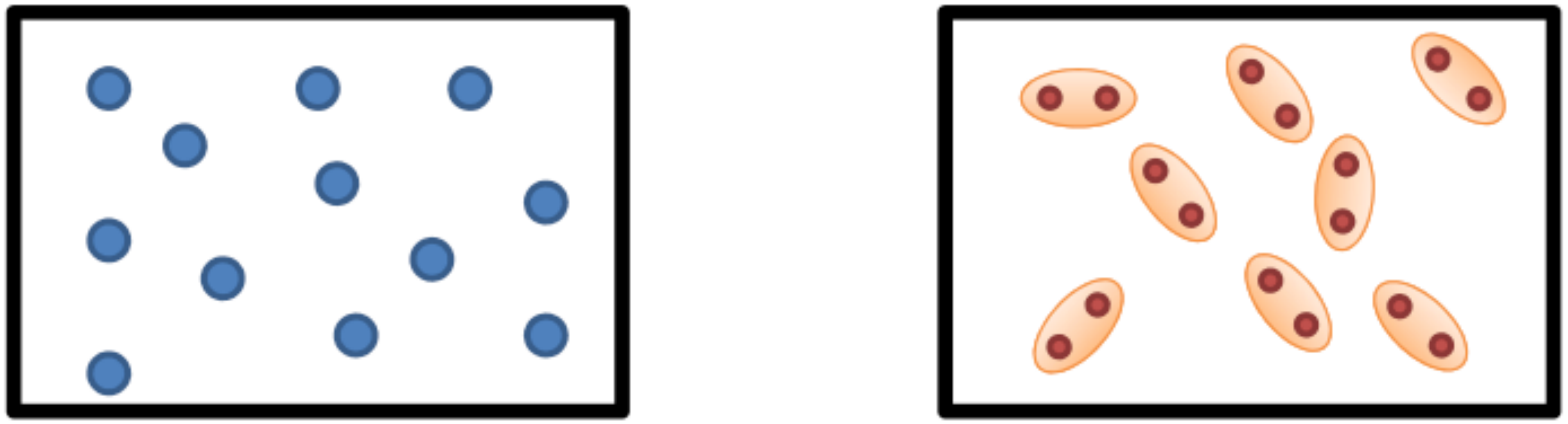
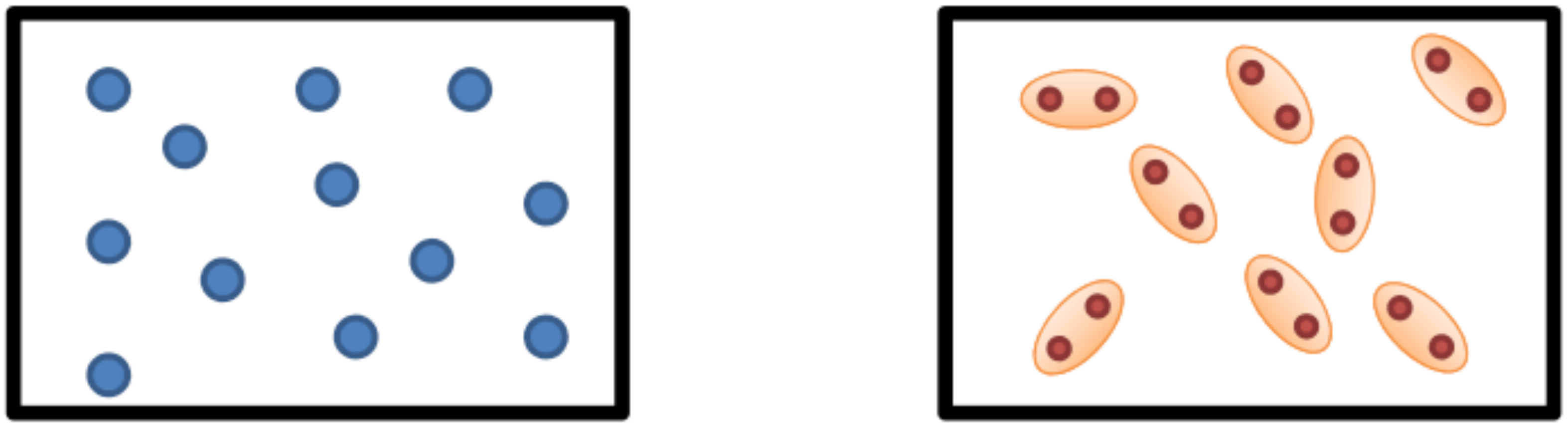
Ignoring interactions, the energy of each gas is all in translational kinetic energy. Since temperature is the average translational kinetic energy per atom and we have the same number of atoms, equal temperature increase translates to equal energy increase.



1 mole of helium and 1 mole of molecular hydrogen in identical containers are each heated from 300K to 400K. Compared to the amount of energy required to heat the helium, the amount of energy required to heat the hydrogen is

A) significantly larger
B) significantly smaller
C) approximately the same



1 mole of helium and 1 mole of molecular hydrogen in identical containers are each heated from 300K to 400K. Compared to the amount of energy required to heat the helium, the amount of energy required to heat the hydrogen is

- **A) significantly larger**
- B) significantly smaller
- C) approximately the same

In each case, the increase in translational kinetic energy is the same, but for hydrogen, we also have larger rotational kinetic energy for the molecules when the temperature is higher. Thus, we need to add more heat to increase the temperature of the hydrogen.

How is energy stored in a gas?

Equipartition Theorem

Maxwell

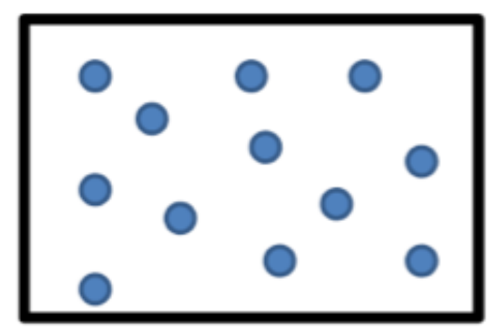
Boltzmann

"... a neglected pioneer of the kinetic theory of gases"

Waterston

Each degree of physical freedom in a system carries the same amount of energy.

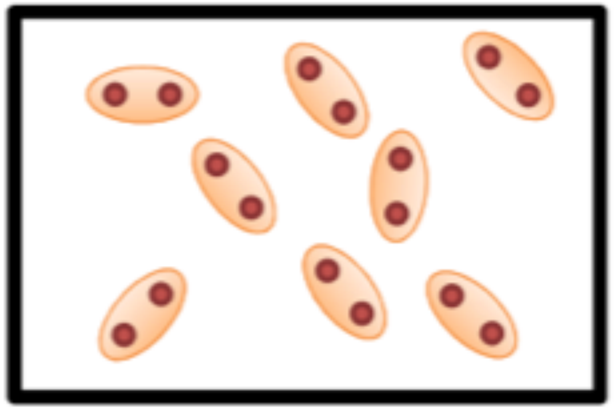
$$E_{\text{degree of freedom}} = \frac{1}{2}nRT$$



A monatomic gas can move in three dimensions, thus three degrees of freedom.

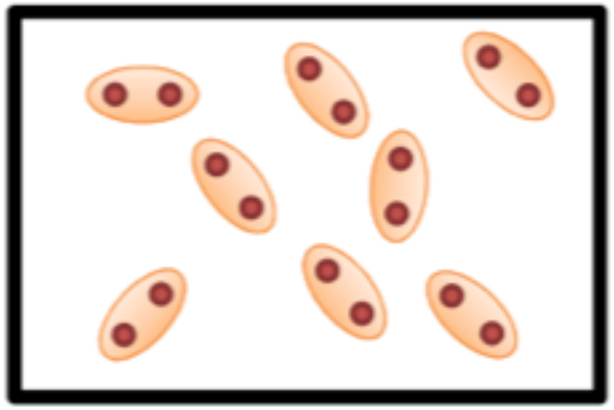
$$E = \frac{3}{2}nRT \quad \Rightarrow \quad C_V = \frac{3}{2}R$$

What about a diatomic gas?



What is the C_V of a diatomic gas?

A) 3/2R
B) 2R
C) 5/2R
D) 3R
E) 7/2R



What is the C_V of a diatomic gas?

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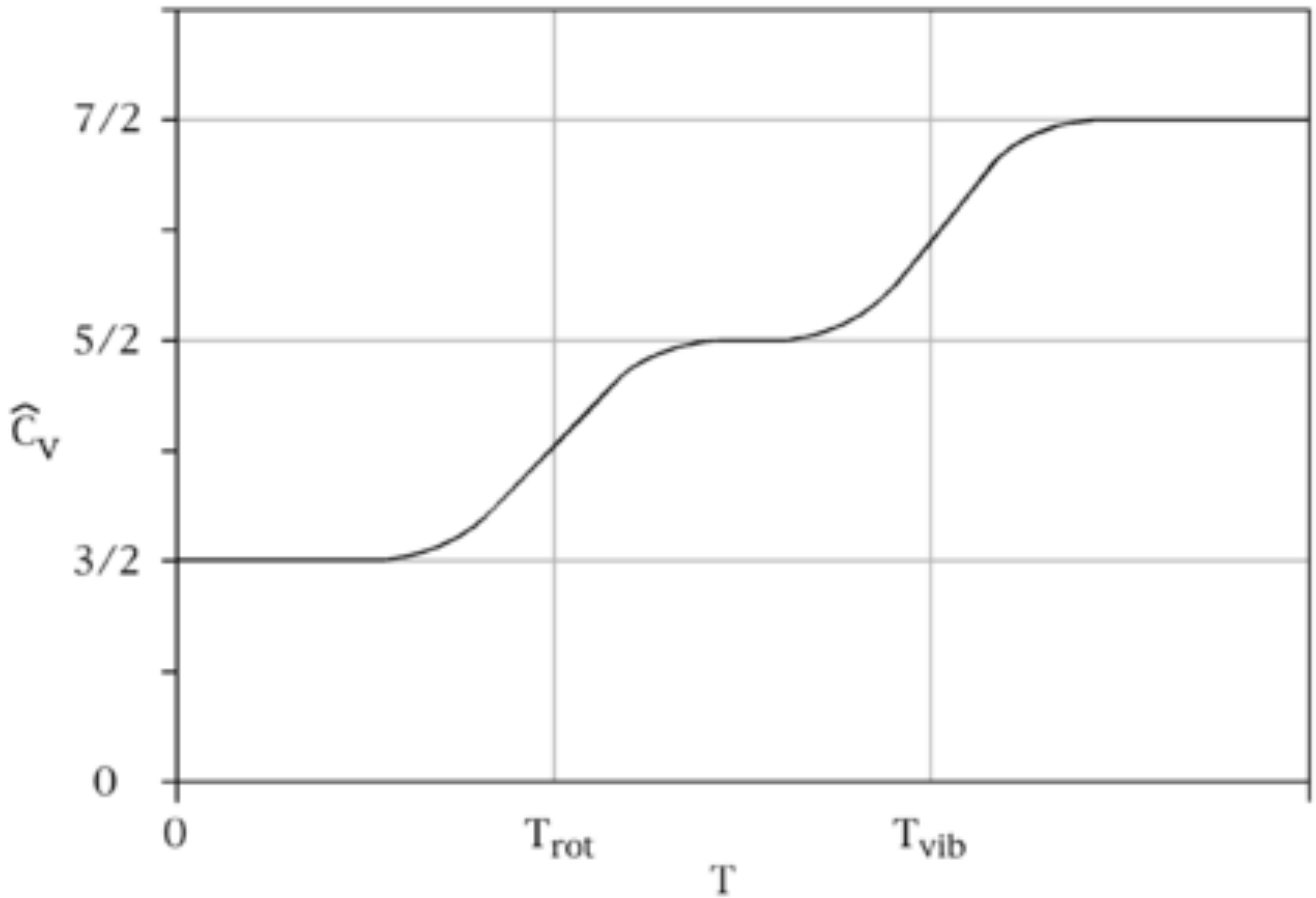
But we have 7 degrees of freedom - 3 translational, 2 vibrational, 2 rotational! Why isn't that what we measure? It turns out that these degrees of freedom aren't always "active".

Equipartition Theorem



Some degrees of freedom are “frozen out” due to quantum effects. The spacing between the energy levels is too high to be accessed at low temperatures.