Thermal Physics

PHYS/BMME 441

Lectures 1 & 2
Chapter 1

When & Where: **Mon. / Wed. 8:45- 10:00 AM, Phillips 247**

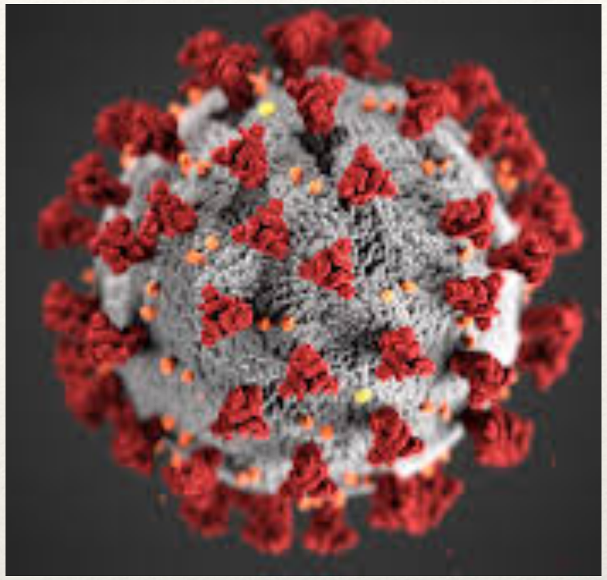
Textbook: An Introduction to Thermal Physics**, Daniel V. Schroeder**

Website: **sakai.unc.edu**

Course Logistics

- All the information is on Sakai
- Textbook: *An Introduction to Thermal Physics*, by Daniel V. Schroeder
- 2 exams and 1 final
- 6 HW sets (~bi-weekly). They will be posted on Sakai. Tentative schedule on Sakai. Pay attention to the due dates on the assignment!
- All the course materials (lecture slides, HW, etc…) will be posted on the Sakai site.
- Office hours: TBA

Covid-19

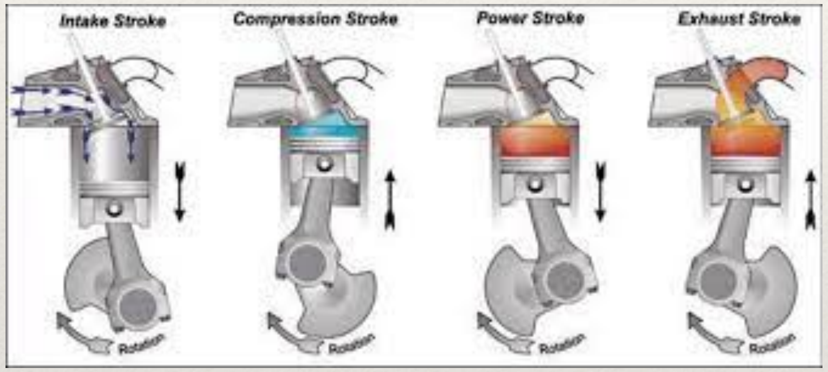


- Health and well being of everyone in the community is a priority
- If you have Covid symptoms (*fever/chills, headaches/body aches, sore throat, cough, congestion, loss of taste or smell*) or feel ill in general ***do not come to class***
- University Approved Excused Absence ***is not required*** for missing class due to health reasons and there will be no penalty.
- In such a case contact me and we'll discuss alternatives to keep up with the coursework. There will be flexibility.
- All the core materials will be posted on the website

General Outlook

Thermodynamics: Physics of macroscopic systems

Temperature and heat, engines, refrigerators, efficiencies, phase transitions…



Not fundamental, but very general: from your teacup to black holes

Statistical Mechanics: How do laws of thermodynamics emerge from microscopic physics?

What happens when you put $\sim 10^{23}$ particles together?

Class #	Date	Chapter	Topics
1	M 8/21	1 - Energy in	1.1 Thermal Equilibrium, 1.2 The Ideal Gas, 1.3 Equipartition of Energy,
	W 8/23		**No Class**
2	M 8/28	Thermal Physics	1.4 Heat and Work , 1.5 Compression Work, 1.6 Heat Capacities
3	W 8/30	2 -The Second Law	2.1 Two-State Systems, 2.2 The Einstein Model of a Solid,
	M 9/4	***Labor Day***	**No Class**
4	W 9/6		2.3 Interacting Systems,
5	M 9/11		2.4 Large Systems, 2.5 The ideal gas
6	W 9/13		2.6 Entropy
7	M 9/18	3 - Interactions and	3.1 Temperature, 3.2 Entropy and Heat
8	W 9/20	Implications	3.3 Paramagnetism, 3.4 Mechanical Equilibrium and Pressure,
	M 9/25	***Well-being Day***	**No Class**
9	W 9/27		3.5 Diffusive Equilibrium and Chemical Potential
10	M 10/2	4 - Engines and	4.1 Heat Engines, 4.2 Refrigerators
11	W 10/4	Refrigerators	4.3 Real Heat Engines
	M 10/9	*Midterm 1*	*Chapters 1-3: Fundamentals*
12	W 10/11		4.4 Real Refrigerators
13	M 10/16	5 - Free Energy and	5.1 Free Energy as Available Work, Thermodynamic Identities,
	W 10/18	***Catch-up Day***	**No Class**
14	M 10/23	Thermodynamics	5.2 Free Energy as a Force toward Equilibrium,
15	W 10/25		5.3- 5.4 Phase Transformations of Pure Substances and Mixtures,
16	M 10/30		*5.5 Dilute solutions 5.6 Phase equilibirum (time permitting)*
17	W 11/01	6 - Boltzmann Statistics	6.1 Boltzmann Factor
18	M 11/6		6.2 Average Values
	W 11/08	*Midterm 2*	*Chapters 4-5: Thermodynamics*
19	M 11/13		6.3 The Equipartition Theorem, 6.4 The Maxwell Speed Distribution,
20	W 11/15		6.5 Partition Functions and Free Energy
21	M 11/20		6.6 Partition Functions for Composite Systems, 6.7 Ideal Gas Revisited
	W 11/22	***Thanksgiving***	**No Class**
22	M 11/27	7 - Quantum Statistics	7.1 The Gibbs Factor, 7.2 Bosons and Fermions,
23	W 11/29		7.3 Degenerate Fermi Gases, 7.4 Blackbody Radiation,
24	M 12/4		7.5 Debye Theory of Solids, 7.6 Bose-Einstein Condensation
25	W 12/6		Review
Friday	*12/15, 4PM*	*Final Exam*	*Including all materials covered*

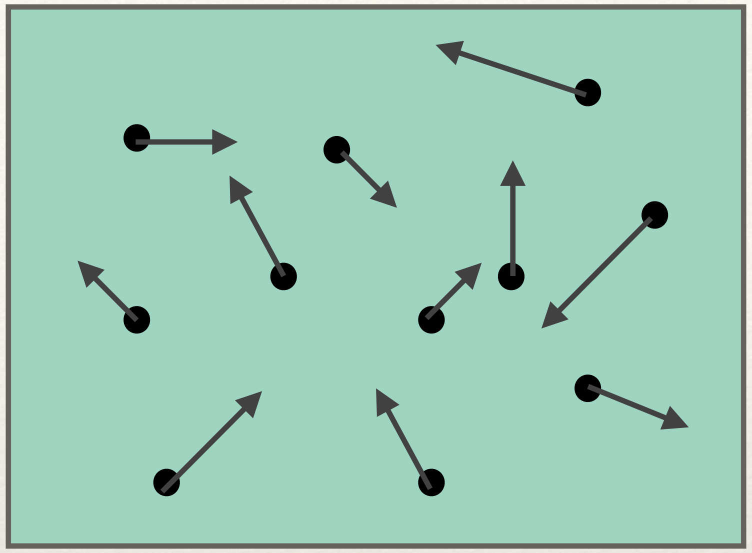
Outlook

Chapter 1: Energy in Thermal Physics (chapters 1.1 to 1.6)

read chapters:

- ***1.1*** Thermal Equilibrium
- ***1.2*** Ideal Gas
- ***1.3*** Equipartition of Energy
- ***1.4*** Heat and Work
- ***1.5*** Compression Work
- ***1.6*** Heat Capacities

Thermal Equilibrium

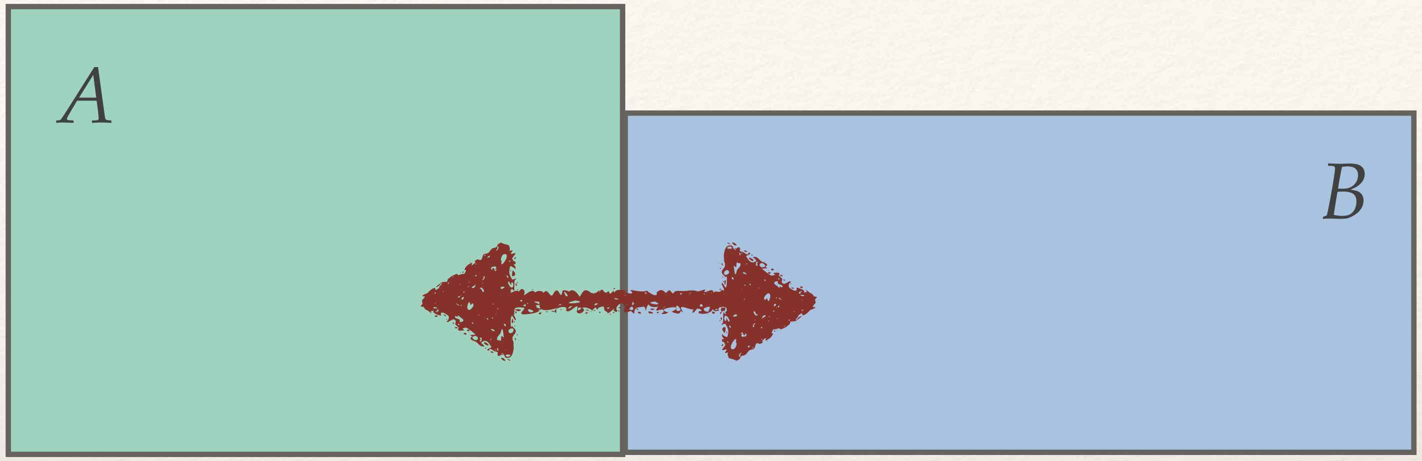


Thermal Equilibrium: An isolated system, after a sufficiently long amount of time, will settle in a state where no further change is observable.

Relaxation Time: Time it takes to reach thermal equilibrium.

State of the system described by a few macroscopic parameters (e. g. pressure, volume for a gas)

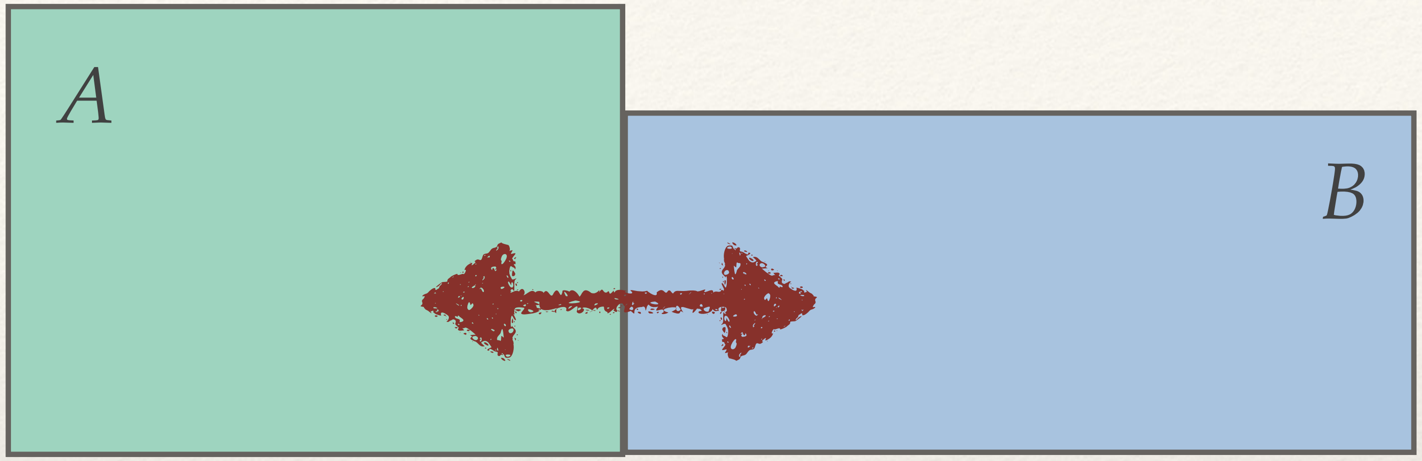
Thermal Equilibrium



After two bodies are in contact for a long time they will be in *thermal equilibrium*

Temperature: The quantity that is the same for two systems when they reach thermal equilibrium

Thermal Equilibrium

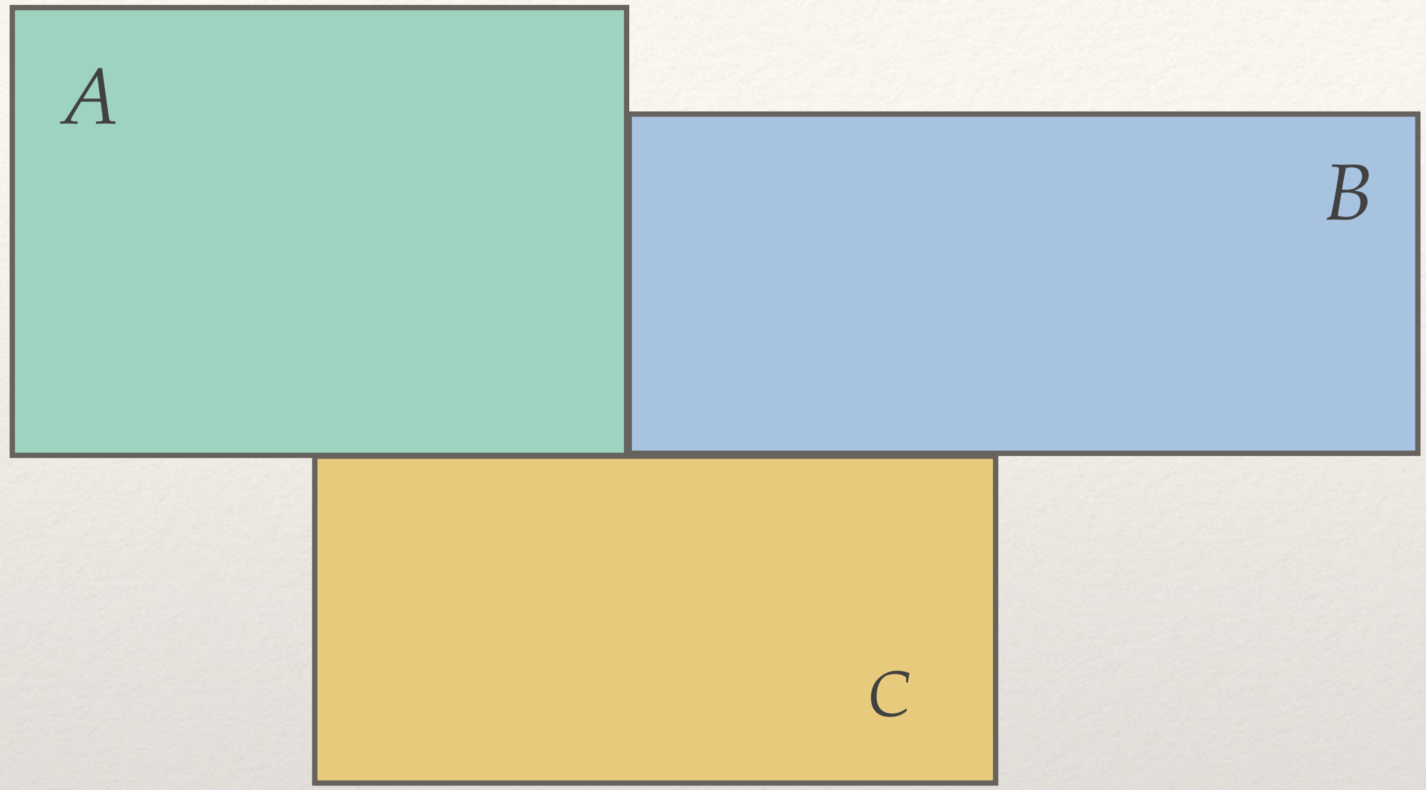


After two bodies are in contact for a long time they will be in *thermal equilibrium*

They will no longer exchange energy.

Temperature: Measure of the tendency of an object to give energy to its surroundings.

Thermal Equilibrium



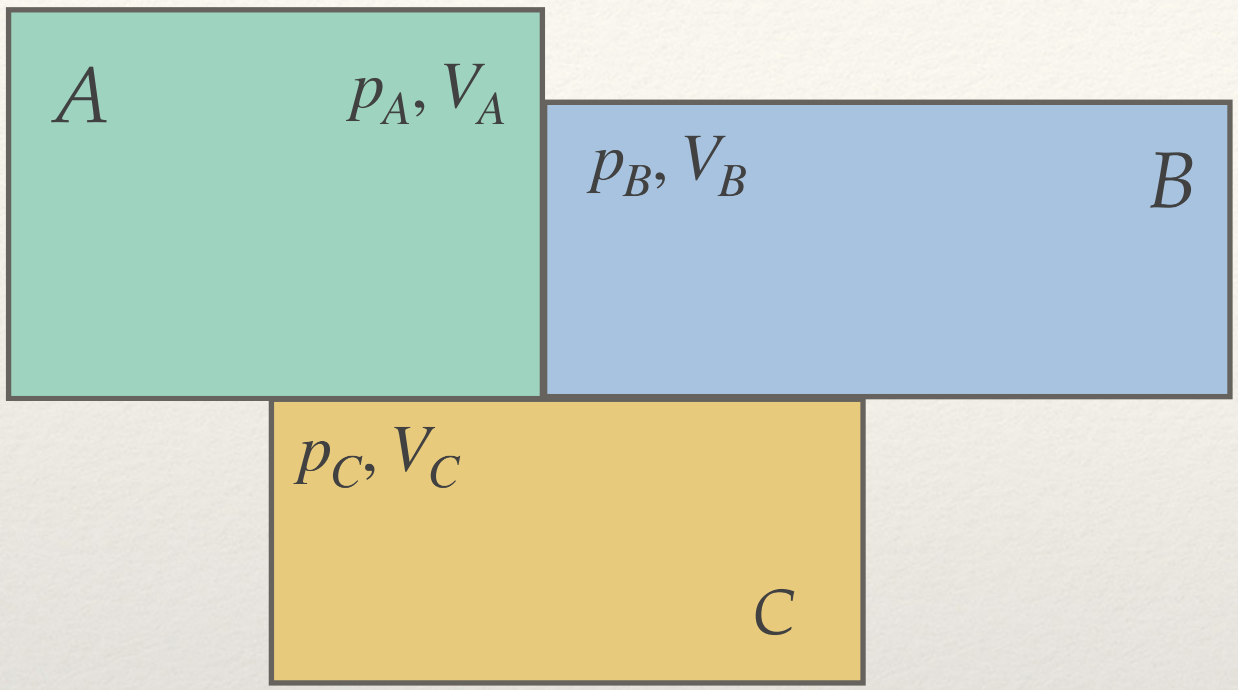
Zeroth Law:

If A and B are in thermal equilibrium with C, they are also in equilibrium with each other.

Definition of temperature follows from the Zeroth Law

[see lecture notes by Tong]

Thermal Equilibrium



Equilibrium relates

(p_A, V_A) to (p_B, V_B)

$$A \leftrightarrow B: \quad p_A = f_{AB}(V_A; p_B, V_B)$$

(p_A, V_A) to (p_C, V_C)

$$A \leftrightarrow C: \quad p_A = f_{AC}(V_A; p_C, V_C)$$

$$\Rightarrow f_{AB}(V_A; p_B, V_B) = f_{AC}(V_A; p_C, V_C)$$

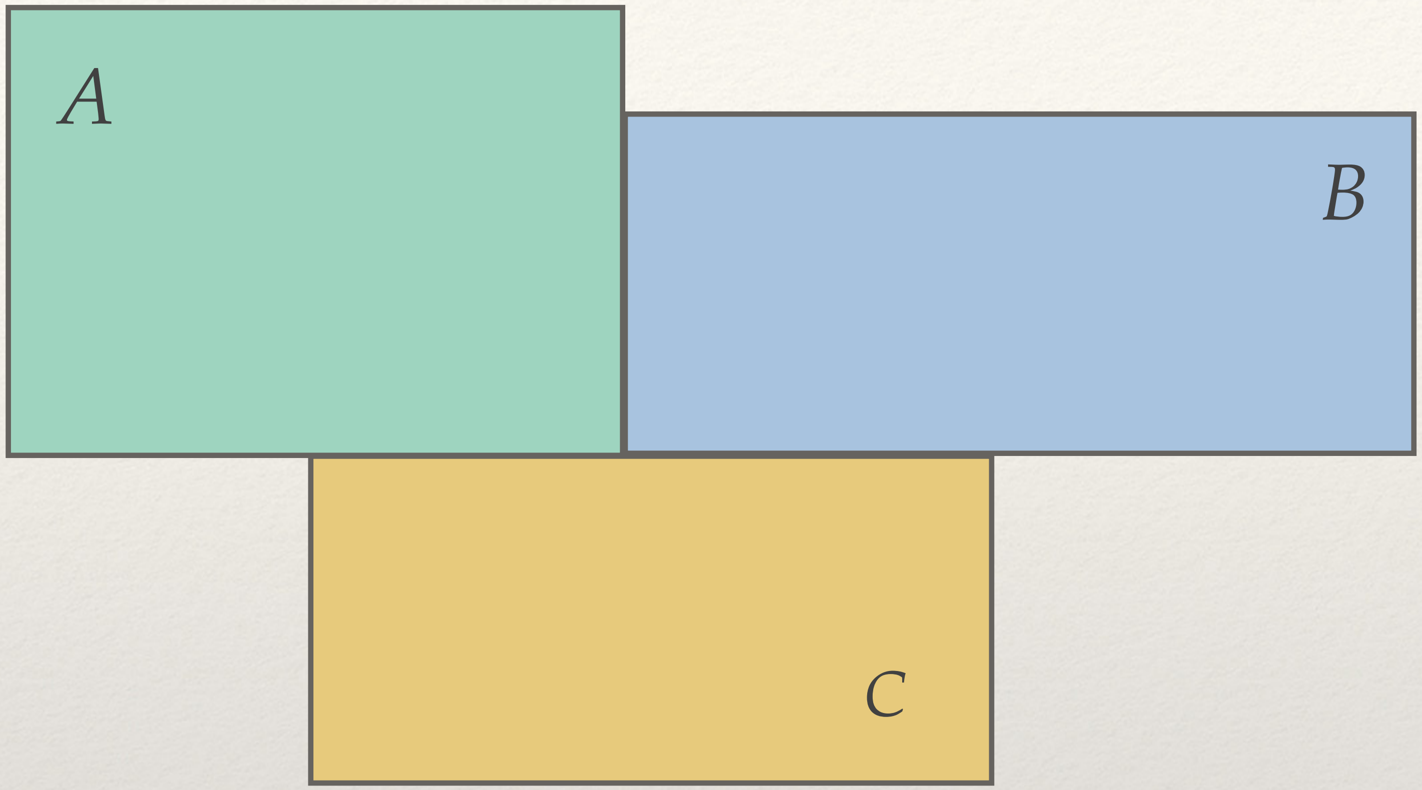
Zeroth Law: *B & C are in equilibrium:* (p_B, V_B) to (p_C, V_C)

$$\Rightarrow f_{AB} = V_A \times f_B(p_B, V_B) \qquad T_B = f_B(p_B, V_B)$$

e.g. ideal gas: $T = \dfrac{pV}{Nk}$

``equation of state''

Thermal Equilibrium



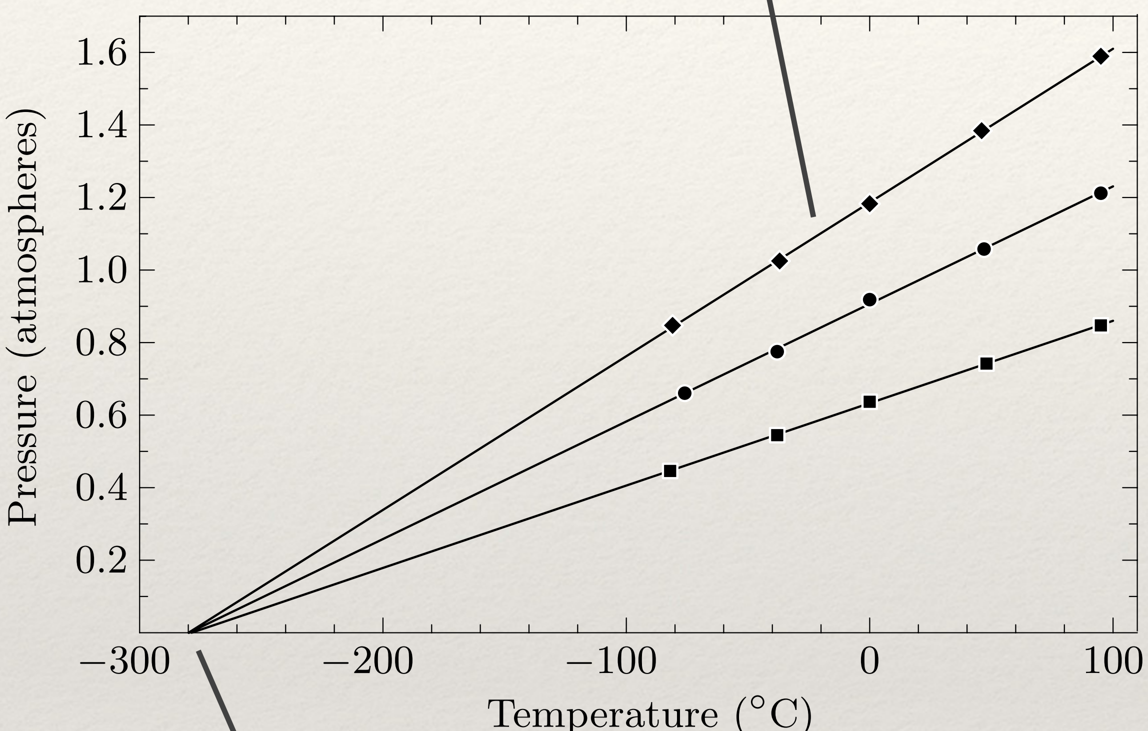
Zeroth Law:

If A and B are in thermal equilibrium with C, they are also in equilibrium with each other.

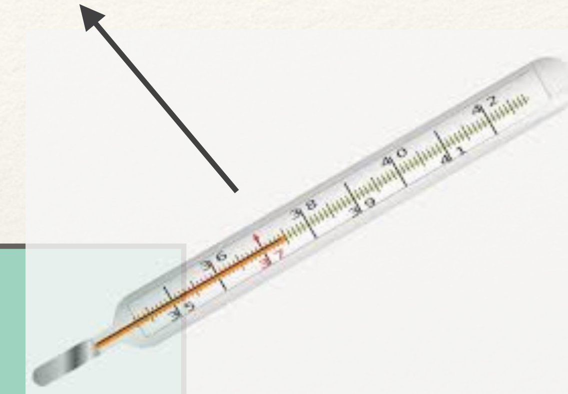
$$T_A = T_B = T_C$$

Temperature: units

Different amounts of gas



Mercury expands when heated



$T_{gas} = T_{thermometer}$

gas at constant V

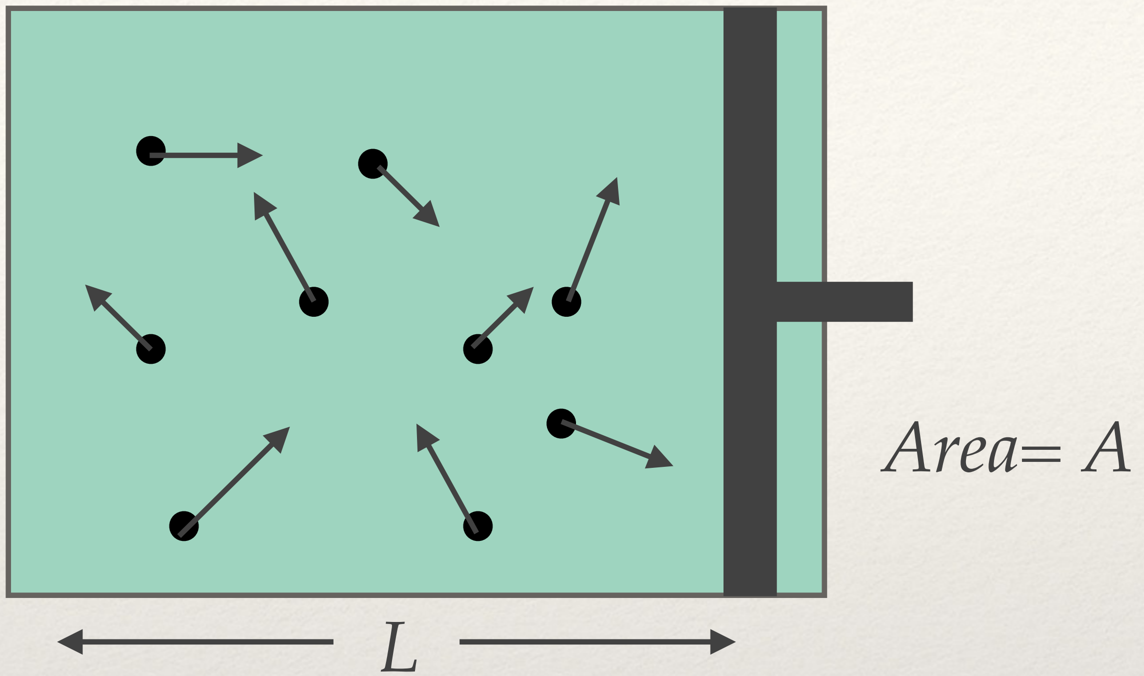
Alternatively: different gases,
electric properties that change with T etc…

-273 °C = 0 K *absolute temperature scale: Kelvin*

pressure goes to zero at *0K*

Temperature and energy: ideal gas

[Daniel Bernoulli, 1783]



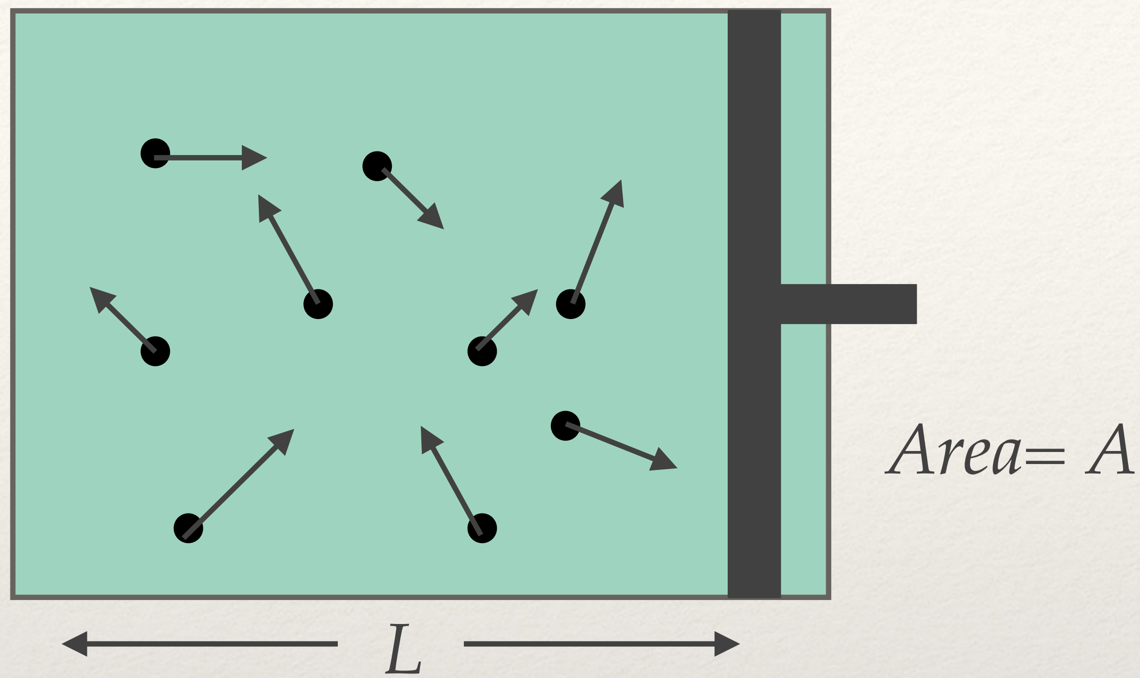
e.o.s

$$P = \frac{\bar{F}_{piston}}{A} = \frac{-mN\frac{\Delta v_x}{\Delta t}}{A} = -\frac{mN}{A}\frac{(-2v_x)}{2L/v_x} = \frac{N}{V}mv_x^2 \qquad PV = Nmv_x^2 = NkT$$

Average translational Kinetic energy: $\bar{K}_{trans.} = \frac{1}{2}m\bar{v}^2 = \frac{1}{2}m(v_x^2 + v_y^2 + v_z^2) = \frac{3}{2}kT$

For a typical gas at room temperature: $\sqrt{\bar{v}^2} \sim ??\ m/s$ guess!

Temperature and energy: ideal gas



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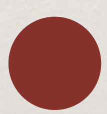
For a typical gas at room temperature: $\sqrt{\bar{v}^2} \sim 10^2 m/s \sim 200\, mph$

Equipartition theorem

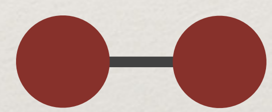
Each degree of freedom contributes $\frac{1}{2}kT$ to the average thermal energy

$$E_{thermal} = N \times N_{dof} \times \frac{1}{2}kT$$



$N_{dof} = 3$

e.g. N_2, O_2, …



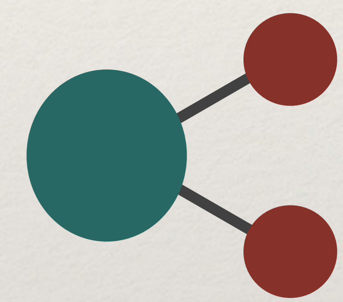
$N_{dof} = 3 + 2 + 2 \times 1$

translation

rotation

vibration

e.g. H_2O



$N_{dof} = 3 + 3 + 2 \times 3$

Note: vibrational d.o.f. typically don't contribute at room temp.
They become relevant at high T. A consequence of quantum mechanics! More on this later…

Boltzmann Constant

Each degree of freedom contributes $\frac{1}{2}kT$ to the average thermal energy

$$E_{thermal} = N \times N_{dof} \times \frac{1}{2}kT$$

$$k \approx 1.38 \times 10^{-23} J/K$$

Typical number of atoms in a macroscopic system: $N_A \approx 6 \times 10^{23}$

$$R \equiv kN_A \approx 8J/K \quad \Rightarrow \quad E_{thermal} = N_{moles} \times N_{dof} \times RT$$

Number of sand grains in the world $\sim 10^{17}$

Number of galaxies in the universe $\sim 10^{22}$

Energy

SI unit for energy: *Joule (J)* $\quad 1\,J = 1\,\dfrac{kg \times m^2}{s^2}$

1 calorie (cal) (energy to raise T of water by 1^{O}C)$\approx$ 4.2J

- Average food intake of a person $\approx 10\,MJ/$day

(*food calorie:* 1 Cal = 1 kcal)

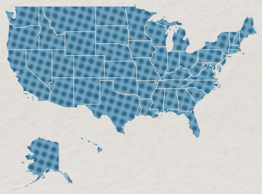
- Energy stored in *1 kg* of

3.5 MJ

44 MJ

- Average per capita energy use

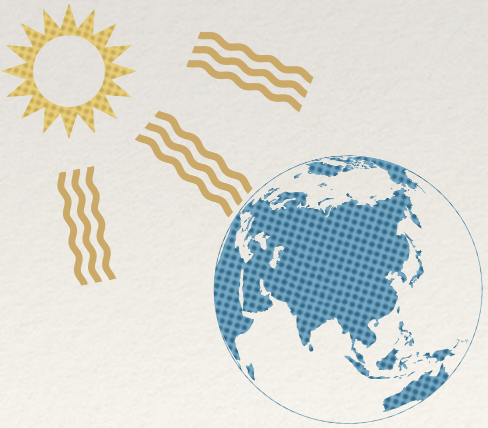
200 MJ / day



1 GJ / day

- Solar energy entering Earth's atmosphere

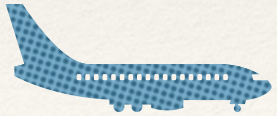
$1.73 \times 10^8\ GJ/$sec



- Mass energy (*1 kg*)

$E = mc^2 \approx 9 \times 10^{10}\ MJ$

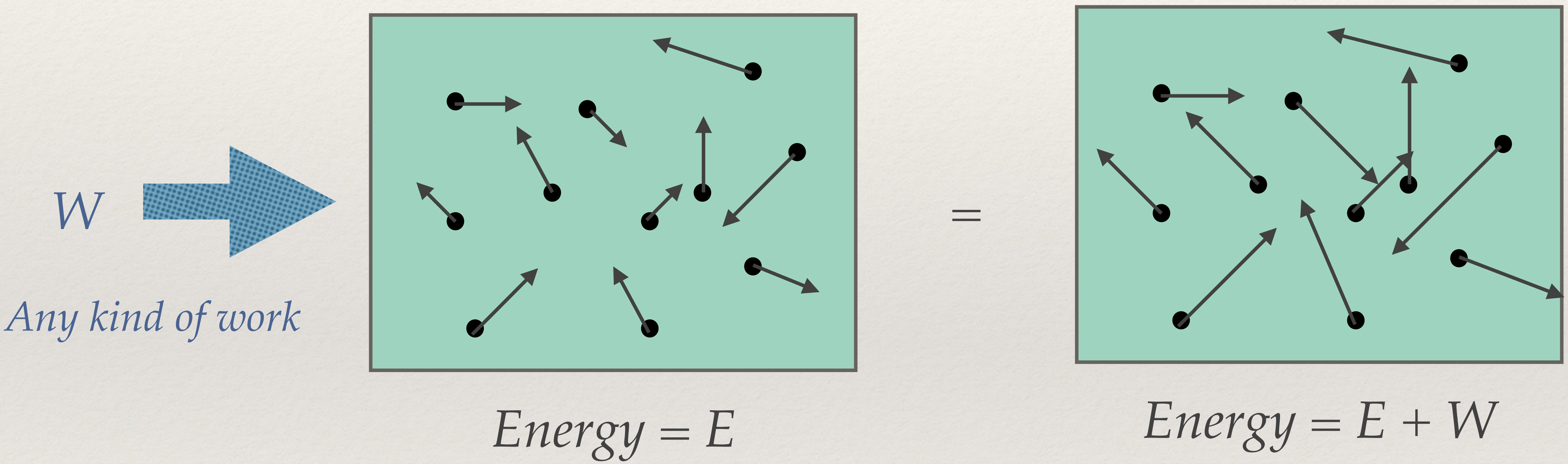
- 747 flight to NYC to LA



$\approx 2000\,GJ$

Heat and Work

Isolated System



$$\Delta E = W$$

Recap

Zeroth Law: A & B, B & C in equilibirium $\Rightarrow$ A & C in equilibirium

Notion of *Temperature*

Average kinetic energy per particle

$$\bar{K} = \frac{3}{2}kT$$

equipartition theorem

$$E_{thermal} = N \times N_{dof} \times \frac{1}{2}kT$$

Thermal energy of the system

First Law:

$$\Delta E = Q + W$$

Change in energy in the system

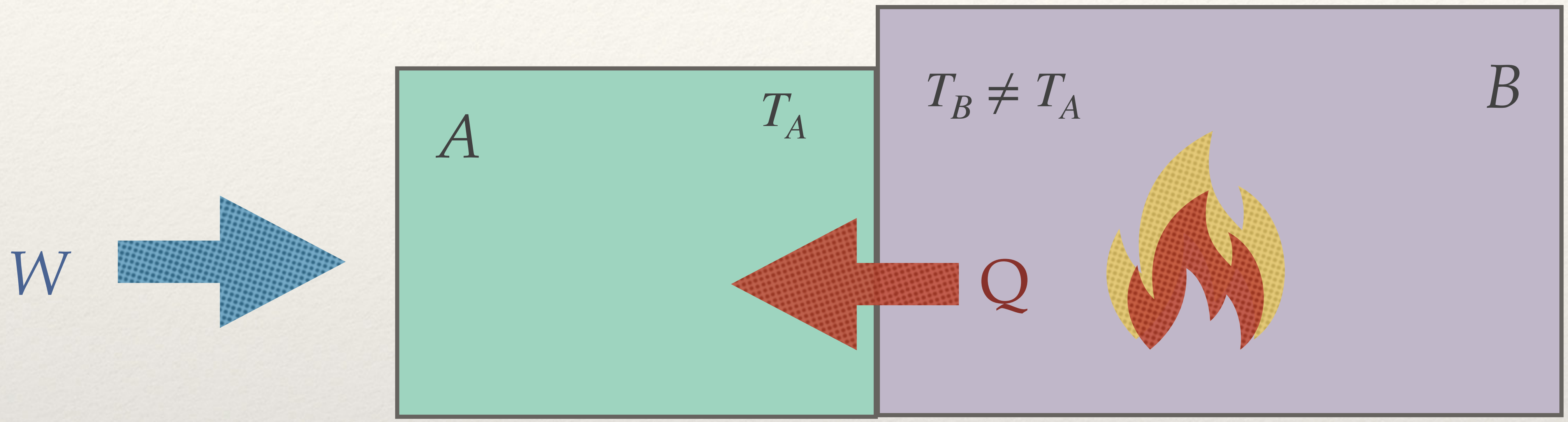
Heat

Spontaneous transfer of energy from surroundings

Work done on the system

Heat and Work

Non- isolated system e.g. in thermal contact with another system



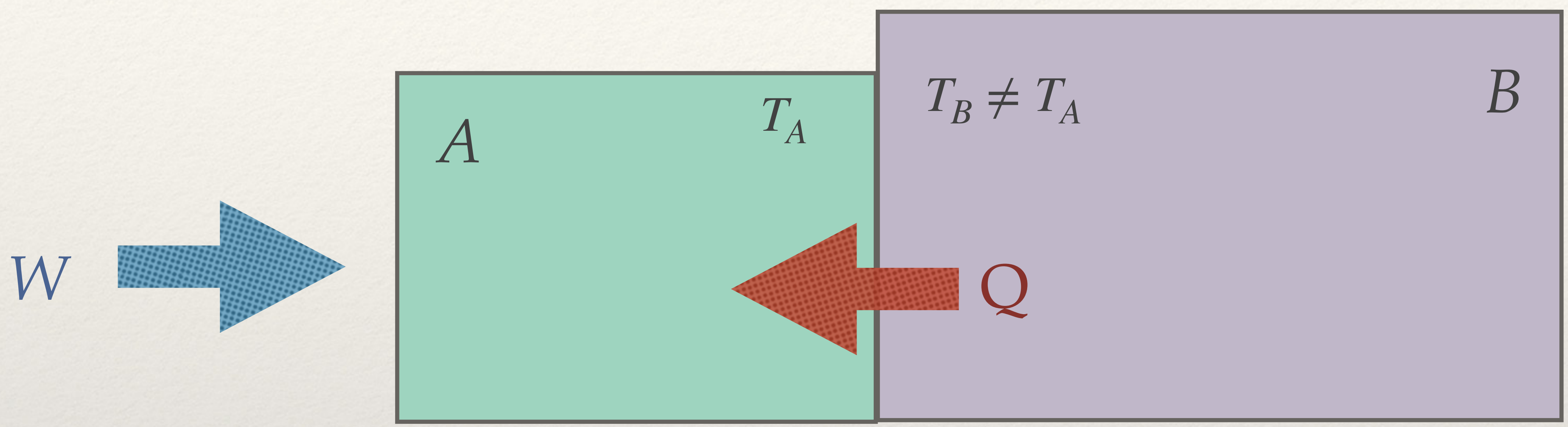
<u>*First Law:*</u> $\Delta E = Q + W$

Heat: Q = *amount of energy transfer that is not work*
(spontaneous transfer of energy due to difference in temperatures)

\+ sign for Q, W: energy *enters* the system
\- sign for Q, W: energy *leaves* the system

Heat and Work

Non- isolated system e.g. in thermal contact with another system



First Law: $\Delta E = Q + W$

Q is *not* a form of energy. It describes *transfer* of energy.

E *is* a state function: $E = E(p, V)$

Q and W are *not* state functions. ``$E = Q + W$'' has no meaning

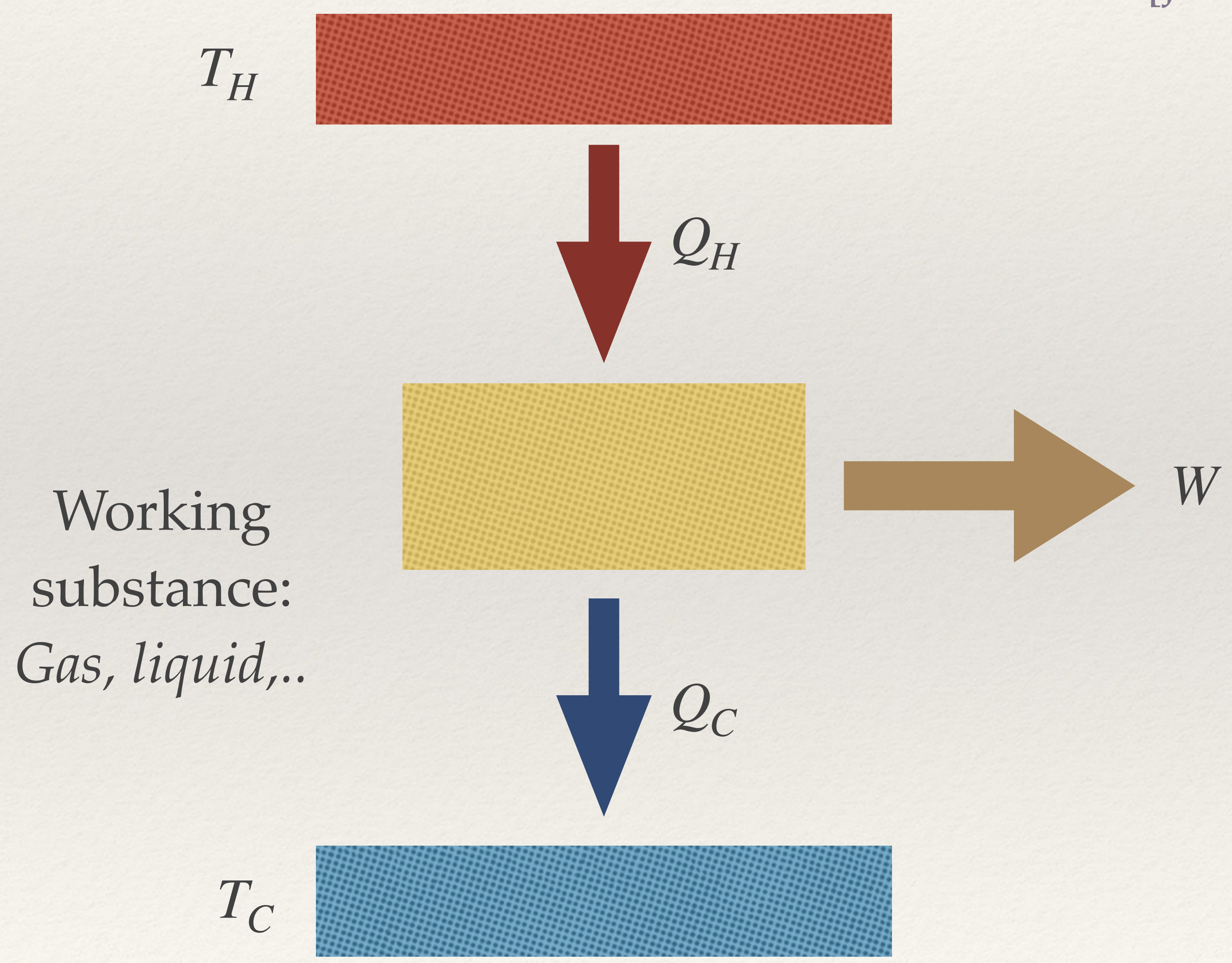
A First Look at Heat Engines

90% US energy includes heat engines one way or another.

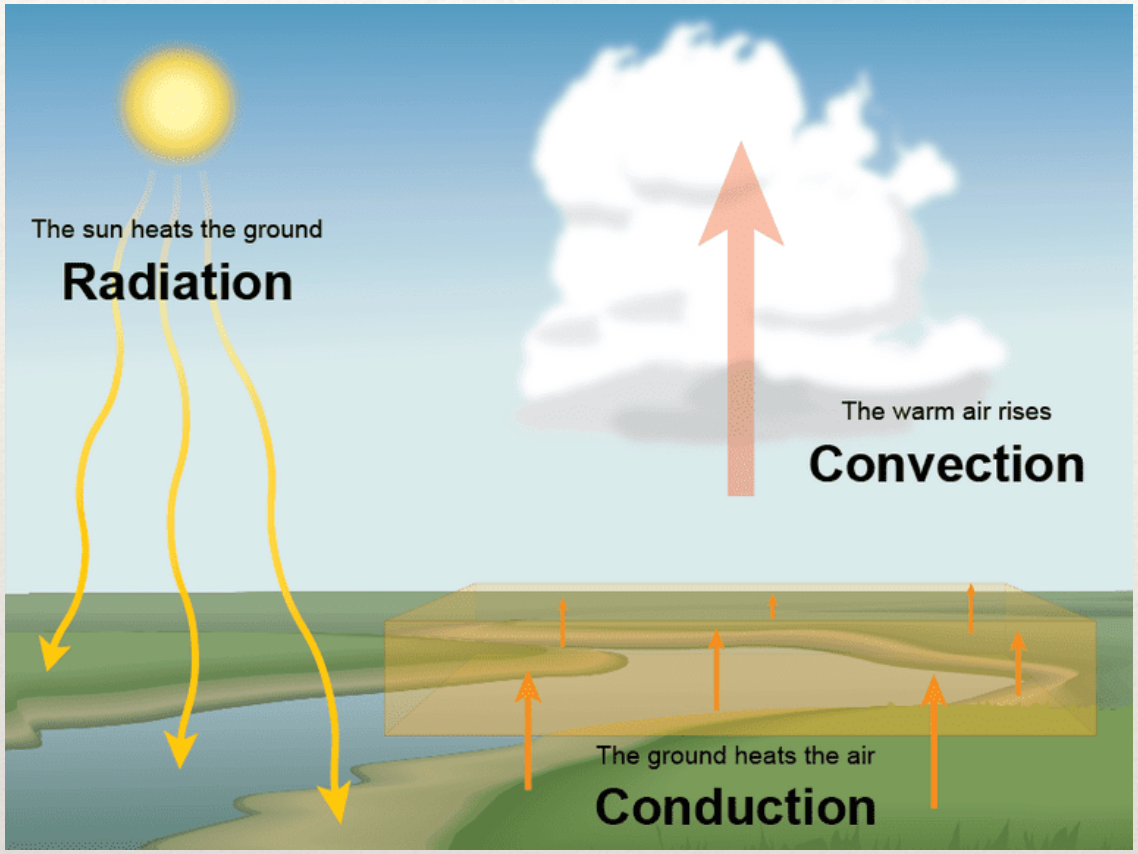
Production: engines, power plants, etc…

Consumption: space heating / cooling, refrigerators,…

[Jaffe, Taylor; Physics of Energy]



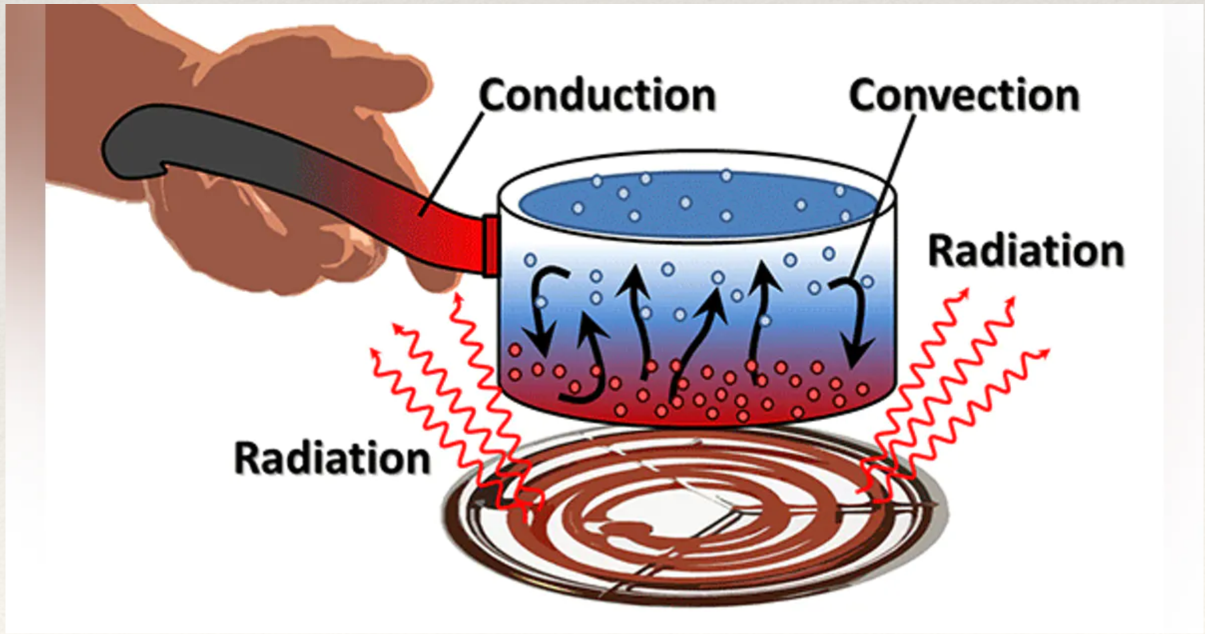
Heat Transfer



Conduction: Direct contact between two objects

Convection: Motion of molecules

Radiation: Electromagnetic waves



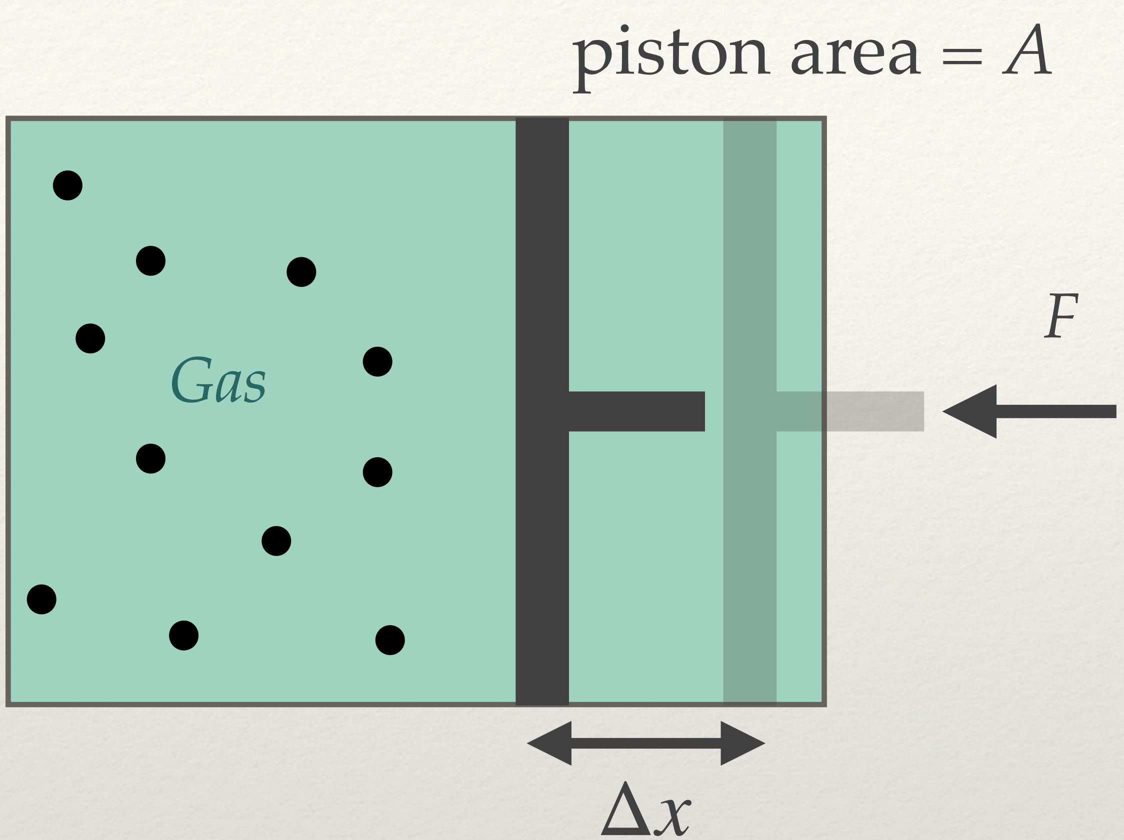
Quasi-Static Processes

So far the energy transfer could be as abrupt as possible, and drive the system out of thermal equilibrium.

We will now be more deliberate and add energy *slowly* such that the system will remain in *equilibrium* during the process.

We can describe the system with the state variables p, V

Compression Work



Work done on gas: $\Delta W = F\Delta x = PA\Delta x = -p\Delta V$

quasi-static

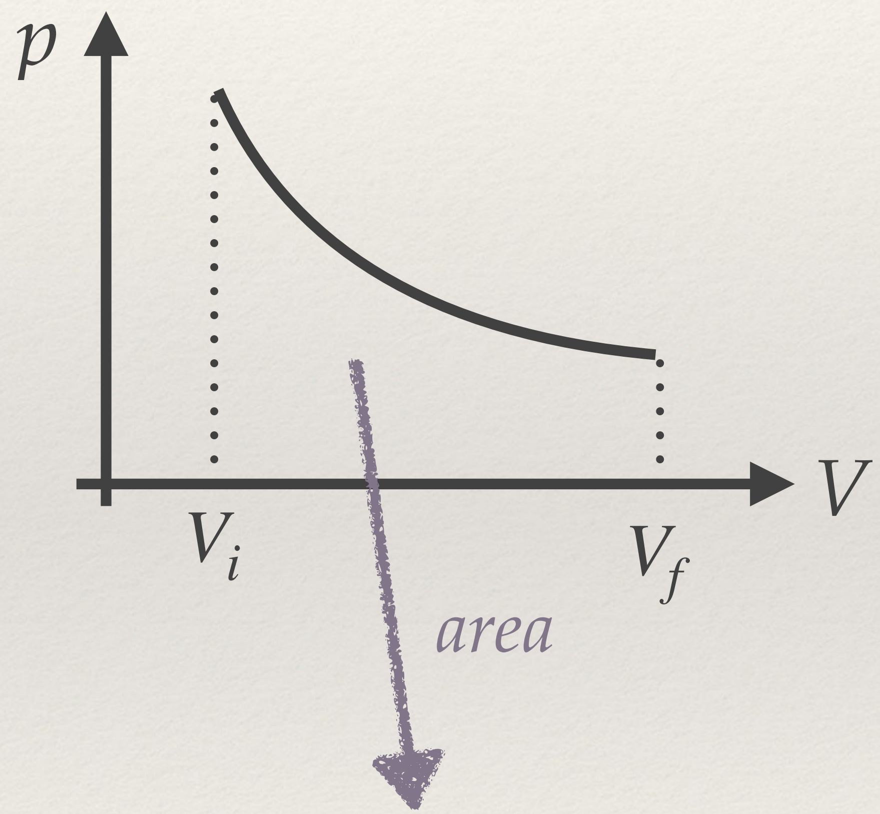
$\Delta W > 0:$ work done on system

$\Delta W < 0:$ system performs work

pV Diagrams

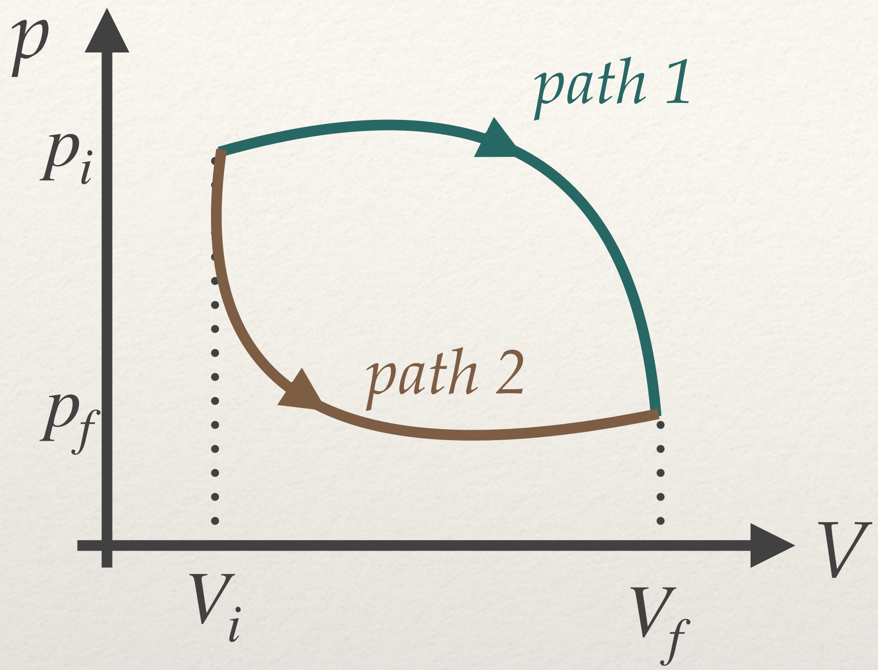
$$\Delta W = -p\Delta V$$

In general pressure changes as volume changes



$$W = -\int_{V_i}^{V_f} p(V)dV$$

pV Diagrams

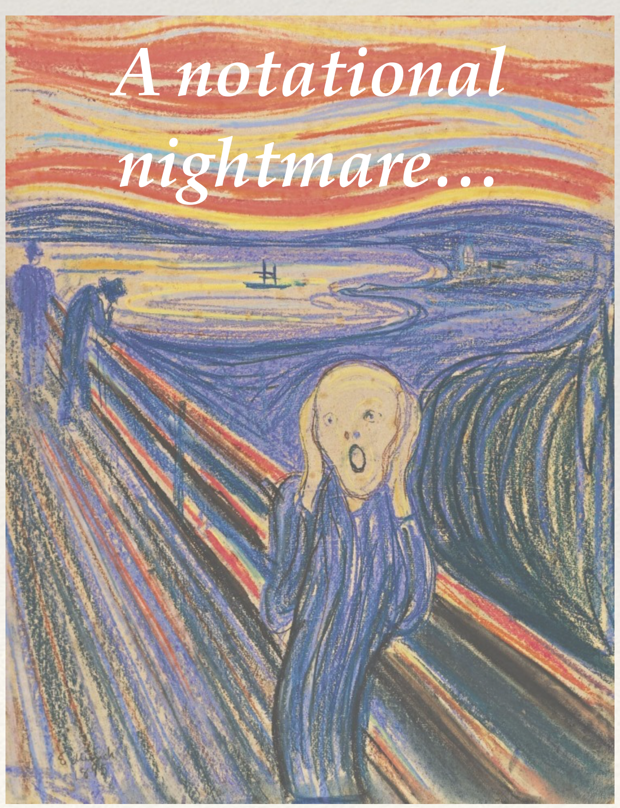


Work is path dependent

$$W_1 = -\int_{path\,1} p(V)dV \neq W_2 - \int_{path\,2} p(V)dV$$

(unlike energy)

$$\int_{path\,1} dE = \int_{path\,2} dE = E(p_f, V_f) - E(p_i, V_i)$$



$$"dE = dQ + dW"$$

not ``changes in Q, W''

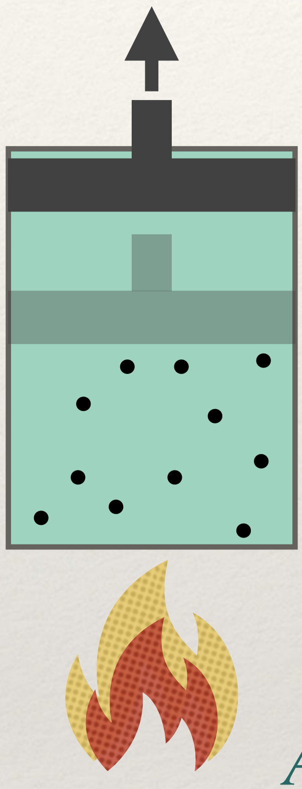
Alternative notations:

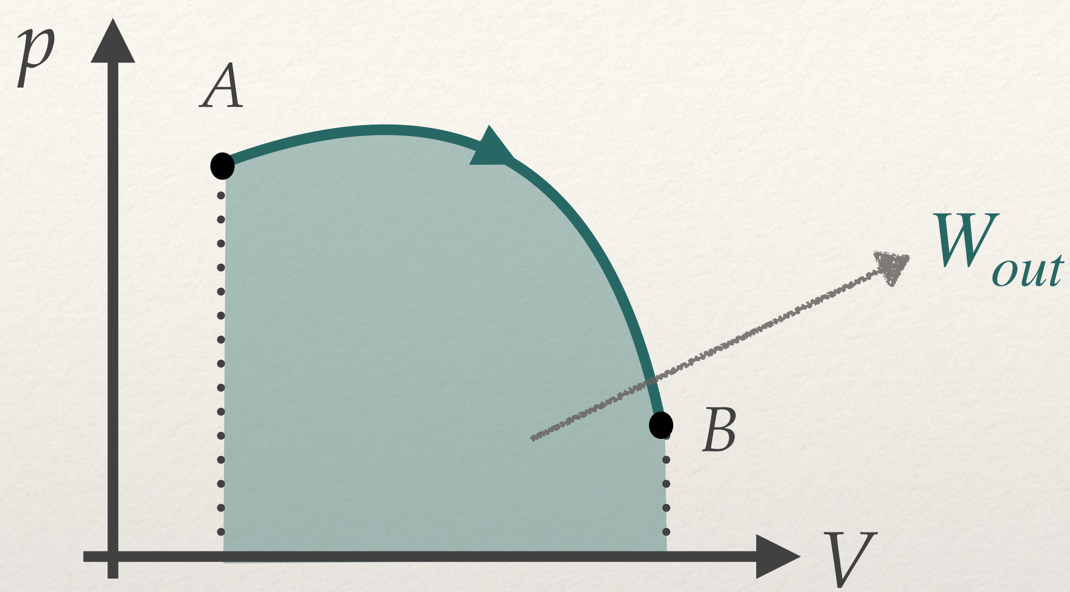
$$dE = \text{đ}Q + \text{đ}W \qquad\qquad dE = Q + W$$

exact

non-exact

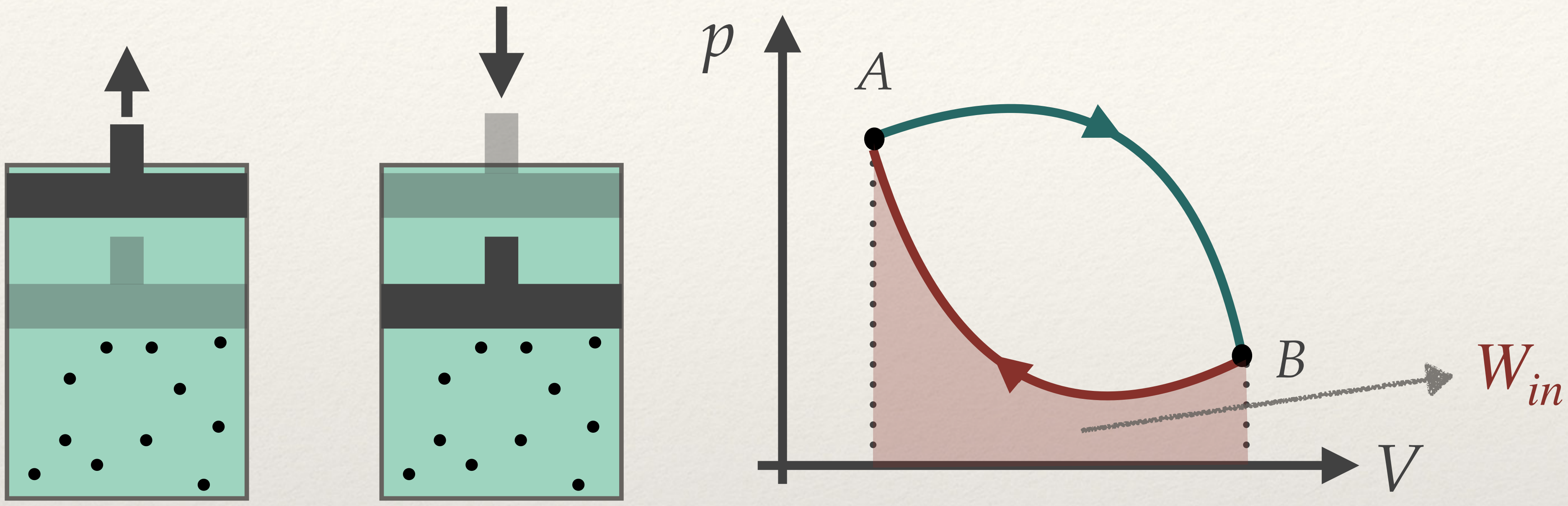
A Second Look at Heat Engines





$A \rightarrow B$ Gas expands, does work on the environment

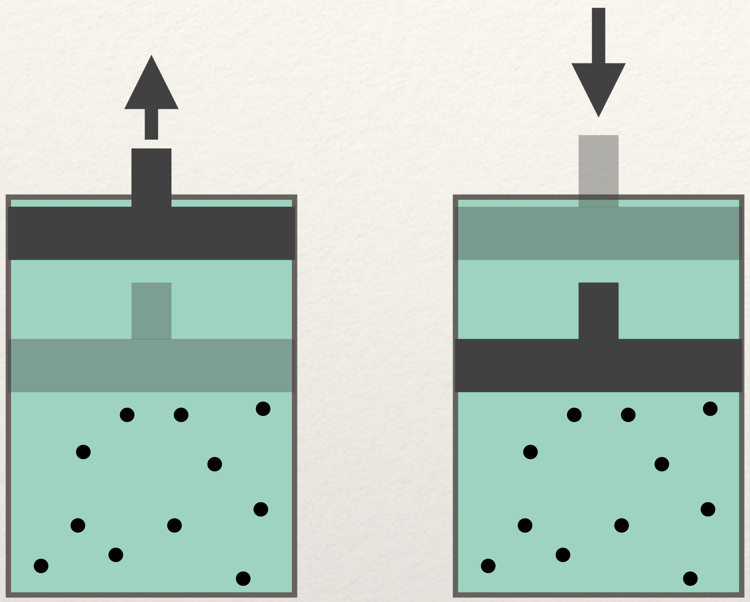
A Second Look at Heat Engines

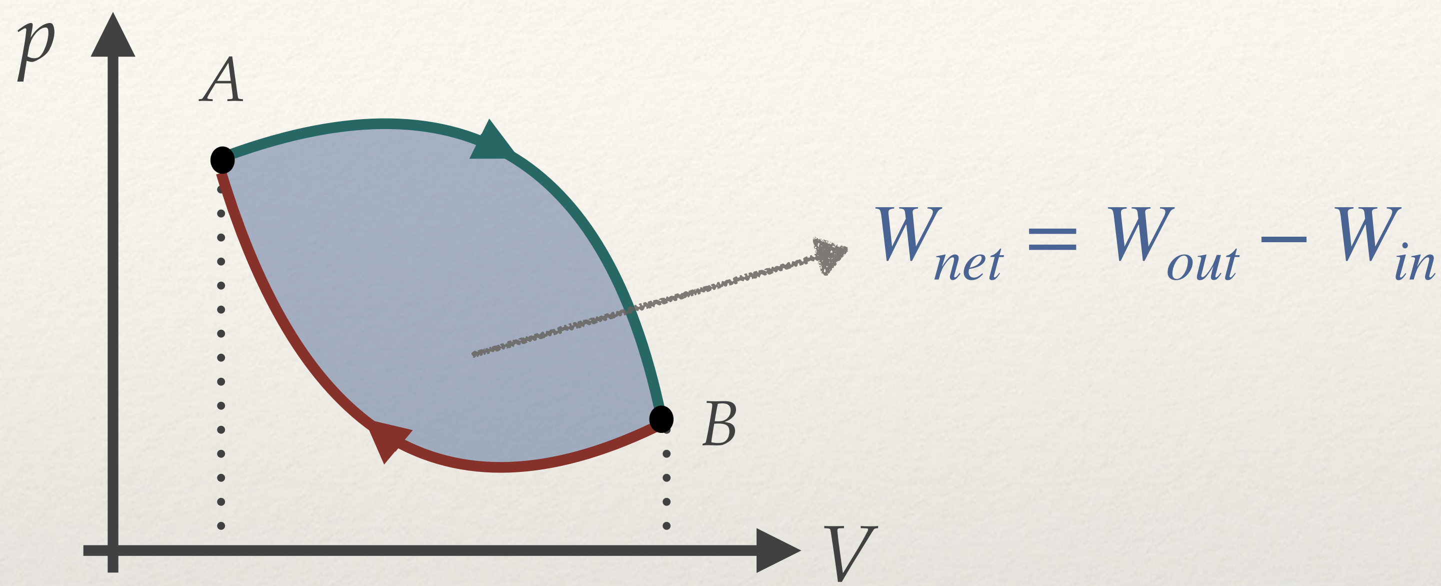


$A \to B$ Gas expands, does work on the environment

$B \to A$ Gas is compressed, environment does work on gas

A Second Look at Heat Engines





$A \rightarrow B$ Gas expands, does work on the environment

$B \rightarrow A$ Gas is compressed, environment does work on gas

Net output: $W_{net} = W_{out} - W_{in}$

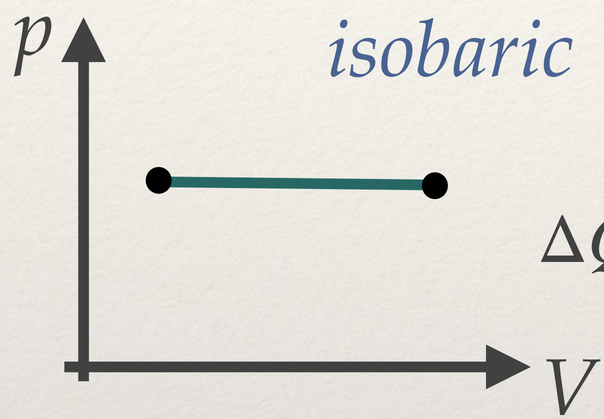
$$\oint dE = 0 \Rightarrow \oint dQ = -\Delta W = W_{net}$$

The net work is provided by heat from a hot reservoir. But not all of it can be converted to work More on this later…

Compression Work

Different ways to heat / compress (moving in the pV plane)

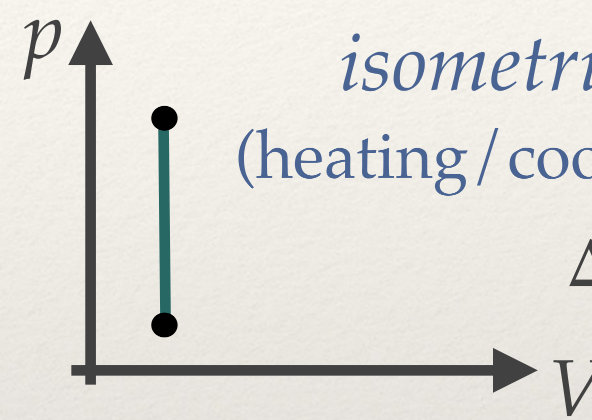
isobaric $\Delta p = 0$



$\Delta Q, \Delta T, \Delta V \neq 0$

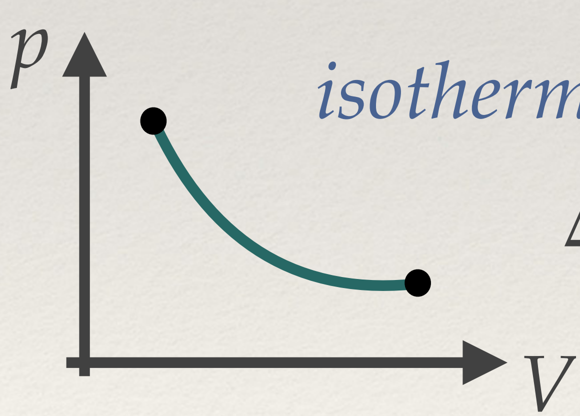
isometric $\Delta V = 0$

(heating / cooling)



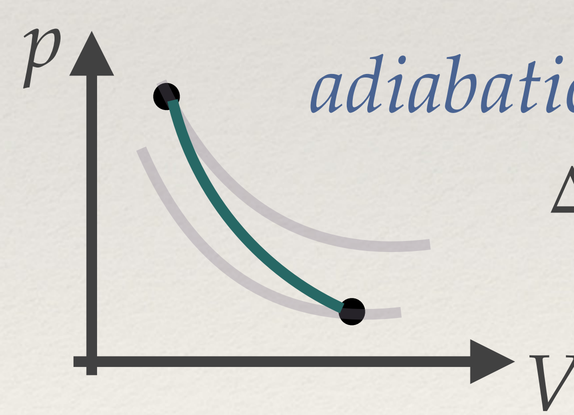
$\Delta Q, \Delta T, \Delta p \neq 0$

isothermal $\Delta T = 0$



$\Delta Q, \Delta V, \Delta p \neq 0$

adiabatic $\Delta Q = 0$

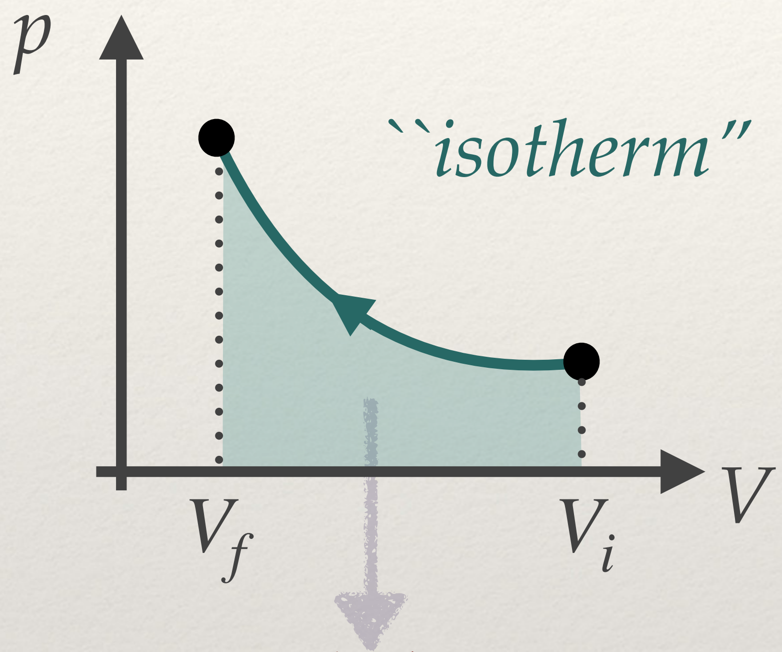


$\Delta p, \Delta T, \Delta V \neq 0$

Isolated system,
no heat exchange with
the environment

Isothermal Compression of Ideal Gas

Equation of state: $pV = NkT$



work done:

$$W = -\int_{V_i}^{V_f} p(V)dV = NkT\log(V_i/V_f)$$

+ for compression

First Law

$$\Delta E = Q + W$$

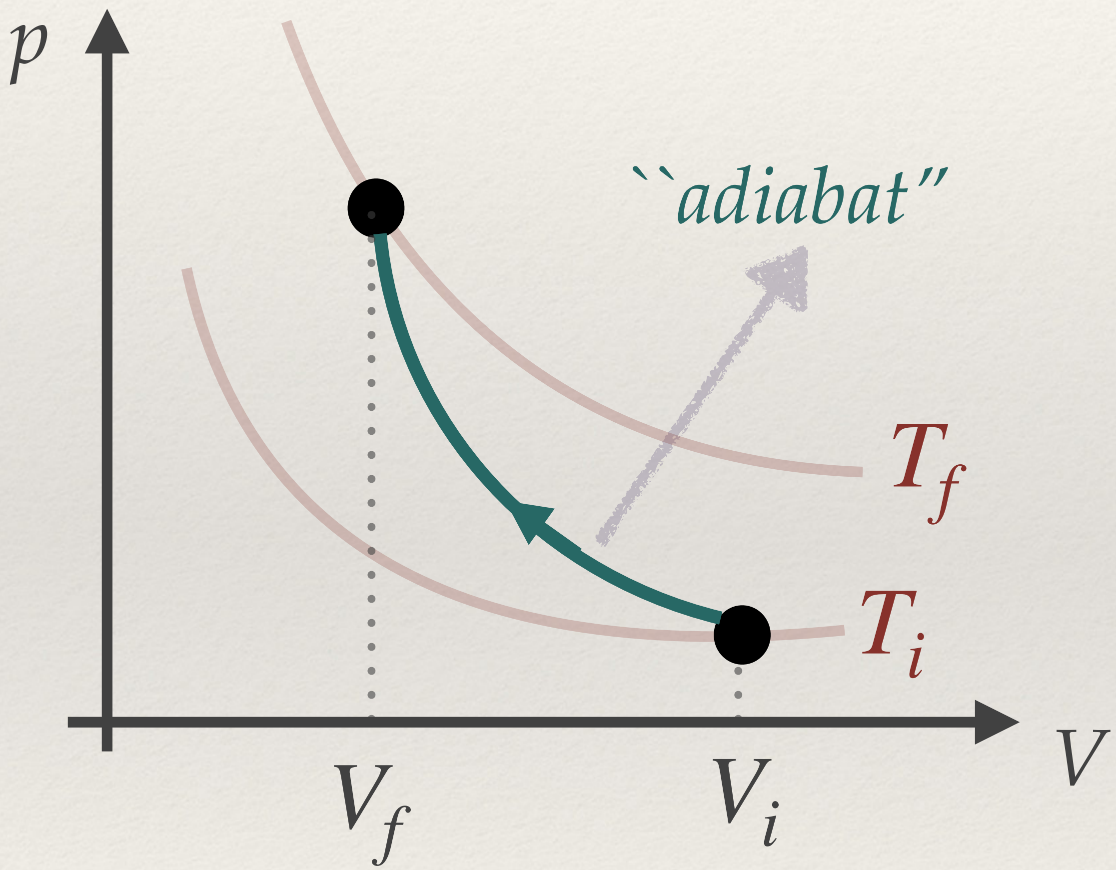
$$\Delta E = 0$$ Why?

$$Q = -W = -NkT\log(V_i/V_f)$$

- for compression,
Heat leaves the system.

Adiabatic Compression of Ideal Gas

Compress fast enough so that no heat escapes the system
(still slow enough to be quasi-static)



First Law

$$dE = \underbrace{\cancel{dQ}}_{=0} + pdV$$

$$\frac{N_{dof}}{2} NkdT \qquad \frac{NkT}{V} dV$$

e.o.s

homework

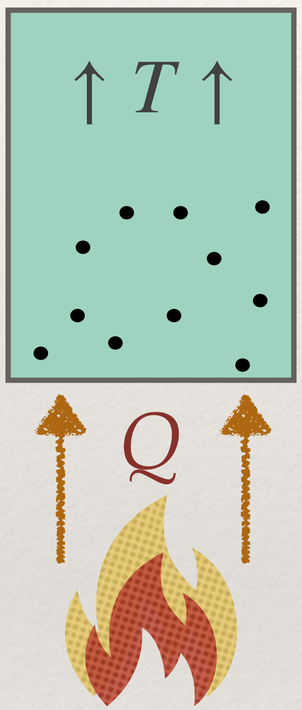
$$p^{\gamma} V = \text{constant}$$

$$\gamma = \frac{N_{dof} + 2}{N_{dof}}$$

``adiabatic index''

Heat Capacities

How much heat is necessary to change the temperature of the system?



Heat capacity: $C \text{ "="} \dfrac{Q}{\Delta T} = \dfrac{\Delta E - W}{\Delta T}$

``Amount of heat need to raise the temperature"

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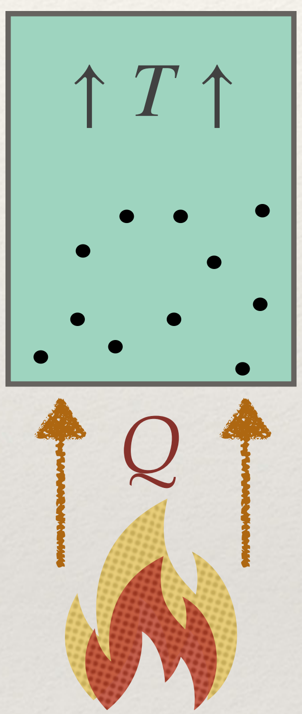

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but wait…
W (hence Q) can be anything, depending on how the system is heated.

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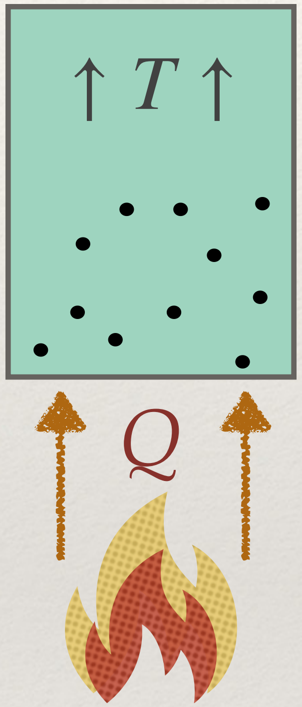

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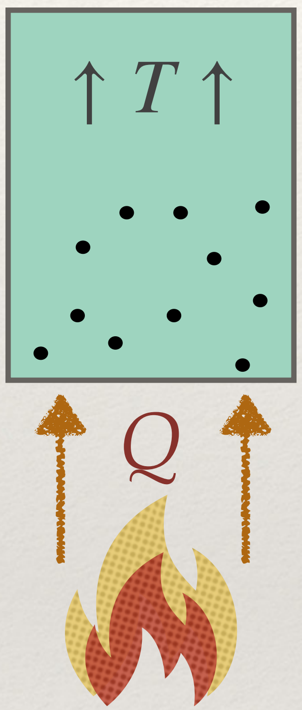


Heat capacity: $C \text{ ''=''} \dfrac{Q}{\Delta T} = \dfrac{\Delta E - W}{\Delta T}$

``Amount of heat need to raise the temperature''

- Keep V fixed: $$C_V = \left(\frac{\Delta E}{\Delta T}\right)_V = \left(\frac{\partial E}{\partial T}\right)_V$$

kept fixed

Typically things expand when heated…

- Keep p fixed: $$C_p = \left(\frac{\Delta E - (-p\Delta V)}{\Delta T}\right)_P = \left(\frac{\partial E}{\partial T}\right)_P + p\left(\frac{\partial V}{\partial T}\right)_P$$

$C_P > C_V$ (system needs energy to expand)

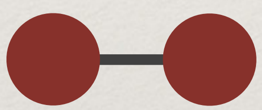
Heat Capacities

$$C_V = \left(\frac{\partial E}{\partial T}\right)_V$$

Ideal gas at ``high enough" T

$$E = \frac{N_{dof}}{2}NkT \longrightarrow C_V = \frac{N_{dof}}{2}Nk = \frac{1}{2}N_{dof}N_{mole}R$$

e.g. N_2



$$N_{dof} = 3 + 2 + 2$$

$$\frac{C_V}{\text{mole}} = \frac{7}{2}R?$$

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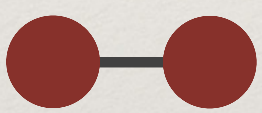

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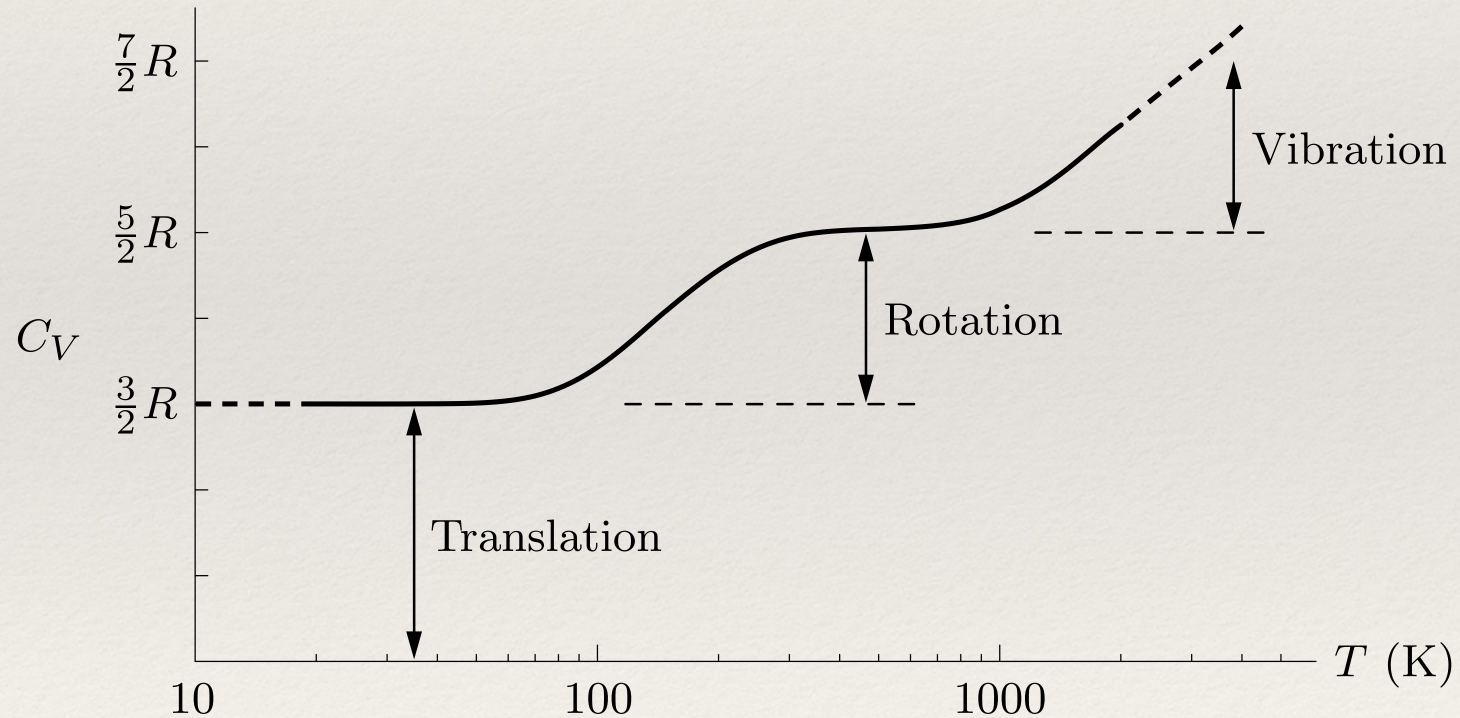

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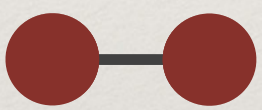

$$N_{dof} = 3 + 2 + 2$$

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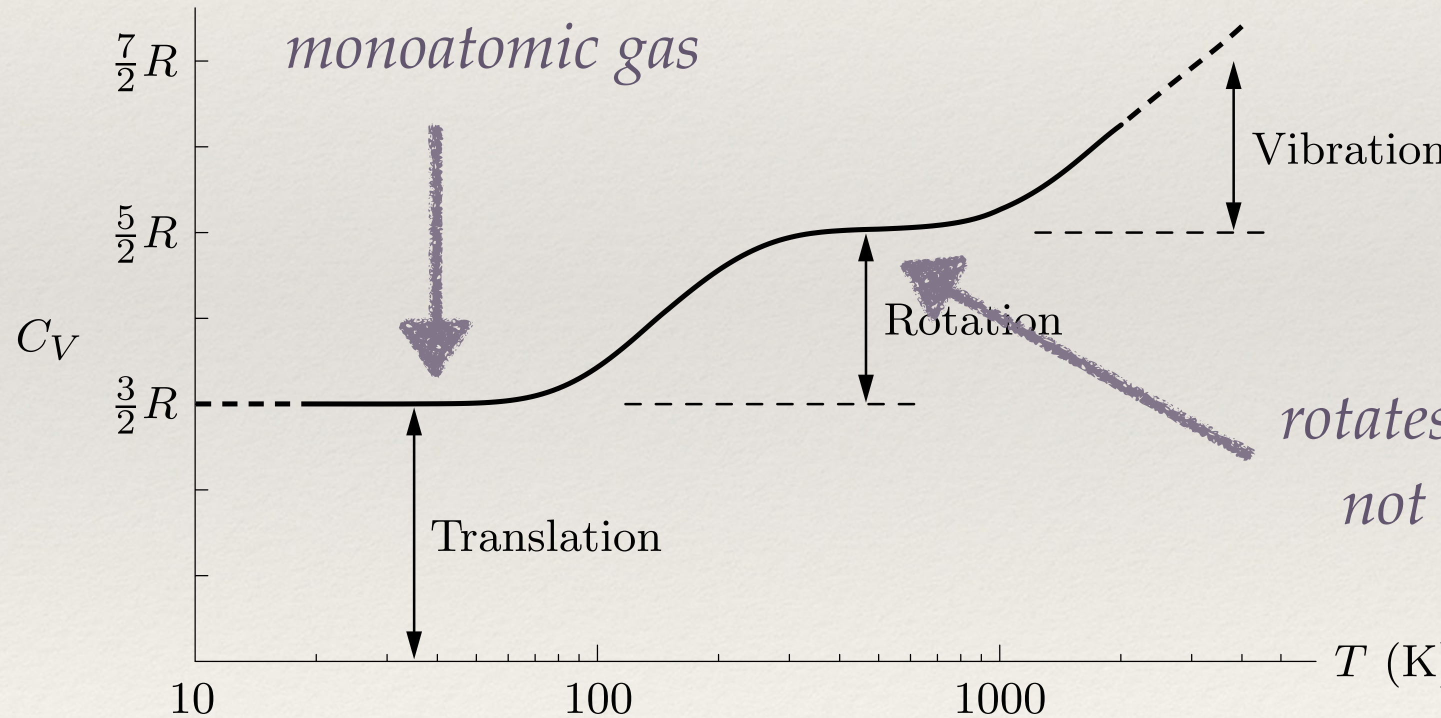
behaves like a monoatomic gas



rotates but does not vibrate

One of the earliest experimental hints of quantum mechanics!

Enthalpy

When dealing with systems at content pressure, it is annoying to keep track of work due to expansion/compression

Enthalpy: $H = E + pV$

energy content of the system

work needed to make room for it

Change in enthalpy

$$dH = dQ + Vdp + dW_{other}$$

At constant pressure:

$$\Delta H = Q + W_{other}$$

When there is no other work:

heat added to the system=change in enthalpy

$$C_p = \left(\frac{\partial H}{\partial T}\right)_p$$

Latent Heat

When ice melts, it absorbs heat but the temperature does not increase before it completely melts

$$C = \frac{Q}{\Delta T} = \infty ???$$ *not very useful…*

Same with water boiling

Latent Heat: Amount of heat required to melt/boil the entire substance

(*Specific Latent Heat:* Latent heat/mass)

e.g. *water:*

melting ice : *333 J/g*

boiling water : *2260 J/g*

heating water from 0 °C to 100 °C : *~418J/g*

Using latent heat is crucial in engine design
More on this later…

A Very Brief History of Thermodynamics


``Caloric theory" *Heat is like a conserved fluid that flows from hot to cold objects.*

Explained many observed phenomena, Carnot engine, etc.

But, alas, was wrong…

Lavoisier 1787,
later Carnot, Clapeyron,… ~1820s

``Mechanical Equivalent of Heat" [Joule 1843]

Benjamin Thompson, 1798
(a.k.a Count Rumford of the Holy Roman Empire)

Heat cannot be conserved, rather has to do with motion

James Prescott Joule, 1843

Mechanical force heat

Julius von Mayer, 1842

also: oxidation, $C_p - C_V = R$

[Joule's heat apparatus, 1845]
Image from wikipedia

Hermann Helmholtz 1847: *Conservation of energy*

William Thompson (a.k.a Lord Kelvin), Rudolf Clausius, … ~1860s: *2nd Law*

No need for Caloric fluid in Carnot's theory

Thank you!