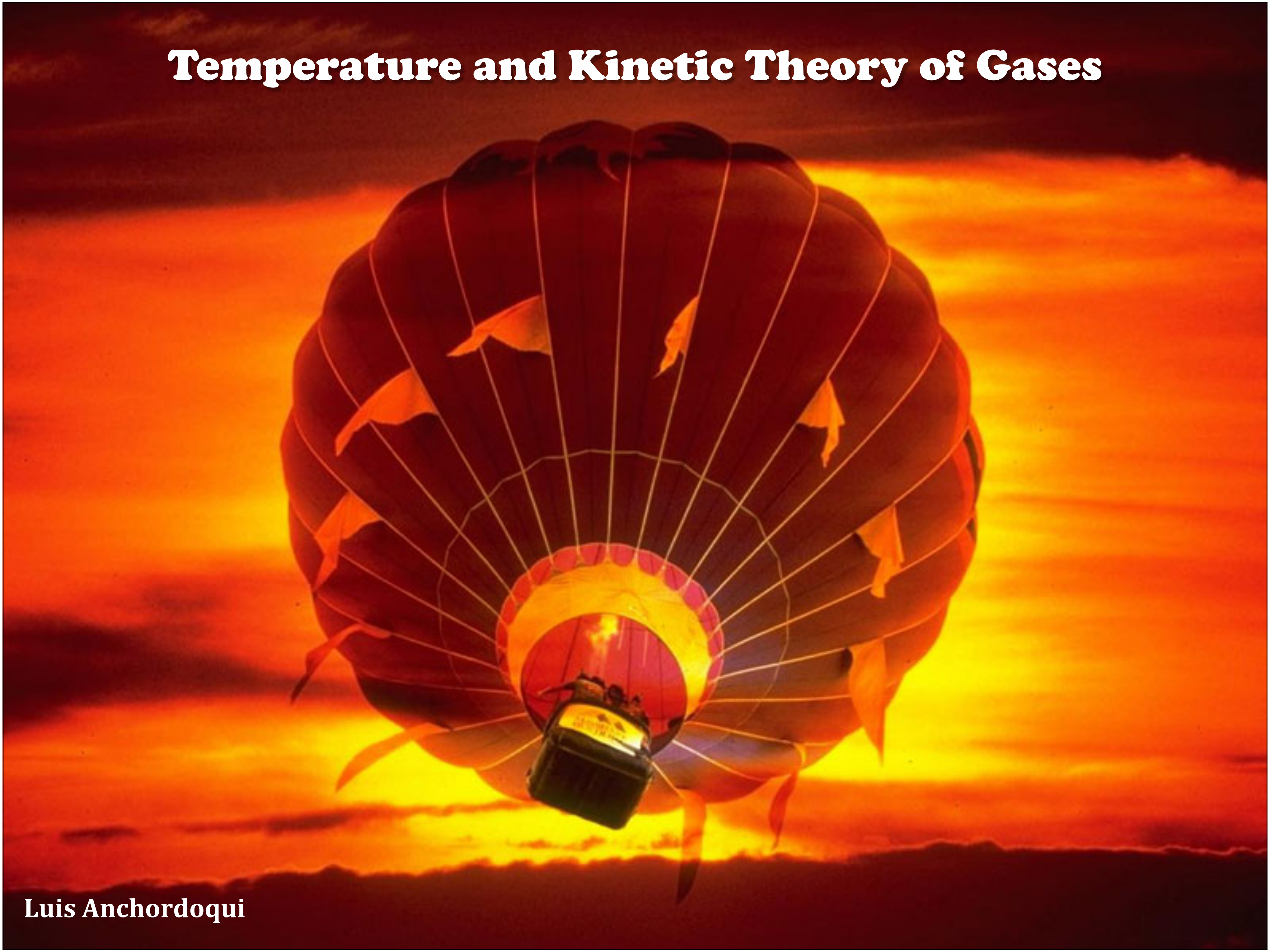
Lesson 12

Physics 168

Luis Anchordoqui

Temperature and Kinetic Theory of Gases



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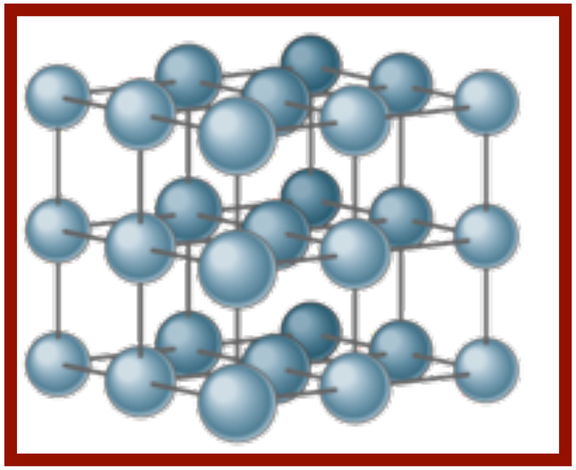
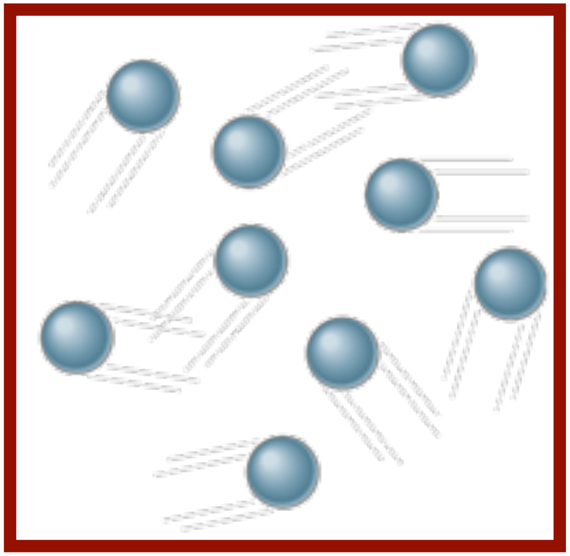
Atomic Theory of Matter

On microscopic scale, arrangements of molecules in solids, liquids, and gases are quite different

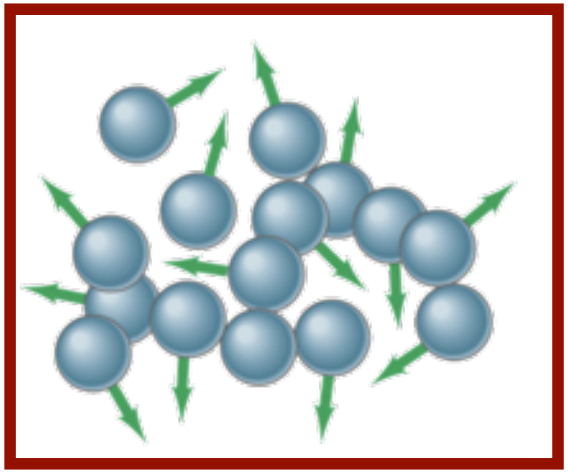
solids



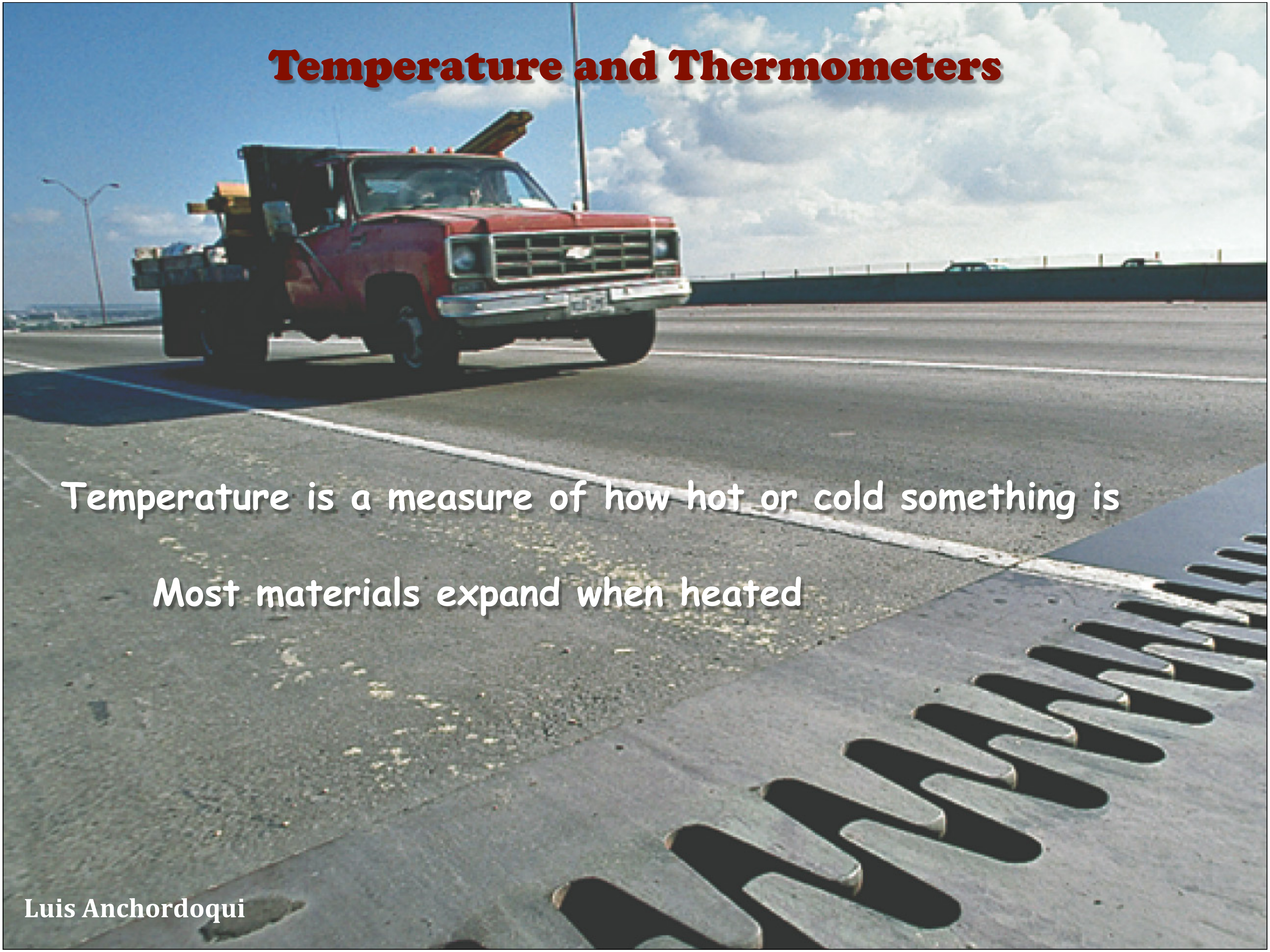
gases



liquids



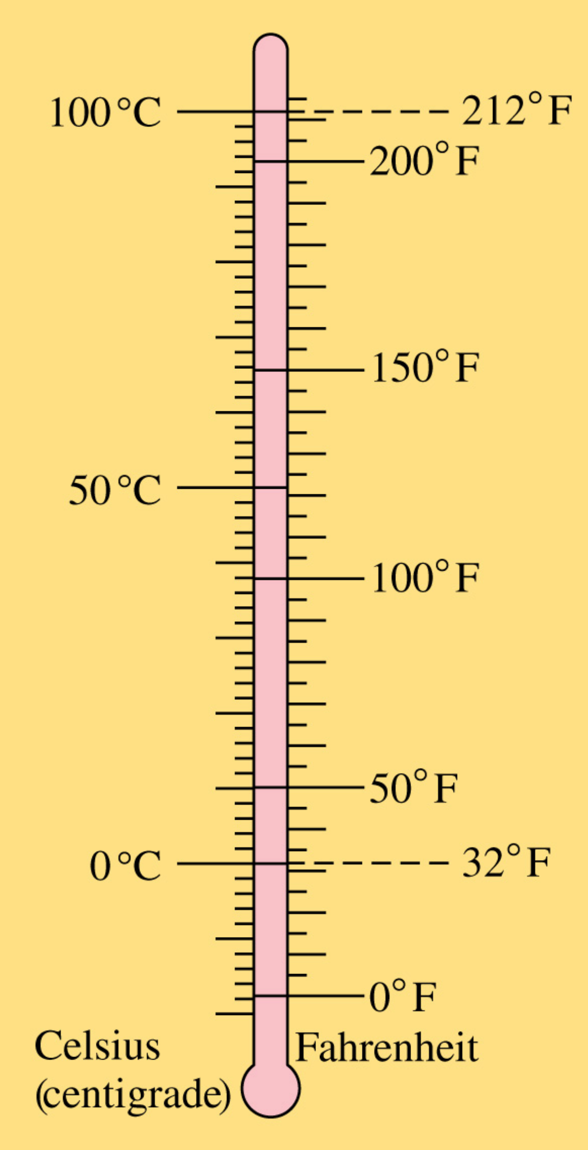
Temperature and Thermometers



Temperature is a measure of how hot or cold something is

Most materials expand when heated

Temperature and Thermometers (cont'd)



Temperature is generally measured using either Fahrenheit or Celsius scale

Freezing point of water is 0°C = 32°F

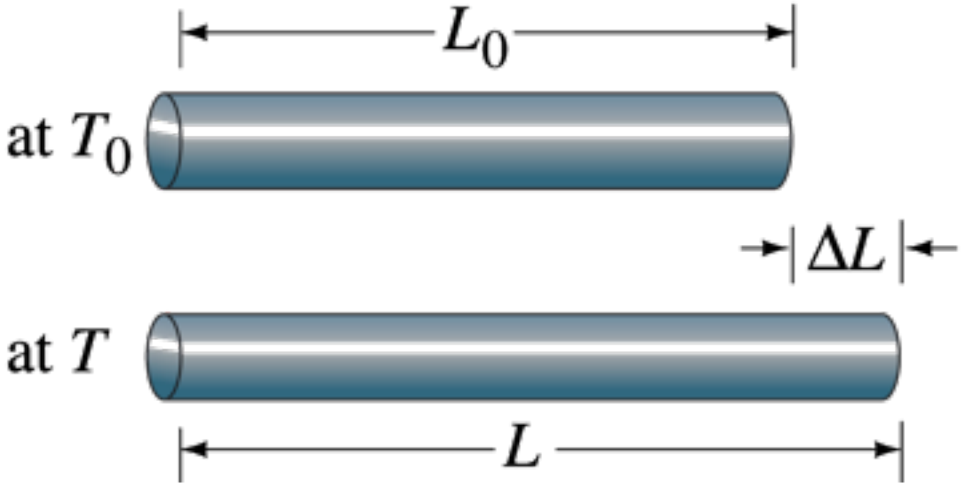
boiling point is 100°C = 212°F

$$T_C = \frac{5}{9}(T_F - 32^\circ)$$

$$T_F = \frac{9}{5}T_C + 32^\circ$$

Thermal Expansion

Linear expansion occurs when an object is heated



$$L = L_0(1 + \alpha\,\Delta\,T)$$

coefficient of linear expansion

Volume expansion is similar except that it is relevant for liquids and gases as well as solids:

$$\underbrace{\Delta V = \beta}_{} V_0 \Delta T$$

$$\Delta V = \underset{\substack{\downarrow \\ \text{coefficient of volume expansion}}}{\beta} V_0 \Delta T$$

For uniform solids ☛ $\beta \approx 3\alpha$

Equation of state and Boyle's law

Relationship between volume, pressure, temperature, and mass of a gas is called an equation of state

We will deal here with gases that are not too dense

Boyle's Law: volume of a given amount of gas is inversely proportional to pressure as long as temperature is constant

$$V \propto \frac{1}{P}$$

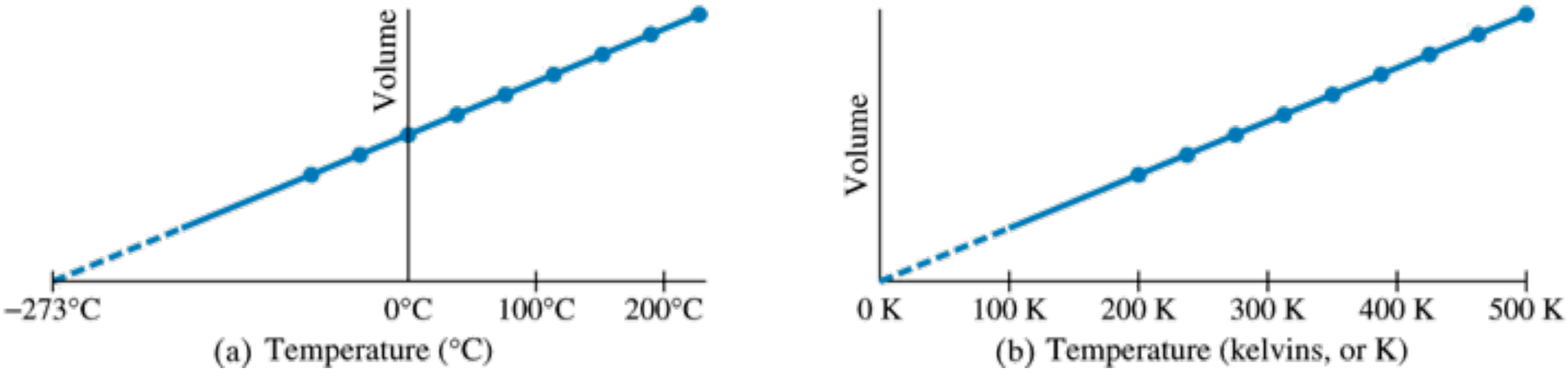
Gas Laws and Absolute Temperature

Volume is linearly proportional to temperature, as long as temperature is somewhat above condensation point and pressure is constant:

$$V \propto T$$

Extrapolating volume becomes zero at -273.15°C

this temperature is called absolute zero



Gas Laws and Absolute Temperature (cont'd)

Concept of absolute zero allows us to define a third temperature scale
absolute (or Kelvin) scale

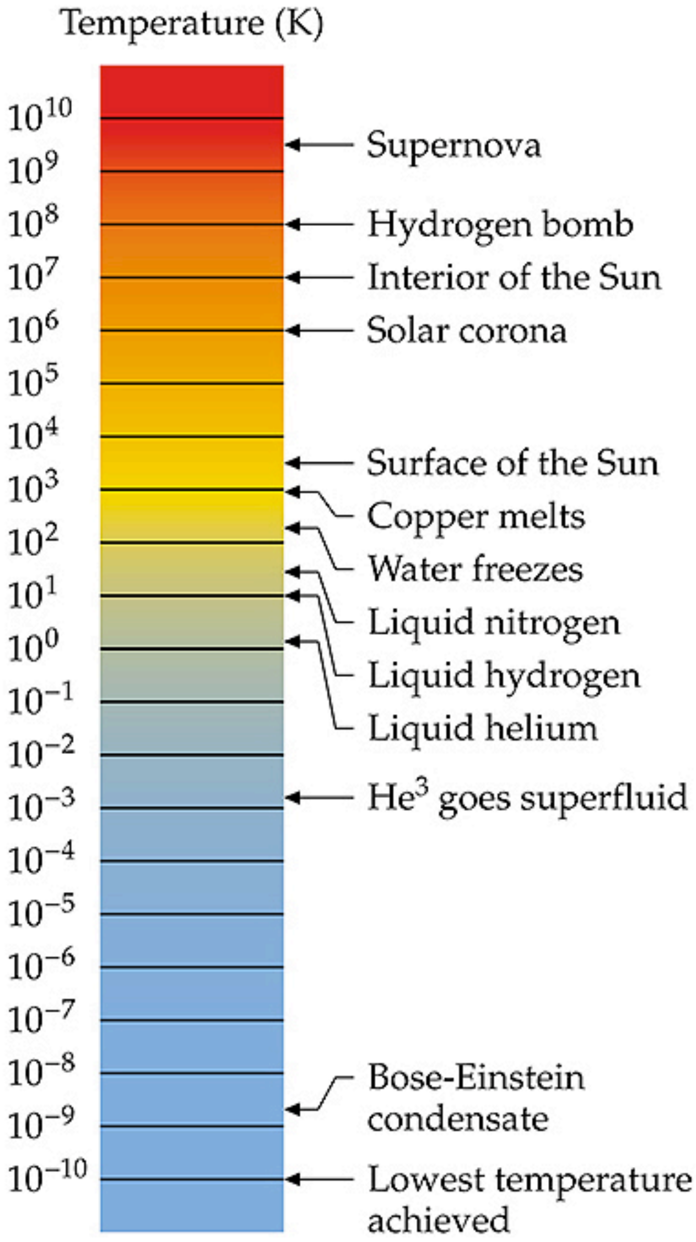
This scale starts with 0 K at absolute zero
but otherwise is same as Celsius scale

Freezing point of water is 273.15 K and boiling point is 373.15 K

When volume is constant ☛ pressure is directly proportional to temperature

$$P \propto T$$

Temperature of various places and phenomena



Ideal Gas Law

We can combine three relations just derived into a single relation:

$$PV \propto T$$

What about amount of gas present?

If temperature and pressure are constant

volume is proportional to amount of gas:

$$PV \propto mT$$

Relative Atomic Mass Scale

What is the Relative Atomic Mass Scale (^{12}Carbon Scale)?

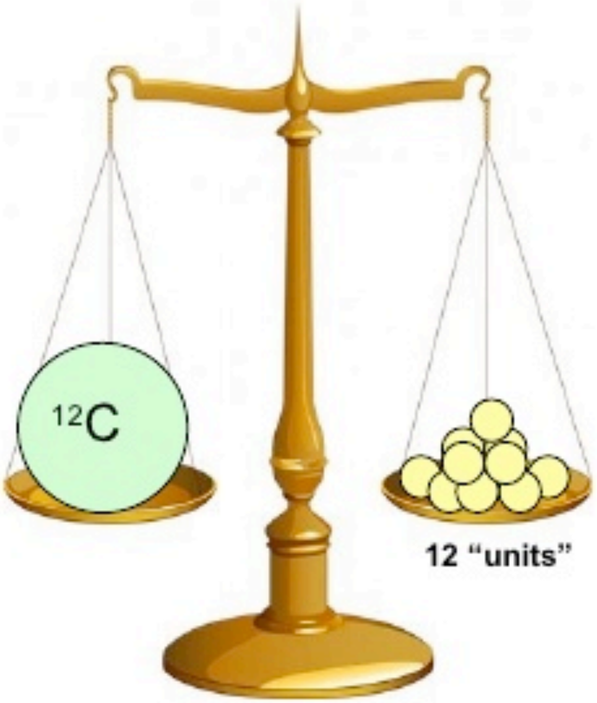
- masses of atoms are expressed relative to the mass of a Carbon-12 atom.

Why is there a need for the scale?

- atoms are too small to be weighed directly.
- inconvenient to express masses of atoms in terms of kg.

So what does it imply?

- 1 atom of ^{12}C = 12
- 1/12 the mass of a ^{12}C atom = 1



microworld $\longrightarrow$ **macroworld**

atoms & molecules → grams

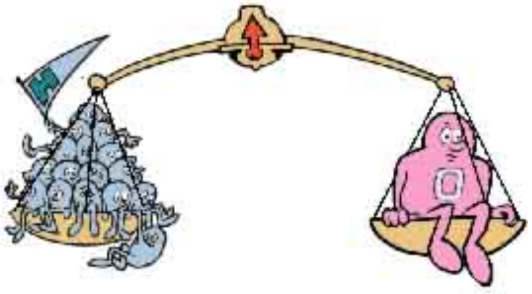
Atomic mass ⇒ mass of an atom in atomic mass units (u)

By definition:
1 atom ^{12}C "weighs" 12 u

On this scale

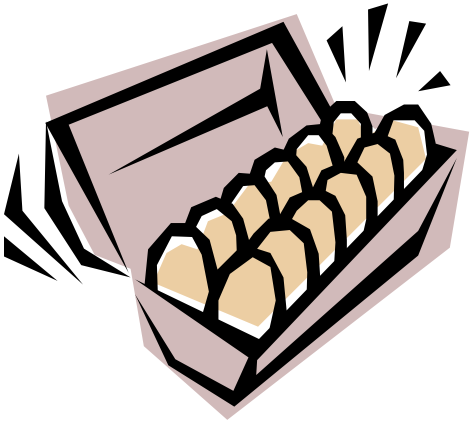
^{1}H = 1.008 u

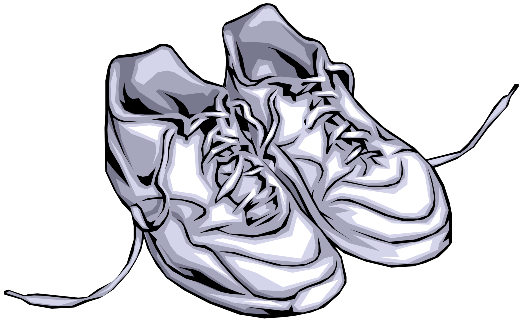
^{16}O = 16.00 u



Mole (mol): A unit to count numbers of particles

Dozen = 12





Pair = 2

Mole ☛ amount of a substance that contains as many elementary entities as there are atoms in exactly 12.00 grams of ^{12}C

$$1 \text{ mol} = N_A = 6.022 \times 10^{23}$$

N_A ☛ Avogadro's number

Molar mass is the mass of 1 mole of eggs / shoes / marbles / atoms in grams

1 mole ^{12}C atoms = 6.022 x 10^{23} atoms = 12.00 g

1 ^{12}C atom = 12.00 u

1 mole ^{12}C atoms = 12.00 g ^{12}C

1 mole lithium atoms = 6.941 g of Li

For any element

atomic mass (u) = molar mass (grams)

1 u = 1.66 x 10^{-24} g or 1 g = 6.022 x 10^{23} u

Ideal Gas Law (cont'd)

A mole (mol) is defined as number of grams of a substance that is numerically equal to molecular mass of substance

$1\,\mathrm{mol}\,\mathrm{H}_2$ has a mass of $2\ \mathrm{g}$

$1\ \mathrm{mol}\ \mathrm{N}_2$ has a mass of $28\ \mathrm{g}$

$1\,\mathrm{mol}\,\mathrm{CO}_2$ has a mass of $44\ \mathrm{g}$

Number of moles in a certain mass of material:

$$\mathrm{n}\,(\mathrm{mol}) = \frac{\mathrm{mass}\,(\mathrm{grams})}{\mathrm{Molecular\,mass}\,(\mathrm{g/mol})}$$

Ideal Gas Law (cont'd)

We can now write ideal gas law ☛

$$PV = \underset{\swarrow}{n}\,\underset{\searrow}{R}T$$

number of moles universal gas constant

$$R = 8.315\,J/(\mathrm{mol}\cdot K)$$

$$= 0.0821\,(L\cdot \mathrm{atm})/(\mathrm{mol}\cdot K)$$

$$= 1.99\,\mathrm{calories}/(\mathrm{mol}\cdot K)$$

Ideal Gas Law in Terms of Molecules: Avogadro's Number

Since gas constant is universal

number of molecules in one mole is same for all gases
That number is called Avogadro's number:

$$N_A = 6.02 \times 10^{23}$$

Number of molecules in a gas is number of moles times Avogadro's number:

$$N = n\, N_A$$

Ideal Gas Law in Terms of Molecules: Avogadro's Number

Therefore we can write:

$$PV = N\underset{\substack{\big\downarrow \\ \text{Boltzmann's constant}}}{k}T$$

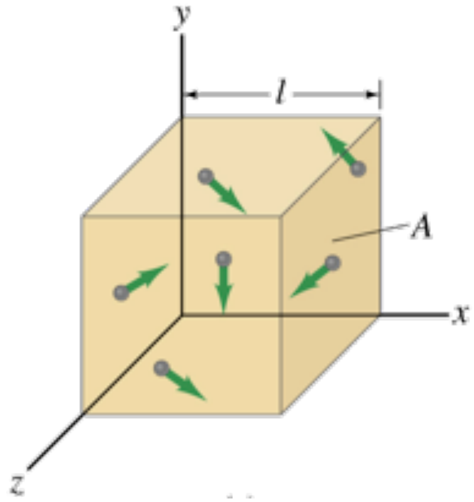
$$k = \frac{R}{N_A} = \frac{8.315\,J/\mathrm{mol}\cdot K}{6.02\times 10^{23}\,/\,\mathrm{mol}} = 1.38\times 10^{-23}\,J/K$$

Kinetic Theory and Molecular Interpretation of Temperature

Assumptions of kinetic theory:

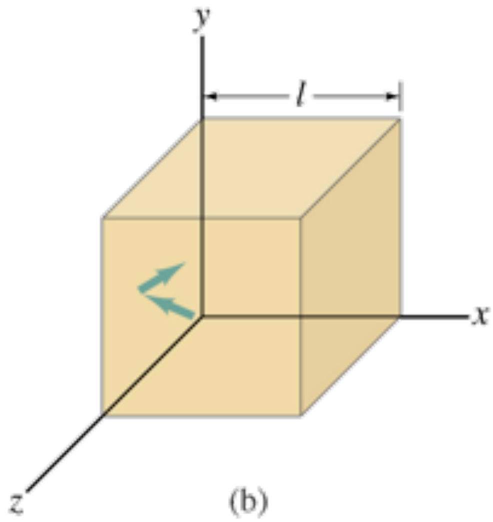
- Large number of molecules moving in random directions with a variety of speeds
- Molecules are far apart on average
- Molecules obey laws of classical mechanics and interact only when colliding
- Collisions are perfectly elastic

Kinetic Theory and Molecular Interpretation of Temperature



Force exerted on the wall by collision of one molecule is

$$F = \frac{\Delta(mv)}{\Delta t} = \frac{2\, m v_x}{2l/v_x} = \frac{m v_x^2}{l}$$



Then force due to all N molecules colliding with that wall is

$$F = \frac{m}{l} N \, \bar{v}_x^2$$

Kinetic Theory and Molecular Interpretation of Temperature

Averages of squares of speeds in all three directions are equal:

$$F = \frac{m}{l} N \frac{\bar{v}^2}{3}$$

So pressure is ☛ $$P = \frac{F}{A} = \frac{1}{3} \frac{N m\bar{v}^2}{Al} = \frac{1}{3} \frac{N m\bar{v}^2}{V}$$

Rewriting ☛ $$PV = \frac{2}{3} N (\frac{1}{2} m\bar{v}^2)$$

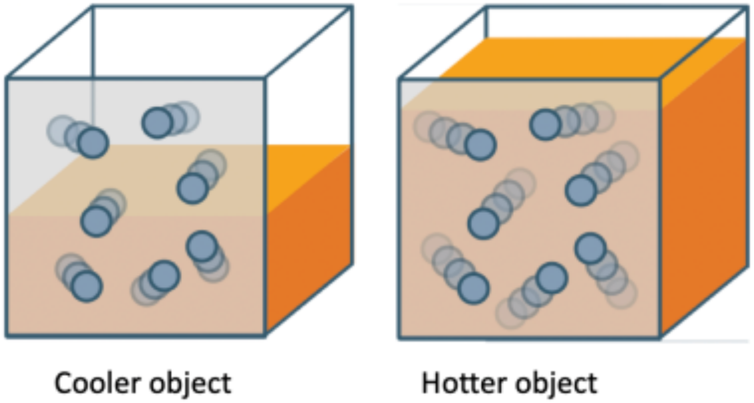
so ☛ $$\langle KE \rangle = \overline{KE} = \frac{1}{2} m\bar{v}^2 = \frac{3}{2} kT$$

Average translational kinetic energy of molecules in an ideal gas is directly proportional to temperature of gas

Kinetic Theory and Molecular Interpretation of Temperature

We can invert this to find average speed of molecules in a gas as a function of temperature

$$v_{\mathrm{rms}} = \sqrt{\bar{v}^2} = \sqrt{\frac{3kT}{m}}$$



Mean Free Path

Average speed of molecules in a gas at normal pressure
is several hundred meters per second
yet if somebody across room from you opens a perfume bottle
you do not detect odor for several minutes

Reason for time delay ☛ perfume molecules do not travel directly towards you
but instead travel a zigzag path due to collisions with air molecules

Average distance λ travelled by a molecule between collisions
is called its mean free path

Mean Free Path (cont'd)

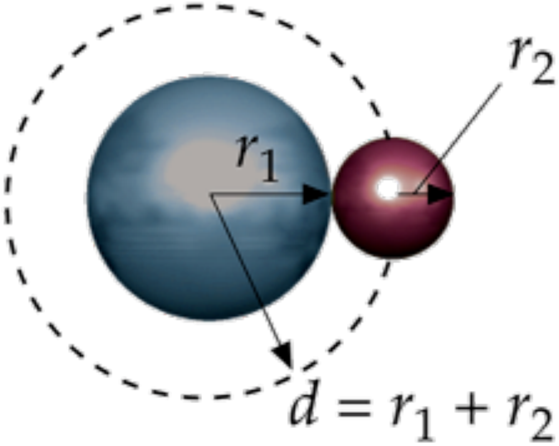
Consider one gas molecule of radius r moving with speed v

through a region of stationary molecules

Moving molecule will collide with another molecule of radius r

if centers of two molecules come within a distance $d = r_1 + r_2$

from each other



If all molecules are of same type

⇩

$$d = \text{molecular diameter}$$

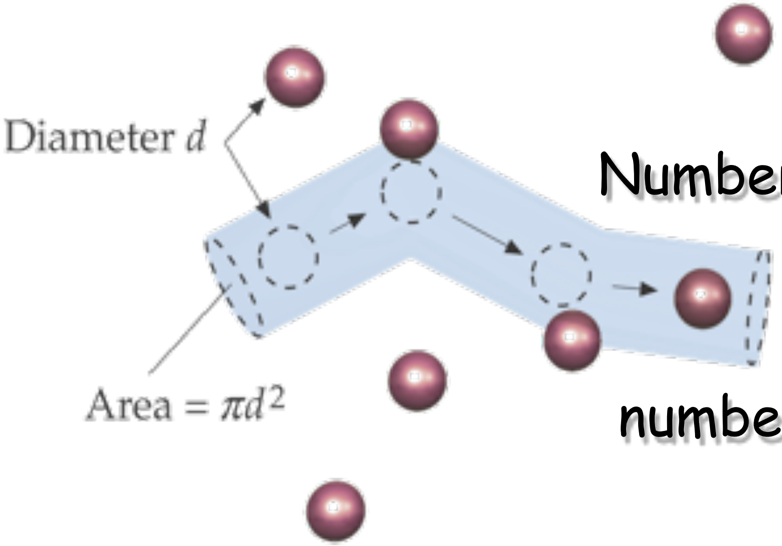
Mean Free Path (cont'd)

As molecule moves it will collide with any molecule

whose center is in a circle of radius d

After a time t molecule moves a distance vt and collides with every molecule

in cylindrical volume $\pi\, d^2\, vt$



Number of molecules in this volume is ☚ $\underbrace{n_V}_{\text{Number density}}\, \pi\, d^2\, vt$

number of molecules per unit volume ☚ $n_V = N/V$

After each collision direction of molecule changes so path actually zigs and zags

Total path length divided by number of collisions is mean free path

$$\lambda = \frac{vt}{n_V\, \pi d^2\, vt} = \frac{1}{n_V\, \pi\, d^2}$$

Mean Free Path (cont'd)

Our previous calculation of mean free path assumes that all but one of gas molecules are stationary which is not a realistic situation

When motion of all molecules is taken into account mean free path becomes

$$\lambda = \frac{1}{\sqrt{2}\, n_V\, \pi\, d^2}$$

Average time between collisions is called collision time τ

Reciprocal of collision time is equal to average number of collisions per second

Then if v_{av} is average speed ☛ average distance travelled between collisions is

$$\lambda = \frac{v_{av}}{\tau}$$

What is heat?

- Subjective sensation (familiar to everyone):
  when you touch something hot
  heat goes into your hand and burns it
- We talk of tansfer of heat
  e.g. ☞ from one object (the radiator) to another (your hand)
- Fact that radiator is hot
  expressed as “radiator is @ high *temperature*”
- Heat and work are mutually convertible:
  you can do work and get heat or use heat and do work
  (a car engine or steam engine)
- The catch is ☞ when you go from heat to useful work
  you never achieve 100% conversion
  there is always some heat wasted
- Thermal energy ☞ random motion of molecules in matter sample
  which constitutes form of kinetic energy
- Temperature ☞ measure of average kinetic energy/molecule

So...what is heat?

- Heat is *also* kinetic energy ☞ moving energy
- It is kinetic energyof zillions of atoms in material
  vibrating (in solid) or zapping around (in gas) in a random way
  going nowehere in particular
- *Heat is random energy of atoms and molecules*
- This is very different from organized energy
- It is like throwing bag of hot popcorn (organized energy)
  vs.
  letting them pop around inside popper (random energy)
- Popcorn popping very quickly has a lot of heat energy

So...what is heat?

- Total heat energy of material depends on two things:
  ① how many atomos there are
  ② how energetic each atom is (on average)

- Heat energy Q can be expressed as product of number of atoms N and average randon energy of single atom $\langle\varepsilon\rangle$

$$Q = N\langle\varepsilon\rangle$$

- *Temperature* ☞ direct measurement of second factor: average energy of an atom

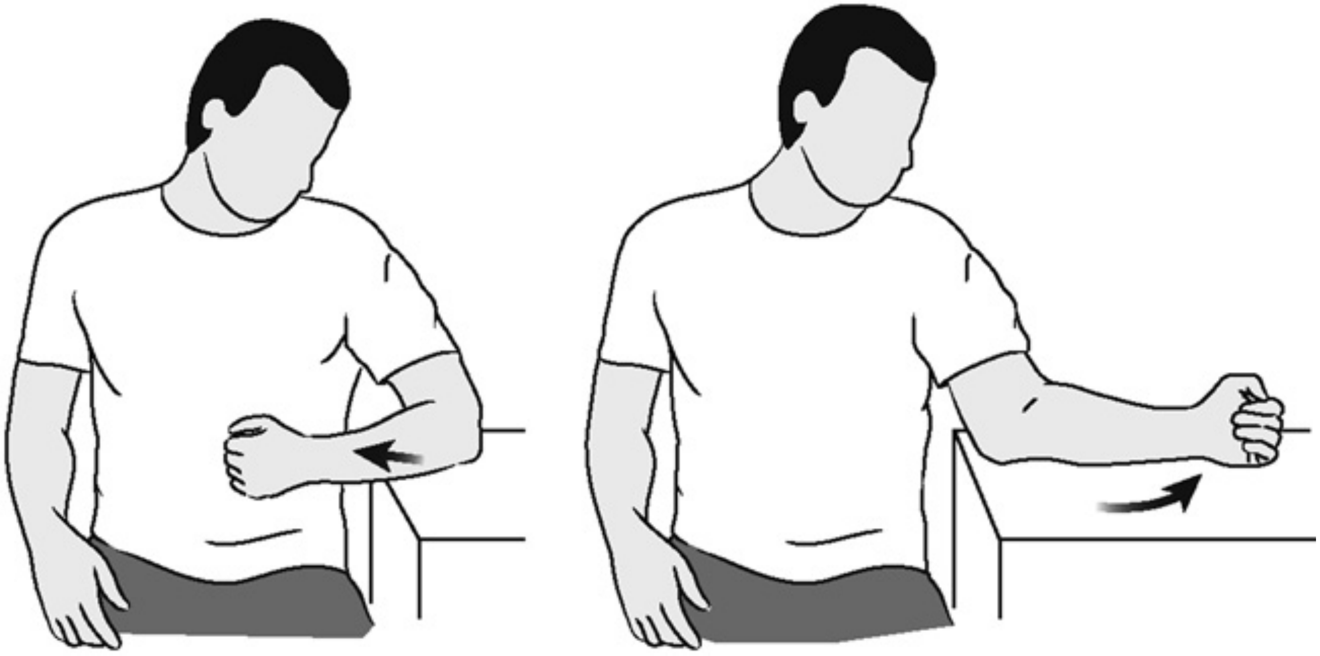
- High temperature means peppy atoms

- Low temperature means sluggish atoms

- When atoms are so sluggish that they don’t move you’ve hit bottom in temperature scale ☞ *absoulte zero* which is $T = -273^\circ$C ☞ very cold!

So...what is heat?

- When you push your hand across table
  (mechanical energy)
- you set molecules at surface of table
  into random vibration and create heat energy
  (friction in this case)



So...what is heat?

- You touch hot poker ☞ madly vibrrating atoms on surface of poker bombard molecules of your skin
- these start vibrating ☞ causing all kind of neurochemical reactions which tell your brain: “It’s hot, let go!”

Laws of Thermodynamics

Via work or via heat, every energy transfer happen can!

THERMODYNAMICS ☛ (developed in 19th century)

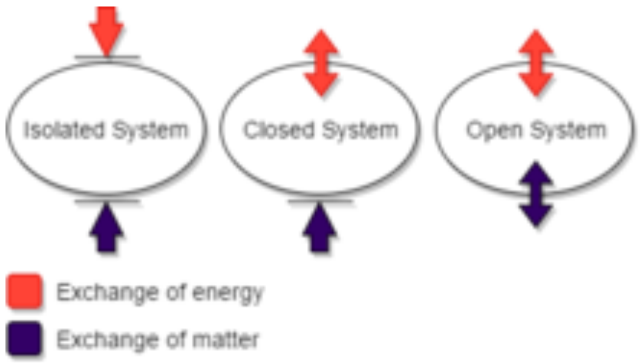
Phenomenological theory

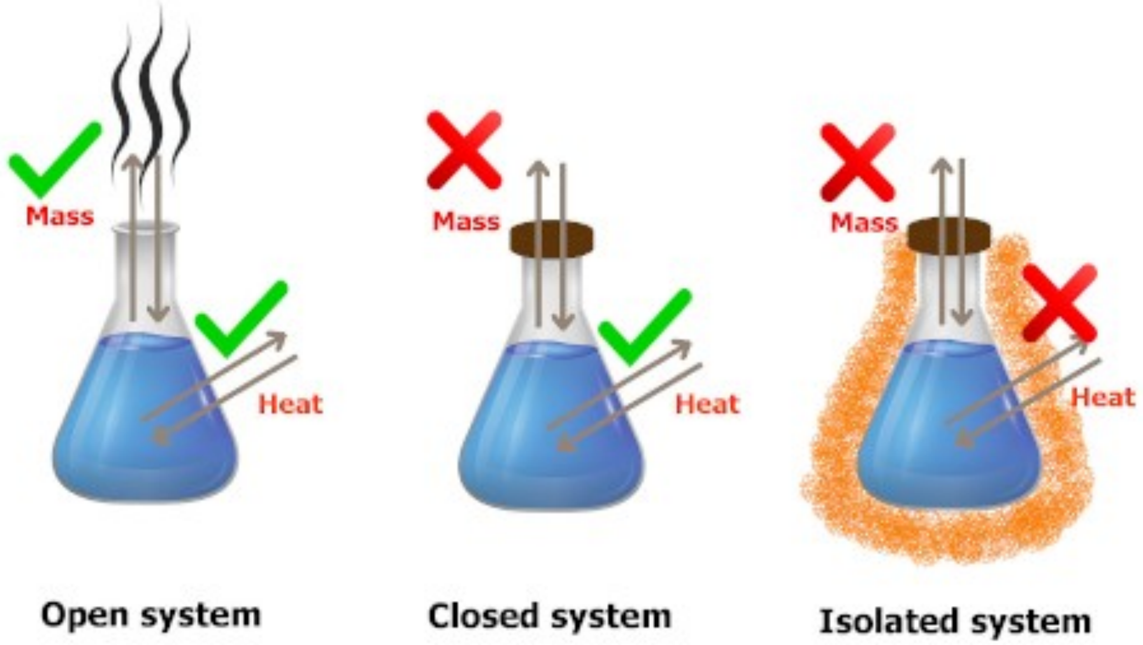
to describe equilibrium properties of macroscopic systems

based on few macroscopically measurable quantities

A thermodynamic system

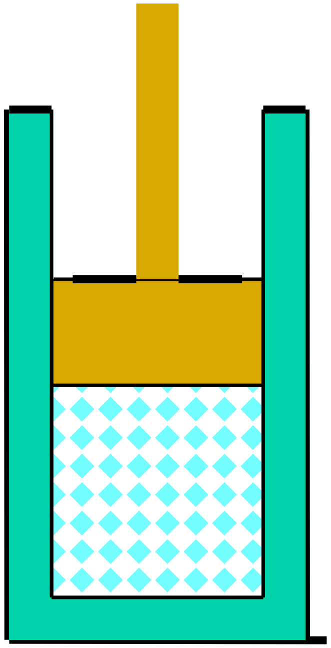
A thermodynamic system is a body of matter and/or radiation confined in space by walls (with defined permeabilities) which separate it from its sourroundings





Internal energy of system

Internal energy U of a system is total of all kinds of energy possessed by particles that make up system



Usually internal energy consists of sum of potential and kinetic energies of working gas molecules

Thermal Equilibrium: Zeroth Law of Thermodynamics

Two objects placed in thermal contact
will eventually come to same temperature
When they do we say they are in thermal equilibrium

In the process of reaching thermal equilibrium
heat Q is transferred from one body to the other

Zeroth law of thermodynamics says that
if 2 objects are each in equilibrium with a 3 object
they are also in thermal equilibrium with each other

Thermodynamics

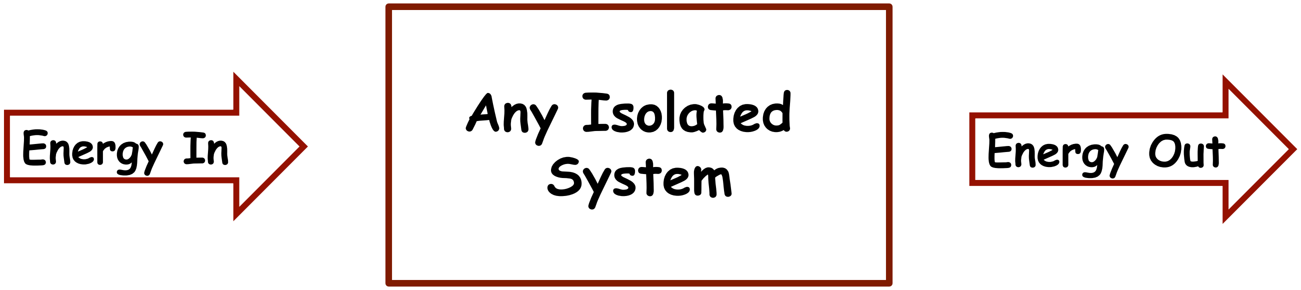
First Law

Energy of universe is constant

Energy can be neither created nor destroyed

so while energy can be converted to another form

total energy remains constant



Energy In

Any Isolated System

Energy Out

This is merely a statement of conservation of energy

First law of Thermodynamics

First law of thermodynamics is application of conservation of energy principle to heat and thermodynamic processes

Change in internal energy of a system
is equal to heat added to system minus work done by system

$$\Delta U = Q - W$$

Change in internal energy	heat added to system	work done by system

Energy can't be created or destroyed ☛ it can only change form

First law of thermodynamics

Net heat put into a system is equal to change in internal energy of system plus work done BY system

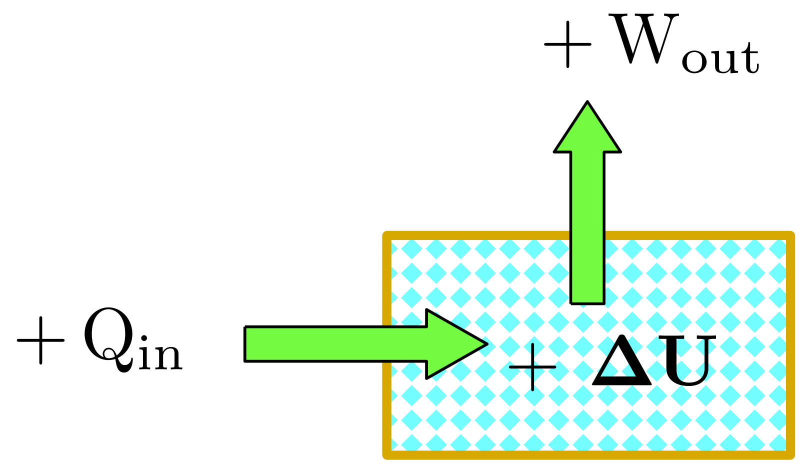
$$\Delta \mathrm{Q} = \Delta \mathrm{U} + \Delta \mathrm{W} \qquad \Delta = (\mathrm{final} - \mathrm{initial})$$

Conversely, work done ON a system is equal to change in internal energy plus heat lost in process

Sign conventions for first law

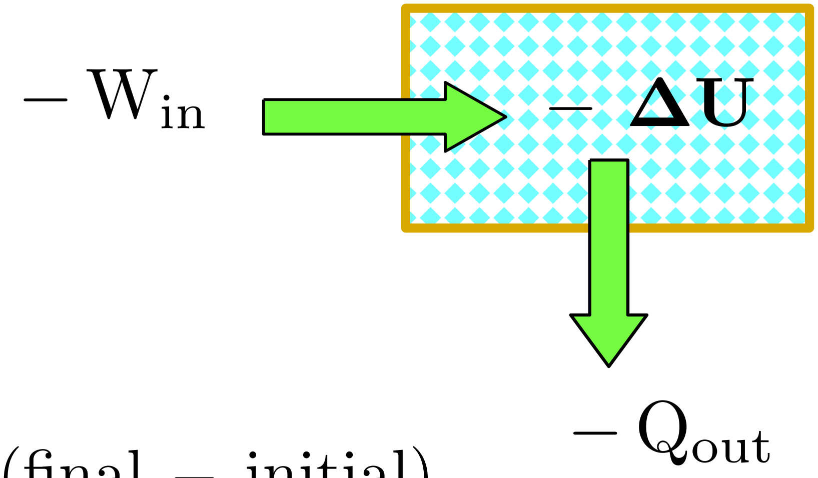
Heat Q INPUT is positive

Work BY a gas is positive



Heat OUT is negative

Work ON a gas is negative



$\Delta \mathrm{Q} = \Delta \mathrm{U} + \Delta \mathrm{W}$ $\qquad$ $\Delta = (\mathrm{final} - \mathrm{initial})$

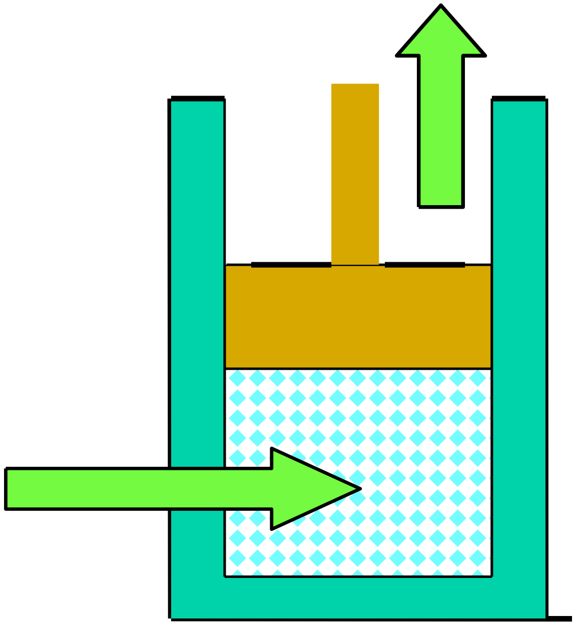
Application of first law of thermodynamics

Example 1

In figure, gas absorbs 400 J of heat and at same time does 120 J of work on piston

What is change in internal energy of system?

$$\mathrm{W_{out}} = 120\,\mathrm{J}$$



$$\mathrm{Q_{in}} = 400\,\mathrm{J}$$

Apply First Law:

$$\Delta \mathrm{Q} = \Delta \mathrm{U} + \Delta \mathrm{W}$$

Application of first law of thermodynamics (cont'd)

ΔQ **is positive** ☛ $+400\,J$ **Heat In**

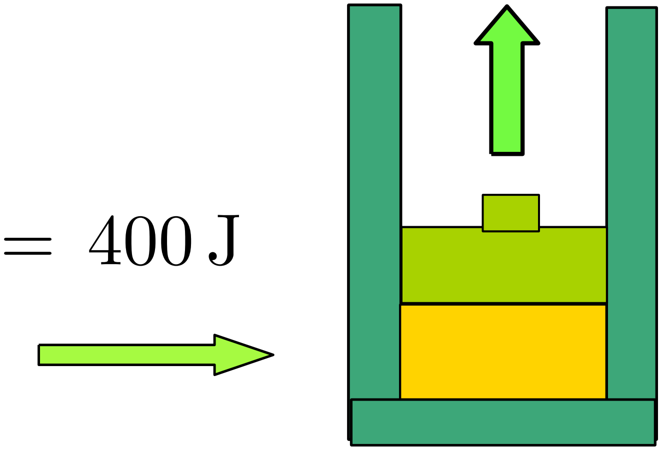
ΔW **is positive** ☛ $+120\,J$ **Work Out**

$$\Delta Q = \Delta U + \Delta W$$

$$\Delta U = \Delta Q - \Delta W$$

$$W_{out} = 120\,J$$

$$Q_{in} = 400\,J$$



$$\Delta U = +280\,J$$

$$\begin{aligned} \Delta U &= \Delta Q - \Delta W \\ &= (+400\,J) - (+120\,J) \\ &= +280\,J \end{aligned}$$

Application of first law of thermodynamics (cont'd)

Energy is conserved

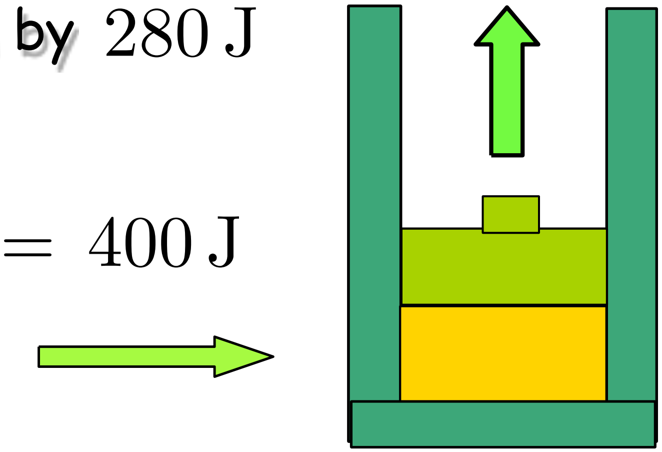
$400\,\mathrm{J}$ of input thermal energy is used to perform

is used to perform $120\,\mathrm{J}$ of external work,

$$\mathrm{W_{out}} = 120\,\mathrm{J}$$

increasing internal energy of system by $280\,\mathrm{J}$

$$\mathrm{Q_{in}} = 400\,\mathrm{J}$$



Increase in internal energy is ☛ $\Delta\mathrm{U} = +280\,\mathrm{J}$

It is common experience that a cup of hot coffee left in a cooler room eventually cools off

This process satisfies first law of thermodynamics since amount of energy lost by coffee is equal to amount gained by surrounding air

Now let us consider reverse process-hot coffee getting even hotter in a cooler room as a result of heat transfer from room air

We all know that this process never takes place

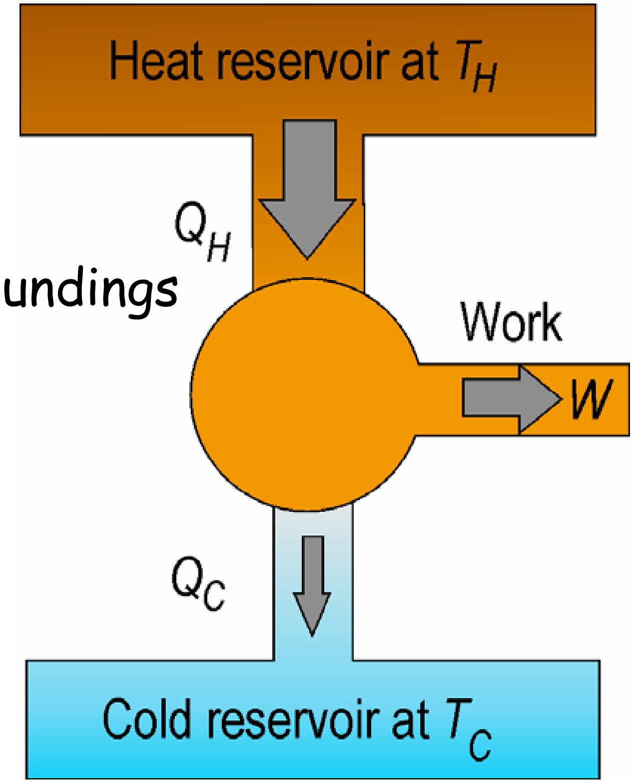
Yet, doing so would not violate first law as long as amount of energy lost by air is equal to amount gained by coffee

It is clear from above that processes proceed in a certain direction and not in reverse direction

First law places no restriction on direction of a process, but satisfying first law does not ensure that that process will actually occur

This inadequacy of first law to identify whether a process can take place is remedied by introducing another general principle, second law of thermodynamic

Second Law of Thermodynamics



Kelvin statement:
No system can absorb heat from a single reservoir
and convert it entirely to work
without additional net changes in the system or its surroundings

Clausius statement:
A process whose only net result is to absorb heat from a cold reservoir
and release the same amount of heat to a hot reservoir is impossible

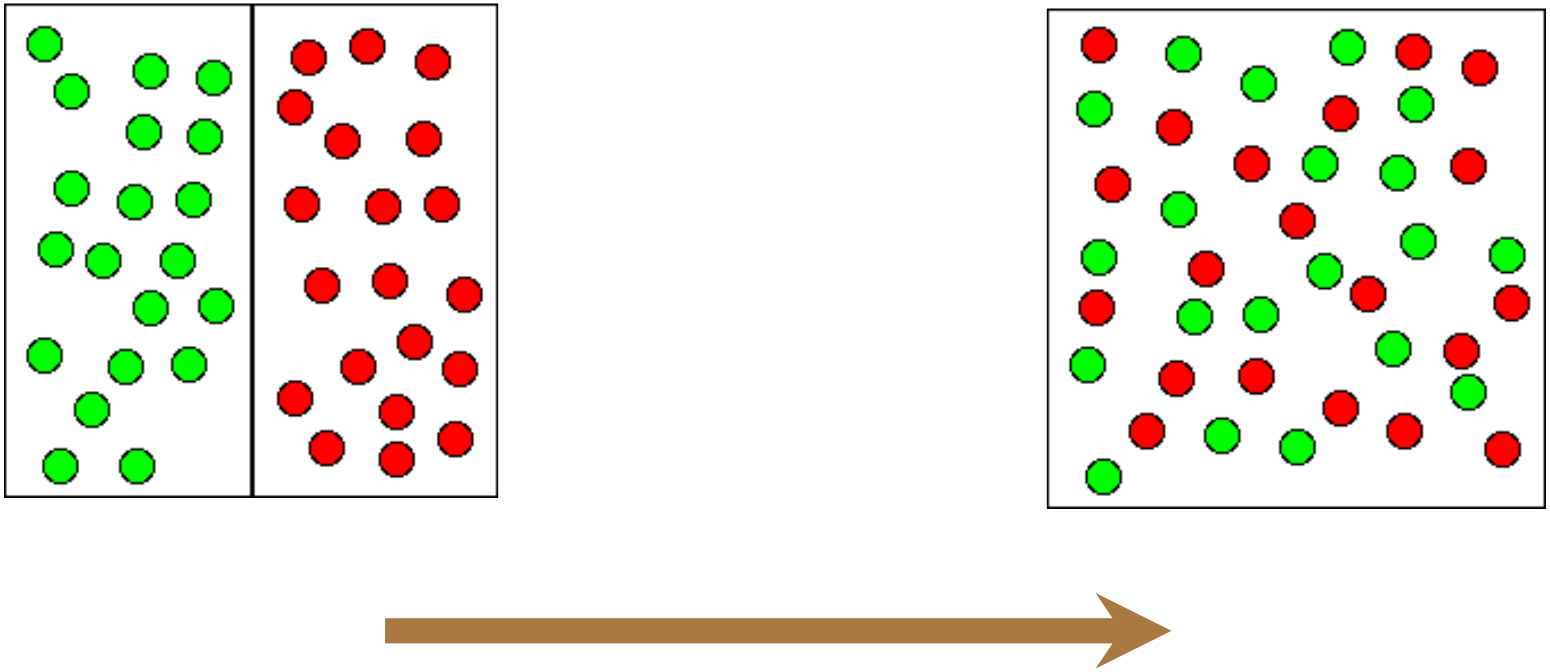
Entropy

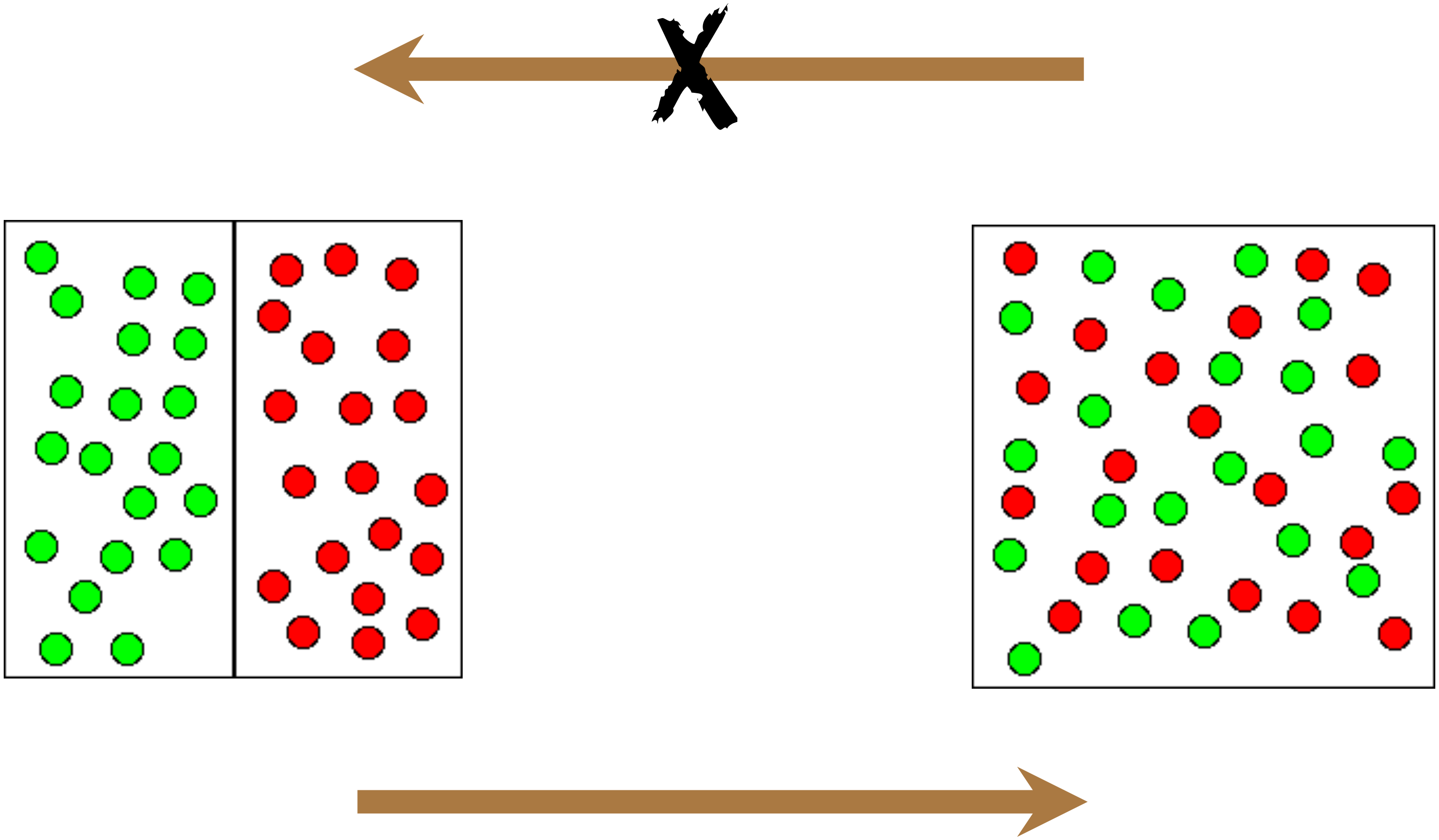
A measure of disorder of a system

Systems tend to change from a state of low entropy

to a state of higher entropy

☛ If left to themselves, systems tend to increase their entropy





Four thermodynamic processes

Isochoric Process ☛ $\Delta V = 0, \quad \Delta W = 0$

Isobaric Process ☛ $\Delta P = 0$

Isothermal Process ☛ $\Delta T = 0, \quad \Delta U = 0$

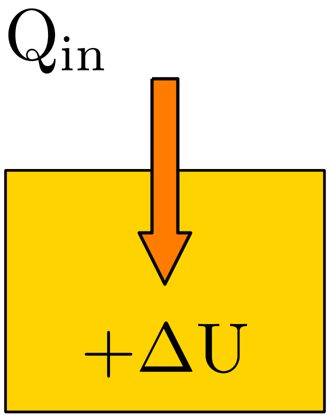
Adiabatic Process ☛ $\Delta Q = 0$

$$\Delta Q = \Delta U + \Delta W$$

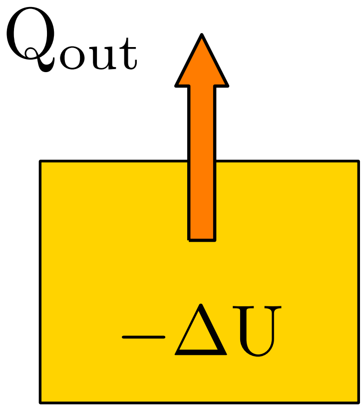
Isochoric process

CONSTANT VOLUME $\Delta V = 0,\ \Delta W = 0$

$\Delta Q = \Delta U + \cancelto{0}{\Delta W}$ so that $\Delta Q = \Delta U$



No Work Done

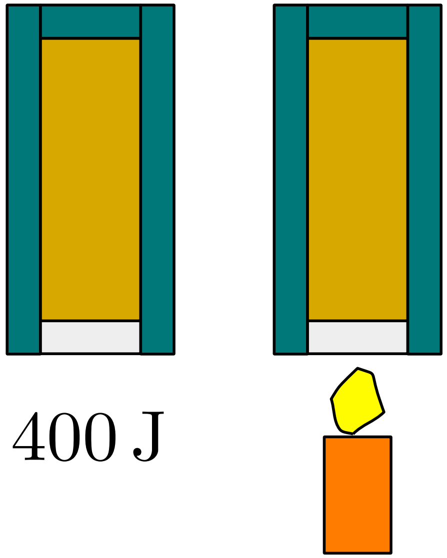


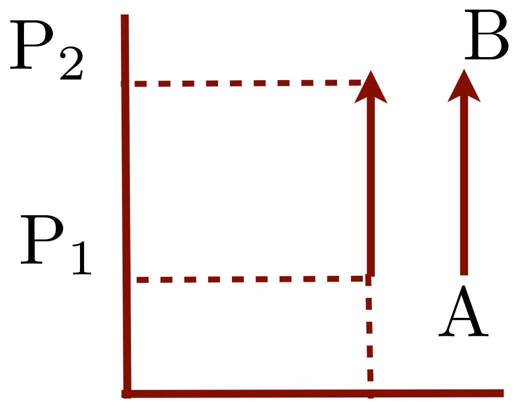
HEAT IN = INCREASE IN INTERNAL ENERGY

HEAT OUT = DECREASE IN INTERNAL ENERGY

Isochoric example

No Change in volume





$$V_1 = V_2$$

$$\frac{P_A}{T_A} = \frac{P_B}{T_B}$$

Heat input increases P with constant V

$400\,J$ heat input increases internal energy by $400\,J$ and zero work is done

Isobaric process

CONSTANT PRESSURE $\Delta P = 0$

$\Delta Q = \Delta U + \Delta W$ But $\Delta W = P\,\Delta V$

Q_{in}

Work Out

$+\Delta U$

Q_{out}

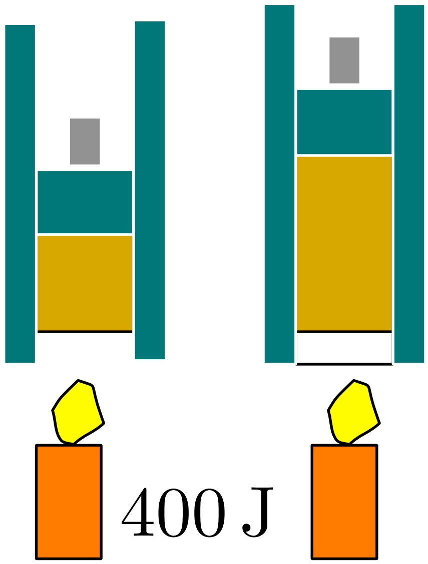
Work In

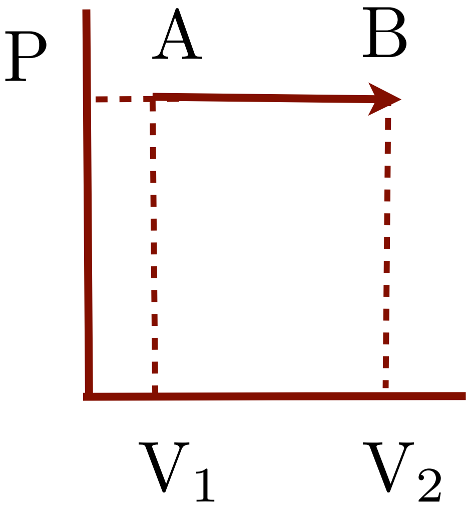
$-\Delta U$

HEAT IN $= W_{out} +$ INCREASE IN INTERNAL ENERGY

HEAT OUT $= W_{out} +$ DECREASE IN INTERNAL ENERGY

Isobaric example



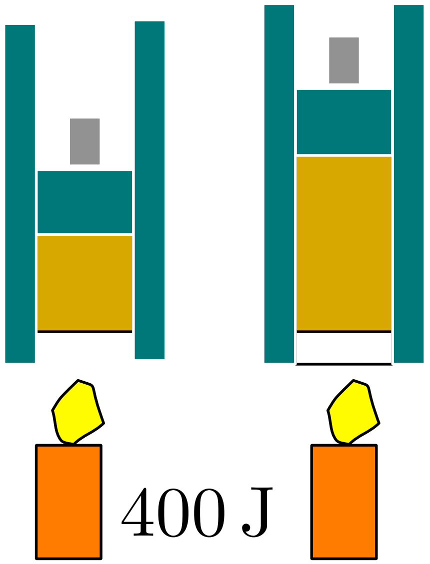


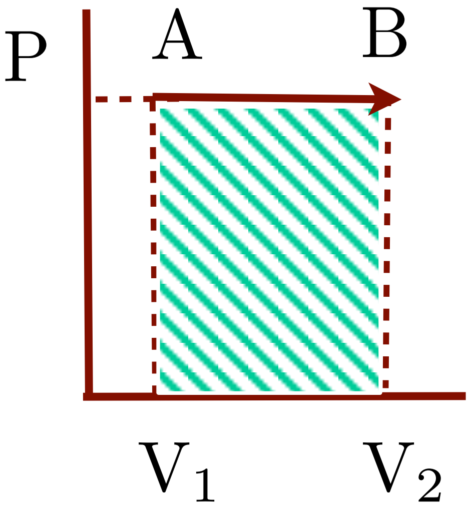
$$\frac{V_A}{T_A} = \frac{V_B}{T_B}$$

Heat input increases V with constant P

$400\,\mathrm{J}$ heat does $120\,\mathrm{J}$ of work increases internal energy by $280\,\mathrm{J}$

Isobaric work





$$\frac{V_A}{T_A} = \frac{V_B}{T_B}$$

Heat input increases V
with constant P

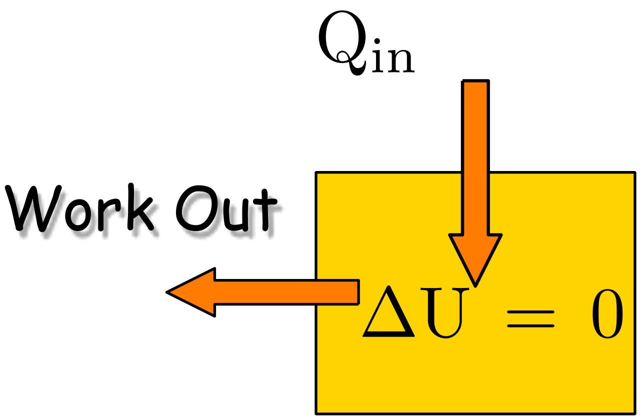
Work = Area under PV curve

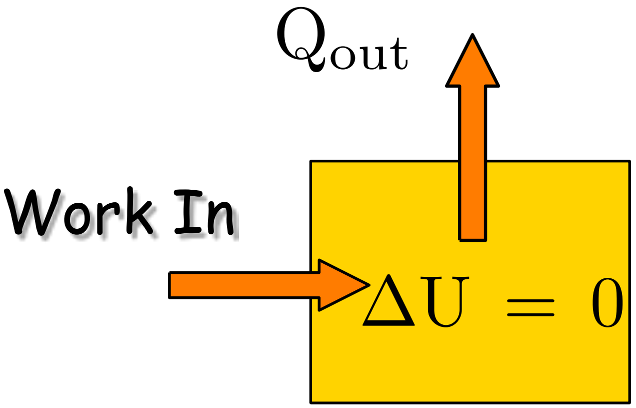
$$\text{Work} = P\,\Delta V$$

Isothermal process

CONSTANT TEMPERATURE $\Delta T = 0, \Delta U = 0$

$\Delta Q = \cancel{\Delta U} + \Delta W$ and $\Delta Q = \Delta W$

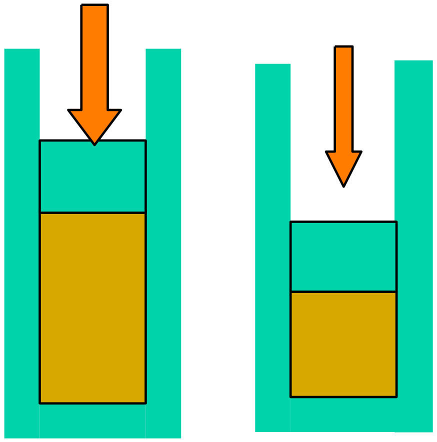


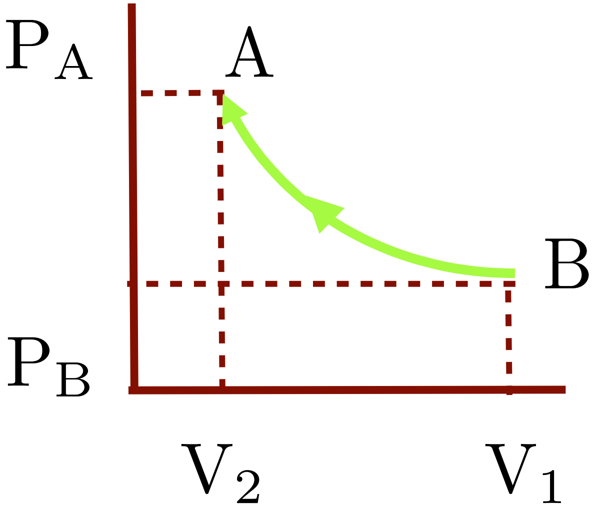


NEAT HEAT INPUT = WORK OUTPUT

WORK INPUT = NET HEAT OUT

Isothermal example ☛ Constant T





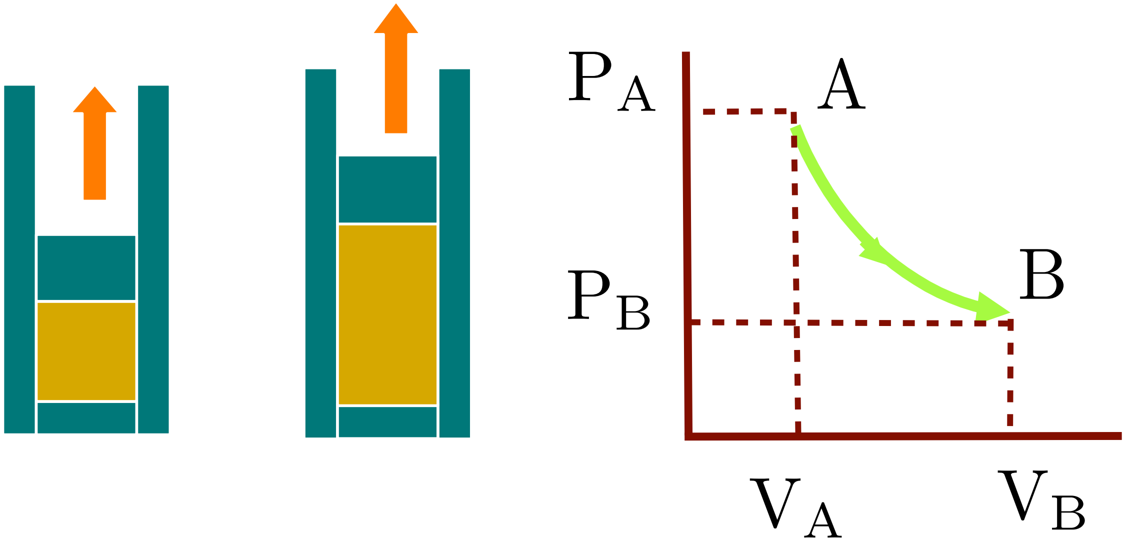
$$\Delta U = \Delta T = 0$$

Slow compression at constant temperature

☛ No change in U

$$P_A V_A = P_B V_B$$

Isothermal expansion ☛ Constant T



$$P_A V_A = P_B V_B$$

$$T_A = T_B$$

$$\Delta U = \Delta T = 0$$

$400\,\mathrm{J}$ of energy is absorbed by gas

as $400\,\mathrm{J}$ of work is done on gas

$$\Delta T = \Delta U = 0$$

Isothermal Work

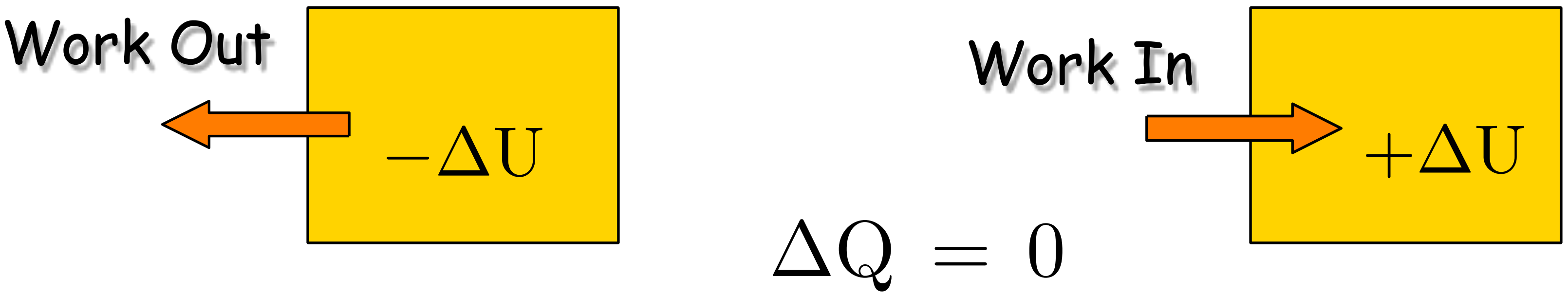
$$W = nRT \ln \frac{V_B}{V_A}$$

Adiabatic process

NO HEAT EXCHANGE $\Delta \mathrm{Q} = 0$

$$\cancelto{}{\Delta \mathrm{Q}} = \Delta \mathrm{U} + \Delta \mathrm{W}; \quad \Delta \mathrm{W} = -\Delta \mathrm{U} \quad \text{or} \quad \Delta \mathrm{U} = -\Delta \mathrm{W}$$

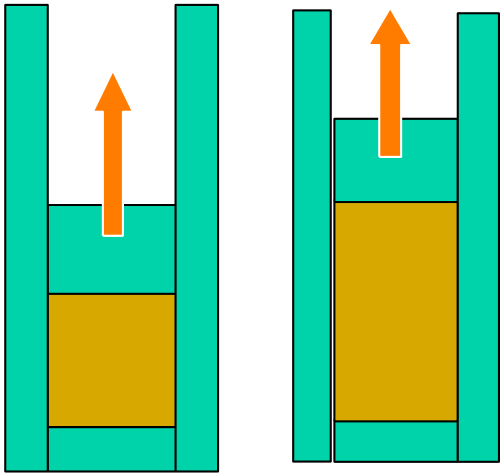
$$\Delta \mathrm{W} = -\Delta \mathrm{U} \qquad \Delta \mathrm{U} = -\Delta \mathrm{W}$$

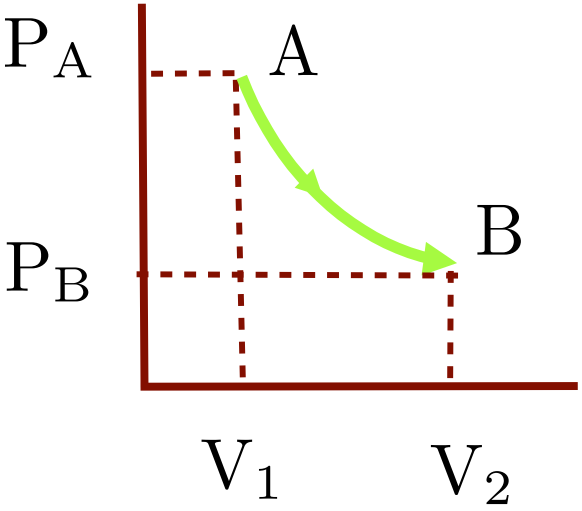


Work done at EXPENSE of internal energy

INPUT Work INCREASES internal energy

Adiabatic example





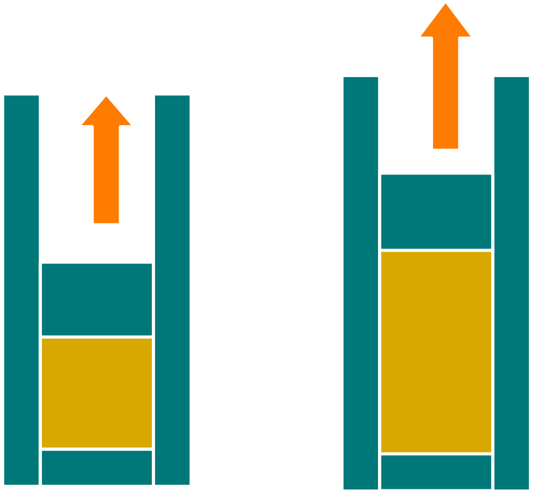
Insulated Walls

$$\Delta Q = 0$$

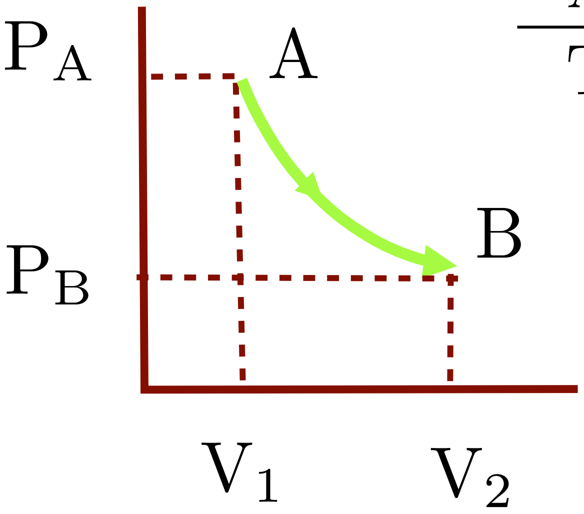
Expanding gas does work
with zero heat loss

$$\text{Work} = -\Delta U$$

Adiabatic expansion



$$\Delta Q = 0$$



$$\frac{P_A V_A}{T_A} = \frac{P_B V_B}{T_B}$$

$400\,\mathrm{J}$ of WORK is done,

DECREASING internal energy by $400\,\mathrm{J}$

Net heat exchange is ZERO

$$\Delta Q = 0$$

$$P_A V_A^{\gamma} = P_B V_B^{\gamma}$$

REMEMBER, FOR ANY PROCESS INVOLVING AN IDEAL GAS

$$PV = nRT$$

$$\frac{P_A V_A}{T_A} = \frac{P_B V_B}{T_B}$$

$$\Delta Q = \Delta U + \Delta W$$

$$\Delta U = nC_V \Delta T$$

Example problem

A $2-\mathrm{L}$ sample of Oxygen gas has an initial temperature and pressure of $200\,\mathrm{K}$ and $1\,\mathrm{atm}$

Gas undergoes four processes:

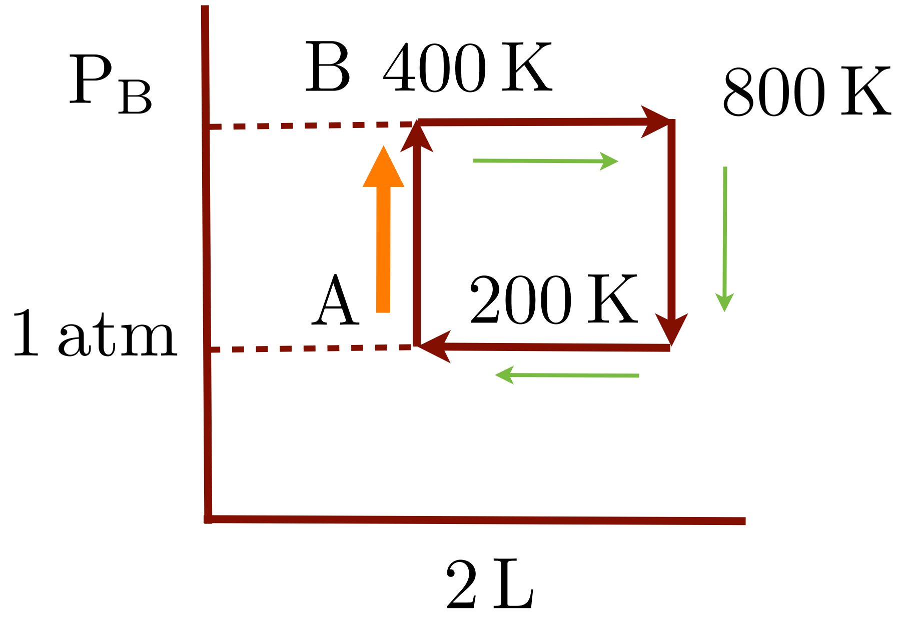
- ✓ AB : Heated at constant V to $400\,\mathrm{K}$
- ✓ BC : Heated at constant P to $800\,\mathrm{K}$
- ✓ CD : Cooled at constant V back to $1\,\mathrm{atm}$
- ✓ DA : Cooled at constant P back to $200\,\mathrm{K}$

PV-DIAGRAM FOR PROBLEM

How many moles of O_2 are present?

Consider point A

$$PV = nRT$$

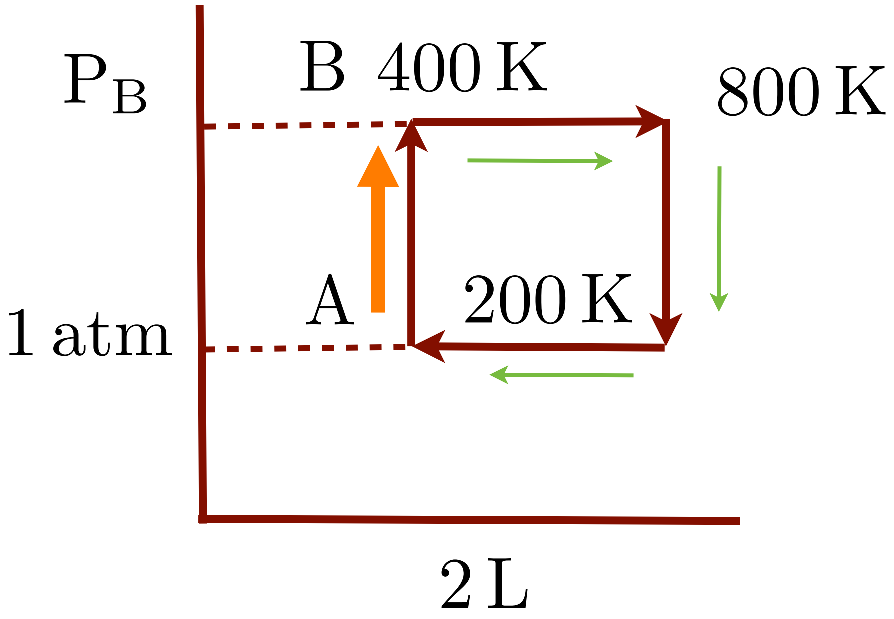


$$n = \frac{PV}{RT} = \frac{(101.300\,\mathrm{Pa})(0.002\,\mathrm{m}^3)}{(8.314\,\mathrm{J/mol\cdot K})(200\,\mathrm{K})} = 0.122\,\mathrm{mol}$$

Process AB ☛ Isochoric

What is pressure at point B ?

$$\frac{P_A}{T_A} = \frac{P_B}{T_B}$$



$$\frac{1\,\mathrm{atm}}{200\,\mathrm{K}} = \frac{P_B}{400\,\mathrm{K}}$$

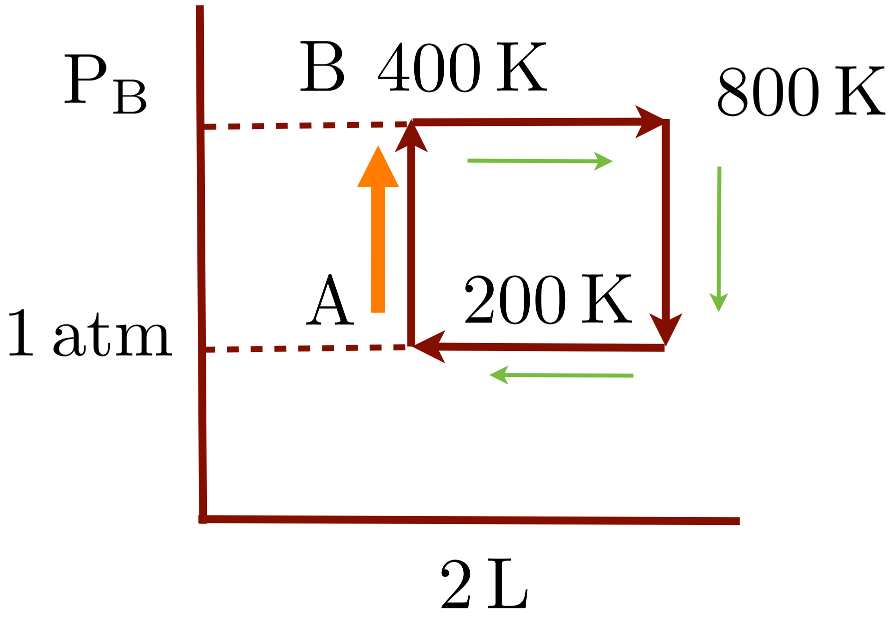
$P_B = 2\,\mathrm{atm}$ or $203\,\mathrm{kPa}$

Process AB ☛ $\Delta Q = \Delta U + \Delta W$

Analyze first law for **ishochoric** process AB

$$\Delta W = 0$$

$$\Delta Q = \Delta U = n\, C_V\, \Delta T$$



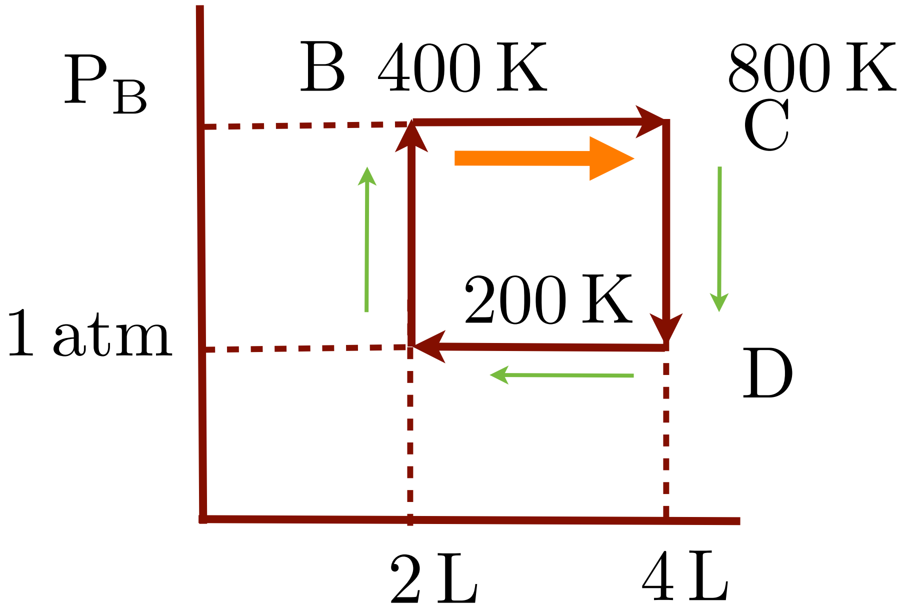
$$\Delta U = (0.122\,\mathrm{mol})(21.1\,\mathrm{J/mol\,K})(400\,\mathrm{K} - 200\,\mathrm{K})$$

$\Delta Q = +514\,\mathrm{J}$ $\Delta U = +514\,\mathrm{J}$ $\Delta W = 0$

Process BC ☛ Isobaric

What is Volume at point C (& D) ?

$$\frac{\mathrm{V_B}}{\mathrm{T_B}} = \frac{\mathrm{V_C}}{\mathrm{T_C}}$$



$$\frac{2\mathrm{L}}{400\,\mathrm{K}} = \frac{\mathrm{V_C}}{800\,\mathrm{K}}$$

$$\mathrm{V_C} = \mathrm{V_D} = 4\,\mathrm{L}$$

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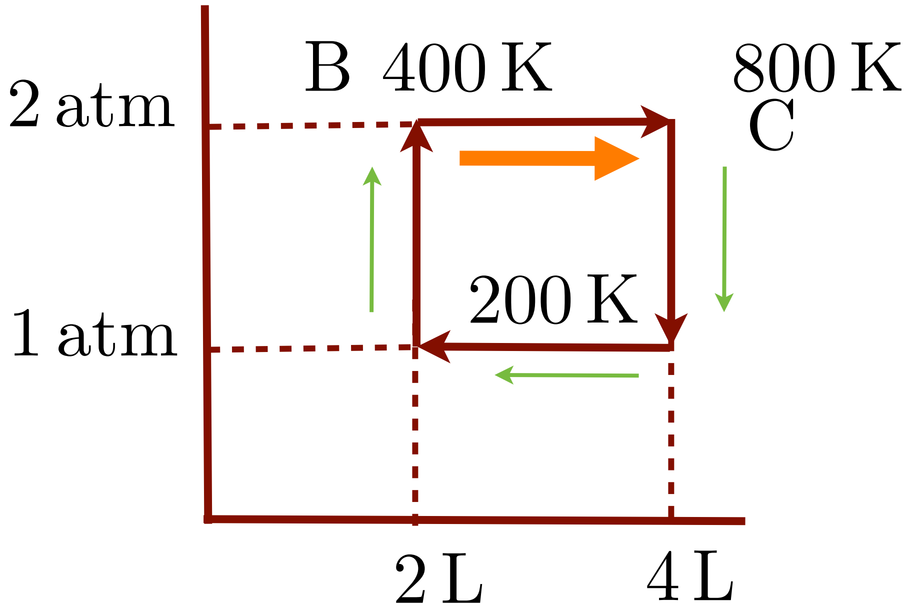
Finding ΔU for process BC

Process BC is ISOBARIC

$$\Delta P = 0$$



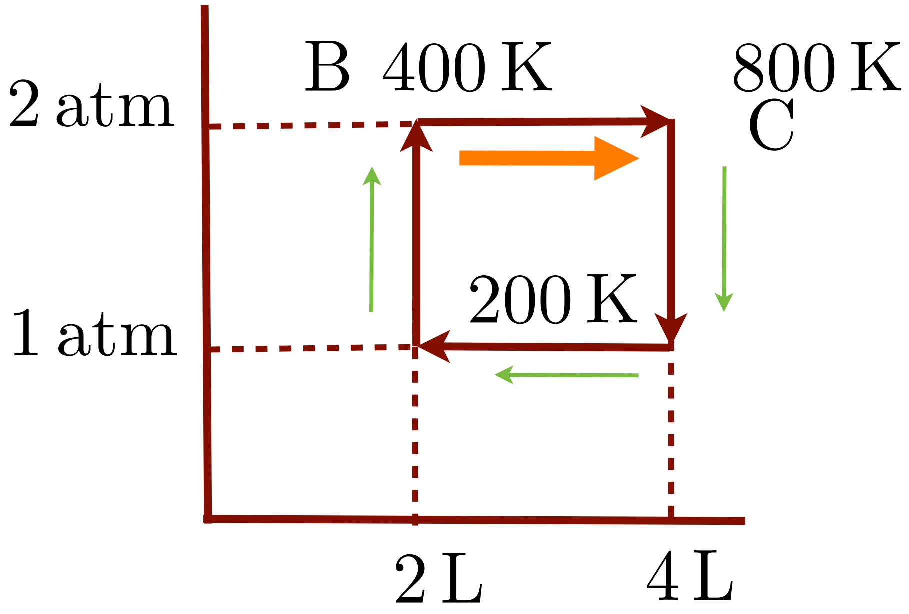
$$\Delta U = n\, C_V\, \Delta T$$

$$\Delta U = (0.122\,\mathrm{mol})(21.1\,\mathrm{J/mol\,K})(800\,\mathrm{K} - 400\,\mathrm{K})$$

$$\Delta U = +1028\,\mathrm{J}$$

Finding ΔW for process BC

Work depends on change in V



$$\Delta \mathrm{P} = 0$$

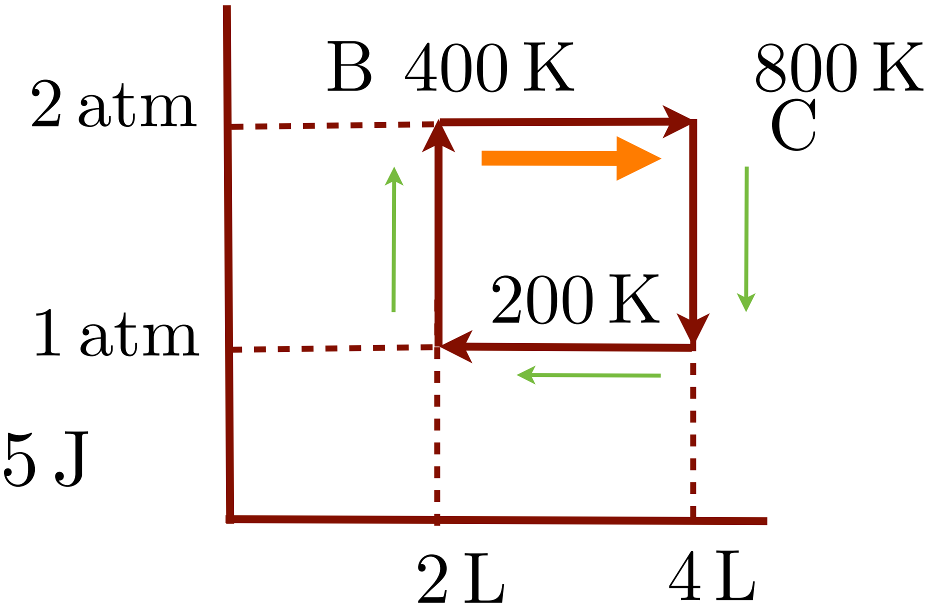
$$\mathrm{Work} = \mathrm{P}\,\Delta \mathrm{V}$$

$$\Delta \mathrm{W} = (2\,\mathrm{atm})(4\,\mathrm{L} - 2\,\mathrm{L}) = 4\,\mathrm{atm\,L} = 405\,\mathrm{J}$$

$$\Delta \mathrm{W} = +405\,\mathrm{J}$$

Finding ΔQ for process BC

Analyze first law for BC



$$\Delta Q = \Delta U + \Delta W$$

$$\Delta Q = +1028\,\mathrm{J} + 405\,\mathrm{J}$$

$$\Delta Q = +1433\,\mathrm{J}$$

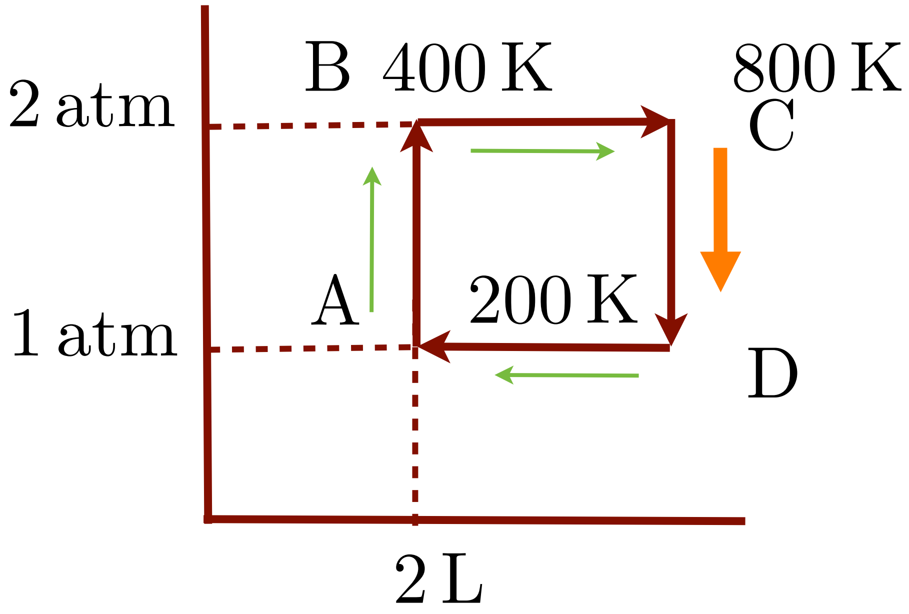
$$\Delta Q = 1433\,\mathrm{J} \qquad \Delta U = 1028\,\mathrm{J} \qquad \Delta W = +405\,\mathrm{J}$$

Process CD ☛ Isochoric

What is temperature at point D?

$$\frac{\mathrm{P_C}}{\mathrm{T_C}} = \frac{\mathrm{P_D}}{\mathrm{T_D}}$$

$$\frac{2\,\mathrm{atm}}{800\,\mathrm{K}} = \frac{1\,\mathrm{atm}}{\mathrm{T_D}}$$



$$\mathrm{T_D} = 400\,\mathrm{K}$$

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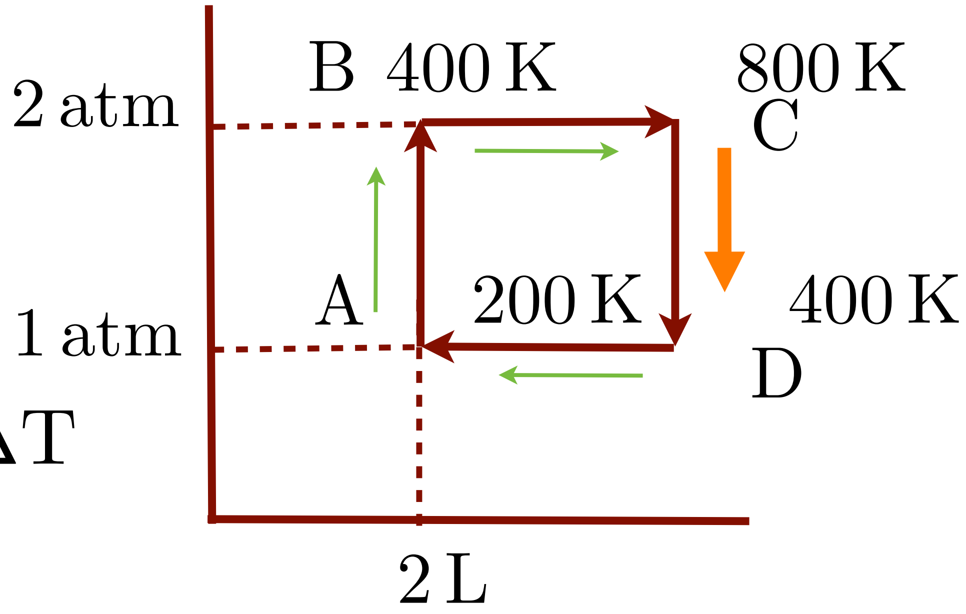
Process CD ☛ $\Delta Q = \Delta U + \Delta W$

Analyze first law for **ISOCHORIC** process CD



$$\Delta W = 0$$

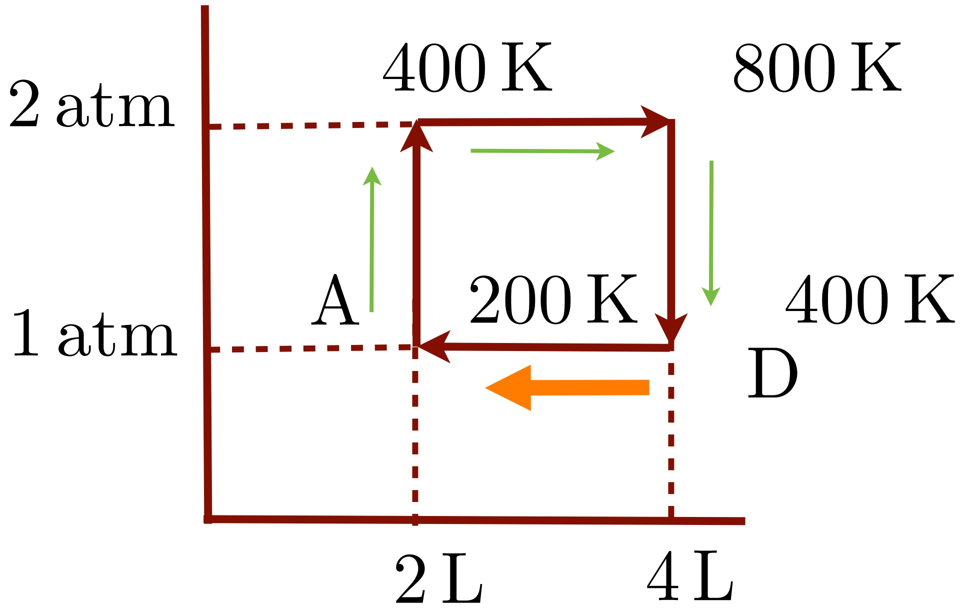
$$\Delta Q = \Delta U = n\,C_V\,\Delta T$$

$$\Delta U = (0.122\,\mathrm{mol})(21.1\,\mathrm{J/mol\,K})(400\,\mathrm{K} - 800\,\mathrm{K})$$

$$\Delta Q = -1028\,\mathrm{J} \qquad \Delta U = -1028\,\mathrm{J} \qquad \Delta W = 0$$

Finding ΔU for process DA

Process DA is **ISOBARIC**



$$\Delta P = 0$$

$$\Delta U = n\,C_V\,\Delta T$$

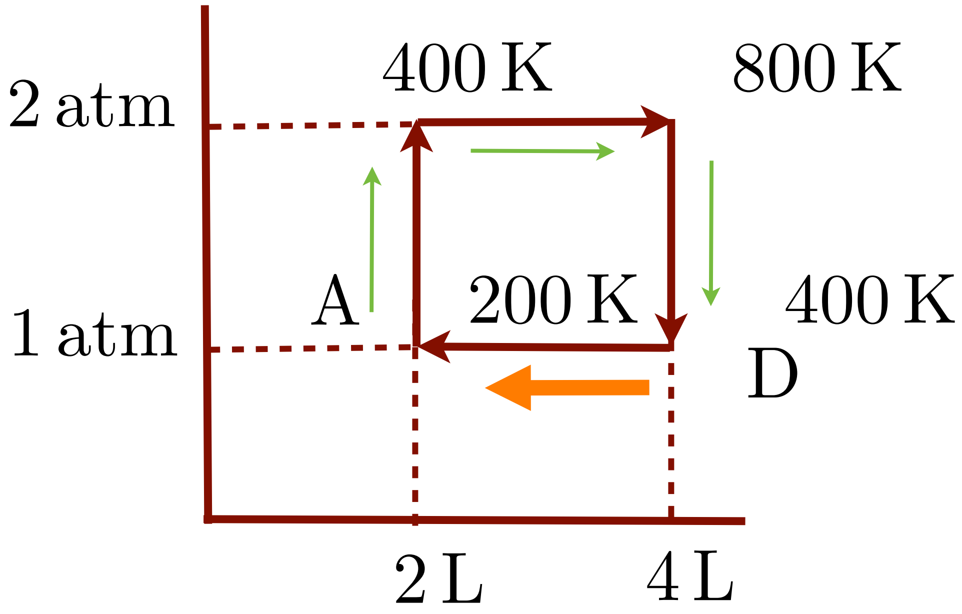
$$\Delta U = (0.122\,\mathrm{mol})(21.1\,\mathrm{J/mol\,K})(400\,\mathrm{K} - 200\,\mathrm{K})$$

$$\Delta U = -514\,\mathrm{J}$$

Finding ΔW for process DA

Work depends on change in V

$$\Delta P = 0$$



$$\text{Work} = P\,\Delta V$$

$$\Delta W = (1\,\text{atm})(2\,\text{L} - 4\,\text{L}) = -2\,\text{atm}\,\text{L} = -203\,\text{J}$$

$$\Delta W = -203\,\text{J}$$

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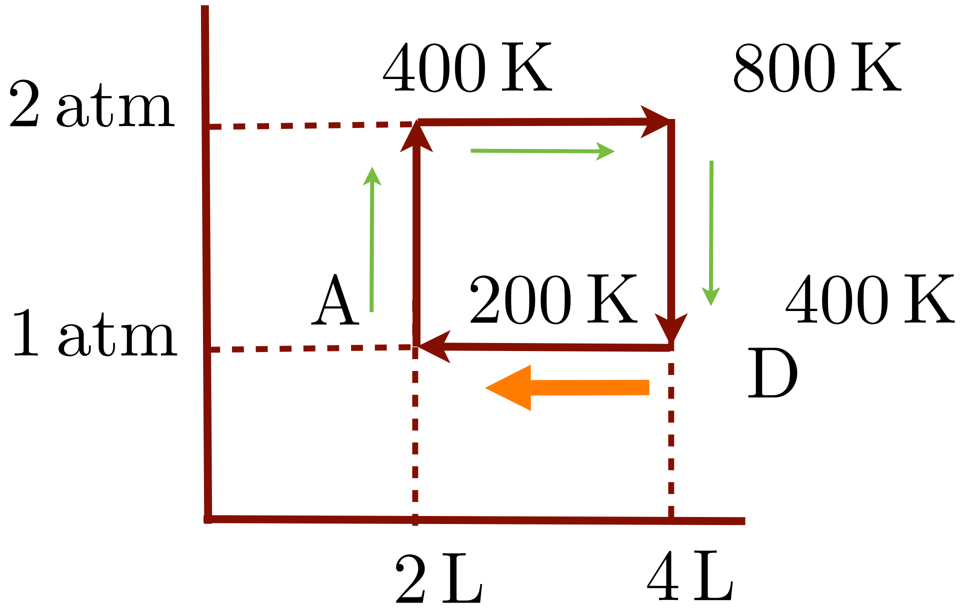
Finding ΔQ for process DA

Analyze first law for DA

$$\Delta Q = \Delta U + \Delta W$$

$$\Delta Q = -514\,\mathrm{J} - 203\,\mathrm{J}$$

$$\Delta Q = -717\,\mathrm{J}$$



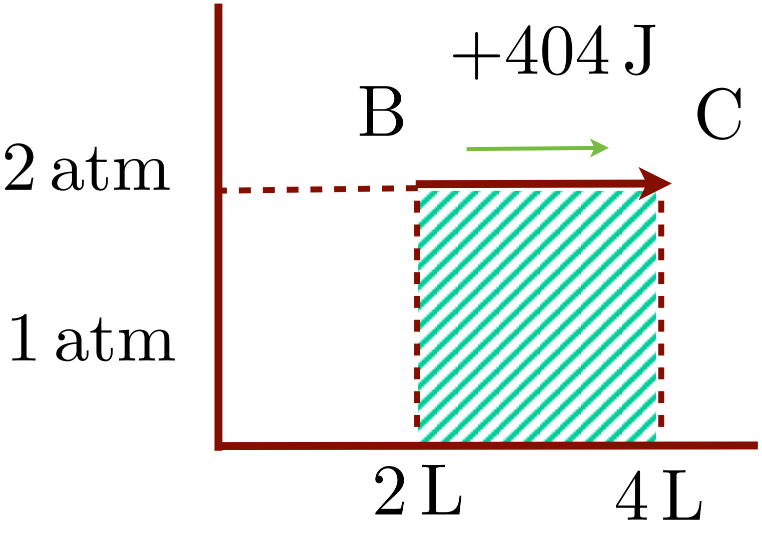
$$\Delta Q = -717\,\mathrm{J} \qquad \Delta U = -514\,\mathrm{J} \qquad \Delta W = -203\,\mathrm{J}$$

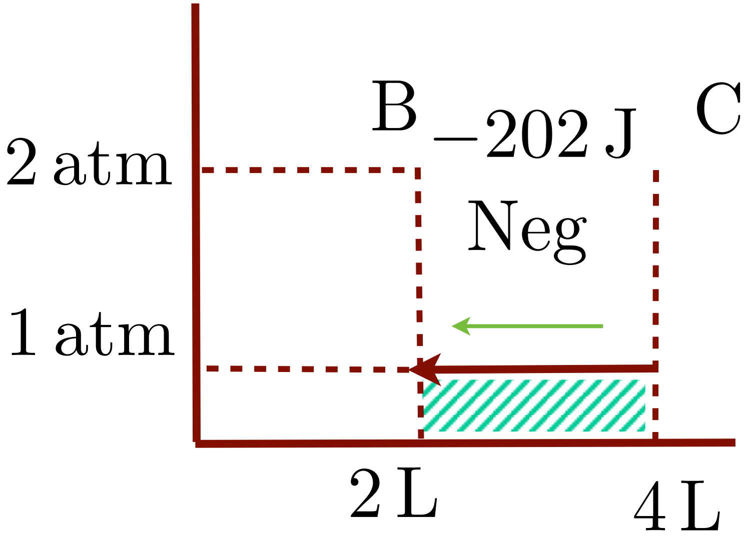
PROBLEM SUMMARY

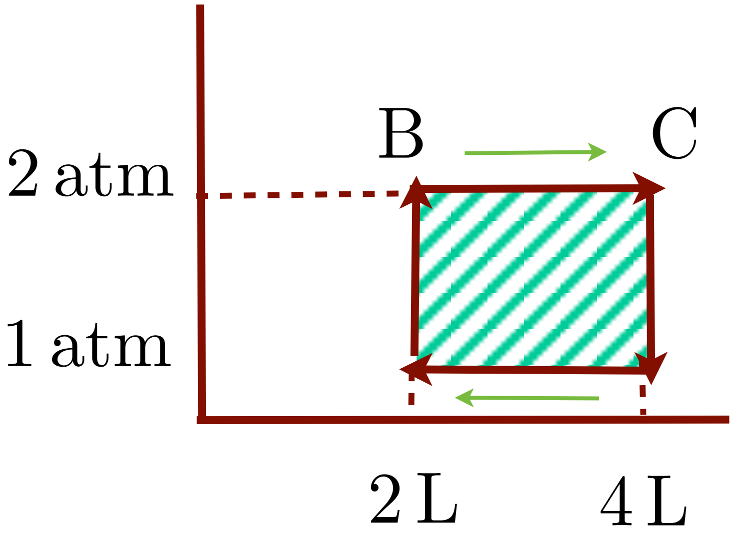
For all processes ☛ $\Delta Q = \Delta U + \Delta W$

PROCESS	ΔQ	ΔU	ΔW
AB	$514\,\mathrm{J}$	$514\,\mathrm{J}$	0
BC	$1433\,\mathrm{J}$	$1028\,\mathrm{J}$	$405\,\mathrm{J}$
CD	$-1028\,\mathrm{J}$	$-1028\,\mathrm{J}$	0
DA	$-717\,\mathrm{J}$	$-514\,\mathrm{J}$	$-203\,\mathrm{J}$
TOTALS	$202\,\mathrm{J}$	0	$202\,\mathrm{J}$

NET WORK FOR COMPLETE CYCLE IS ENCLOSED AREA







$$\text{Area} = (1\,\text{atm})(2\,\text{L})$$

$$\text{Net work} = 2\,\text{atm L} = 202\,\text{J}$$