

Thermodynamics is the branch of physics that deals with the way in which a system interacts with its surroundings.



Thermodynamic System substance - usually an ideal gas

Surroundings everything else – walls of container, outside environment

State of the system for a gas, a particular set of values of P , V , n , and T

Internal energy: total potential energy and random kinetic energy of the molecules of a substance

Symbol: U

Units: J

Work: product of force and displacement in the direction of the force

Symbol: W

Units: J

Thermal Energy (Heat): the transfer of energy between two substances by non-mechanical means – conduction, convection and radiation

Symbol: Q

Units: J

The internal energy of a system can change by . . .

Heating

$Q = +400 J$

$Q_{in} = +400 J$

$Q_{out} = -400 J$

Block of ice

Cooling

$Q = -400 J$

$Q_{in} = -400 J$

$Q_{out} = +400 J$

Expansion

$W = +100 J$

$W_{by} = +100 J$

$W_{on} = -100 J$

Compression

$W = -100 J$

$W_{by} = -100 J$

$W_{on} = +100 J$

Definitions: Q = thermal energy added to system

W = work done by the system

ΔU = change in internal energy of the system

1. A sample of gas is heated with a Bunsen burner and allowed to expand. If 400 J of thermal energy are transferred to the gas during heating and the gas does 100 J of work by expanding, what is the resulting change in the internal energy of the gas?

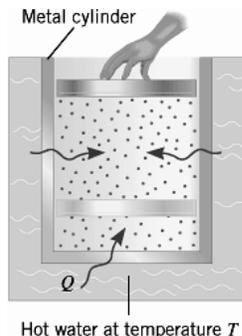


$$\Delta U = 400 \text{ J} - 100 \text{ J}$$

$$\Delta U = 300 \text{ J}$$

$$\Delta U = Q - W$$

2. A sample of gas is warmed by placing it in a bath of hot water, adding 400 J of thermal energy. At the same time, 100 J of work is done compressing the gas manually. What is the resulting change in the internal energy of the gas?



$$\Delta U = 400 \text{ J} - (-100 \text{ J})$$

$$\Delta U = 500 \text{ J}$$

$$\Delta U = Q - W$$

Formula:

$$\Delta U = Q - W$$

$$Q = \Delta U + W$$

First Law of Thermodynamics: The thermal energy transferred to a system from its surroundings is equal to the work done by the system plus the change in internal energy of the system.

NOTE: The First Law is a statement of . . . the principle of conservation of energy.

In each case, determine the change in the internal energy of the gas.

- a) A gas gains 1500 J of heat from its surroundings, and expands, doing 2200 J of work on the surroundings.

$$Q = \Delta U + W$$

Heating a hot air balloon
and let canvas expand

$$1500 \text{ J} = \Delta U + (+2200 \text{ J})$$

$$\Delta U = -700 \text{ J}$$

- b) A gas gains 1500 J of heat at the same time as an external force compresses it, doing 2200 J of work on it.

$$Q = \Delta U + W$$

$$1500 \text{ J} = \Delta U + (-2200 \text{ J})$$

$$\Delta U = +3700 \text{ J}$$

Heating and
stirring water

Internal energy of many substances depends on . . . temperature and intermolecular bonds

Internal energy of an ideal gas depends on . . .

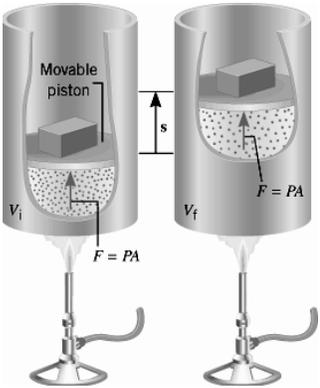
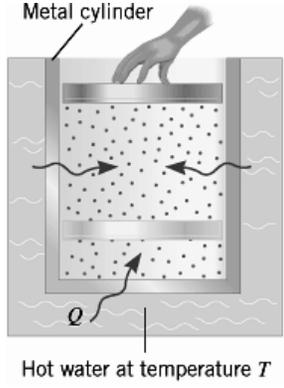
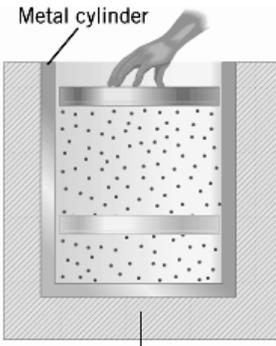
- only on temperature since there are no intermolecular bonds

$$U \propto T \text{ so } \Delta U \propto \Delta T$$

U increases if T increases, if +ΔT then +ΔU

- the change in internal energy of ideal gas is path independent

1. An isobaric process is one that occurs **at constant pressure.** $\Delta P = 0$
2. An isochoric (isovolumetric) process is one that occurs **at constant volume.** $\Delta V = 0$
3. An isothermal process is one that occurs **at constant temperature.** $\Delta T = 0$
4. An adiabatic process is one that occurs **without the transfer of thermal energy.** $\Delta Q = 0$

Isobaric Process	Isochoric (Isovolumetric) Process	Isothermal Process	Adiabatic Process
 <p>The gas in the cylinder is expanding isobarically because the pressure is held constant by the external atmosphere and the weight of the piston and the block. Heat can enter or leave through the non-insulating walls.</p>	 <p>The gas in the cylinder is being heated isochorically since the volume of the cylinder is held fixed by the rigid walls. Heat can enter or leave through the non-insulating walls.</p>	 <p>The gas in the cylinder is being allowed to expand isothermally since it is in contact with a water bath (heat reservoir) that keeps the temperature constant. Heat can enter or leave through the non-insulating walls.</p>	 <p>The gas in the cylinder is being compressed adiabatically since the cylinder is surrounded by an insulating material.</p>

Isobaric Process

Work Involved in a Volume Change at Constant Pressure

How is pressure held constant? **weight of brick, piston, and atmosphere constant**

How is work done by the gas?

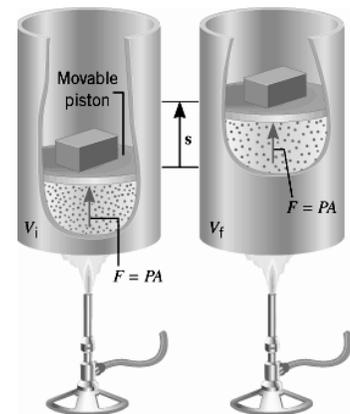
Molecules strike piston and transfer momentum and KE to it causing it to move upward/outward - as KE decreases, so does internal energy and T

How much work is done by the gas if it expands at constant pressure?

$$W = F s \cos \theta$$

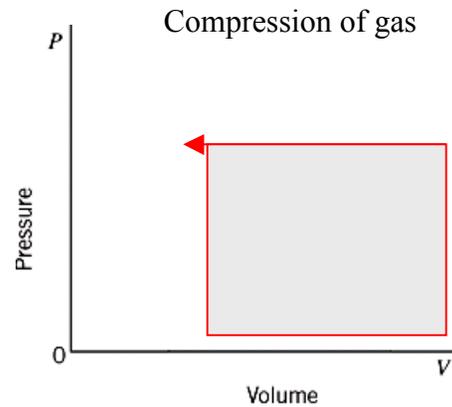
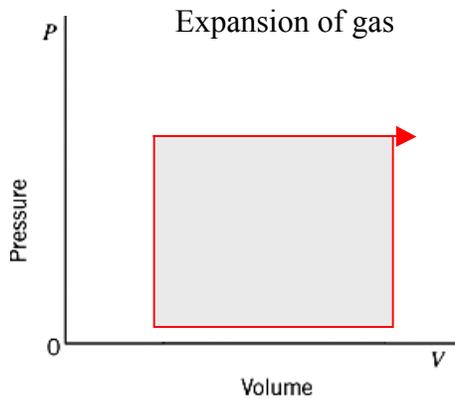
$$W = p A s \cos \theta^0$$

$$W = p \Delta V$$



animation: Serway: chap 12: TDM06AN1

What does an isobaric process look like on a diagram of pressure vs. volume (P-V diagram)?

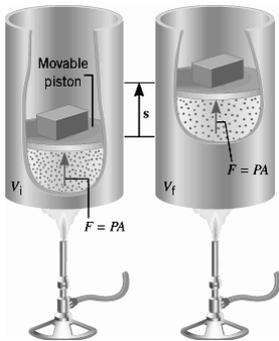


How can the amount of work done by a gas during a process be determined from a P-V diagram?

work done by the gas = area underneath curve
 arrow to right = positive work done by the gas = expansion
 arrow to left = negative work done by the gas = compression

Isobaric Processes and the First Law of Thermodynamics

Expansion at constant pressure



Gas laws:

$$PV/T = PV/T$$

$$P_1 = P_2$$

$$\text{So } V_1/T_1 = V_2/T_2$$

If V increases, so does T

1st law:

If gas is ideal, U increases when T increases

$$U \propto T$$

So + ΔT means + ΔU

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

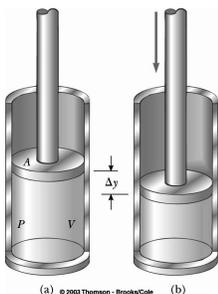
$$(+)= (+) + (+)$$

More heat is added than work done if isobaric

Example: A gas is allowed to expand isobarically by adding 1000 J of thermal energy, causing the gas to increase its internal energy by 200 J. How much work is done by the gas in expanding?

800 J

Compression at constant pressure



Gas laws:

$$PV/T = PV/T$$

$$P_1 = P_2$$

$$\text{So } V_1/T_1 = V_2/T_2$$

If V decreases, so does T

1st law:

If gas is ideal, U decreases when T decreases

$$U \propto T$$

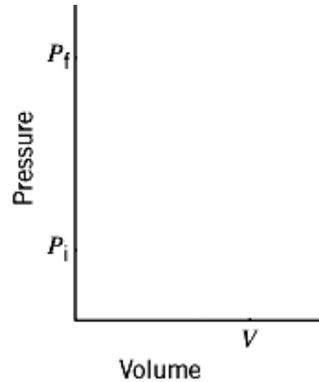
So - ΔT means - ΔU

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

$$(-) = (-) + (-)$$

More heat is removed than work done if isobaric - Heat leaves system

Isochoric (Isovolumetric) Process



Work:

Means $\Delta V = 0$

If $\Delta V = 0$
then $W = 0$

No area = no work

Gas law:

$$PV/T = PV/T$$

$$V = V$$

$$\text{So } P/T = P/T$$

If T increases, P increases

1st law:

$$Q = \Delta U + W$$

If $Q+$, then $\Delta U+$

$$\Delta U = Q$$

(since $W = 0$)

If ideal gas, $\Delta U \propto \Delta T$
so $\Delta T+$

1. One mole of an ideal gas is heated at a constant volume of $2.0 \times 10^{-3} \text{ m}^3$ from an initial pressure of $1.0 \times 10^5 \text{ Pa}$ to a final pressure of $5.0 \times 10^5 \text{ Pa}$.

a) Determine the initial and final temperatures of the gas.

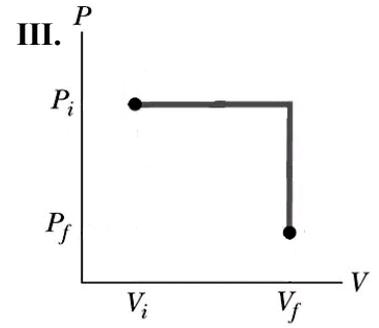
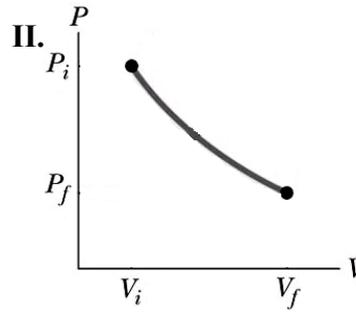
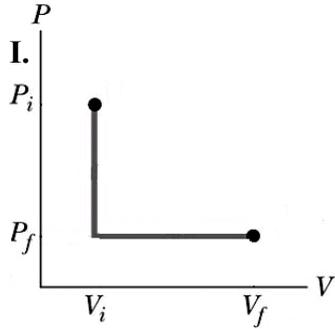
b) Does the internal energy of the gas increase or decrease? Justify your answer.

c) Determine the work done by the gas during this process.

c) If the change in internal energy of the gas is 1200J, determine the amount of thermal energy added to the gas.

2. In each case shown below, an ideal gas at $5.0 \times 10^5 \text{ Pa}$ and $1.0 \times 10^{-3} \text{ m}^3$ expands to $4.0 \times 10^{-3} \text{ m}^3$ at a pressure of $1.0 \times 10^5 \text{ Pa}$ by a different process or series of processes.

IB 12



a) Compare the change in internal energy of the gas as a result of each process. Justify your answer.

b) Compare the work done by (or on) the gas during each process. Justify your answer.

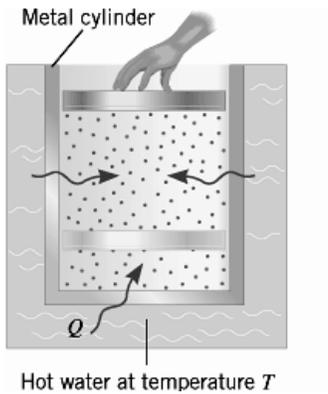
c) Compare the thermal energy added to or removed from the gas during each process. Justify your answer.

d) If the change in internal energy in each case is 500 J, calculate the work done and thermal energy exchanged in each case.

Conclusions:

- 1) Change in internal energy does not depend on the path taken – only on the change in temperature – path independent.
- 2) Work done and thermal energy transferred depend on the path taken between the initial and final states.

Heat reservoir: **hot or cold water bath that maintains constant temperature of gas by supplying or removing thermal energy**



Gas Law:

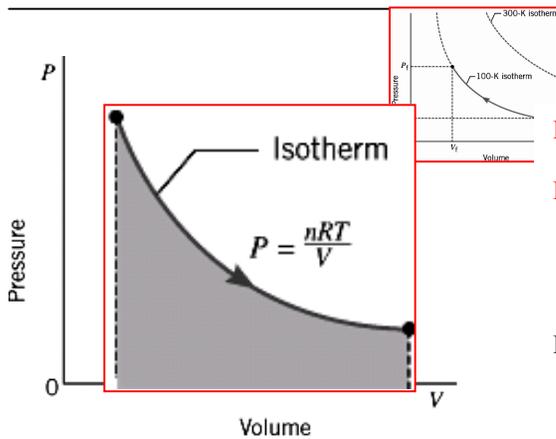
- $PV/T = PV/T$
- $T = T$
- $PV = PV$
- If V increases, P decreases

1st Law:

- $Q = \Delta U + W$
- If ideal gas, $\Delta U \propto \Delta T$
so $\Delta T = 0$ means $\Delta U = 0$
- so $Q = W$

Expansion: **thermal energy flows in at same rate as work is done by gas**

Compression: **thermal energy flows out at same rate as work done on the gas**



Ideal Gas Equation of State

- $PV = nRT$
- $P = nRT/V$ - hyperbola for fixed T

$P_1V_1/T_1 = P_2V_2/T_2$

$P_1V_1 = P_2V_2$ on one isotherm

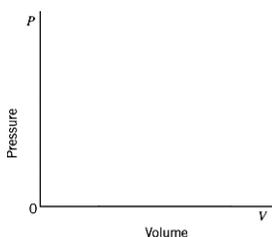
Isotherm: hyperbola of constant temperature

Conclusions:

- 1) **all states on one isotherm have same U since have same $T - \Delta T$ and $\Delta U = 0$ moving along same isotherm**
- 2) **isotherms further from origin - higher T so higher U**
- 3) **ΔU between two isotherms is path independent - same ΔU since same ΔT and $\Delta U \propto \Delta T$**

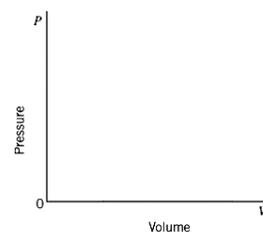
Expansion

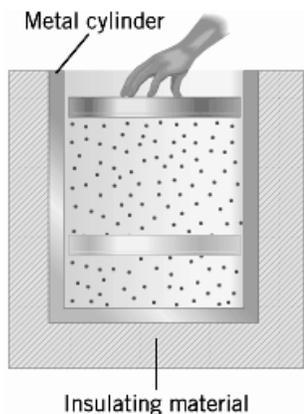
arrow to right
work done by +
 Q added +



Compression

arrow to left
work done on -
 Q removed -





Adiabatic walls: **insulating walls so no thermal energy can enter or leave system**

NOTE: **Rapid expansion or compression of gas is approximately adiabatic**

1st Law:

$$Q = \Delta U + W$$

$$Q = 0$$

$$\text{so } \Delta U = -W$$

If ideal gas, $\Delta U \propto \Delta T$

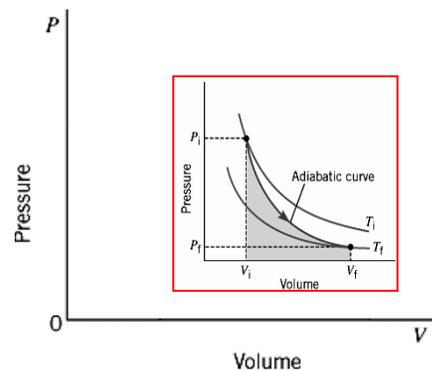
$$\text{so } W \propto -\Delta T$$

Expansion: **work done by gas cools gas down as it loses internal energy**

$$PV/T = PV/T$$

P decreases and V increases and T decreases

$W \propto -\Delta T$ so $+W$ means temperature goes down = jumps to lower isotherm = gas cools down

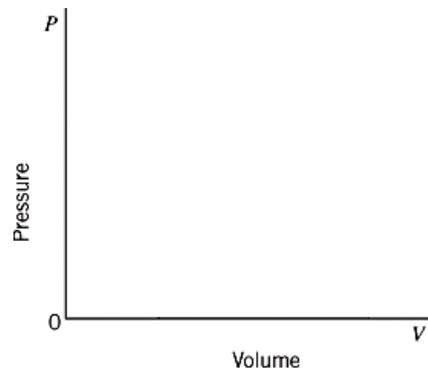


Compression: **work done on gas heats gas up as it gains internal energy**

$$PV/T = PV/T$$

P increases and V decreases and T increases

$W \propto -\Delta T$ so $-W$ means temperature goes up = jumps to higher isotherm = gas gets hotter



1. If 410 J of heat energy are added to an ideal gas causing it expand at constant temperature,

a) what is the change in internal energy of the gas?

b) how much work is done by the gas?

c) how much work is done on the gas?

2. If an ideal gas is allowed to expand adiabatically, the internal energy of the gas changes by 2500 J.

a) Does the internal energy of the gas increase or decrease? Justify your answer.

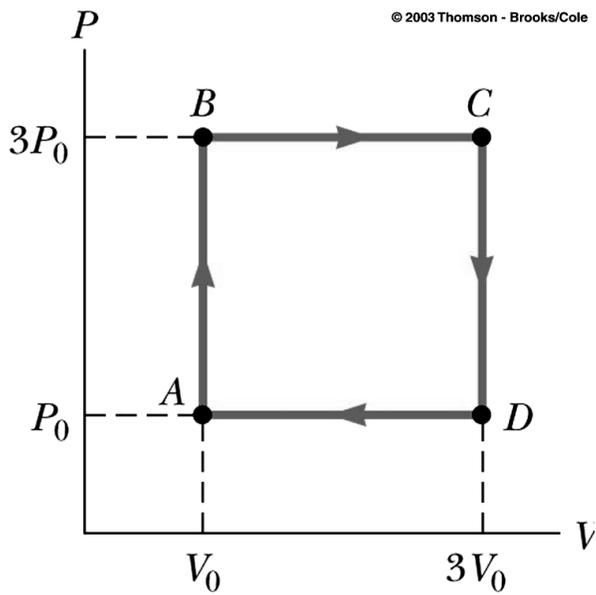
Determine:

b) the thermal energy added or removed from the gas.

c) the work done by the gas.

Cycle: a series of processes that returns a gas to its initial state

The cycle shown below represents processes performed on an ideal gas initially at $P_0 = 1.0 \times 10^5 \text{ Pa}$ and $V_0 = 2.0 \times 10^{-3} \text{ m}^3$.

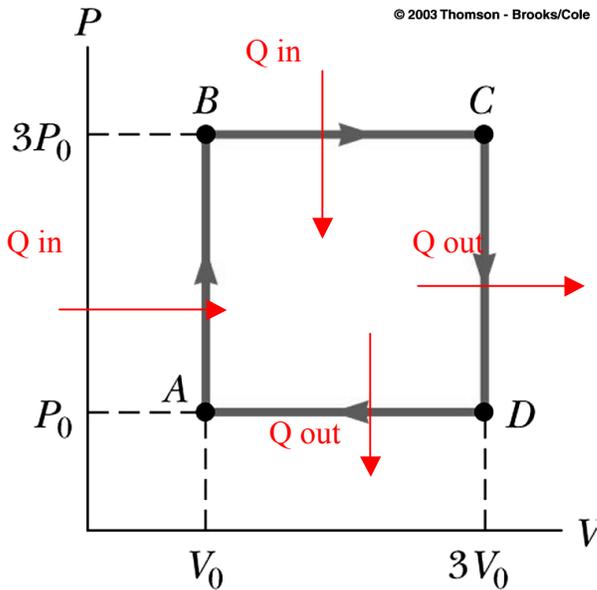


	Q	ΔU	W
A → B			
B → C			
C → D			
D → A			
Cycle			

1. Compare the temperatures at each state A, B, C, and D.

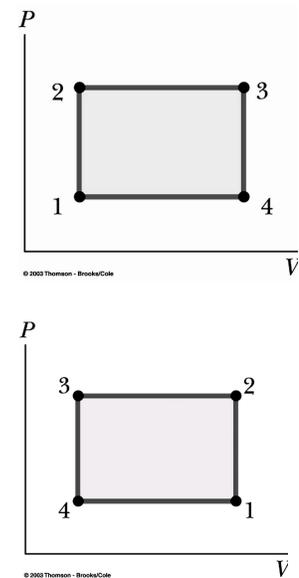
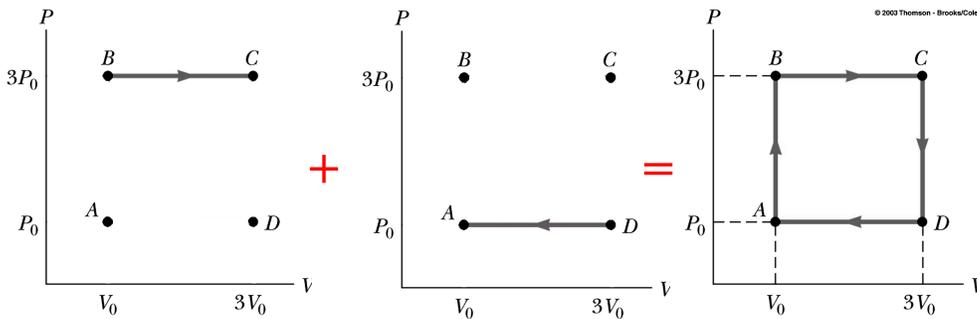
2. During process A→B, 600 J of thermal energy were added to the gas. Complete the chart.

Properties of the individual thermal processes



A → B	isochoric, temperature increase, U increase, W = 0, Q+ = Q in
B → C	isobaric expansion, +W, temperature increase, U increase, Q in (more than W by)
C → D	isochoric, temperature decrease, U decrease, W = 0, Q- = Q out
D → A	isobaric compression, -W, temperature decrease, U decrease, Q out (more than W on)

Net Work for a Cycle



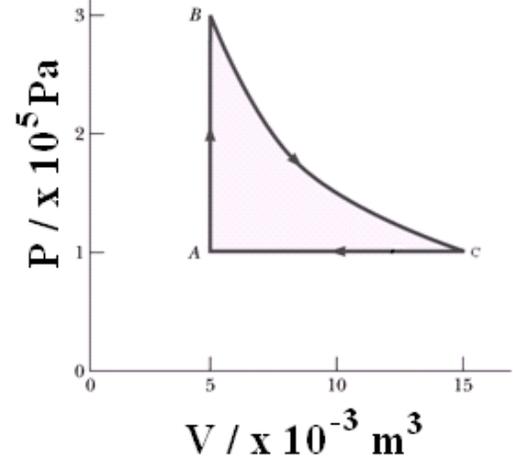
Properties of the entire cycle

- 1) gas returns to same P, V, and T
- 2) $\Delta T = 0$ so $\Delta U = 0$ (for all ideal gases)
- 3) $\Delta U = 0$ so net Q = netW
- 4) net W = area enclosed by figure so positive area enclosed means positive net work = work done by gas = net work out
- 5) net Q = W so Q+ so more heat added than removed during cycle = net heat in

An ideal gas is confined in a cylinder with a movable piston. The gas starts at 300 K in state A and proceeds through the cycle shown in the diagram

a) Find the temperatures at B and at C.

900 K isothermal



of the three processes and for the entire cycle.

	Q	ΔU	W
A → B			
B → C			
C → A			
Cycle			

A to B: $W = 0, Q +, \Delta U +$

B to C: $\Delta U = 0, W +$ and $Q +$

C to A: $Q -, W -, \Delta U -$

Cycle: $\Delta U = 0, Q +, W +$

A to B. 1700 J of heat are added to the gas for the entire cycle.

	Q	ΔU	W
A → B			
B → C			
C → A			
Cycle			

A to B: $Q = \Delta U = 1520 \text{ J}, W = 0$

B to C: $\Delta U = 0, Q = +1700 \text{ J}, W = +1700 \text{ J}$

C to A: $\Delta U = -1520 \text{ J}, W = -1000 \text{ J}, Q = -2520 \text{ J}$

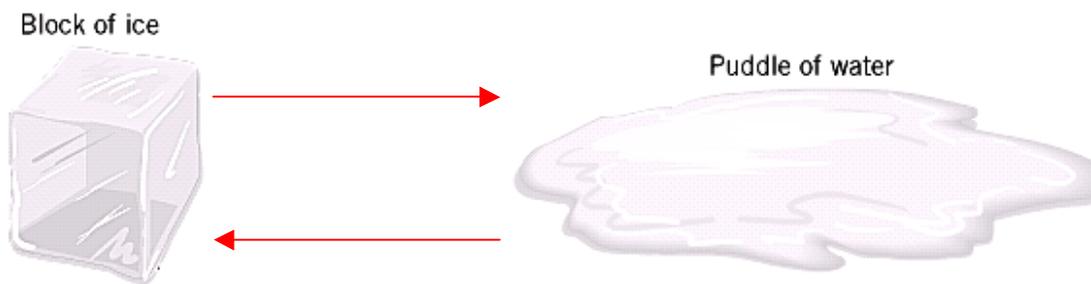
Cycle: $\Delta U = 0, Q = 700 \text{ J}, W = 700 \text{ J}$

The Second Law of Thermodynamics implies that . . . **thermal energy cannot spontaneously transfer from a region of low temperature to a region of high temperature.**

Entropy: a system property that expresses the degree of disorder in the system.

Second Law of Thermodynamics:

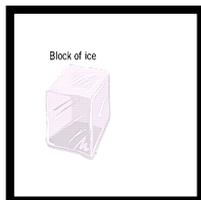
- 1) the overall entropy of the universe is increasing
- 2) all natural processes increase the entropy of the universe



Although local entropy can decrease, any process will increase the total entropy of a system *and* its surroundings (the universe).

1. Discuss this statement for the case of a puddle of water freezing into a block of ice.

2. A block of ice is placed in a thermally insulated room initially at room temperature. Discuss any changes in the total energy, total entropy, and temperature of the room.



3. An operating refrigerator with its door open is placed in a thermally insulated room. The refrigerator operates for a long period of time. Discuss any changes in the total energy, total entropy, and temperature of the room.

