AE1241 Physics
Chapter 19 Heat and the First Law of Thermodynamics

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What do we focus on?
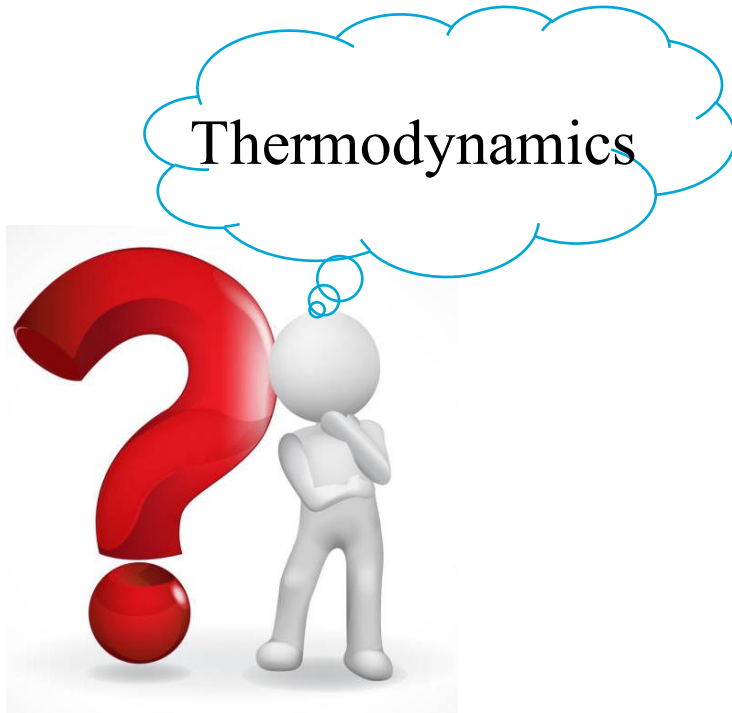
Thermodynamics

- ~~0th law of thermodynamics (temperature) – Chapter 17~~

- ~~Kinetic Theory of Gases (temperature) – Chapter 18~~

- **1st law of thermodynamics (energy) – Chapter 19**

- 2nd law of thermodynamics (entropy) – Chapter 20

- 3rd law of thermodynamics (absolute 0 K) – Chapter 20

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Structure of today's lecture

- ~~Heat, work and energy transfer~~

- ~~Specific Heat~~

- ~~Latent Heat~~

- ~~The 1st law of thermodynamics~~

- ~~Various thermodynamic processes: isothermal; adiabatic; isobaric and isovolumetric~~

- ~~Molar specific heat for gases and the equipartition of energy~~

- Heat Transfer: conduction and radiation

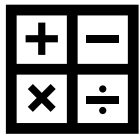**T**U Delft

Learning objective

After today's lecture, you should be able to:

- Understand the concept of **heat, work, internal energy, specific heat, molar specific heat, latent heat**

- Analyze different **thermodynamic processes,** using the 1$^{\text{st}}$ law of thermodynamics

- Calculate thermodynamic parameters, e.g., **heat, work, internal energy, temperature, pressure, volume, etc.** for given processes

- Calculate the **heat transfer rate** by conduction and radiation

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Prior Knowledge

- Chapter 18 of this lecture $\rightarrow$ gas kinetics

- Basic mechanics (e.g., *Newton's law*)

- Calculus (e.g., *differential and integral*)

- Statistics (e.g., *distribution, probability*)

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19-1 Energy Transfer

Two means of energy transfer: **Heat transfer** vs. **Work**

Heat is ***energy*** *transferred* from one object to another because of a ***temperature difference***.

cal (calorie)	J (Joule)	BTU (British Thermal Unit)
Heat is required to raise the temperature of 1 g of water by 1 °C	1 cal = 4.186 J	Heat is required to raise the temperature of 1 lb of water by 1 °F
14.5°C→15.5°C		1 BTU = 252 cal =1055J

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Exercise 1: How many joules and calories are generated when braking a 1500 kg car from 120 km/h to rest?

120 km/h = 33.33 m/s

$$E_k = \frac{1}{2} \, m \, \Delta v^2 = \frac{1}{2} \times 1500 kg \times (33.33 \, m/s)^2 \approx 833333 \, J$$

Total calories: $E_k / (4.186 \, J/cal) = 199076 \, cal$

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Different concepts

- **Temperature**: a measure of average kinetic energy of individual molecules

$$\bar{K} = \frac{3}{2}kT = \frac{1}{2}m\overline{v^2}$$

- **Internal energy**: total energy of all the molecules within the object

- **Heat**: energy transfer from one object to another due to temperature difference

- **Work**: energy transfer not due to temperature differences

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https://www.andrew.cmu.edu/course/09-105/RotationalSpectra.html
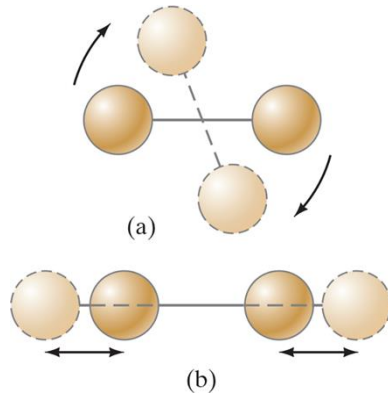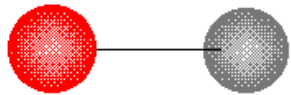
Degree of freedom

$$i_{total} = i_{trans} + i_{rotation} + i_{vibration}$$

Monatomic gas:
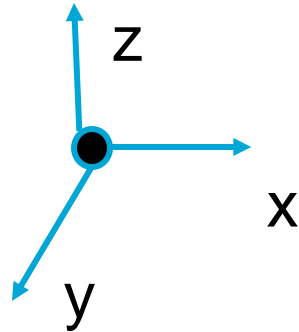
3 translational degrees of freedom

Diatomic gas:

3 translational degrees of freedom
2 rotational degrees of freedom
1 vibrational degree of freedoms

Total: 6 degrees of freedom

More complex gas: 3N with N equals to the total number of atoms
→ 3 translational; 3 rotational, and the rest is vibrational

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19-2 Internal energy

Energy of a molecule is equally shared by each degree of freedom:

$$\frac{1}{2}kT$$

Everage energy per molecule is:

$$\bar{e} = \frac{i_{total}}{2}kT$$

Internal energy of all number of molecules (N):

$$E_{int} = N\frac{i_{total}}{2}kT$$

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19-2 Internal energy of monatomic gas

$$E_{int} = N\left(\frac{1}{2}m\overline{v^2}\right) = \frac{3}{2}NkT = \frac{3}{2}nRT$$

$$\overset{E_{int}}{= \frac{3}{2}NkT = \frac{3}{2}nRT}$$

T: temperature in K

N: number of molecules

k: Boltzmann constant, $1.38 * 10^{-23}$ J/K

R: universal gas constant, 8.314 J/mol/K

n: mole of gas, mole

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19-3 Specific Heat

The mount of **heat** (Q) required to change the temperature of a **given material** is proportional to the **mass** (m) of the material and the **temperature change** (ΔT in °C).

$$Q = mc\Delta T$$

Where, specific heat c (J/kg·°C) is a quantity characteristic of the material.
The value c depends on T (and P), but often considered constant

![TU Delft logo]

Substance	Specific heat	
	$\dfrac{J}{kg \cdot °C}$	$\dfrac{kcal}{kg \cdot °C}$ $(\dfrac{cal}{g \cdot °C})$
Aluminum	900	0.22
Alcohol (ethyl)	2400	0.58
Copper	390	0.093
Water		
Ice (-5°C)	2100	0.5
Liquid (15 °C)	4186	1.00
Steam (110 °C)	2010	0.48
Human Body (average)	3470	0.83

Reproduced from Table 19-1 Specific Heats (at 1 atm constant pressure and 20 °C unless other wise stated)

Heat capacity vs specific heat capacity
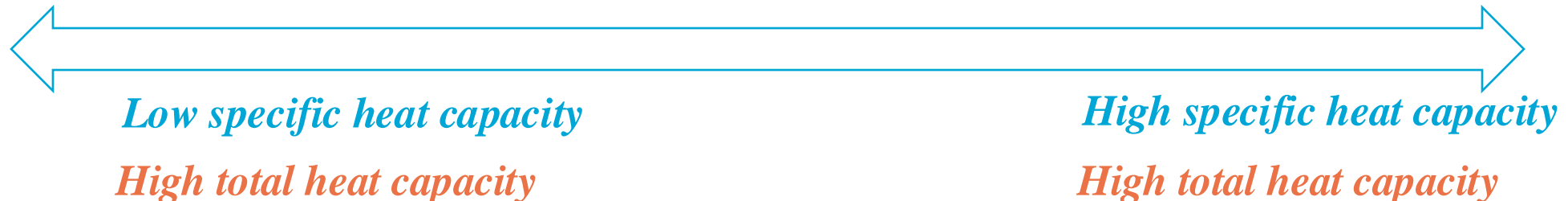
Copper 2 kg

Specific heat:
390 J/kg/℃

Aluminum 0.5 kg

Specific heat:
900 J/kg/℃

Water 0.1 kg

Specific heat:
4186 J/kg/℃

Low specific heat capacity

High specific heat capacity

High total heat capacity

High total heat capacity

Heat capacity depends on specific heat and the mass of the substance.

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Exercise 2: specific heat

Samples of copper, aluminum, and water experience the same temperature increase when they absorb the same amount of heat. What is their mass ratio?

$$Q_{copper} = Q_{aluminum} = Q_{water}$$

$$Q_{copper} = m_{copper} c_{copper} \Delta T$$

$$Q_{aluminum} = m_{aluminum} c_{aluminum} \Delta T$$

$$Q_{water} = m_{water} c_{water} \Delta T$$

$$\frac{m_{aluminum}}{m_{copper}} = \frac{c_{copper}}{c_{aluminum}} = 0.43$$

$$\frac{m_{water}}{m_{copper}} = \frac{c_{copper}}{c_{water}} = 0.093$$

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19-4 System definition

Surroundings

System

Boundary

Energy in/out

Closed System

No mass in/out

Energy in/out

Open System

Mass in/out

No energy & mass in/out

Isolated System

- **Isolated system**: Heat lost by one part of an isolated system = heat gained by another part of an isolated system (i.e. **conservation of energy**)
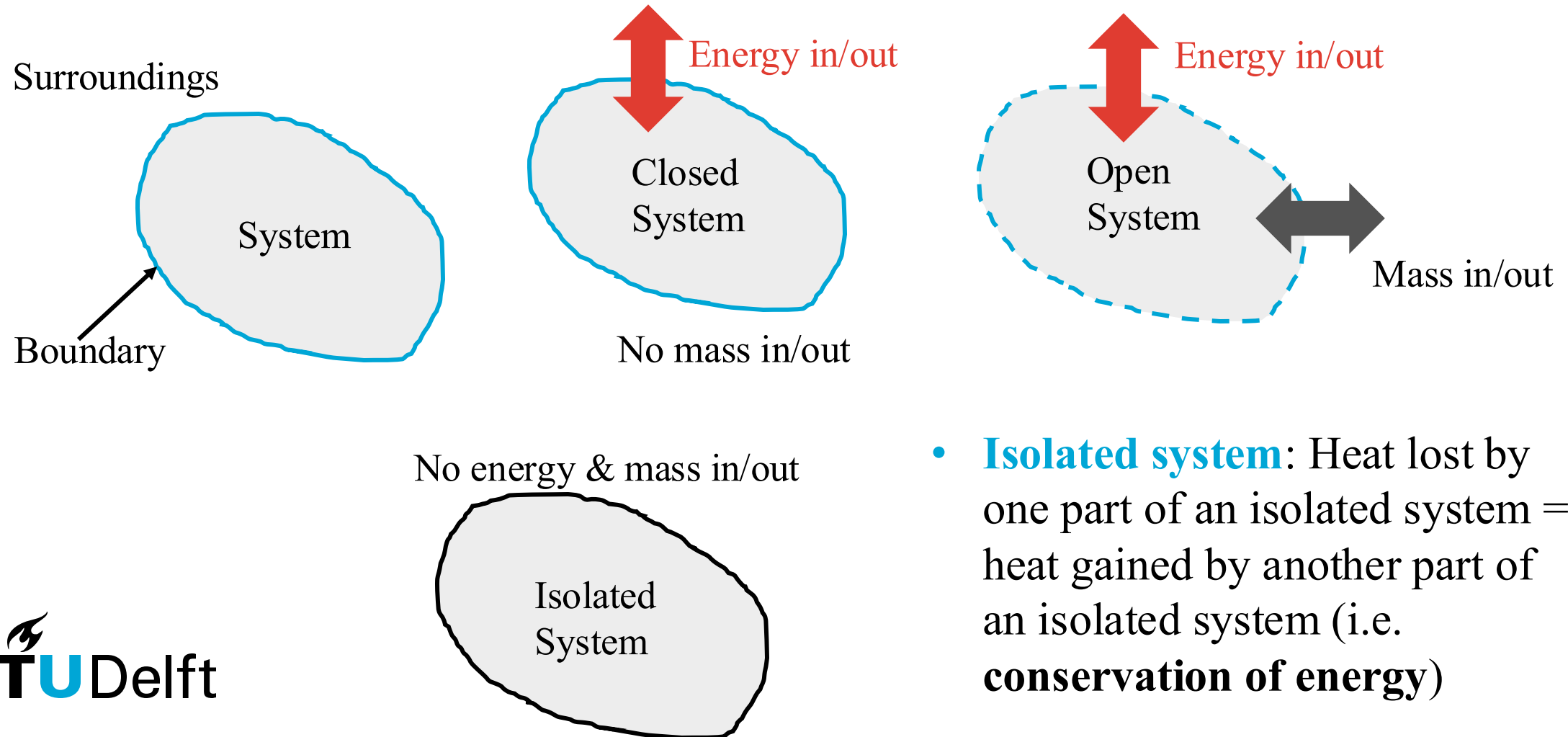
19-4 Calorimetry − determining specific heat

- <u>Calorimetry</u>: quantitative measurement of heat exchange

- <u>Calorimeter</u>: determination of specific heat by the "method of mixtures"

(a sample of a substance is heated to a known temperature → quickly placed in the cool water of calorimeter (temperature known) →the final temperature of the mixture is measured → specific heat of the substance is calculated)

Figure 19-4 simple water calorimeter

$$m_{sub}c_{sub}\Delta T_{sub} = m_{H_2O}c_{H_2O}\Delta T_{H_2O} + m_{cup}c_{cup}\Delta T_{cup}$$
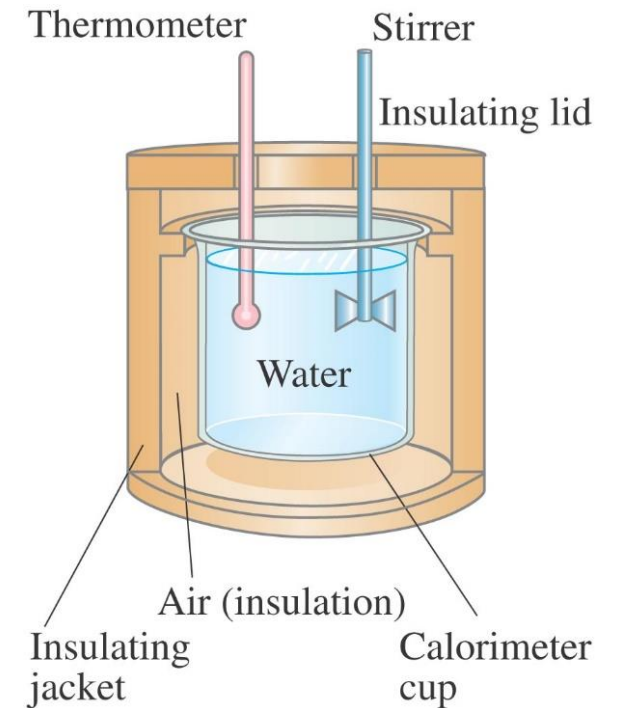
Unknown quantity (specific heat of substance)

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Exercise 3: unknown specific heat determined by calorimetry

When a 290-gram piece of iron at 180 °C is placed in a 95-gram aluminium calorimeter cup containing 250 grams of glycerin at 10 °C. The final temperature is observed to be 38 °C. Estimate the specific heat of glycerin.

$$Q = mc\Delta T$$

$$m_{iron}c_{iron}\left(T_{iron,i} - T_{iron,f}\right) = m_{gly}c_{gly}\left(T_{gly,f} - T_{gly,i}\right) + m_{cup}c_{cup}(T_{cup,f} - T_{cup,i})$$
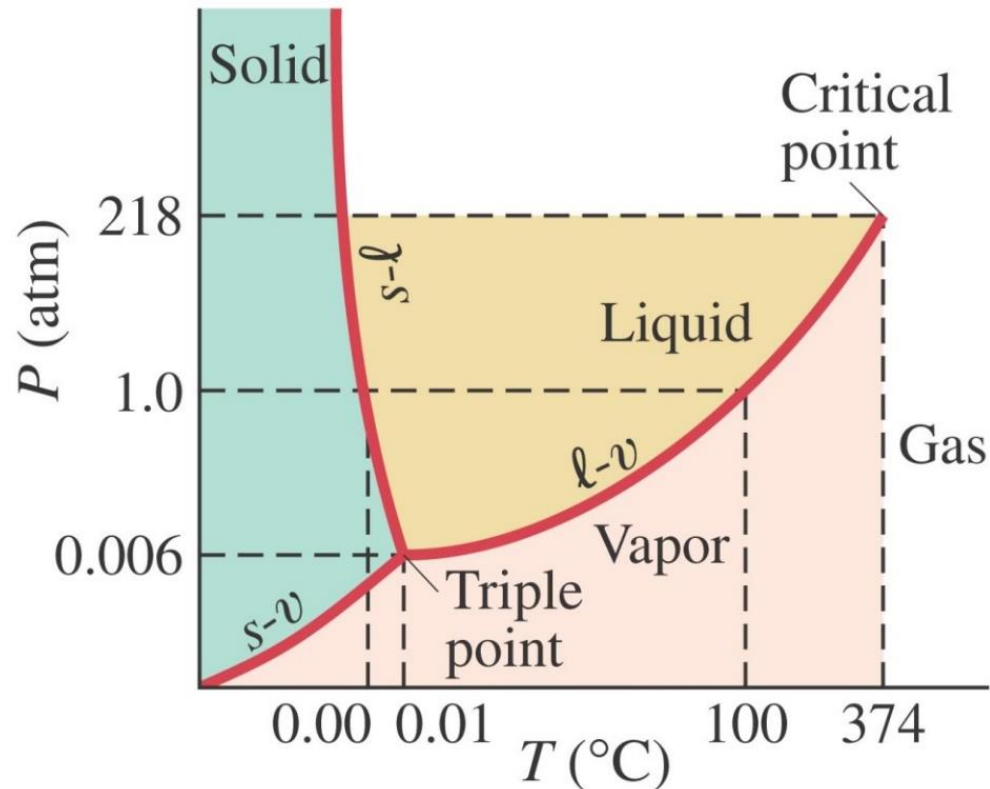
$$(0.29kg)\left(450\frac{J}{kg \cdot °C}\right)(180°C - 38°C)$$

$$= (0.25kg)\left(c_{gly}\right)(38°C - 10°C) + (0.095kg)\left(900\frac{J}{kg \cdot °C}\right)(38°C - 10°C)$$

$$c_{gly} = 2305\frac{J}{kg \cdot °C}$$

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19-5 Latent Heat

The amount of energy involved during **change of phases**

Step 1: 1 kg of ice at -40 °C is heated to 0 °C

Step 2: Continue heating until all ice melts

Step 3: all liquid water is heated up to 100 °C

Step 4: Continue heating to evaporate

Step 5: all liquid water is changed into steam

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19-5 Latent Heat
---Change of phase

Specific heat		
Substance	$\dfrac{J}{kg \cdot °C}$	$\dfrac{kcal}{kg \cdot °C}$ $\left(\dfrac{cal}{g \cdot °C}\right)$
Ice (-5°C)	2100	0.5
Liquid (15 °C)	4186	1.00
Steam (110 °C)	2010	0.48

$c_{\text{ice}} = 0.50$ kcal/kg·°C

Latent heat

Heat of fusion

$L_F = 79.7$ kcal/kg

$c_{\text{water}} = 1.00$ kcal/kg·°C

Heat of vaporization

$L_v = 539$ kcal/kg

$c_{\text{steam}} = 0.48$ kcal/kg·°C
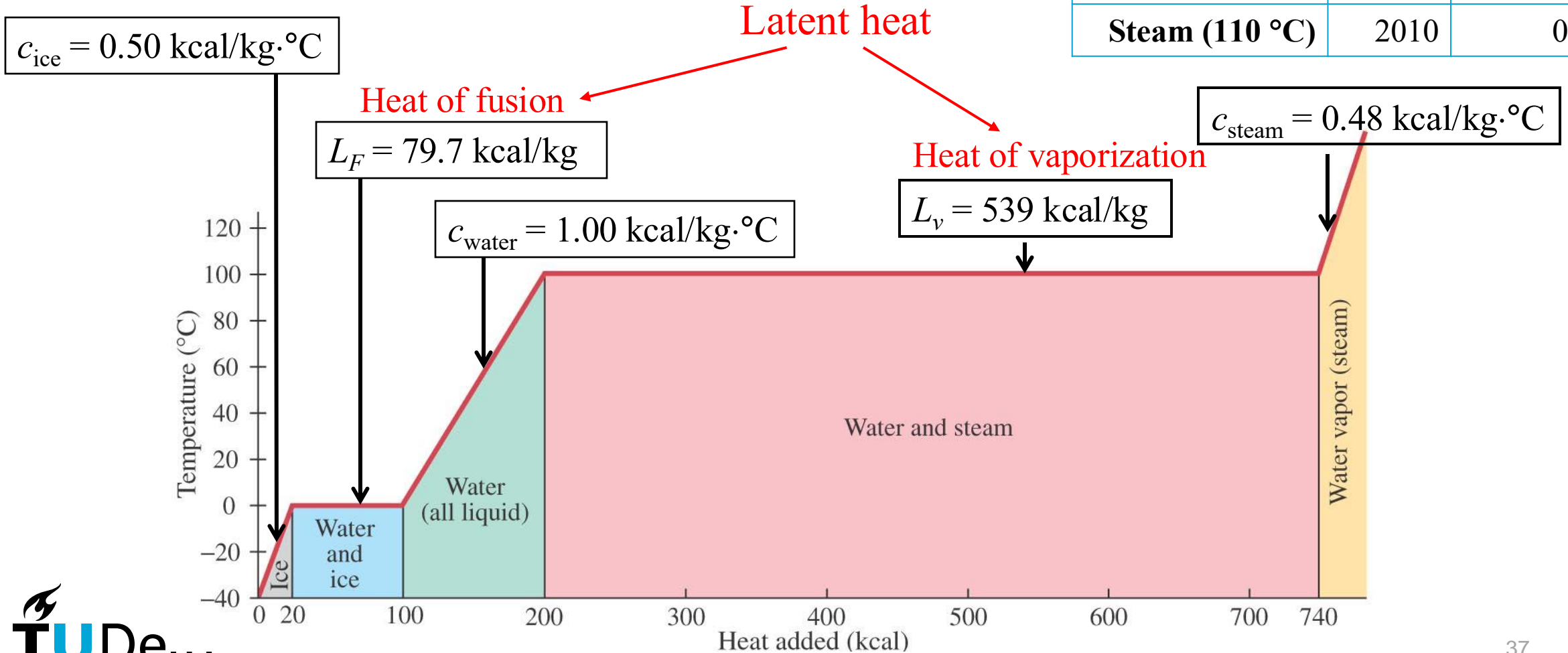
Figure 19-5 Temperature as a function of the heat added to bring 1 kg of ice at -40 °C to steam above 100 °C.

19-5 Latent Heat

- Heat of fusion: the mount of heat released by a substance when changing from a **liquid to a solid**;

- Heat of vaporization: the mount of heat released by a substance when changing from a **gas to a liquid**.

$$Q = mL$$

2260 kJ/kg

Heat release

- *L:* the latent heat, J/kg;
- *m:* is the mass of the substance, kg;
- Q: is the heat added or released during the phase change, J.
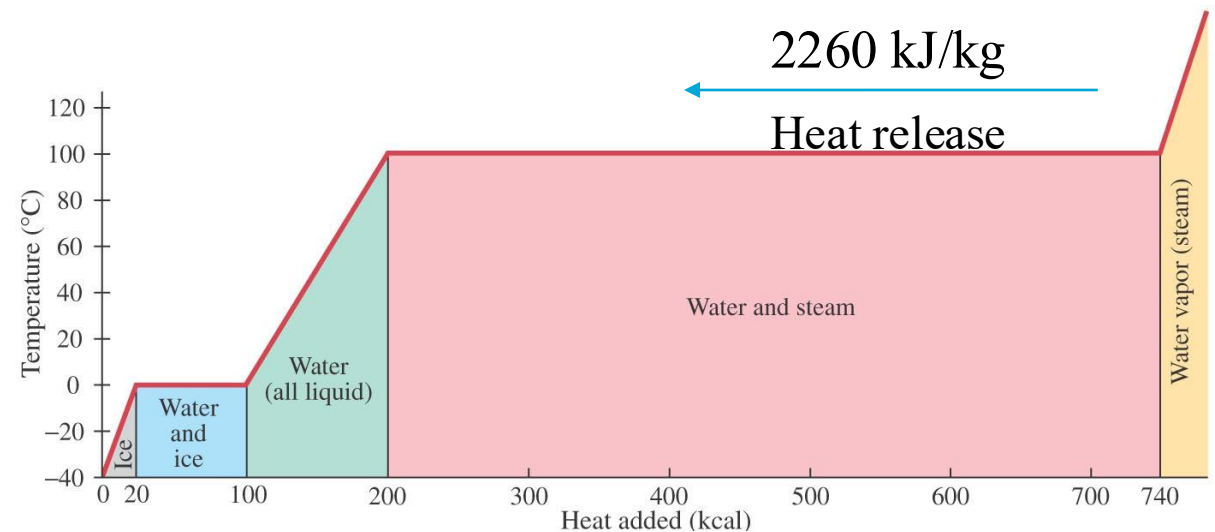
19-5 Latent Heats & Kinetic Theory of Latent Heats

TABLE 19–2 Latent Heats (at 1 atm)

Substance	Melting Point (°C)	Heat of Fusion		Boiling Point (°C)	Heat of Vaporization	
		kcal/kg†	kJ/kg		kcal/kg†	kJ/kg
Oxygen	−218.8	3.3	14	−183	51	210
Nitrogen	−210.0	6.1	26	−195.8	48	200
Ethyl alcohol	−114	25	104	78	204	850
Ammonia	−77.8	8.0	33	−33.4	33	137
Water	0	79.7	333	100	539	2260
Lead	327	5.9	25	1750	208	870
Silver	961	21	88	2193	558	2300
Iron	1808	69.1	289	3023	1520	6340
Tungsten	3410	44	184	5900	1150	4800

†Numerical values in kcal/kg are the same in cal/g.

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- *Note*: latent heat of fusion = *not* to increase the average kinetic energy (temperature) of molecules, but to overcome potential energy associated with forces between molecules

- Similar for the latent heat of vaporization, but this is a more 'violent' process than melting as the distance between molecules is greatly increased.

In general:

$$L_v \gg L_F$$
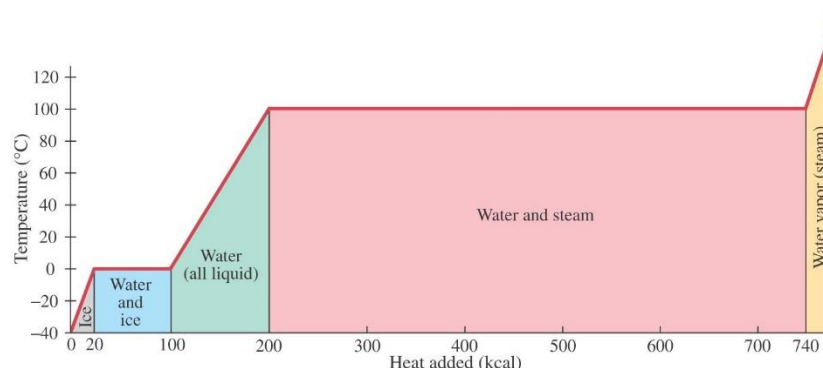
Exercise 4: boiling water

An iron boiler of mass 180 kg contains 710 kg of water at 18 °C. A heater supplies energy at the rate of 58000 kJ/h. How long does it take for water (a) to reach the boiling point, and (b) to all have changed to steam?

$$Q_{supply} = 58000 \; kJ/h$$

$$Q_{water} = m_{water} c_{water} (100°C - 18°C) = (710kg)\left(4.186 \frac{kJ}{kg \cdot °C}\right)(82°C) = 243708.9 \; kJ$$

$$Q_{boiler} = m_{boiler} c_{boiler} (100°C - 18°C) = (180kg)\left(0.45 \frac{kJ}{kg \cdot °C}\right)(82°C) = 6642 \; kJ$$

$$t = \frac{Q_{total}}{Q_{supply}} \approx 4.3 \; h$$

$$Q_{steam} = m_{water} c_L = (710kg)\left(2260 \frac{kJ}{kg \cdot °C}\right) = 1604600 \; kJ \qquad t = \sim 27.6 \; h$$

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Exercise 5: Determining a latent heat of fusion

Determine the latent heat of fusion of mercury using the following calorimeter data: 1.0 kg of solid Hg at its melting point of -39.0 °C is placed in a 0.62-kg aluminum calorimeter with 0.4 kg of water at 12.8 °C . The resulting equilibrium temperature is 5.06 °C. The specific heat of liquid mercury is 140 J/kg/°C.

$$Q(melt\ of\ Hg) = m_{Hg}L_{Hg}$$

$$Q(Liquid\ of\ Hg) = m_{Hg}c_{Hg}(T_{Hg,f} - T_{Hg,i}) = (1.0\ kg)(140\frac{J}{kg°C})(5.06°C - (-39°C)) = 6168.4J$$

$$Q(aluminium) = m_{Al}c_{Al}(T_{Al,f} - T_{Al,i}) = (0.62\ kg)(900\frac{J}{kg°C})(12.8°C - 5.06°C) = 4318.9J$$

$$Q(water) = m_{W}c_{W}(T_{W,f} - T_{W,i}) = (0.4\ kg)(4186\frac{J}{kg°C})(12.8°C - 5.06°C) = 12959.9J$$

$$Q(water) + Q(aluminium) = Q(liquid\ Hg) + Q(melt\ Hg) \qquad L_{Hg} = {\sim}11\ kJ/kg$$

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19-6 The First Law of Thermodynamics

- Study of processes in which energy is *transferred* as <u>heat</u> and as <u>work</u>

transfer of energy due to a temperature difference.

W

Work done
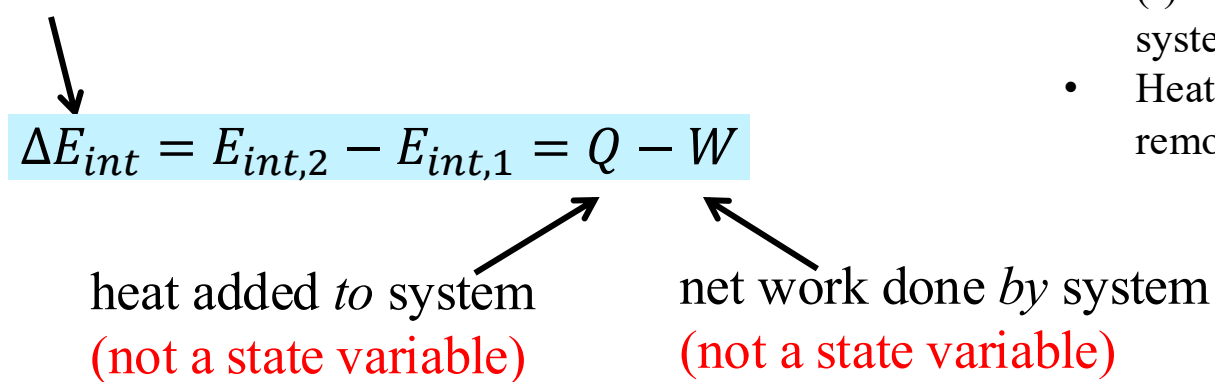
Q

Heat transfer
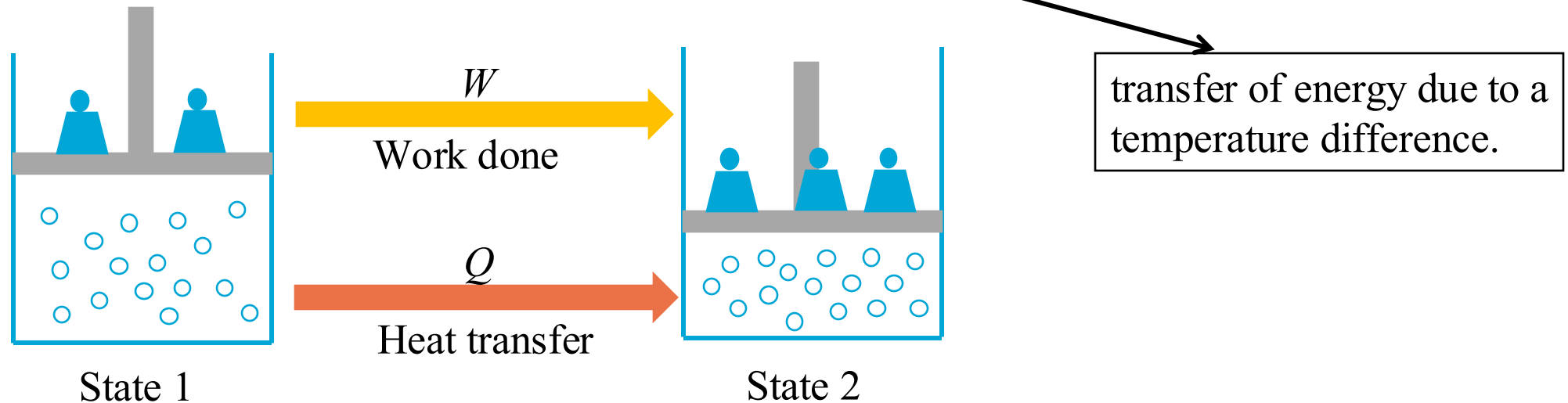
State 1

State 2

- The change in **internal energy (E_{int})** **between any two equilibrium states** is equal to the **energy added to the system by heating** and the **work done by the system on surroundings**.

$$\Delta E_{int} = E_{int,2} - E_{int,1} = Q - W$$ **(Law of energy conservation)**

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19-6 The First Law of Thermodynamics

change in internal energy (is a state variable, like P, V, T, m, n)

$$\Delta E_{int} = E_{int,2} - E_{int,1} = Q - W$$

heat added *to* system
(not a state variable)

net work done *by* system
(not a state variable)

in differential form:
$$dE_{int} = dQ - dW$$

isolated system:
$$W = Q = 0 \implies \Delta E_{int} = 0$$

For a moving system with kinetic energy (K) and suppose there is potential energy (U):

$$\Delta K + \Delta U + \Delta E_{int} = Q - W$$

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19-7 Thermodynamic Processes and the First Law
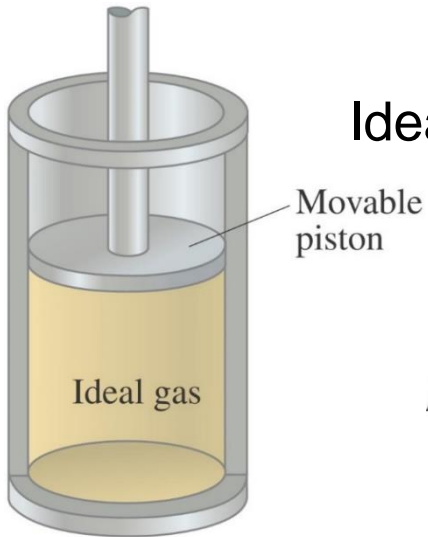## ---isothermal processes($\Delta T=0$)

Ideal gas: $PV = constant$

Movable piston

Ideal gas

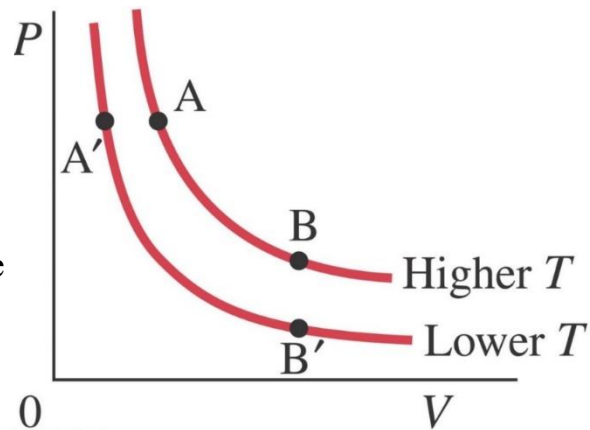Figure 19-6 An ideal gas in a cylinder fitted with a movable piston.

Figure 19-7 PV diagram for ideal gas undergoing isothermal quasi-static process at two different temperatures.

A→B, if an amount of heat Q is added to the gas (at constant T), then gas expands and does an amount of work W on the environment.

$$\Delta E_{int} = \frac{3}{2}nR\Delta T = 0 \longrightarrow W = Q$$

The work done by a gas in an isothermal process equals the heat added to the gas.

Assumptions for idealized approximation to a real process:

- Gas in contact with **heat reservoir** (= body mass so large that its temperature does not change when heat is exchanged with our system)

- Compression/expansion is done **quasi-statically** (i.e. slowly such that gas moves between a series of *equilibrium states*, all at same temperature)

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19-7 Thermodynamic Processes and the First Law
---work done in volume changes

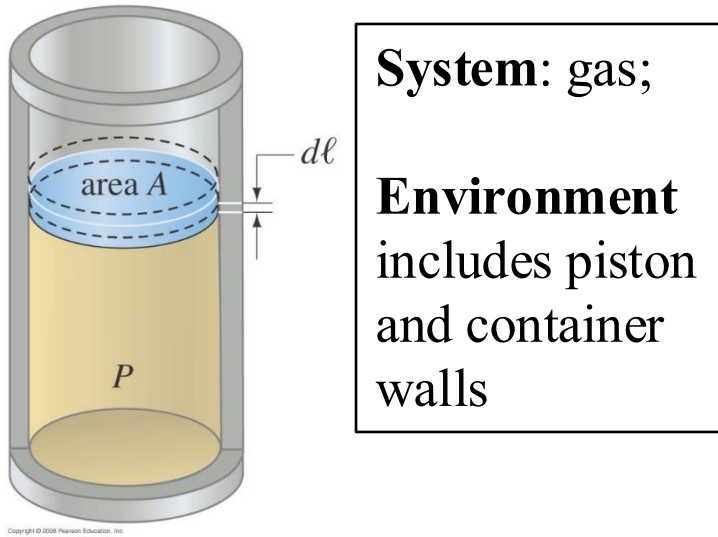| System: gas; |
| Environment includes piston and container walls |

Figure 19-10 The work done by a gas when its volume increases.

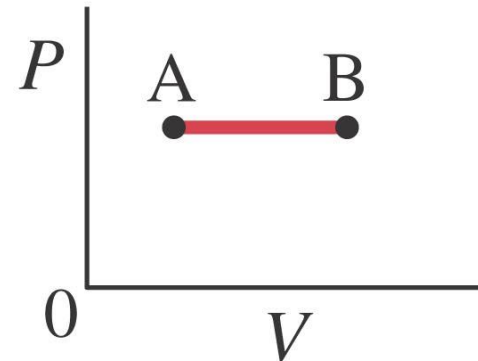Work done by the gas when it expands *quasi-static*:

$$dW = PAdl = PdV$$

If gas compressed then $dV < 0$ and $W < 0$ (i.e. *positive* work done *on* the gas)
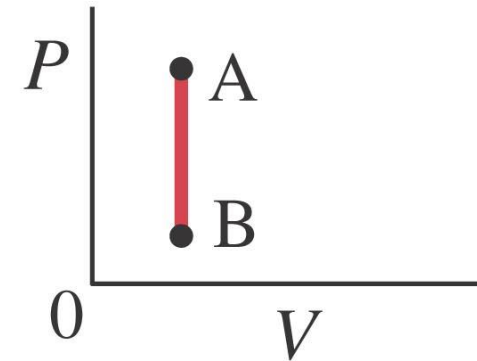
For a finite change in volume we have:

$$W = \int_{V_A}^{V_B} PdV$$

19-7 Thermodynamic Processes and the First Law
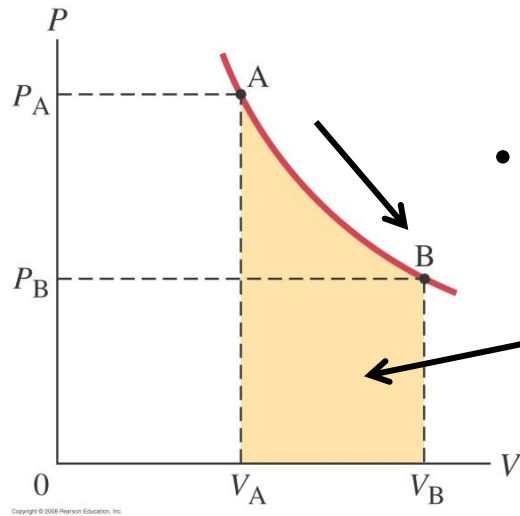## ---isobaric and isovolumetric processes

Figure 19-9 (a) Isobaric ("same pressure") process. (b) Isovolumetric (isochoric, "same volume") process.

19-7 Thermodynamic Processes and the First Law
---work done in volume changes

Figure 19-11 work done by an ideal gas in an isothermal process equals the area under the PV curve.

- Quasi-static isothermal expansion of *ideal gas*

$$W_{AB} = \int_{V_A}^{V_B} P \, dV = \int_{V_A}^{V_B} \frac{nRT}{V} \, dV = nRT \ln\left(\frac{V_B}{V_A}\right)$$

Isovolumetric: $W_{AD} = 0 \ (dV = 0)$

Isobaric: $W_{DB} = P_B(V_B - V_A) = nRT_B\left(1 - \frac{V_A}{V_B}\right)$
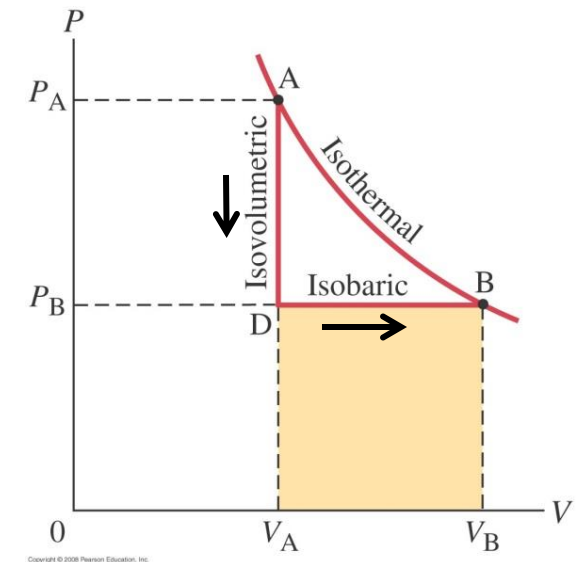
Figure 19-12 Process ADV consists of an isovolumetric (AD) and an isobaric process (DB).

Work depends not only on initial and final states, but also on type of processes, i.e. *the path;* the same holds for heat (*W* and *Q* are not *state variables*)

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19-7 Thermodynamic Processes and the First Law
---adiabatic processes ($Q=0$)

If a system is *extremely well insulated* or process happens very quickly (heat flows slowly):

$$Q = 0: \ \Delta E_{int} = -W$$

$$\Delta E_{int} = \frac{3}{2}nR\Delta T$$

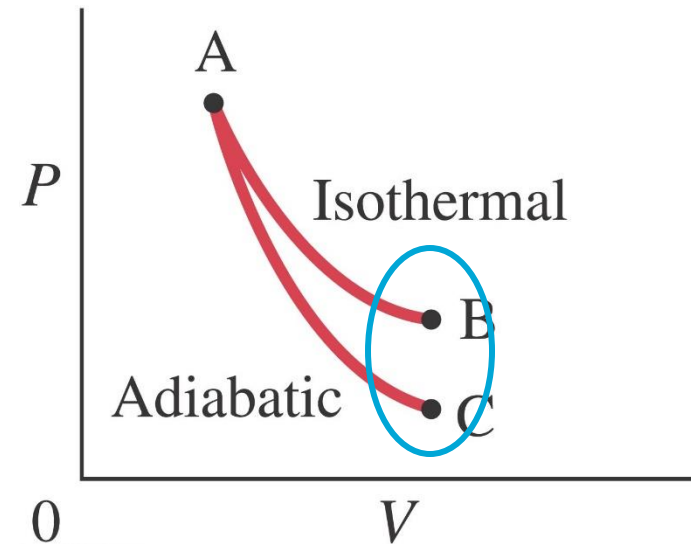When gas expands:
- Internal energy decrease
- T decrease

Figure 19-8 PV diagram for adiabatic (AC) and isothermal (AB) processes on an ideal gas.

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19-7 Thermodynamic Processes and the First Law
---free expansion

When the valve is opened, *no work* is done because the gas does not move any other object

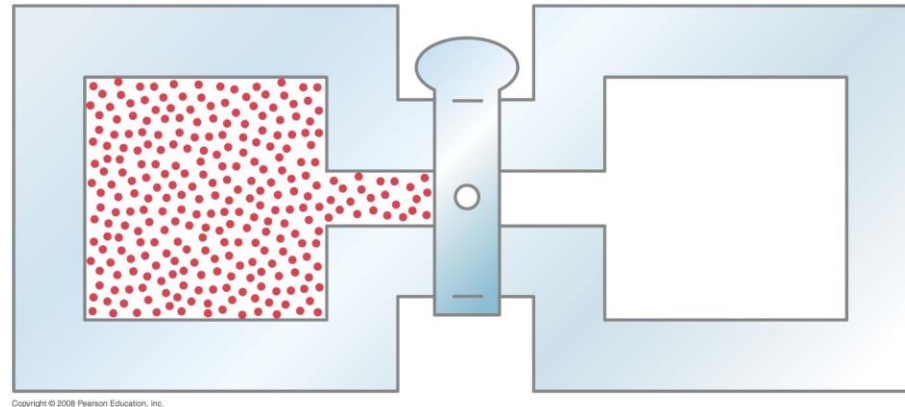Copyright © 2008 Pearson Education, Inc.

Figure 19-14 Free expansion.

$$Q = W = 0 \longrightarrow \Delta E_{int} = 0 \xrightarrow{\text{ideal gas}} \Delta T = 0$$

Notes:
- Free expansion has been used to determine internal energy of real gases
- The free expansion process cannot be plotted on PV diagram, because the process is rapid, not quasi-static

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19-7 Thermodynamic Processes and the First Law
---Summary

TABLE 19–3 Simple Thermodynamic Processes and the First Law

Process	What is constant:	The first law predicts:
Isothermal	T = constant	$\Delta T = 0$ makes $\Delta E_{int} = 0$, so $Q = W$
Isobaric	P = constant	$Q = \Delta E_{int} + W = \Delta E_{int} + P\,\Delta V$
Isovolumetric	V = constant	$\Delta V = 0$ makes $W = 0$, so $Q = \Delta E_{int}$
Adiabatic	$Q = 0$	$\Delta E_{int} = -W$

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Exercise 6: using the first law

An ideal gas expands isothermally, performing 4.3×10^3 J of work in the process. Calculate (a) the change in internal energy of the gas, and (b) the heat absorbed during the expansion

$$\Delta E_{int} = Q - W$$

$$T = constant \qquad \Delta E_{int} = 0$$

$$Q = W = 4.3 \times 10^3 J$$

Exercise 7: first law application

Two liter of an ideal gas is cooled at constant pressure (1.01x10^5 Pa) and room temperature (20°C) until its volume is halved, and then it is allowed to expand isothermally back to its original volume. (a) draw the process on PV diagram; (b) calculate the work done by the gas during the process

Compression: $W_{1\text{-}2}$ = (1.01x10^5 Pa)(-1.0x10^{-3} m^3) = -101 J

Expansion: $W_{2-3} = nRT_2 \ln(V_3/V_2) = P_2 V_2 \ln 2 = 70J$

$$W_{done} = W_{1-2} + W_{2-3} = -41\,J$$

$$T_2 = \frac{V_2}{V_1} T_1 = 10°C \qquad n = \frac{PV}{RT} = 0.08\,mol$$

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19-8 Molar Specific Heats for Gases

- Specific heat of gases depend on how the process is carried out

$$Q = nC_V\Delta T \quad \text{(volume constant)} \qquad C = Mc$$

$$Q = nC_P\Delta T \quad \text{(pressure constant)}$$

C_V and C_P for gases are quite different!

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19-8 Molar Specific Heats for Gases

At constant volume first law gives (no work, $\Delta V = 0$)

$Q_V = \boxed{\Delta E_{int}}$ $\longleftarrow$ increase ΔT is the same

At constant pressure first law gives (work is done)

$Q_P = \boxed{\Delta E_{int}} + P\Delta V > Q_V$

$\left.\begin{array}{c} \\ \\ \\ \end{array}\right\}$ $Q_P - Q_V = P\Delta V$

Combine with <u>ideal gas law</u> (constant P)

$$V = \frac{nRT}{P} \longrightarrow \Delta V = \frac{nR\Delta T}{P}$$

we obtain

$$nC_P\Delta T - nC_V\Delta T = P\frac{nR\Delta T}{P}$$

or $\boxed{C_P - C_V = R}$ $(= 2.0 \text{ cal/(mol·°C)}$
$= 8.3 \text{ J/(mol·°C))}$

19-8 Molar Specific Heats for Gases

- Specific heat of gases depend on how the process is carried out

$$Q = nC_V\Delta T \quad \text{(volume constant)} \qquad C = Mc$$

$$Q = nC_P\Delta T \quad \text{(pressure constant)}$$

C_V and C_P for gases are quite different!

TABLE 19–4 Specific Heats of Gases at 15°C

M (g/mol)	Gas	Specific heats (kcal/kg · K)		Molar specific heats (cal/mol · K)		$C_P - C_V$ (cal/mol · K)	$\gamma = \dfrac{C_P}{C_V}$
		c_V	c_P	C_V	C_P		
	Monatomic						
4	He	0.75	1.15	2.98	4.97	1.99	1.67
20	Ne	0.148	0.246	2.98	4.97	1.99	1.67
	Diatomic						
28	N$_2$	0.177	0.248	4.96	6.95	1.99	1.40
32	O$_2$	0.155	0.218	5.03	7.03	2.00	1.40
	Triatomic						
44	CO$_2$	0.153	0.199	6.80	8.83	2.03	1.30
18	H$_2$O (100°C)	0.350	0.482	6.20	8.20	2.00	1.32

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19-8 Molar Specific Heats for Gases

Monatomic ideal gas at constant V (first law)

$$Q_V = \Delta E_{int}$$

(Section 19-2)

$$nC_V\Delta T = \frac{3}{2}nR\Delta T$$

$$C_V = \frac{3}{2}R \qquad (= 3.0 \text{ cal/(mol·°C)} = 12.5 \text{ J /(mol·°C)}$$

$$C_P = \frac{5}{2}R \qquad (= 5.0 \text{ cal/(mol·°C)} = 20.8 \text{ J /(mol·°C)}$$

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19-8 Molar Specific Heats for Gases
--- Equipartition of energy

Figure 19-15

Ideal gas: $C_P - C_V = R$

| - energy is shared equally among the *active* degrees of freedom |
| - each active degree of freedom has an energy equal to ½ kT |

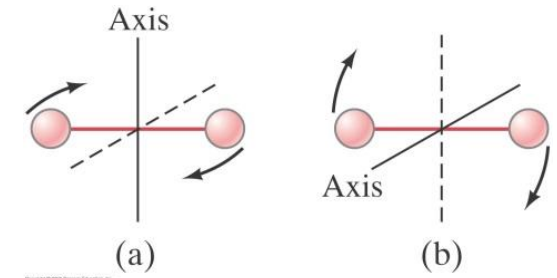Monatomic (ideal) gas:
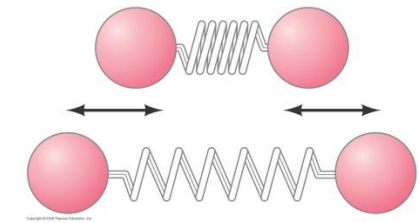
$$C_V = \frac{3}{2} R$$

Figure 19-16

Diatomic (ideal) gas according to 'equipartition of energy':

$$C_V = (3 + 2)\frac{1}{2} R$$

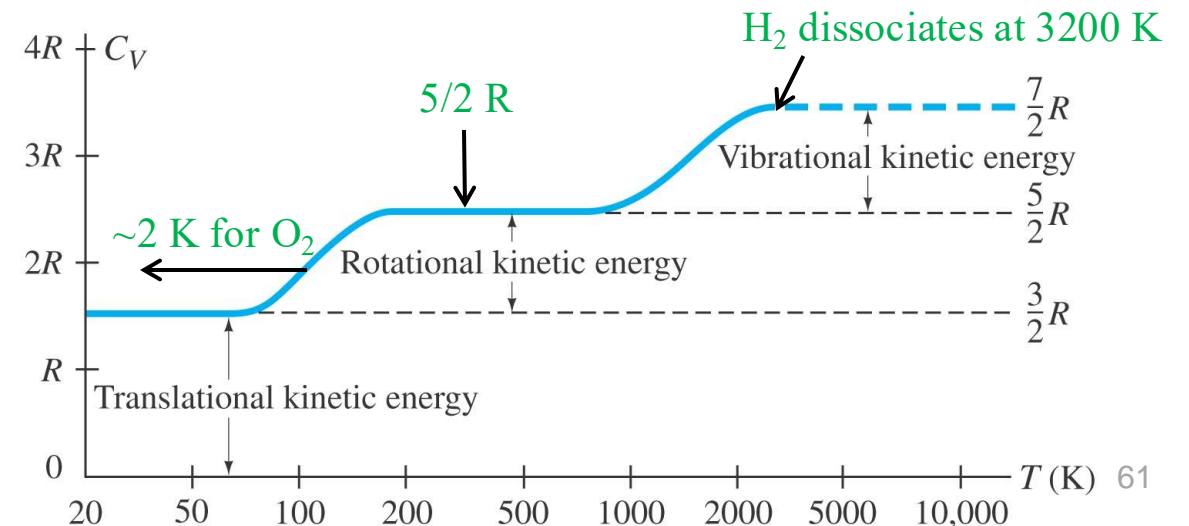$$C_V = \frac{7}{2} R$$ at high temperatures

For H$_2$ molecules, temperature axis is logarithmic scale.

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19-8 Molar Specific Heats for Gases

- Specific heat of gases depend on how the process is carried out

$$Q = nC_V\Delta T \quad \text{(volume constant)} \qquad C = Mc$$

$$Q = nC_P\Delta T \quad \text{(pressure constant)}$$

C_V and C_P for gases are quite different!

TABLE 19–4 Specific Heats of Gases at 15°C

M (g/mol)	Gas	Specific heats (kcal/kg·K)		Molar specific heats (cal/mol·K)		$C_P - C_V$ (cal/mol·K)	$\gamma = \dfrac{C_P}{C_V}$
		c_V	c_P	C_V	C_P		
	Monatomic			3R/2	5R/2	R	
4	He	0.75	1.15	2.98	4.97	1.99	1.67
20	Ne	0.148	0.246	2.98	4.97	1.99	1.67
	Diatomic						
28	N$_2$	0.177	0.248	4.96	6.95	1.99	1.40
32	O$_2$	0.155	0.218	5.03	7.03	2.00	1.40
	Triatomic			5R/2	7R/2		
44	CO$_2$	0.153	0.199	6.80	8.83	2.03	1.30
18	H$_2$O (100°C)	0.350	0.482	6.20	8.20	2.00	1.32

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19-8 Molecular Specific Heats ---solids

Dulong-Petit law:

$$C_V = (3 + 3)\frac{1}{2}R = 3R$$

= 6.0 cal/(mol.°C) = 24.9 J/(mol.°C) at high temperatures
(see also section 19-3)

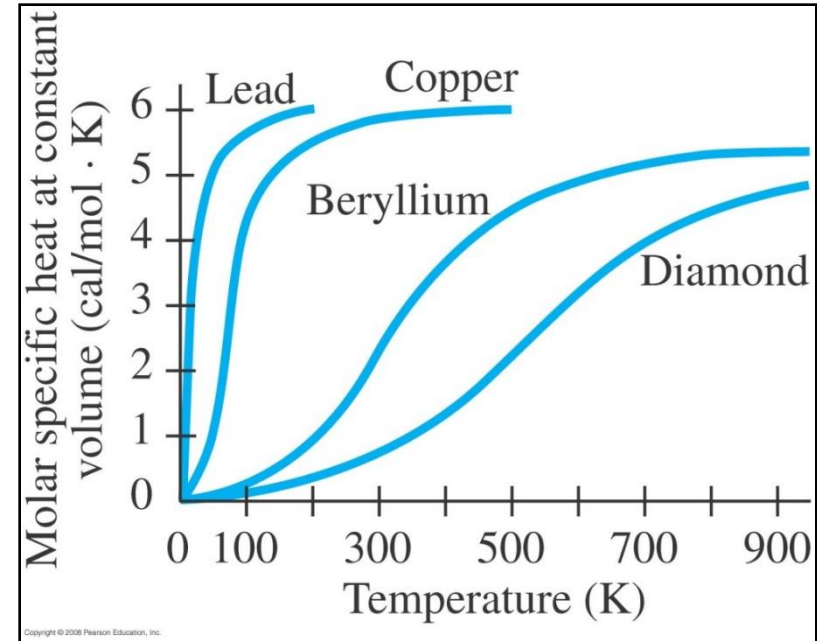6 degrees of freedom, although not all active at low temperatures

Figure 19-18 Molar specific heats of solids as a function of temperature.

19-3 Specific Heat (revisited)

Dulong-Petit law

$$C = 3R = 24.9 \; J/mol \cdot °C$$

(= 6.0 cal/mol·°C)

example Al:
900 (J/(kg·°C)) x 0.027 (kg/mol)

TABLE 19–1 Specific Heats
(at 1 atm constant pressure and 20°C unless otherwise stated)

Substance	Specific Heat, c	
	kcal/kg · C° (= cal/g · C°)	J/kg · C°
Aluminum	0.22	900
Alcohol (ethyl)	0.58	2400
Copper	0.093	390
Glass	0.20	840
Iron or steel	0.11	450
Lead	0.031	130
Marble	0.21	860
Mercury	0.033	140
Silver	0.056	230
Wood	0.4	1700
Water		
Ice (−5°C)	0.50	2100
Liquid (15°C)	1.00	4186
Steam (110°C)	0.48	2010
Human body (average)	0.83	3470
Protein	0.4	1700

C in J/(mol·°C)

24.3 (Al)

24.8 (Cu)
25.2 (Fe)
26.9 (Pb)

24.8 (Ag)

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19-9 Adiabatic Expansion of a Gas

First law (diff. form) $dE_{int} = dQ - PdV = -PdV$

(adiabatic $dQ = 0$) $nC_V dT + PdV = 0$

$dE_{int} = nC_V dT$

Ideal gas law and differentiate $PV = nRT$
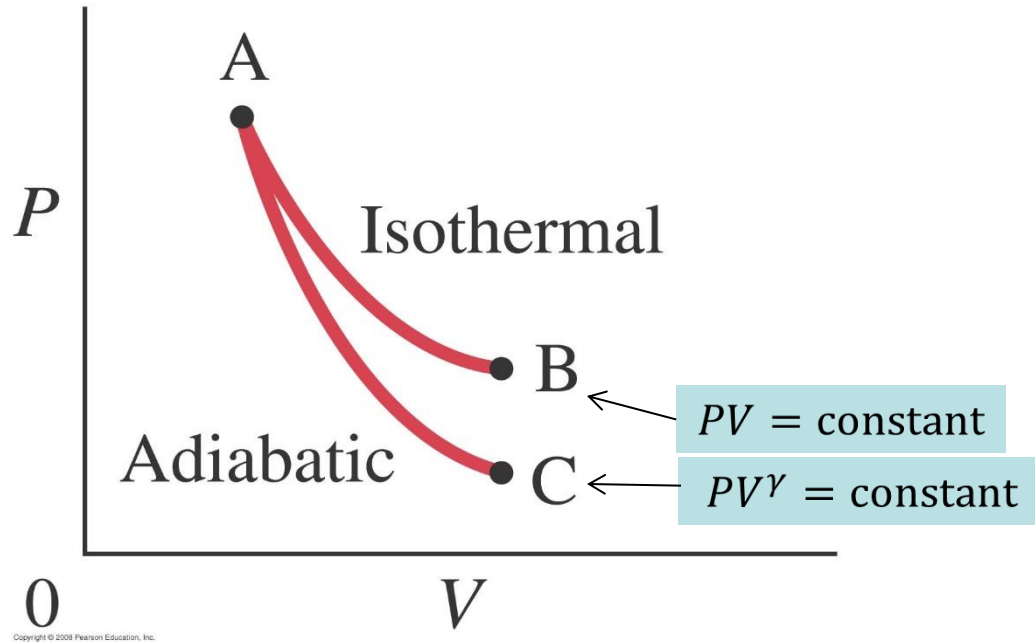
Substitute ndT and rearrange

$PdV + VdP = nRdT$

$\longrightarrow$ $C_V PdV + C_V VdP + RPdV = 0$ $\longrightarrow$ $(C_V + R)PdV + C_V VdP = 0$

C_P

Define $\gamma = C_P / C_V$ and we obtain the differential equation

$\gamma \, dV/V + dP/P = 0$ with solution $\ln P + \ln V^\gamma = $ constant

using properties of "ln" $\longrightarrow$ $\ln(PV^\gamma) = $ constant $PV^\gamma = $ constant

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19-9 Adiabatic Expansion of a Gas

- For adiabatic process ideal gas law still holds, meaning that T is not constant (in the adiabatic process)

- $\gamma = 5/3$ and $7/5$ for monatomic and diatomic gases, respectively)

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Exercise 8: In the process of taking an ideal monatomic gas from state A to state C along the two steps (A to B and B to C) shown in the PV diagram. The temperature at the state A is 310 K.

a. Determine the volume at the state B.
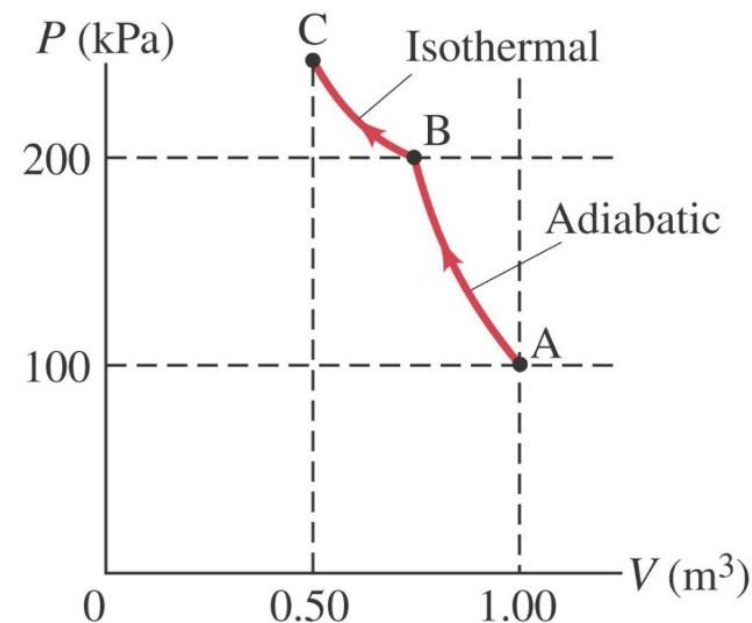b. Calculate the work done on the gas during the whole process.

Adiabatic: $P_A V_A^\gamma = P_B V_B^\gamma$

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

$\gamma = 1.67$

$\ln(V_B) = \dfrac{\ln(\frac{P_A V_A^\gamma}{P_B})}{\gamma} = -0.415$

$V_B = 0.66 m^3$

$W_{AB} = -\Delta E_{AB} = \dfrac{3}{2}(P_B V_B - P_A V_A) = -0.48\,J$

Isothermal: $T_C = T_B$

$W_{BC} = nRT \ln\left(\dfrac{V_C}{V_B}\right) = P_B V_b \ln(0.75) = -0.37J$

$W_{AC} = -0.85J$

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19-10 Heat Transfer

The sun heats the ground
Radiation

The warm air rises
Convection

The ground heats the air
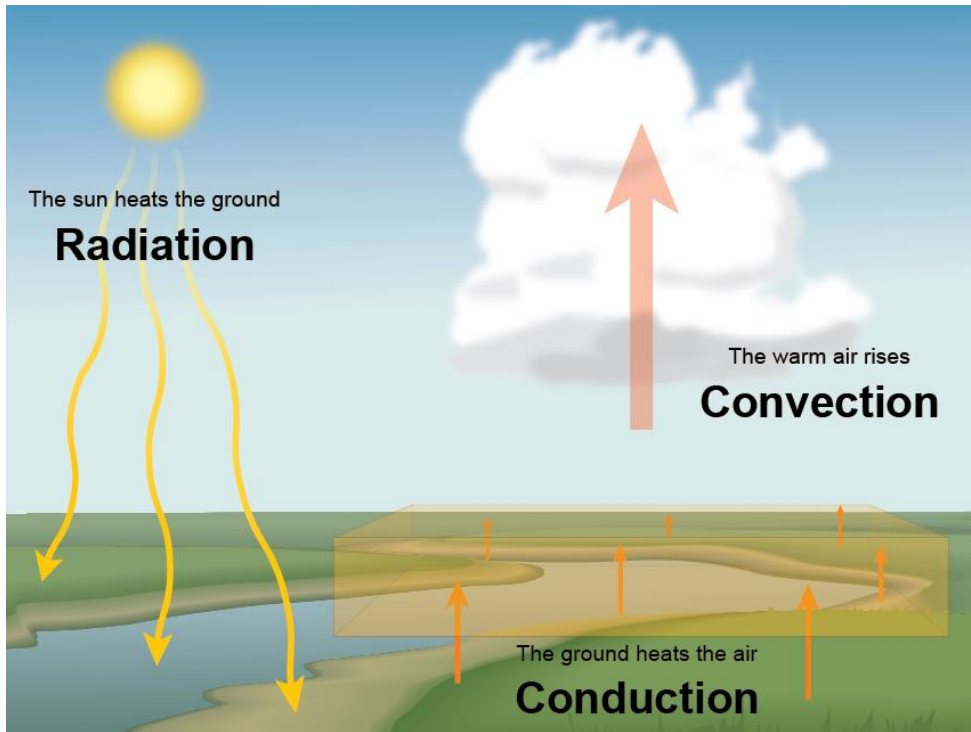**Conduction**

Figure source: https://www.weather.gov/jetstream/heat.
retrieved on 8th March 2023

- **Conduction**: energy is transferred by direct contact → collisions of molecules, more in solids and liquids.

- **Convection**: energy is transferred by a mass motion of molecules over a long distance.

- **Radiation**: energy is transferred through electromagnetic waves.

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19-10 Conduction

Due to molecular **collisions**, i.e. transfer of kinetic energy to slower moving neighbours.
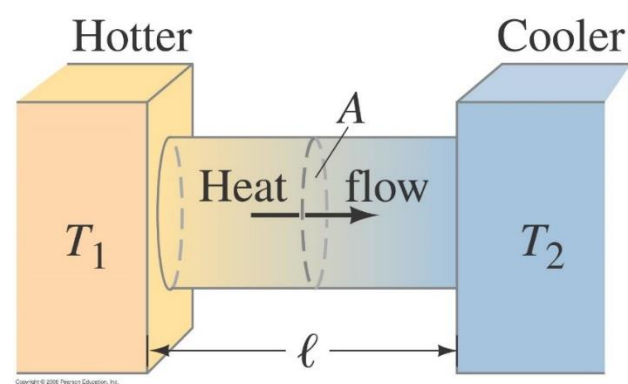

Figure 19-21 Heat conduction between areas at temperature T_1 and T_2.

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{l} \quad (J/s)$$

Thermal conductivity

Cross-sectional area

Temperature gradient

Thermal conductivity k large
→ good thermal conductor

TABLE 19–5
Thermal Conductivities

Substance	Thermal conductivity, k	
	$\frac{kcal}{(s \cdot m \cdot C°)}$	$\frac{J}{(s \cdot m \cdot C°)}$
Silver	10×10^{-2}	420
Copper	9.2×10^{-2}	380
Aluminum	5.0×10^{-2}	200
Steel	1.1×10^{-2}	40
Ice	5×10^{-4}	2
Glass	2.0×10^{-4}	0.84
Brick	2.0×10^{-4}	0.84
Concrete	2.0×10^{-4}	0.84
Water	1.4×10^{-4}	0.56
Human tissue	0.5×10^{-4}	0.2
Wood	0.3×10^{-4}	0.1
Fiberglass	0.12×10^{-4}	0.048
Cork	0.1×10^{-4}	0.042
Wool	0.1×10^{-4}	0.040
Goose down	0.06×10^{-4}	0.025
Polyurethane	0.06×10^{-4}	0.024
Air	0.055×10^{-4}	0.023

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Exercise 9: Conduction

One end of a 64-cm-long copper rod with a diameter of 2.0 cm is kept at 460 °C, and the other is immersed in water at 22 °C. Calculate the heat conduction rate through the rod.
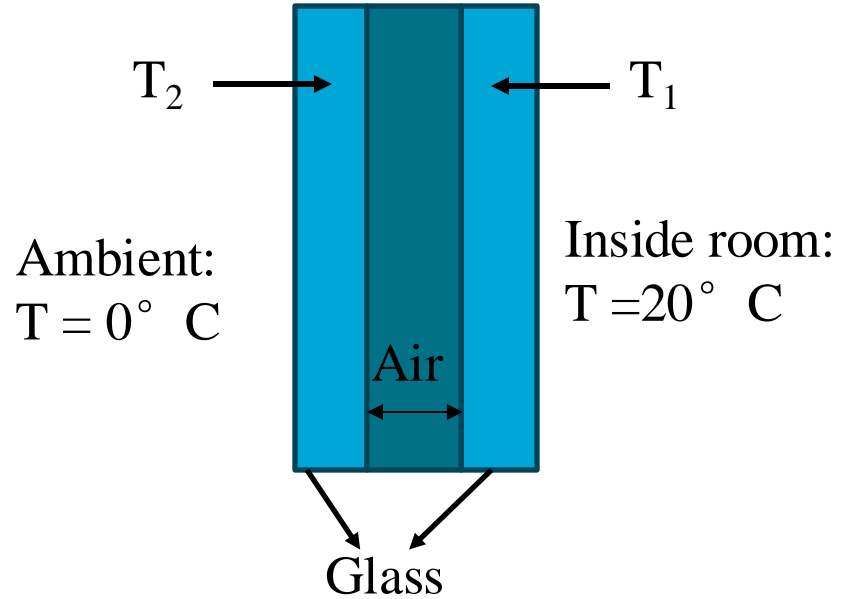
$$\frac{Q}{t} = kA \frac{T_1 - T_2}{l}$$

$$= 380(J/s \cdot m \cdot °C) \times \pi \times \left(\frac{0.02}{2}\right)^2 m^2 \times \frac{(460 - 22)°C}{0.64\ m} = 81.66\ J/s$$

Exercise 10: Conduction

In winter, the room temperature is 20°C and the outside temperature is 0°C. The glass of the window has an area of 2.0 m^2 and a thickness of 5 mm. Ignore the window frame and thermal resistance. The thermal conductivity of glass is 0.84 J/s/m/°C and of the air is 0.023 J/s/m/°C.
1) Determine the heat loss rate through the window;
2) To improve the house insulation, a double layer of glass, separated by air, is considered. The aera and thickness of the glass remain unchanged, and the thickness of the air layer is about 5 mm. Assume that the same amount of heat rate crosses each surface. Determine the heat loss rate by conduction through the window.

$$\frac{Q}{t} = kA\frac{T_1 - T_2}{l}$$

$$= 0.84 \ \frac{J}{°C.\,s.\,m} \times 2 \ m^2 \times \frac{20°C - 0°C}{5 \times 10^{-3} \ m} = 6720 \ J/s$$

$$T_2 \longrightarrow \quad \longleftarrow T_1$$

Ambient:
T = 0° C

Inside room:
T = 20° C

Air

Glass

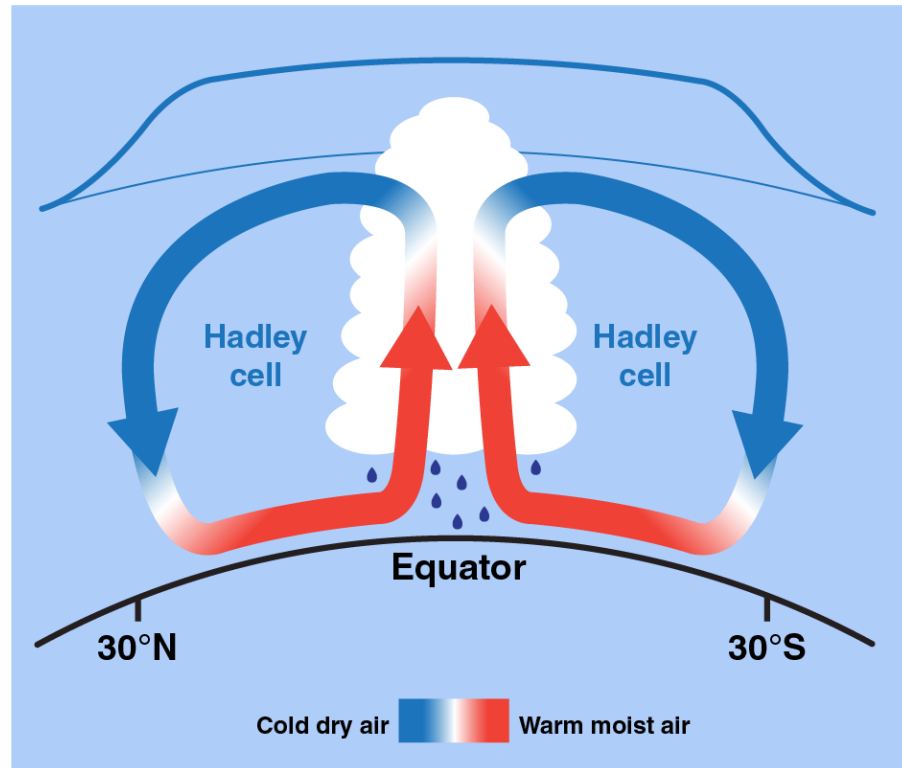$$\frac{Q}{t} = kA \frac{T_1 - T_2}{l}$$

Thermal resistance: $\quad R = \frac{l}{k}$

$$\frac{Q}{t} = A \frac{T_1 - T_2}{R}$$

$$R_{total} = 2R_{glass} + R_{air} = \frac{2 \times 0.005m}{0.84 \, J/(s \cdot m \cdot °C)} + \frac{0.005 \, m}{0.023 J/(s \cdot m \cdot °C)} = 0.119 + 0.217 = 0.336 \, (s \cdot m^2 \cdot °C)/J$$

$$\frac{Q}{t} = A \frac{T_1 - T_2}{R} = 2.0 \, m^2 \times \frac{(20 - 0)°C}{0.336(s \cdot m^2 \cdot °C)/J} = 119 \, J/s$$

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19-10 Convection

convection, i.e. heat flows by the **mass movement** of molecules from one place to another

Natural convection
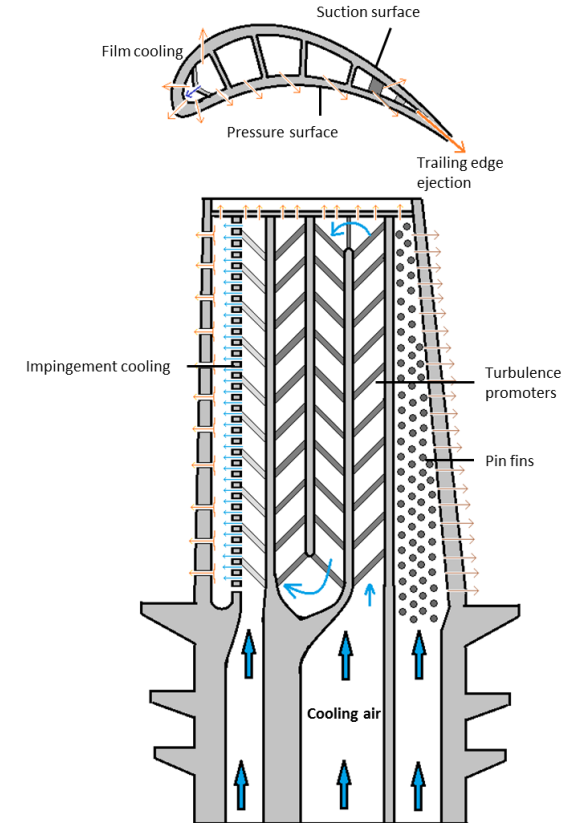
Forced convection

$$Q = hA\Delta T$$

19-10 Radiation

radiation (e.m. waves), i.e. no medium needed (heat is transferred over empty space)
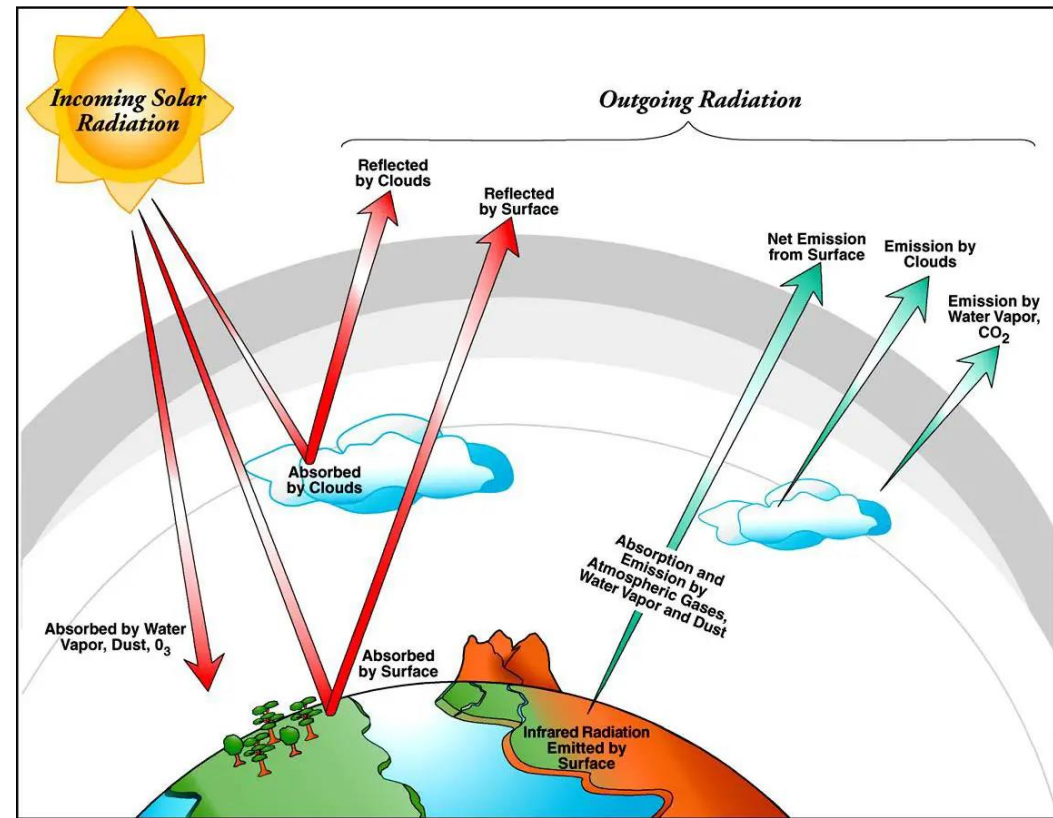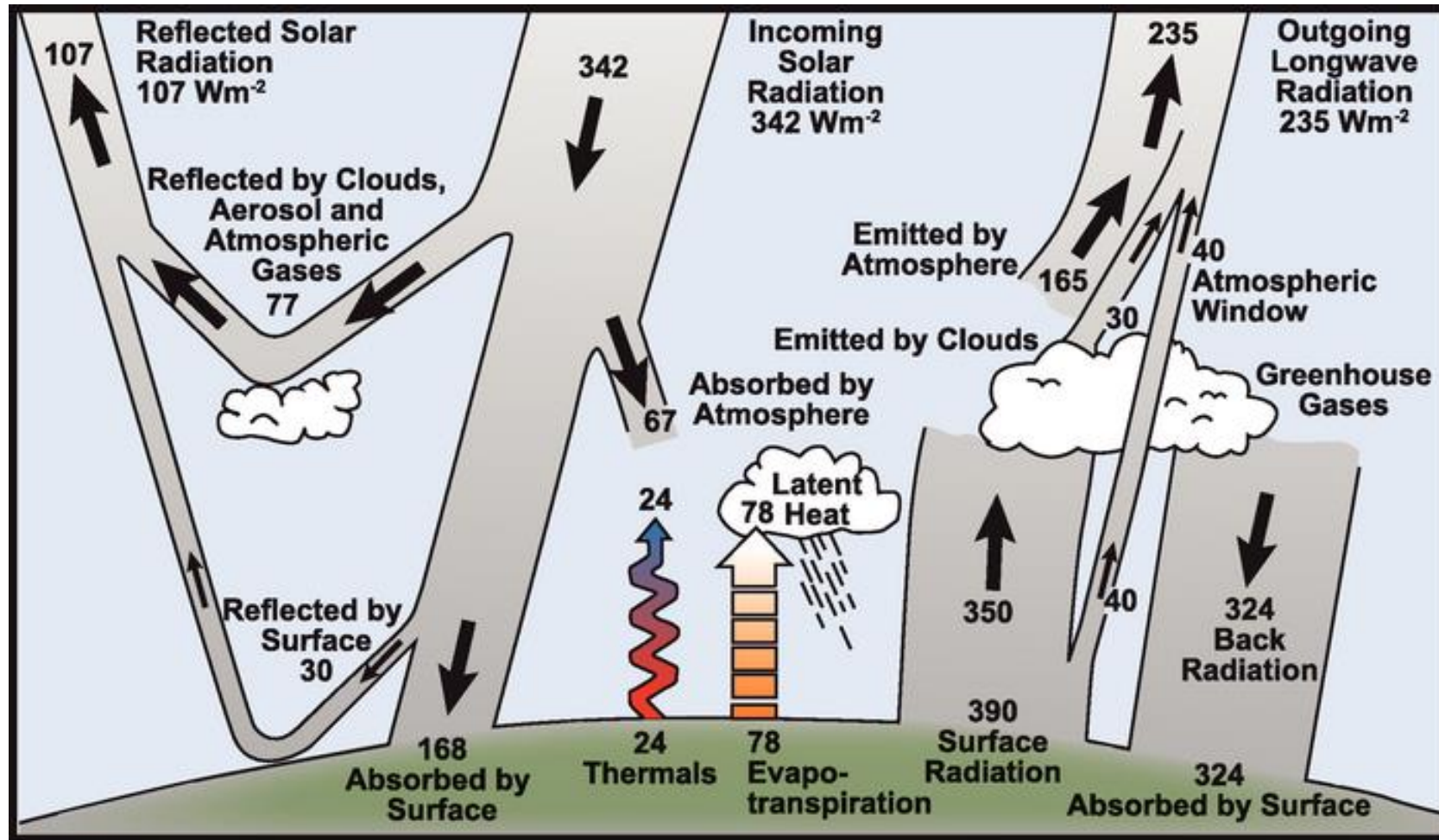example: warmth we receive from a fire is mainly radiant energy

Image courtesy of the U.S. Department of Energy Atmospheric Radiation Measurement (ARM) user facility

Example: Radiation and climate change

Figure source: Kiehl and Trenberth, Earth's Annual Global Mean Energy Budget, 1997.

19-10 Stefan-Boltzmann law

Stefan-Boltzmann law gives rate at which energy leaves the object (in W):

area of emitting object, m²

$$\frac{Q}{t} = \epsilon \sigma A T^4$$

Emissivity
(≈ 1 for very dark objects, ≈ 0 shiny metal surfaces)

Stefan-Boltzmann constant = 5.67x10⁻⁸ (W/m²·K⁴)

net rate from object   object   environment

$$\frac{Q}{t} = \epsilon \sigma A (T_1^4 - T_2^4)$$

Kirchoff's law: good emitter = good absorber (very dark objects)

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Example: radiation

What is the net power that a person with a surface area of 1.5 m^2 radiates if his/her emissivity is 0.9, the skin temperature is 37 ˚C, and the room temperature is 20 ˚C?
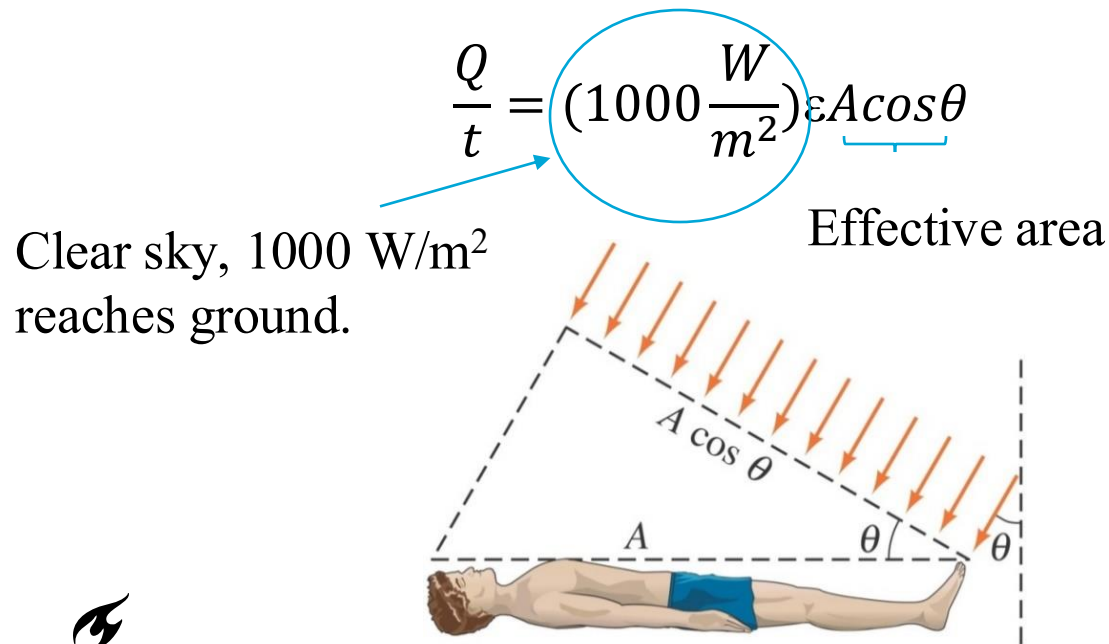
(A) 707 W

(B) 143 W ✓

(C) 165 W

(D) 147 W

$$\frac{Q}{t} = \epsilon\sigma A(T_1^4 - T_2^4) = 0.9 \times 5.67 \times 10^{-8}\ W/(m^2 \cdot K^4) \times 1.5m^2 \times (310K^4 - 293K^4)$$
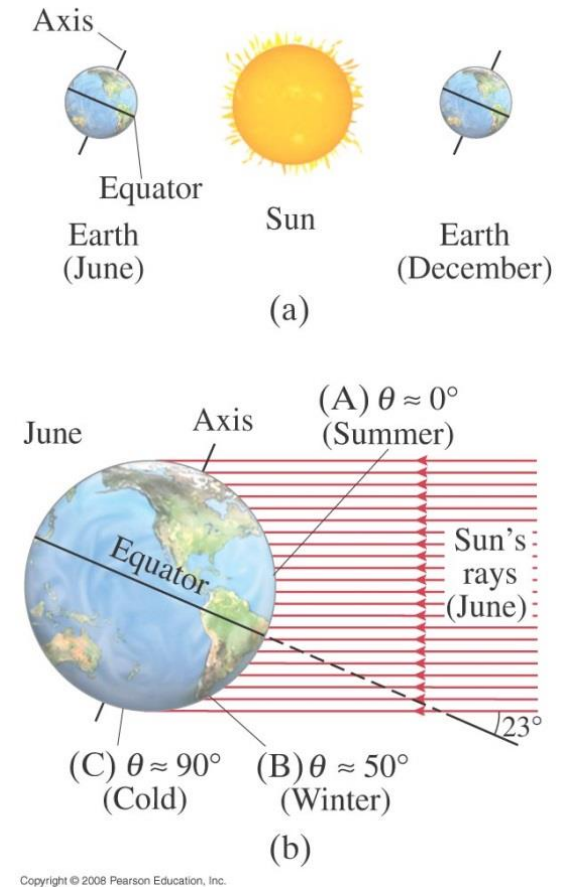
$$= 143\ W$$

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19-10 Heat Transfer

heating of objects by radiation from the Sun (visible light, UV and IR)

- Solar constant = 1350 W/m²
- Namely 1350 J of energy falls on atmosphere per second and per m²
- Object facing the Sun absorbs energy at a rate

$$\frac{Q}{t} = \left(1000\,\frac{W}{m^2}\right)\varepsilon A\cos\theta$$

Clear sky, 1000 W/m² reaches ground.

Effective area

explains the seasons! (not the result of how close is Earth to Sun)
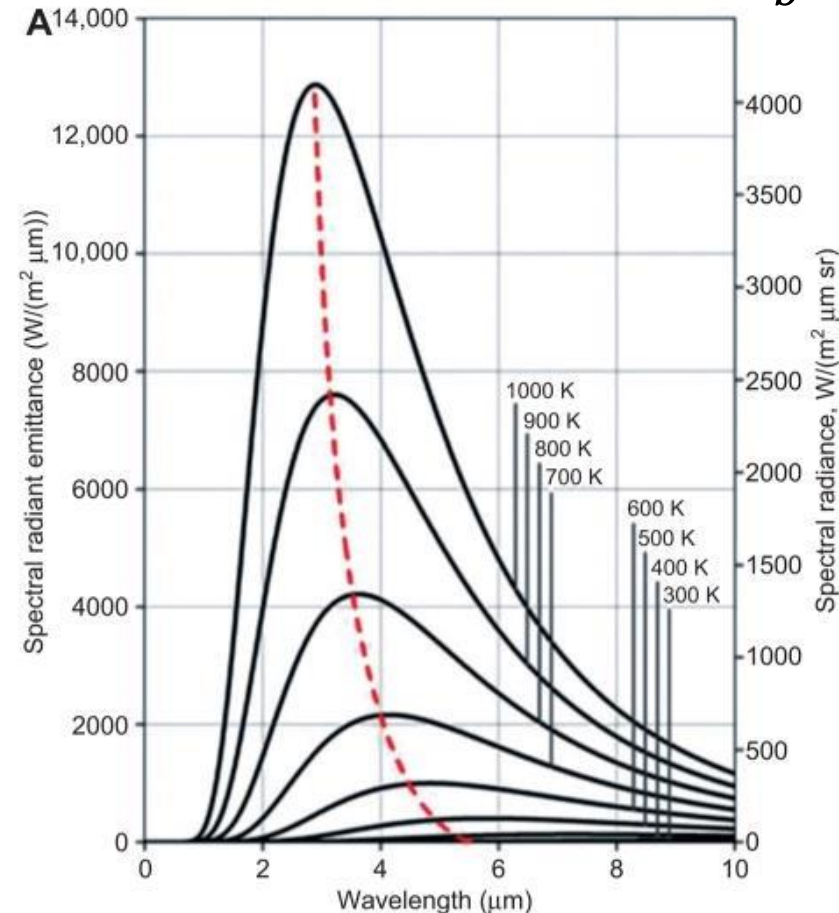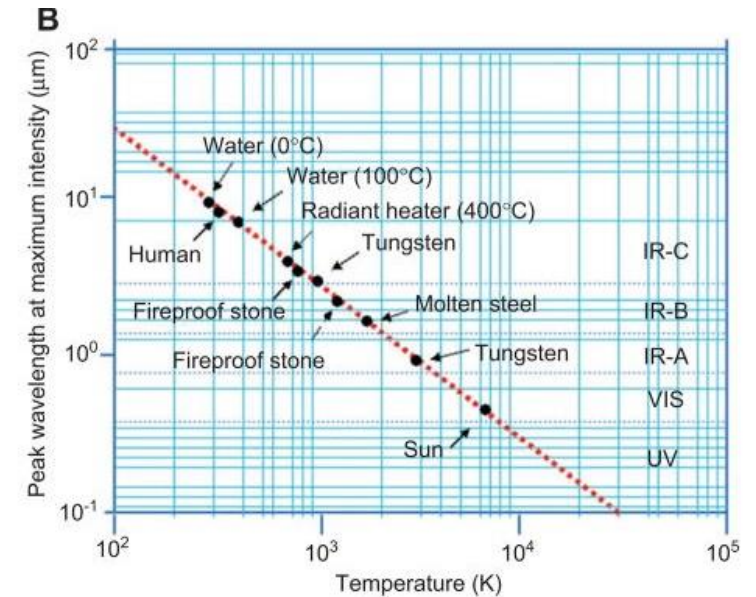
Figure 19-26

Figure 19-25

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Black body radiation

$$B(v, T) = \frac{2hv^3}{c^2} \frac{1}{e^{\frac{hv}{k_b T}} - 1}$$

Black body: The object absorbs all radiation falling on it at all wavelength

Blackbody radiation distribution depends on temperature.

Planck's law

Wien's displacement law

Application: Thermography
---imaging surface temperature

Before smoking

after smoking (lower temperature due to impaired blood circulation)

(a)   (b)

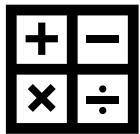**Color cold (blue) to hot (white)**

Wrap up: revisit learning objective

After today's lecture, you should be able to:

- Understand the concept of **heat, work, internal energy, specific heat, molar specific heat** (*Cp=Cv+R*)**, latent heat**

- Analyze different **thermodynamic processes,** using 1st law of thermodynamics: *isothermal (T=c); isobaric (p=c); isovolumetric (V=c); adiabatic ($\Delta Q = 0$)*

- Calculate thermodynamic parameters, e.g., **heat, work, internal energy, temperature, pressure, volume, etc.** for given processes

$$E_{int} = N \frac{i_{total}}{2} kT = Q - W$$

- Calculate the **heat transfer rate** by conduction and radiation

$$Conduction: \frac{Q}{t} = kA \frac{T_1 - T_2}{l} \qquad Radiation: \frac{Q}{t} = \epsilon \sigma A(T_1^4 - T_2^4)$$

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Thank you for your attention!

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