

Lecture 40

(Ch. 12: 3-4)

Chapter 12: Thermodynamics

Thermodynamic Processes

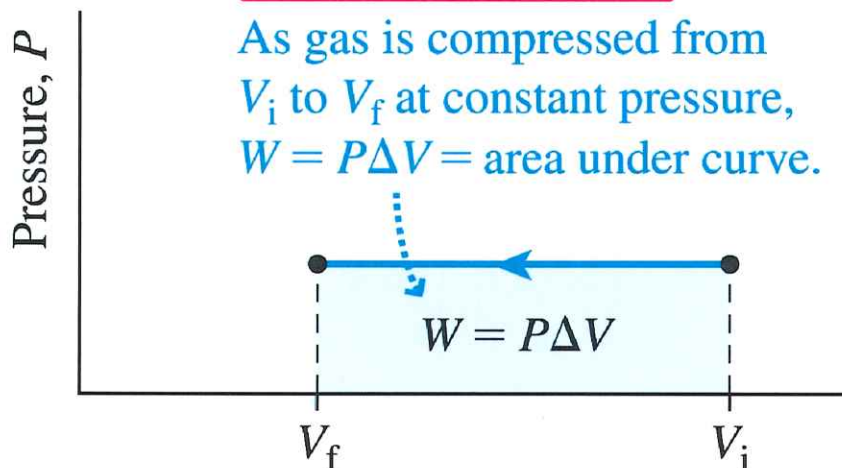
A state variable describes the state of a system at time t , but it does not reveal how the system was put into that state. Examples of state variables: pressure, temperature, volume, number of moles, and internal energy.

Constant-Pressure Processes (isobaric)

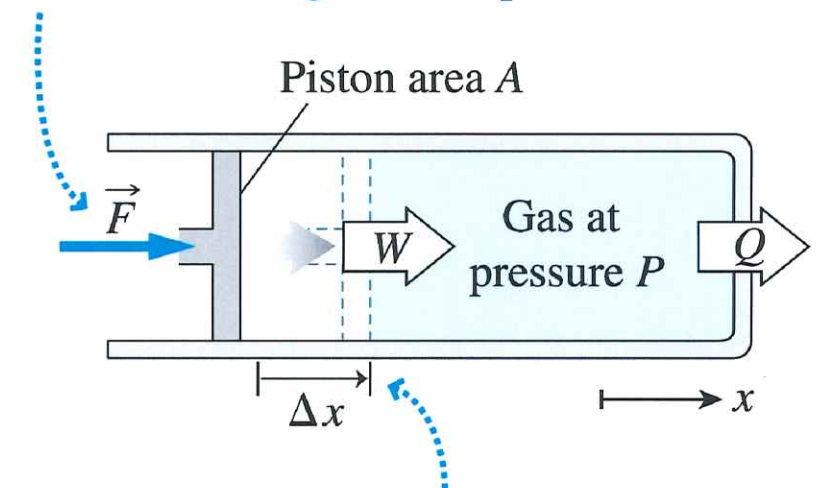
$$W = F_x \Delta x = PA \Delta x$$

$$V = Ax \Rightarrow \Delta V = -A \Delta x$$

$$W = -P \Delta V$$



A constant force is applied and heat is allowed to escape, so the pressure remains constant as the gas is compressed.



The piston moves through displacement Δx , resulting in work $W = F_x \Delta x = PA\Delta x$.

Example: A flexible container contains $2.42 \times 10^{-5} \text{ m}^3$ of fluid at room T. Somebody pushes on the container, maintaining a constant 1-atm pressure, and reduces its volume by 25 %. How much work is done on the fluid ?

$$\begin{aligned} \text{Work done on the fluid: } W &= -P(V_f - V_i) = 1\text{atm}(0.75V_i - V_i) = \\ &= - (1.013 \times 10^5 \text{ Pa})V_i(-0.25) = (1.013 \times 10^5 \text{ Pa})(2.42 \times 10^{-5} \text{ m}^3)(0.25) = 0.61 \text{ J} \end{aligned}$$

19. A gas is enclosed in a container fitted with a piston of cross-sectional area 0.150 m^2 . The pressure of the gas is maintained at $6.00 \times 10^3 \text{ Pa}$ as the piston moves inward 20.0 cm .

a. Calculate the work done by the gas.

Answer ↓

b. If the internal energy of the gas decreases by 8.00 J , find the amount of energy removed from the system by heat during the compression.

12.19 (a) The change in the volume occupied by the gas is

$$\Delta V = V_f - V_i = A(L_f - L_i) = (0.150 \text{ m}^2)(-0.200 \text{ m}) = -3.00 \times 10^{-2} \text{ m}^3$$

and the work done by the gas is

$$W_{\text{env}} = +P(\Delta V) = (6\,000 \text{ Pa})(-3.00 \times 10^{-2} \text{ m}^3) = \boxed{-180 \text{ J}}$$

(b) The first law of thermodynamics is $\Delta U = Q + W = -Q_{\text{output}} - W_{\text{env}}$.

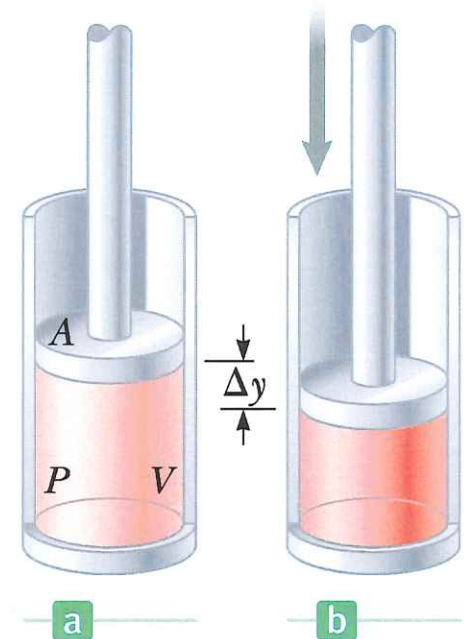
Thus, if $\Delta U = -8.00 \text{ J}$, the energy transferred out of the system by heat is

$$Q_{\text{output}} = -\Delta U - W_{\text{env}} = -(-8.00 \text{ J}) - (-180 \text{ J}) = \boxed{+188 \text{ J}}$$

Topic Summary

- **Work in Thermodynamic Processes**

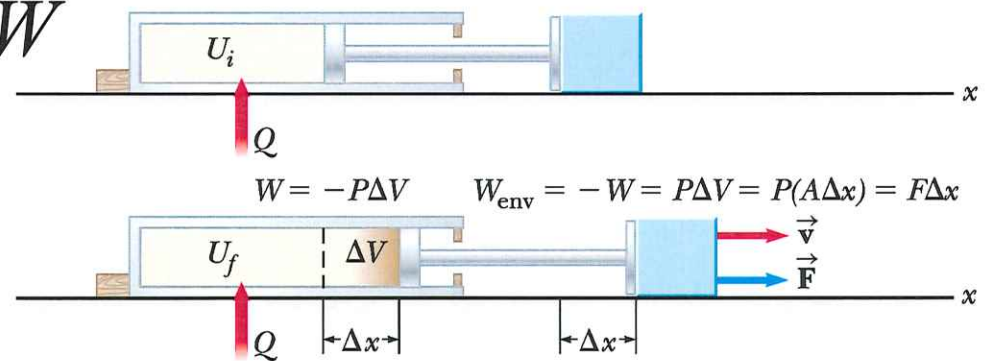
$$W = -P\Delta V$$



- **The First Law of Thermodynamics**

$$\Delta U = U_f - U_i = Q + W$$


$$\Delta U = nC_v\Delta T$$





Calorimetry reveals that a Milky Way® candy bar contains more energy than a stick of dynamite. The candy bar contains 200 food Calories. That's 200,000 physicist calories or about 840,000 joules! Nearly a megajoule! A megajoule of energy from a candybar can perform enough work to lift an average 70-kilogram human being 1200 meters in the air. That's higher than the cliff face of Yosemite's El Capitan. No stick of dynamite can do that! In fact, an ounce of dynamite produces only one-quarter as many calories when it explodes as an ounce of sugar does when it burns.

$$1 \text{ Cal} = 1000 \text{ cal} = 4186 \text{ J}$$


 80 kcal per km = ΔU | $Q = ?$ (lost) if 120g H_2O lost
 - 330 kJ | $W = ?$

The heat lost is the heat of vaporization of 120 g of perspiration, which we can approximate with 120 g of water. The heat of vaporization of water is listed in Table 13.3. The work done is the internal energy change minus the heat lost.

Known: $\Delta U = 80$ kcal; $m = 120$ g; $L_v = 2.26 \times 10^6$ J/kg.

SOLVE (a) The heat lost is given by Equation 13.6:

$$Q = mL_v = (120 \text{ g})(2.26 \times 10^6 \text{ J/kg}) = 271 \text{ kJ}$$

(b) The work done by the runner is:

$$W = \Delta U - Q = (80 \text{ kcal}) - (271 \text{ kJ}) = 64 \text{ kJ}$$

330 kJ

$\Delta U = Q + W$
 $1 \text{ cal} = 4.18 \text{ J}$

REFLECT The heat of vaporization for water has a slight dependency on temperature. The value in Table 13.3 is the heat of vaporization at the boiling point. At body temperature, the heat of vaporization of water is slightly larger, 2.4×10^6 J/kg, but the value used above is an acceptable approximation.

1st Law | $\Delta U = Q + W$

Chapter 12: Thermodynamics

The First Law of Thermodynamics

Neglecting the potential energy of the gas molecules, in the case of an ideal gas, the change of internal energy is the change of thermal energy.

$$\Delta U = \Delta E_{th} = \frac{3}{2} N k_B \Delta T, \text{ for monoatomic gases}$$

$$\Delta U = \Delta E_{th} = \frac{5}{2} N k_B \Delta T, \text{ for diatomic gases}$$

Work in Thermodynamic Processes

$$U = \frac{3}{2}nRT$$

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

$$C_v \equiv \frac{3}{2}R \quad (\text{monatomic gases})$$

$$\Delta U = nC_v\Delta T$$

$$C_v \equiv \frac{5}{2}R \quad (\text{diatomic gases})$$

Work in Thermodynamic Processes

Table 12.1 Molar Specific Heats of Various Gases

Gas	Molar Specific Heat (J/mol · K) ^a			
	C_p	C_v	$C_p - C_v$	$\gamma = C_p/C_v$
Monatomic Gases				
He	20.8	12.5	8.33	1.67
Ar	20.8	12.5	8.33	1.67
Ne	20.8	12.7	8.12	1.64
Kr	20.8	12.3	8.49	1.69
Diatomic Gases				
H ₂	28.8	20.4	8.33	1.41
N ₂	29.1	20.8	8.33	1.40
O ₂	29.4	21.1	8.33	1.40
CO	29.3	21.0	8.33	1.40
Cl ₂	34.7	25.7	8.96	1.35
Polyatomic Gases				
CO ₂	37.0	28.5	8.50	1.30
SO ₂	40.4	31.4	9.00	1.29
H ₂ O	35.4	27.0	8.37	1.30
CH ₄	35.5	27.1	8.41	1.31

^aAll values except that for water were obtained at 300 K.

11. A balloon holding 5.00 moles of helium gas absorbs 925 J of thermal energy while doing 102 J of work expanding to a larger volume.

a. Find the change in the balloon's internal energy.

Answer ↓

b. Calculate the change in temperature of the gas.

12.11 (a) Substitute the values $Q = 925 \text{ J}$ and $W = -102 \text{ J}$ (the work done on the gas is negative because it expands to a larger volume) into the first law of thermodynamics to find

$$\Delta U = Q + W = 925 \text{ J} - 102 \text{ J} = \boxed{823 \text{ J}}$$

(b) Use the relation $\Delta U = nC_v\Delta T$ to find the change in temperature.

Here, $n = 5.00$ moles and $C_v = \frac{3}{2}R$ (for helium, a monatomic gas) to

find

$$\Delta U = nC_v\Delta T = \frac{3}{2}nR\Delta T \rightarrow \Delta T = \frac{\Delta U}{\frac{3}{2}nR}$$
$$\Delta T = \frac{823 \text{ J}}{\frac{3}{2}(5.00 \text{ mol})(8.31 \text{ J/mol}\cdot\text{K})} = \boxed{13.2 \text{ K}}$$

17. **v** A gas is compressed at a constant pressure of 0.800 atm from 9.00 L to 2.00 L. In the process, 400. J of energy leaves the gas by heat.

a. What is the work done *on* the gas?

Answer ↓

b. What is the change in its internal energy?

$$\begin{aligned} 12.17 \quad (a) \quad W &= -P(\Delta V) = -(0.800 \text{ atm})(-7.00 \text{ L}) \left(\frac{1.013 \times 10^5 \text{ Pa}}{1 \text{ atm}} \right) \left(\frac{10^{-3} \text{ m}^3}{1 \text{ L}} \right) \\ &= \boxed{567 \text{ J}} \end{aligned}$$

$$(b) \quad \Delta U = Q + W = -400 \text{ J} + 567 \text{ J} = \boxed{167 \text{ J}}$$

Isobaric Processes

$$\Delta U = nC_v\Delta T \qquad PV = nRT$$

$$Q = \Delta U - W = \Delta U + P\Delta V$$

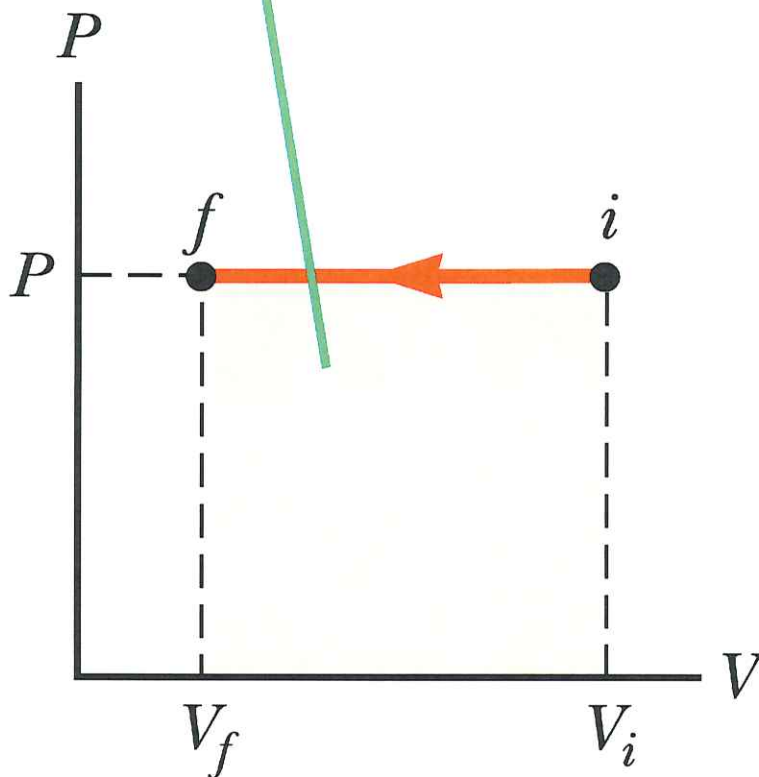
$$Q = \frac{3}{2}nR\Delta T + nR\Delta T = \frac{5}{2}nR\Delta T$$

$$Q = nC_p\Delta T$$

$$C_p = C_v + R$$

Isobaric Processes

The shaded area represents the work done on the gas.



$$W_{\text{env}} = P\Delta V$$

$$W = -P\Delta V$$

24. **S** An ideal gas expands at constant pressure.

- Show that $P\Delta V = nR\Delta T$.
- If the gas is monatomic, start from the definition of internal energy and show that $\Delta U = \frac{3}{2}W_{\text{env}}$, where W_{env} is the work done by the gas on its environment.
- For the same monatomic ideal gas, show with the first law that $Q = \frac{5}{2}W_{\text{env}}$.
- Is it possible for an ideal gas to expand at constant pressure while exhausting thermal energy? Explain.

12.24 (a) From the ideal gas law, $P_i V_i = nRT_i$ and $P_f V_f = nRT_f$. Thus, if $P_i = P_f$

$= P$, subtracting these two expressions gives $PV_f - PV_i = nRT_f -$

nRT_i , or $\boxed{P(\Delta V) = nR(\Delta T)}$.

(b) For a monatomic, ideal gas containing N gas atoms, the internal

energy is $U = N\left(\frac{1}{2}m\overline{v^2}\right) = (nN_A)\left(\frac{3}{2}k_B T\right) = \frac{3}{2}nRT$. Thus, the change in

internal energy of this gas in a thermodynamic process is

$\Delta U = \frac{3}{2}nR(\Delta T)$. But, using the result of part (a) above, we have,

for an isobaric process involving an monatomic ideal gas,

$$\Delta U = \frac{3}{2}nR(\Delta T) = \frac{3}{2}P(\Delta V) = \boxed{\frac{3}{2}W_{\text{env}}}$$

- (c) We recall that the work done on the gas is $W = -W_{\text{env}}$, and use the first law of thermodynamics to find that the energy transferred to the gas by heat to be

$$Q = \Delta U - W = \Delta U + W_{\text{env}} = \frac{3}{2}W_{\text{env}} + W_{\text{env}} \quad \text{or} \quad \boxed{Q = \frac{5}{2}W_{\text{env}}}$$

- (d) In an isobaric *expansion* ($\Rightarrow \Delta V > 0$), the work done on the environment is $W_{\text{env}} = P(\Delta V) > 0$. Thus, from the result of part (c) above, the energy transfer as heat is $Q > 0$, meaning that the energy flow is *into* the gas. Therefore, it is **impossible for the gas to exhaust thermal energy in an isobaric expansion**.

Figure 14.8

Adiabatic Process

$$Q = 0 \quad \Delta U = W + Q$$

Insulation prevents heat flow between gas and surroundings.

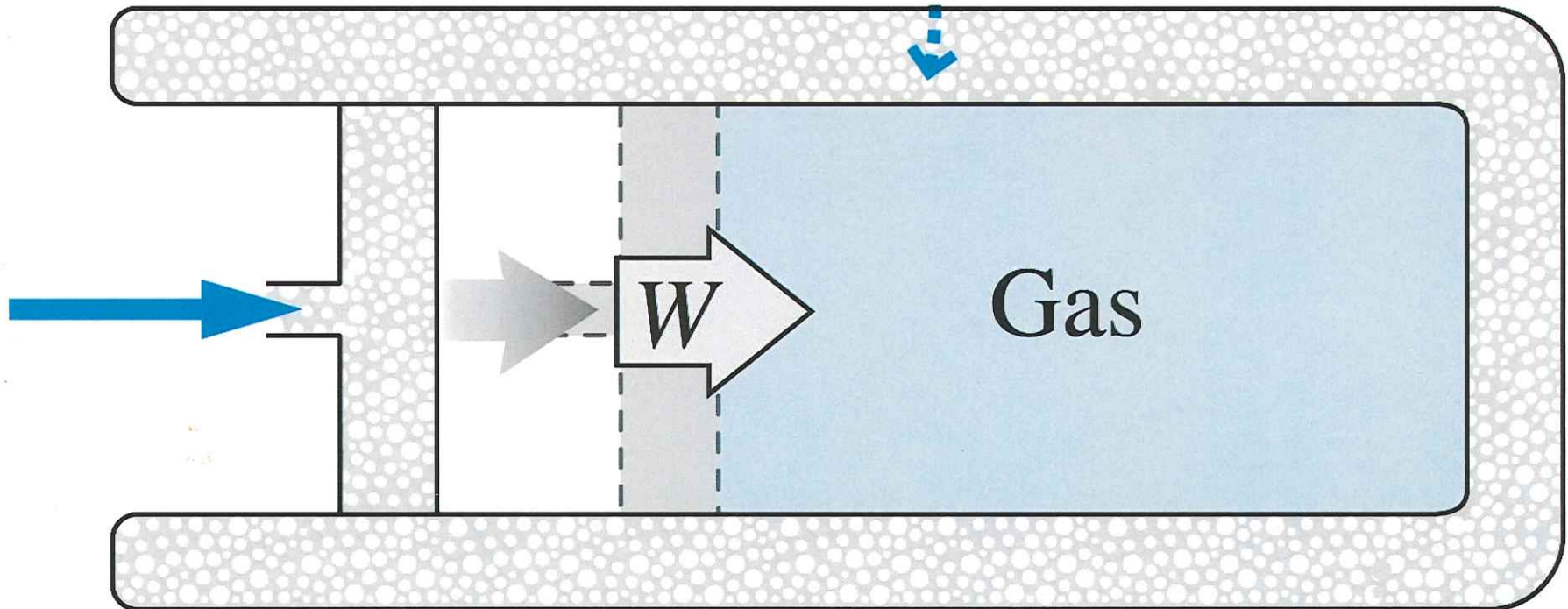
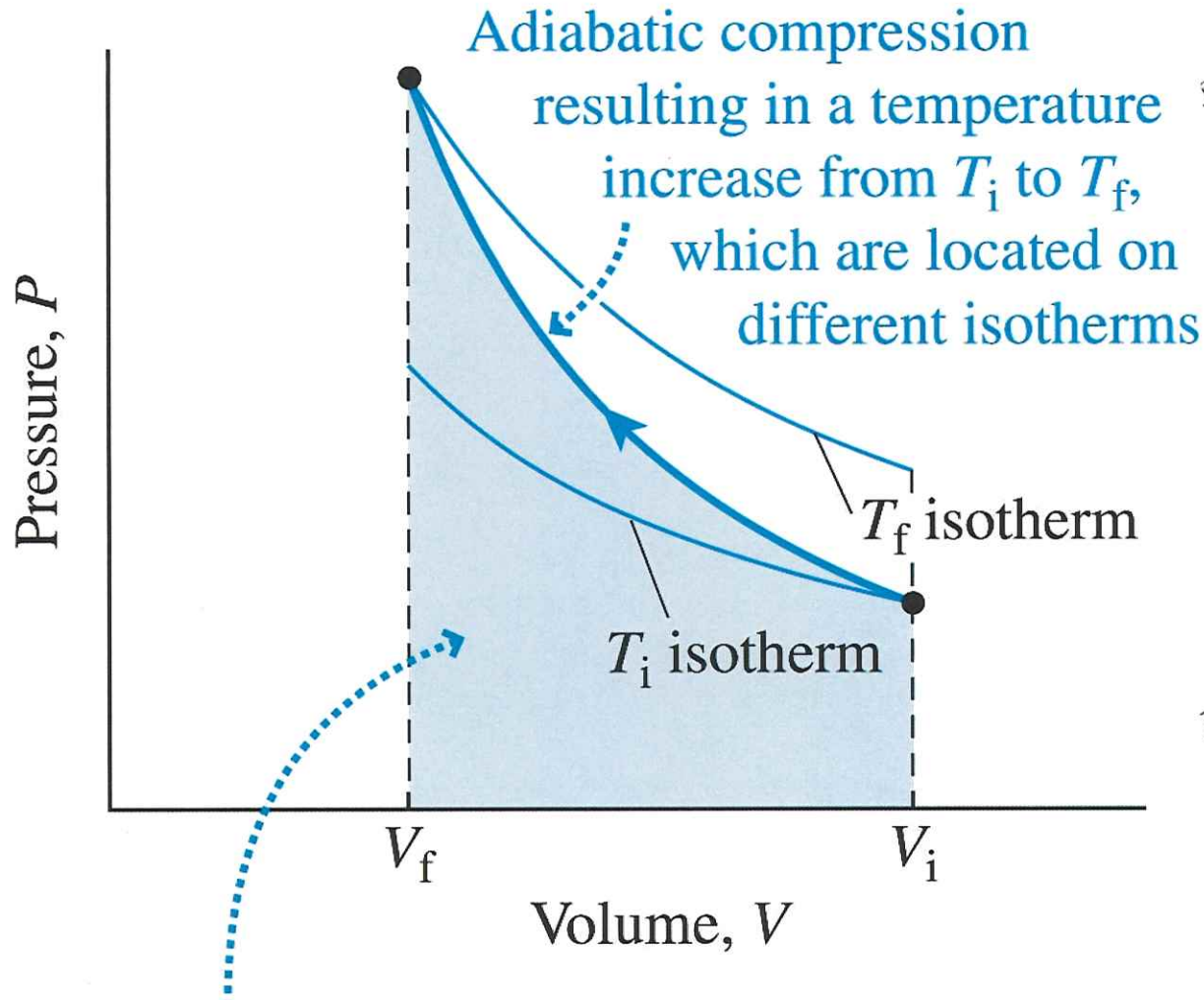


Figure 14.9



$Q = 0$
so from $\Delta U = Q + W$
 $\Delta U = W$
But $U \sim E_{TR} \sim T$
so if $(+w)$ then $T \uparrow$

$$P \cdot V^\gamma = \text{const}$$

$$\gamma = C_p / C_v$$

For monatomic

$$C_p = 5R/2$$

$$C_v = 3R/2$$

$$\text{so } \gamma = 5/3$$

For diatomic

$$\gamma = 7/5$$

$$W = \frac{P_f V_f - P_i V_i}{\gamma - 1}$$

Adiabatic Processes

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CH ₄	35.5	27.1	8.41	1.31

^aAll values except that for water were obtained at 300 K.

Adiabatic Processes

$$Q = 0 \rightarrow \Delta U = W \quad (\text{Adiabatic Process})$$

$$PV^\gamma = \text{constant} \quad \gamma = \frac{C_p}{C_v}$$

26. An ideal diatomic gas expands adiabatically from 0.750 m^3 to 1.50 m^3 . If the initial pressure and temperature are $1.50 \times 10^5 \text{ Pa}$ and 325 K , respectively, find

- the number of moles in the gas,
- the final gas pressure,
- the final gas temperature, and
- the work done on the gas.

12.26 The gas expands adiabatically with $V_i = 0.750 \text{ m}^3$, $V_f = 1.50 \text{ m}^3$, $P_i = 1.50 \times 10^5 \text{ Pa}$, and $T_i = 325 \text{ K}$.

(a) Use the ideal gas law to find the number of moles:

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{P_i V_i}{RT_i} = \frac{(1.50 \times 10^5 \text{ Pa})(0.750 \text{ m}^3)}{(8.31 \text{ J/mol} \cdot \text{K})(325 \text{ K})} = \boxed{41.7 \text{ mol}}$$

(b) For an adiabatic process, $P_i V_i^\gamma = P_f V_f^\gamma$ where $\gamma = C_p/C_v = 7/5$ is the adiabatic index for a diatomic gas. Solve for the final gas pressure to find

$$P_f = P_i \left(\frac{V_i}{V_f} \right)^\gamma = (1.50 \times 10^5 \text{ Pa}) \left(\frac{0.750 \text{ m}^3}{1.50 \text{ m}^3} \right)^{\frac{7}{5}} = \boxed{5.68 \times 10^4 \text{ Pa}}$$

(c) To find the final gas temperature, take the ratio of the ideal gas

law for quantities in the final and initial states, and solve for T_f :

$$\frac{P_f V_f}{P_i V_i} = \frac{nRT_f}{nRT_i}$$

$$T_f = \left(\frac{P_f V_f}{P_i V_i} \right) T_i = \left(\frac{(5.68 \times 10^4 \text{ Pa})(1.50 \text{ m}^3)}{(1.50 \times 10^5 \text{ Pa})(0.750 \text{ m}^3)} \right) (325 \text{ K}) = \boxed{246 \text{ K}}$$

(d) From the first law of thermodynamics, the work done on a

diatomic gas during an adiabatic ($Q = 0$) process is

$$\Delta U = Q + W \rightarrow W = \Delta U = nC_v \Delta T$$

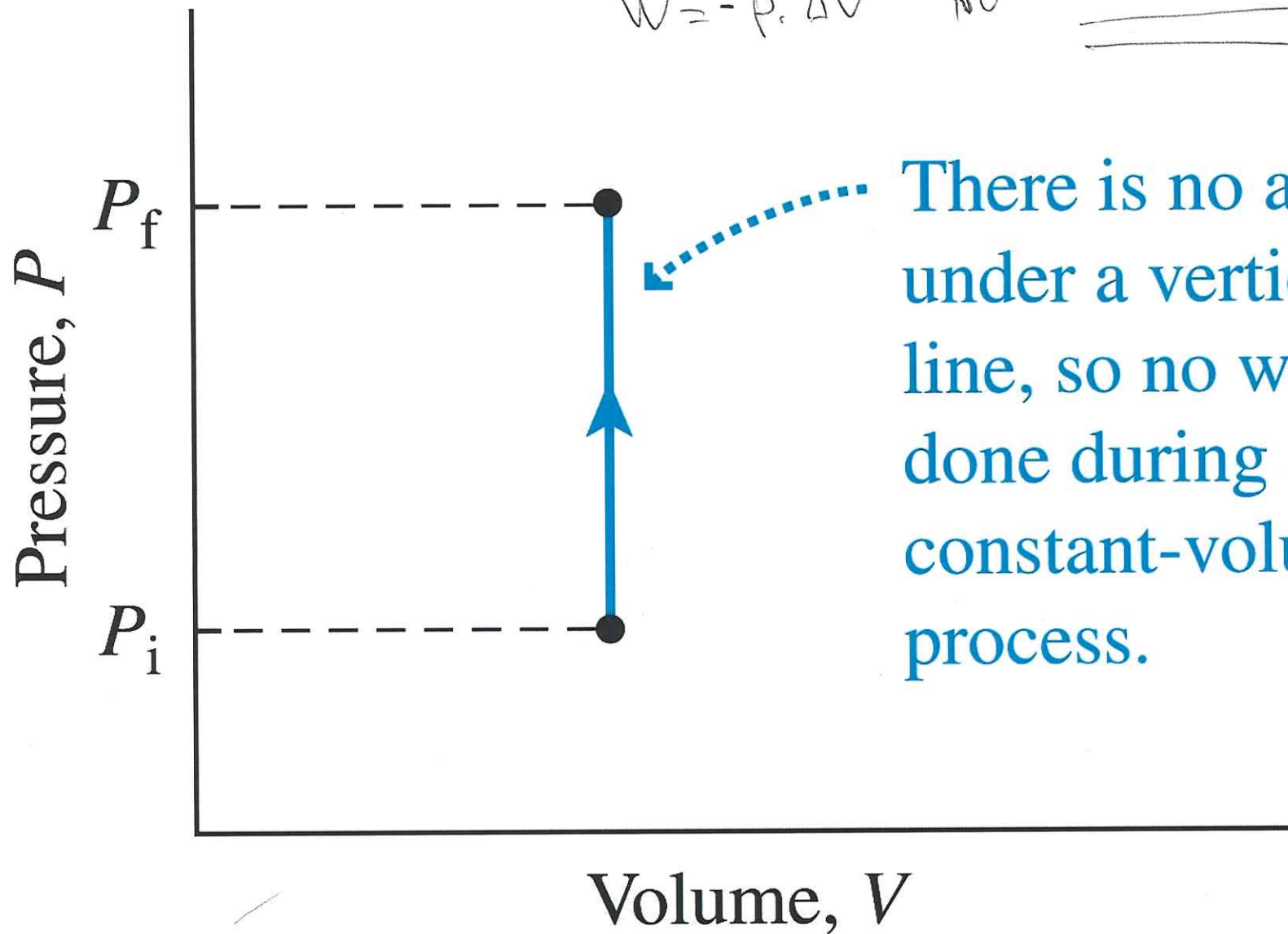
$$W = (41.7 \text{ mol}) \left(\frac{5}{2} R \right) (246 \text{ K} - 325 \text{ K}) = \boxed{-6.84 \times 10^4 \text{ J}}$$

Figure 14.7

Constant Volume

$$W = -P \cdot \Delta V \quad \text{no}$$

$$\Delta U = Q + W$$
$$W = 0$$



There is no area under a vertical line, so no work is done during a constant-volume process.

Isovolumetric Processes

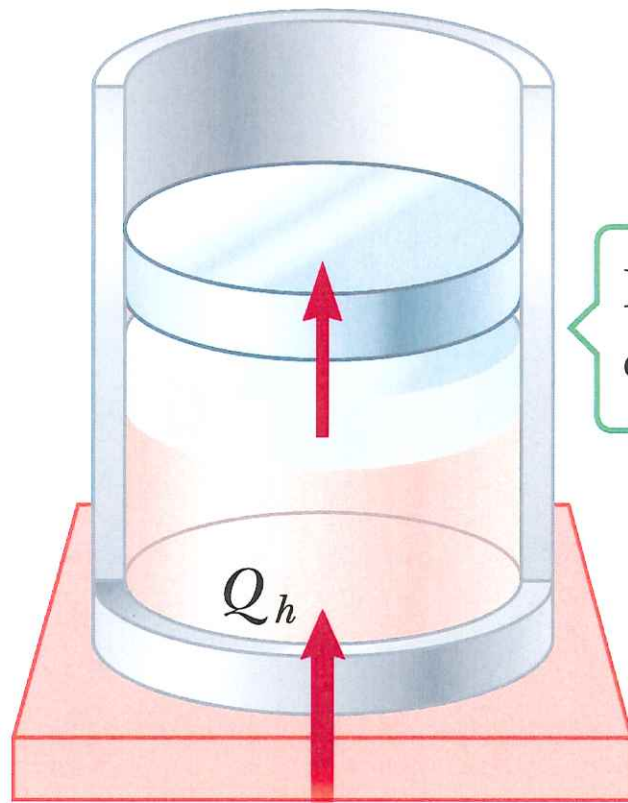
$W = 0 \rightarrow \Delta U = Q$ isovolumetric process

$$Q = nC_v\Delta T$$

Isothermal Processes

$$\Delta U = 0 \text{ because } \Delta T = 0$$

$$W = -Q \text{ (Isothermal Process)}$$

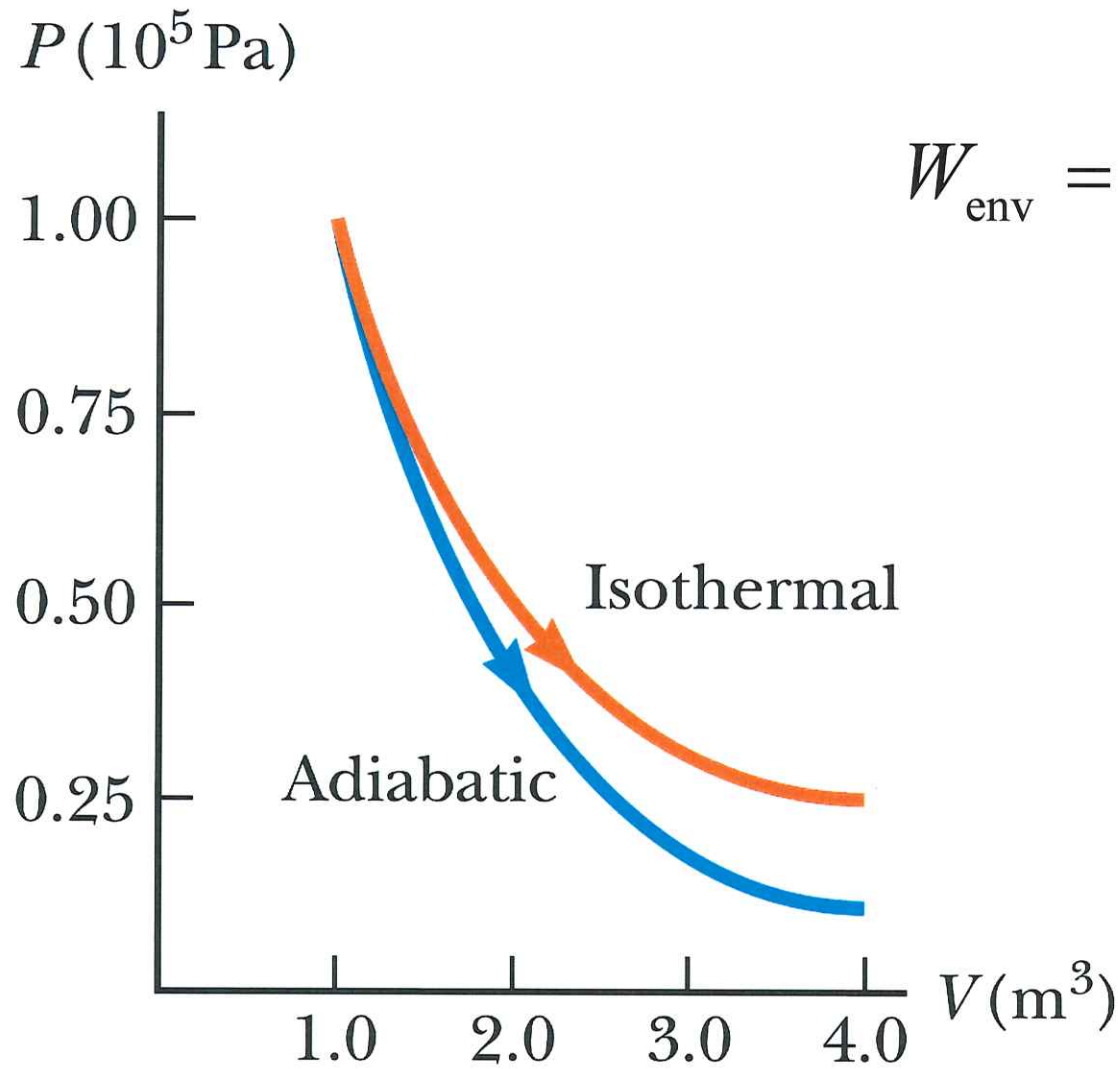


Isothermal
expansion

Energy reservoir at T_h

$$P = \frac{nRT}{V}$$

Isothermal Processes



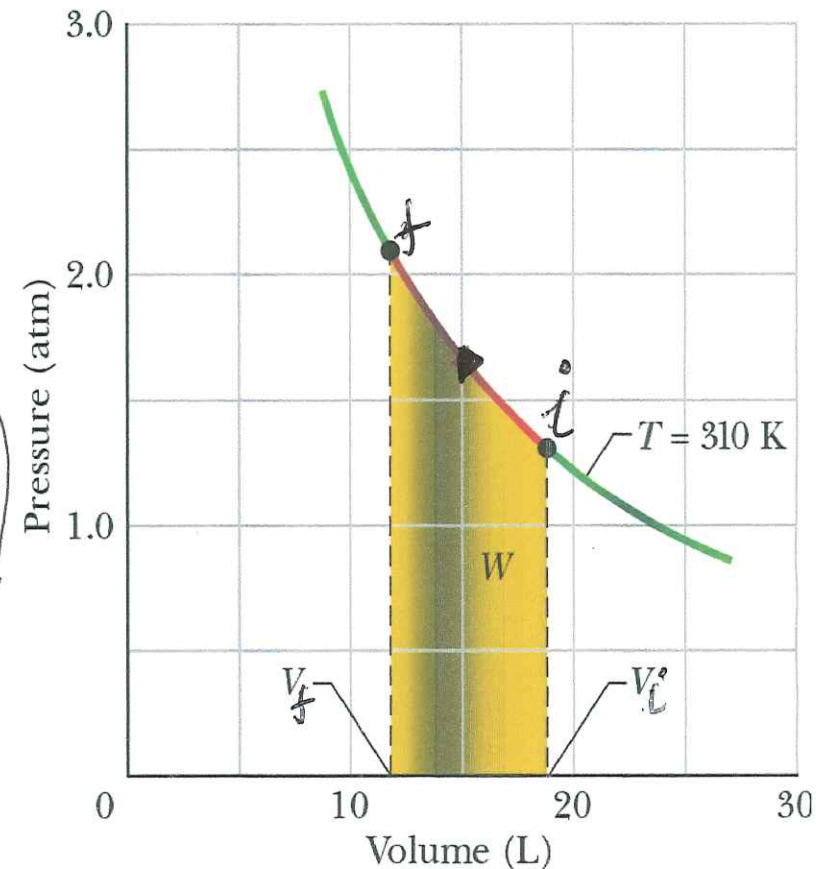
$$W_{\text{env}} = nRT \ln \left(\frac{V_f}{V_i} \right)$$

Sample Problem 19-2

One mole of oxygen (assume to be an ideal gas) *shrinks* at a constant temperature T of 310 K from an initial volume V_i of 19 L to a final volume V_f of 12 L. How much work is done *on* the gas during the *shrinking*?

$$\begin{aligned} W &= &= nRT \ln \frac{V_i}{V_f} \\ &= (1 \text{ mol})(8.31 \text{ J/mol} \cdot \text{K})(310 \text{ K}) \ln \left(\frac{19 \text{ L}}{12 \text{ L}} \right) \\ &= 1183 \text{ J} \end{aligned}$$

$$1 \text{ L} = 10^{-3} \text{ m}^3$$



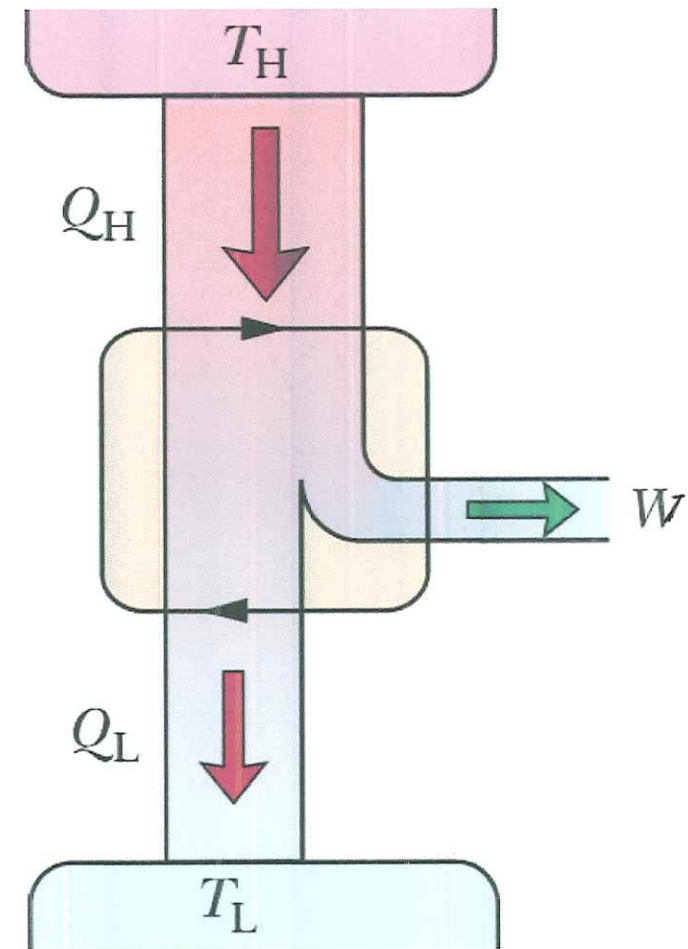
Thermal Processes in Gases

Table 12.2 The First Law and Thermodynamic Processes (Ideal Gases)

Process	ΔU	Q	W
Isobaric	$nC_v \Delta T$	$nC_p \Delta T$	$-P \Delta V$
Adiabatic	$nC_v \Delta T$	0	ΔU
Isovolumetric	$nC_v \Delta T$	ΔU	0
Isothermal	0	$-W$	$-nRT \ln \left(\frac{V_f}{V_i} \right)$
General	$nC_v \Delta T$	$\Delta U - W$	(PV Area)

Heat Engines

- Elements of an engine (
 - Heat Q_H is transferred from the hot reservoir of temperature T_H to the working substance
 - Heat Q_L is transferred from the working substance to the cold reservoir T_L



$$\epsilon = \frac{W}{Q_H}$$

- **Thermal efficiency (or efficiency):**

$$\varepsilon = \frac{\text{energy we get}}{\text{energy we pay for}} = \frac{|W|}{|Q_H|}$$

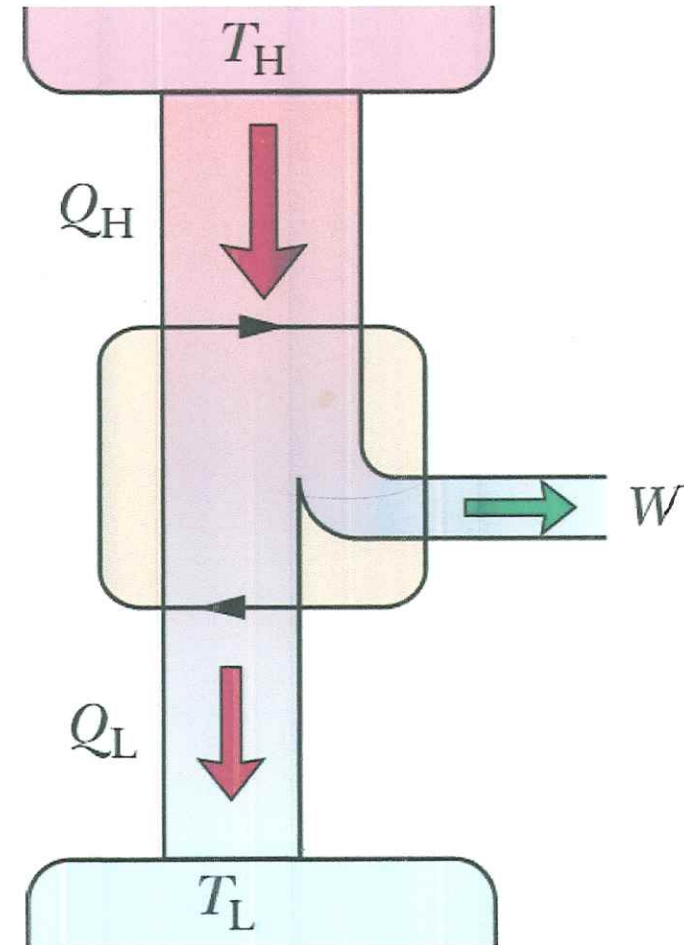
- Engines work in **cycles**; if reversible

Energy is conserved

$$\Delta E_{\text{int}} = 0 = (|Q_H| - |Q_L| - W)$$

$$W = Q_H - Q_L$$

$$\varepsilon = \frac{|W|}{|Q_H|} = \frac{|Q_H| - |Q_L|}{|Q_H|} = 1 - \frac{|Q_L|}{|Q_H|}$$



35. The work done by an engine equals one-fourth the energy it absorbs from a reservoir.

a. What is its thermal efficiency?

Answer ↓

b. What fraction of the energy absorbed is expelled to the cold reservoir?

12.35 (a) The efficiency of a heat engine is $e = W_{\text{eng}}/|Q_h|$, where W_{eng} is the work done by the engine and $|Q_h|$ is the energy absorbed from the higher temperature reservoir. Thus, if $W_{\text{eng}} = |Q_h|/4$, the efficiency is $e = 1/4 = \boxed{0.25 \text{ or } 25\%}$.

(b) From conservation of energy, the energy exhausted to the lower temperature reservoir is $|Q_c| = |Q_h| - W_{\text{eng}}$. Therefore, if $W_{\text{eng}} = |Q_h|/4$, we have $|Q_c| = 3|Q_h|/4$ or $\boxed{|Q_c|/|Q_h| = 3/4}$.

36. In each cycle of its operation, a heat engine expels 2 400 J of energy and performs 1 800 J of mechanical work.

a. How much thermal energy must be added to the engine in each cycle?

b. Find the thermal efficiency of the engine.

12.36 (a) The work done by a heat engine equals the net energy absorbed

by the engine, or $W_{\text{eng}} = |Q_h| - |Q_c|$. Thus, the energy absorbed

from the high temperature reservoir is

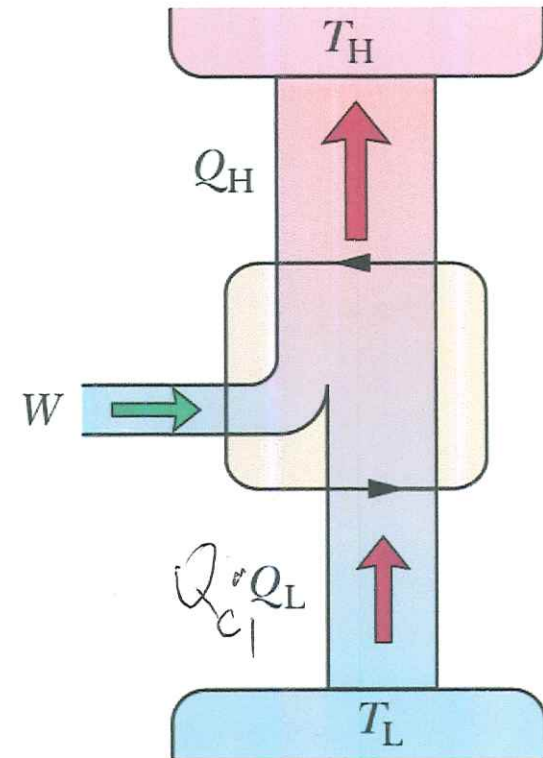
$$|Q_h| = W_{\text{eng}} + |Q_c| = 1\,800\text{ J} + 2\,400\text{ J} = \boxed{4\,200\text{ J}}$$

(b) The efficiency of the heat engine is

$$e \equiv \frac{W_{\text{eng}}}{|Q_h|} = \frac{1\,800\text{ J}}{4\,200\text{ J}} = 0.43 \text{ or } \boxed{43\%}$$

Refrigerators

- **Refrigerator:** device that uses work to transfer thermal energy from the low-temperature reservoir to the high-temperature reservoir (Fig 20-13)
- **Ideal refrigerator:** processes involved in the refrigerator's operations are reversible



$$W = Q_H - Q_L$$

$$Q_L + W = Q_H$$

conservation

- Coefficient of performance:

$$\text{COP} = K = \frac{\text{what we want}}{\text{what we pay for}} = \frac{|Q_L| \rightarrow \text{benefit}}{|W| \rightarrow \text{Energy cost}}$$

$$|W| = |Q_H| - |Q_L| \quad \leftarrow \quad Q_L + W = Q_H$$

$$K = \frac{|Q_L|}{|Q_H| - |Q_L|} = \frac{1}{\frac{Q_H}{Q_L} - 1} \quad (\text{from } 2 \sim 4)$$

- Carnot (ideal) refrigerator:

$$K_C = \frac{T_L}{T_H - T_L}$$

$$K = \frac{\text{what we want}}{\text{what we pay for}} = \frac{|Q_L|}{|W|}$$

$$|W| = |Q_H| - |Q_L|$$

$$K = \frac{|Q_L|}{|Q_H| - |Q_L|}$$

Problem

A Carnot refrigerator does 200 J of work to remove 600 J from its cold compartment. (a) What is the refrigerator's coefficient of performance? (b) How much energy per cycle is exhausted to the kitchen as heat?

34. (a) We use Eq. 21-12,

$$\frac{|Q_L|}{|Q_H| - |Q_L|} = K = \frac{|Q_L|}{|W|} = \frac{600 \text{ J}}{200 \text{ J}} = 3.$$

(b) Energy conservation for a refrigeration cycle requires $|Q_L| + |W| = |Q_H|$, so that the result is 800 J.

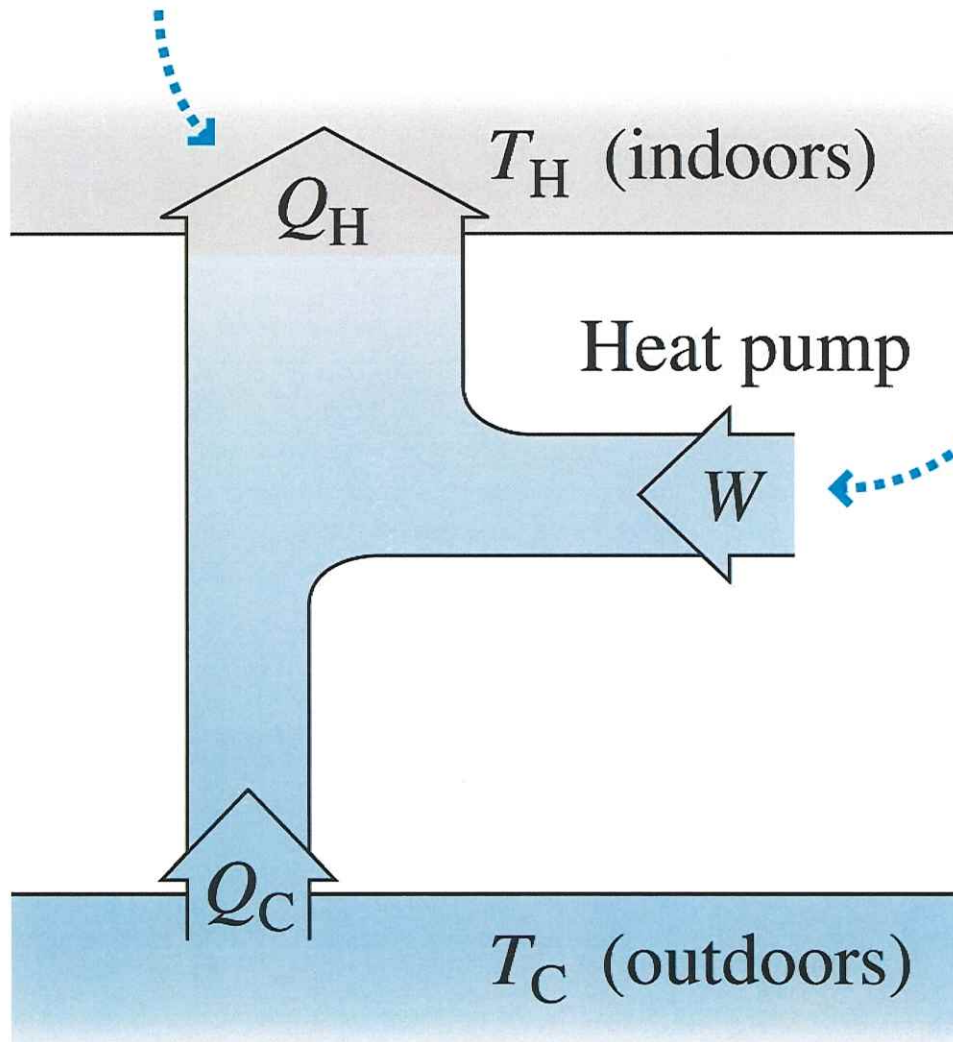
$$Q_H - Q_L = W$$

$$Q_H = W + Q_L = 200 \text{ J} + 600 \text{ J} = 800 \text{ J}$$

Air conditioners and Heat pumps

Figure 14.15

Heat pump uses energy W to extract heat Q_C from outdoors and deposit heat Q_H indoors. By conservation of energy, $Q_H = W + Q_C$.



Handwritten notes:

$$COP = \frac{Q_H}{W} = \frac{W + Q_C}{W}$$

$$= 1 + \frac{Q_C}{W}$$

Heat pump

Note air conditioners

$$COP = \frac{Q_C}{W}$$

like refrigerator

40. A heat pump has a coefficient of performance of 3.80 and operates with a power consumption of $7.03 \times 10^3 \text{ W}$.

- How much energy does the heat pump deliver into a home during 8.00 h of continuous operation?
- How much energy does it extract from the outside air in 8.00 h?

12.40 (a) The coefficient of performance of a heat pump is $\text{COP} = |Q_h|/W$,

where $|Q_h|$ is the thermal energy delivered to the warm space

and W is the work input required to operate the heat pump.

Therefore,

$$\begin{aligned} |Q_h| &= W \cdot \text{COP} = (P \cdot \Delta t) \cdot \text{COP} \\ &= \left[\left(7.03 \times 10^3 \frac{\text{J}}{\text{s}} \right) \left(8.00 \text{ h} \right) \left(\frac{3600 \text{ s}}{1 \text{ h}} \right) \right] 3.80 = \boxed{7.69 \times 10^6 \text{ J}} \end{aligned}$$

(b) The energy extracted from the cold space (outside air) is

$$|Q_c| = |Q_h| - W = |Q_h| - \frac{|Q_h|}{\text{COP}} = |Q_h| \left(1 - \frac{1}{\text{COP}} \right)$$

$$\text{or } |Q_c| = (7.69 \times 10^6 \text{ J}) \left(1 - \frac{1}{3.80} \right) = \boxed{5.67 \times 10^6 \text{ J}}$$