



Chapter 2. Heat and the First Law of Thermodynamics

PHYS242: Thermal Physics
Dr. Salem Marhaba
Faculty of Science
Department of Physics

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Chapter Outline

- **Introduction**
- **Thermodynamic system**
- **Internal Energy**
- **Heat**
- **Mechanical Equivalent of Heat**
- **Heat Capacity and Specific Heat**
- **Calorimetry**
- **Phase Changes**
- **Mechanisms of Energy Transfer**
- **State Variables and Transfer Variables**
- **The First Law of Thermodynamics**
- **Adiabatic Process, Isobaric Processes, Isovolumetric Processes and Isothermal Process**

Thermodynamics – Historical Background

Thermodynamics and mechanics were considered to be distinct branches of physics.

- Until about 1850
- Experiments by James Joule and others showed a connection between them.

A connection was found between the transfer of energy by heat in thermal processes and the transfer of energy by work in mechanical processes.

The concept of energy was generalized to include internal energy.

The principle of conservation of energy emerged as a universal law of nature.

Thermodynamics – Chapter Overview

Will discuss internal energy, the first law of thermodynamics, and applications of the first law

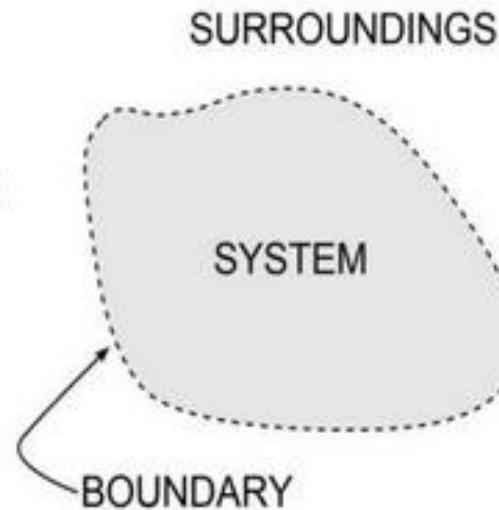
The first law of thermodynamics describes systems in which the only energy change is that of internal energy.

The transfers of energy are by heat and work.

Will consider work done on deformable systems

Thermodynamic systems

- System:-Anything which are kept in consideration is called System.
- Surrounding:-Everything external to the system is known as Surrounding.
- Boundary:-The thermodynamic system and surroundings are separated by an envelope called Boundary of a system.
- Universe:-Thermodynamic system+surrounding=Universe

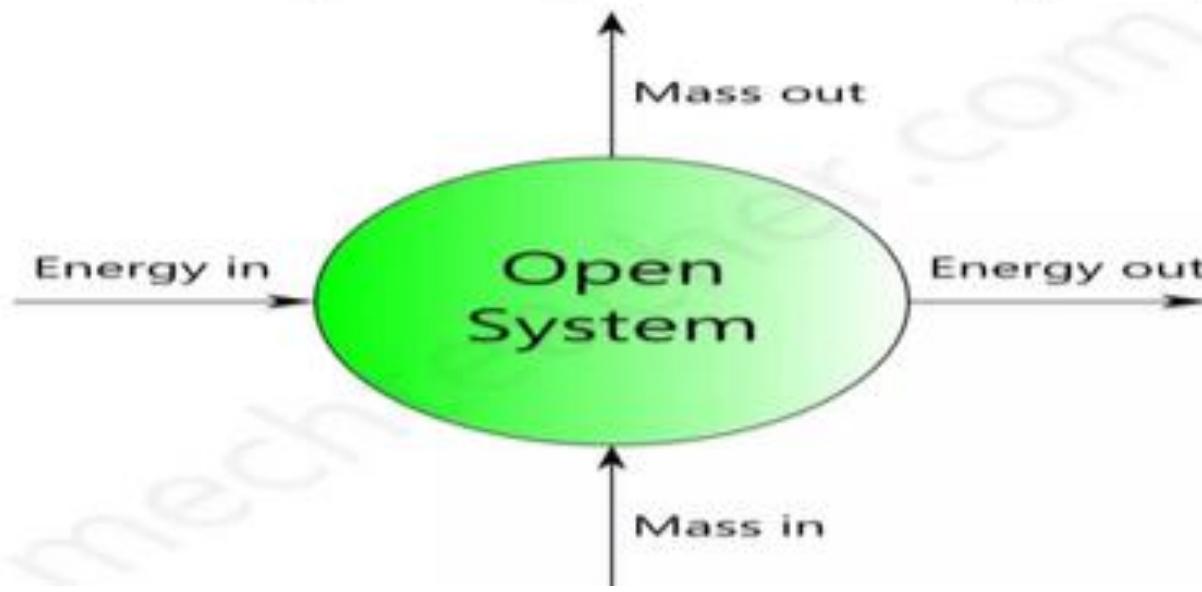


Thermodynamic systems

- open system
- closed system
- isolated system

Thermodynamic systems

- Open system:-in open system, the mass as well as energy transfer may take place between system and its surroundings.
- Most of Engineering devices are open system.



Thermodynamic systems

Example of open system:

- Internal combustion engines
- Air compressor
- Water pump
- Steam engine
- Boiler

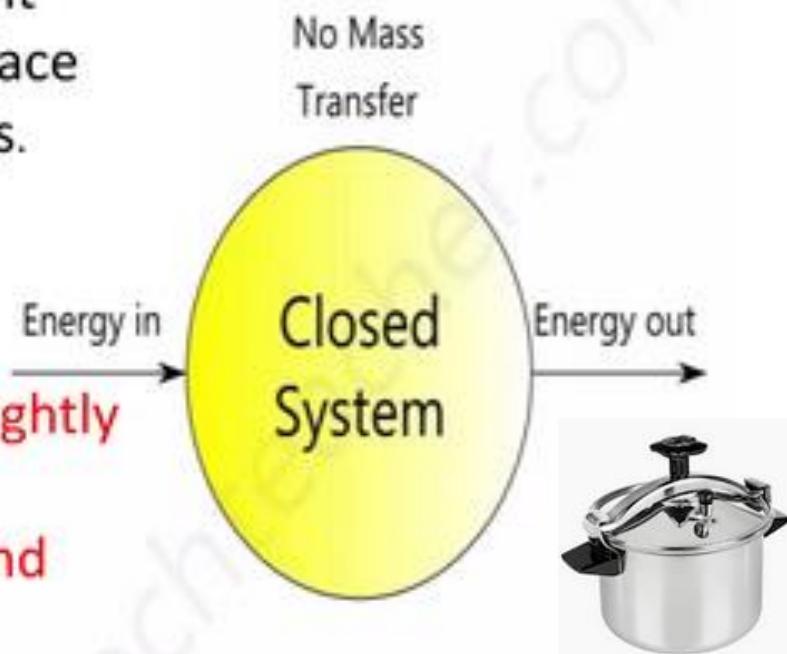
Thermodynamic systems

Closed system:-

*In closed system, the mass within the boundary of the system remains constant and only the energy transfer may take place between the system and its surroundings.

Examples of closed system

1. Pressure cooker
2. A rubber balloon filled with air and tightly closed
3. The gas confined between a piston and cylinder



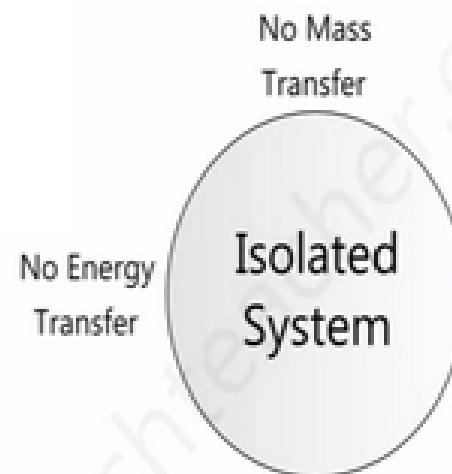
Thermodynamic systems

Isolated system:-

*In an isolated system, neither mass nor energy transfer takes place between the system and its surroundings.

Examples of isolated system

- 1.Thermos flask
- 2.The universe



Internal Energy

Internal energy is all the energy of a system that is associated with its microscopic components.

- These components are its atoms and molecules.
- The system is viewed from a reference frame at rest with respect to the center of mass of the system.

Internal Energy and Other Energies

The kinetic energy due to its motion through space is not included.

Internal energy does include kinetic energies due to:

- Random translational motion
- Rotational motion
- Vibrational motion

Internal energy also includes potential energy between molecules.

Heat

Heat is defined as the transfer of energy across the boundary of a system due to a temperature difference between the system and its surroundings.

The term heat will also be used to represent the amount of energy transferred by this method.

There are many common phrases that use the word “heat” incorrectly.

Heat, internal energy, and temperature are all different quantities.

- Be sure to use the correct definition of heat.
- You cannot talk about the “heat of a system,” you can refer to heat only when energy has been transferred as a result of a temperature difference.

Units of Heat

Historically, the **calorie** was the unit used for heat.

- One calorie is the amount of energy transfer necessary to raise the temperature of 1 g of water from 14.5°C to 15.5°C.
 - The “Calorie” used for food is actually 1 kilocalorie.

The standard in the text is to use Joules.

James Prescott Joule

1818 – 1889

British physicist

Largely self-educated

- Some formal education from John Dalton

Research led to establishment of the principle of conservation of energy

Determined the amount of work needed to produce one unit of energy



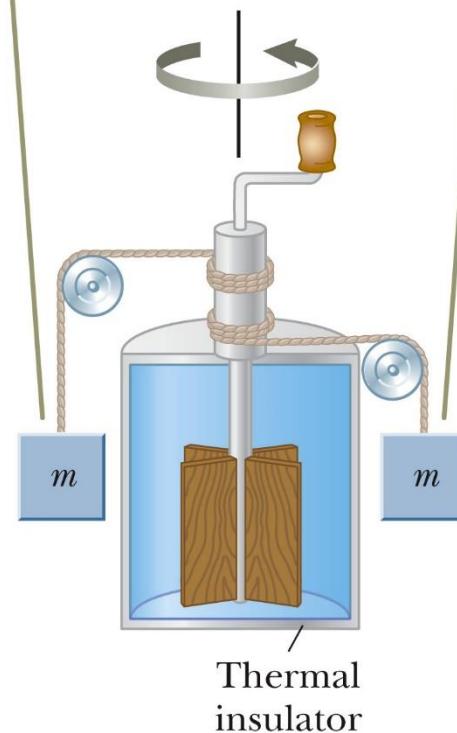
Mechanical Equivalent of Heat

Joule established the equivalence between mechanical energy and internal energy.

His experimental setup is shown at right.

The decrease in potential energy associated of the system as the blocks fall equals the work done by the paddle wheel on the water.

The falling blocks rotate the paddles, causing the temperature of the water to increase.



Mechanical Equivalent of Heat, cont

Joule found that it took approximately 4.18 J of mechanical energy to raise the water 1°C.

Later, more precise, measurements determined the amount of mechanical energy needed to raise the temperature of water from 14.5°C to 15.5°C.

$$1 \text{ cal} = 4.186 \text{ J}$$

- This is known as the **mechanical equivalent of heat**.

A student eats a dinner rated at 2 000 Calories. He wishes to do an equivalent amount of work in the gymnasium by lifting a 50.0-kg barbell. How many times must he raise the barbell to expend this much energy? Assume that he raises the barbell 2.00 m each time he lifts it and that he regains no energy when he lowers the barbell.

Solution Because 1 Calorie = 1.00×10^3 cal, the total amount of work required to be done on the barbell–Earth system is 2.00×10^6 cal. Converting this value to joules, we have

$$W = (2.00 \times 10^6 \text{ cal})(4.186 \text{ J/cal}) = 8.37 \times 10^6 \text{ J}$$

The work done in lifting the barbell a distance h is equal to mgh , and the work done in lifting it n times is $nmgh$. We equate this to the total work required:

$$W = nmgh = 8.37 \times 10^6 \text{ J}$$

$$n = \frac{W}{mgh} = \frac{8.37 \times 10^6 \text{ J}}{(50.0 \text{ kg})(9.80 \text{ m/s}^2)(2.00 \text{ m})}$$

$$= 8.54 \times 10^3 \text{ times}$$

Heat Capacity

The **heat capacity**, C , of a particular sample is defined as the amount of energy needed to raise the temperature of that sample by 1°C .

If energy Q produces a change of temperature of ΔT , then $Q = C \Delta T$.

Specific Heat

Specific heat, c , is the heat capacity per unit mass.

If energy Q transfers to a sample of a substance of mass m and the temperature changes by ΔT , then the specific heat is

$$c \equiv \frac{Q}{m \Delta T}$$

The specific heat is essentially a measure of how thermally insensitive a substance is to the addition of energy.

- The greater the substance's specific heat, the more energy that must be added to a given mass to cause a particular temperature change.

The equation is often written in terms of Q : $Q = m c \Delta T$

Some Specific Heat Values

TABLE 20.1*Specific Heats of Some Substances at 25°C and Atmospheric Pressure*

Substance	Specific Heat (J/kg · °C)	Substance	Specific Heat (J/kg · °C)
<i>Elemental solids</i>			
Aluminum	900	Brass	380
Beryllium	1 830	Glass	837
Cadmium	230	Ice (-5°C)	2 090
Copper	387	Marble	860
Germanium	322	Wood	1 700
Gold	129	<i>Liquids</i>	
Iron	448	Alcohol (ethyl)	2 400
Lead	128	Mercury	140
Silicon	703	Water (15°C)	4 186
Silver	234	<i>Gas</i>	
		Steam (100°C)	2 010

Note: To convert values to units of cal/g · °C, divide by 4 186.

1- Imagine you have 1 kg each of iron, glass, and water, and that all three samples are at 10° C. Rank the samples from lowest to highest temperature after 100 J of energy is added to each sample.

2- Considering the same samples as in question 1, rank them from least to greatest amount of energy transferred by heat if each sample increases in temperature by 20° C.

Internal Energy Revisited

The change in internal energy of a system can be identified with $m c \Delta t$.

- If we ignore any thermal expansion or contraction of the system.

Then $\Delta E_{\text{int}} = Q$

The internal energy of a system can be changed by transferring energy into the system by any mechanism.

- This also indicates temperature is related to the energy of the molecules of a system.

Specific Heat Varies With Temperature

Technically, the specific heat varies with temperature.

The corrected equation is $Q = m \int_{T_i}^{T_f} c \, dT$.

However, if the temperature intervals are not too large, the variation can be ignored and c can be treated as a constant.

- For example, for water there is only about a 1% variation between 0° and 100°C.
- These variations will be neglected unless otherwise stated.

Specific Heat of Water

Water has the highest specific heat of common materials.

This is in part responsible for many weather phenomena:

- Moderate climates near large bodies of water
- Global wind systems
- Land and sea breezes

Calorimetry

One technique for measuring specific heat involves heating a material, adding it to a sample of water, and recording the final temperature.

This technique is known as **calorimetry**.

- A calorimeter is a device in which this energy transfer takes place.

The system of the sample and the water is isolated.

Conservation of energy requires that the amount of energy that leaves the sample equals the amount of energy that enters the water.

- Conservation of Energy gives a mathematical expression of this:

$$Q_{\text{cold}} = -Q_{\text{hot}}$$

Sign Conventions

If the temperature increases:

- Q and ΔT are positive
- Energy transfers into the system

If the temperature decreases:

- Q and ΔT are negative
- Energy transfers out of the system

The negative sign in the calorimetry equation is critical for consistency with the sign convention.

Calorimetry, final

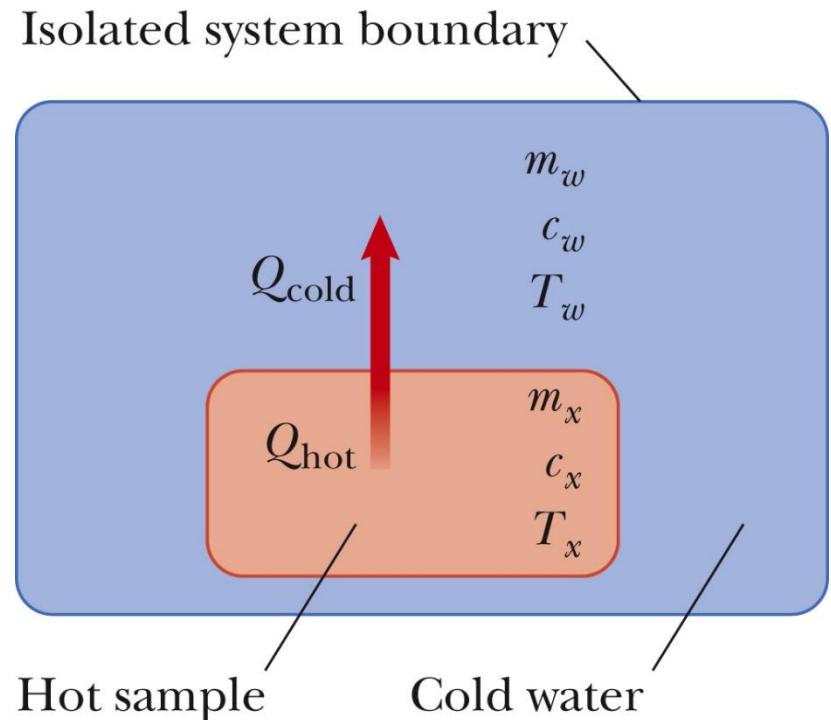
T_f is the final temperature after the system comes to equilibrium.

The subscript w represent values for water and x represents the values for the sample whose specific heat is to be determined.

Since each $Q = mc\Delta T$, the calorimetry equation can be expressed as

$$m_w c_w (T_f - T_w) = -m_x c_x (T_f - T_x)$$

- This can be solved for the unknown specific heat.
- Technically, the mass of the container should be included, but if $m_w \gg m_{\text{container}}$ it can be neglected.



Calorimetry, Example

A 0.050 0-kg ingot of metal is heated to 200.0°C and then dropped into a beaker containing 0.400 kg of water initially at 20.0°C. If the final equilibrium temperature of the mixed system is 22.4°C, find the specific heat of the metal.

$$m_w c_w (T_f - T_w) = - m_x c_x (T_f - T_x)$$

$$\begin{aligned}(0.400 \text{ kg}) (4186 \text{ J/kg} \cdot ^\circ\text{C}) (22.4^\circ\text{C} - 20.0^\circ\text{C}) \\= - (0.050 \text{ 0 kg}) (c_x) (22.4^\circ\text{C} - 200.0^\circ\text{C})\end{aligned}$$

$$c_x = 453 \text{ J/kg} \cdot ^\circ\text{C}$$

Phase Changes

A **phase change** is when a substance changes from one form to another.

- Two common phase changes are
 - Solid to liquid (melting)
 - Liquid to gas (boiling)

During a phase change, there is no change in temperature of the substance.

- For example, in boiling the increase in internal energy is represented by the breaking of the bonds between molecules, giving the molecules of the gas a higher intermolecular potential energy.

Latent Heat

Different substances react differently to the energy added or removed during a phase change.

- Due to their different internal molecular arrangements

The amount of energy also depends on the mass of the sample.

- The *higher-phase material* is the material existing at the higher temperature.
 - Example, water is the higher-phase material in an ice-water mixture
- The initial amount of the higher-phase material in a system is m_i .

If an amount of energy Q is required to change the phase of a sample is

$$L \equiv Q / \Delta m$$

- $\Delta m = m_f - m_i$ is the change in mass of the higher-phase material

Latent Heat, cont.

The quantity L is called the **latent heat** of the material.

- Latent means “hidden”.
- The value of L depends on the substance as well as the actual phase change.

The energy required to change the phase is $Q = L \Delta m$.

- Δm refers to the higher-phase material.
- If the entire amount of the lower-phase material undergoes a phase change, the change in mass of the higher-phase material is equal to initial mass of the lower-phase material□

Latent Heat, final

The *latent heat of fusion* is used when the phase change is from solid to liquid.

The *latent heat of vaporization* is used when the phase change is from liquid to gas.

If energy enters the system:

- This will result in melting or vaporization
- The amount of the higher-phase material will increase
- Δm and Q are positive

If energy is extracted from the system:

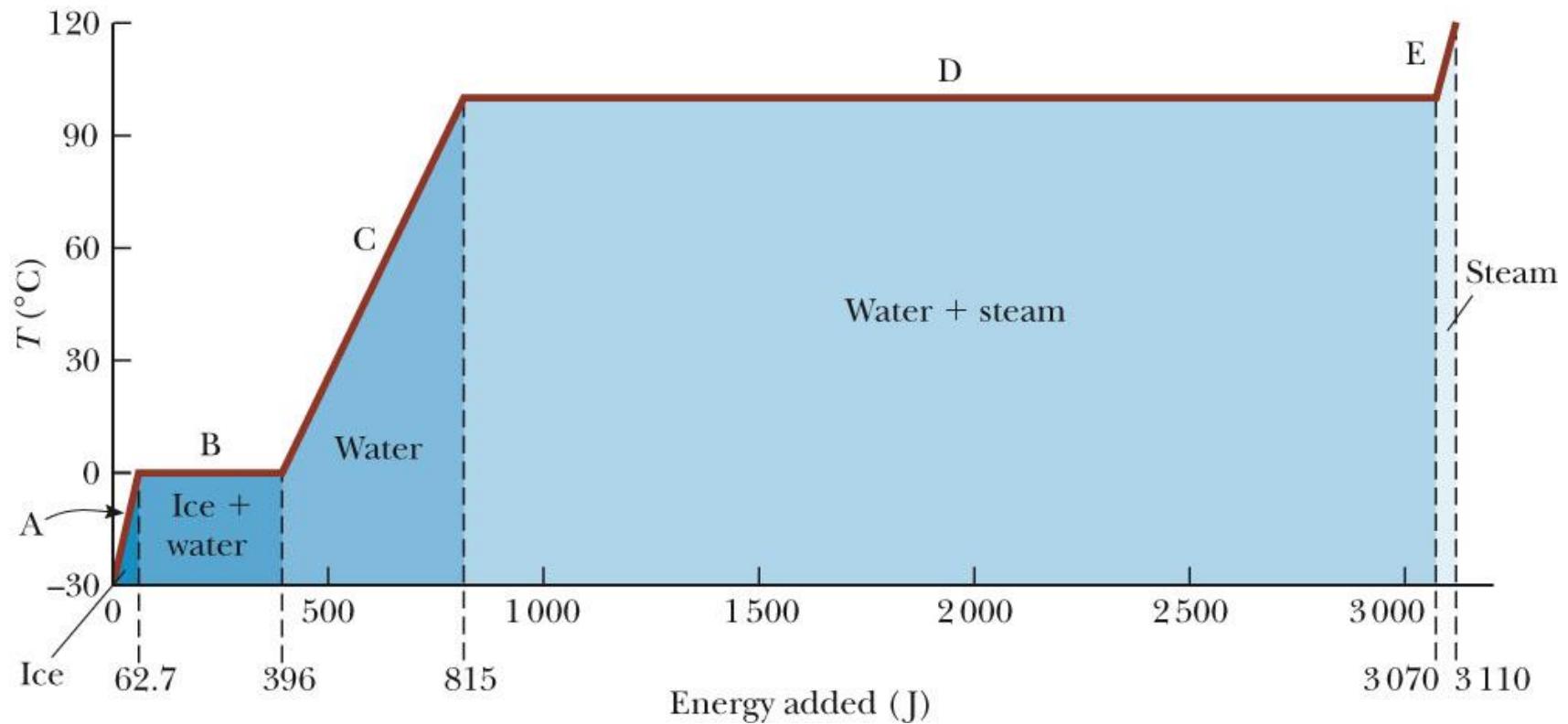
- This will result in freezing or condensation
- The amount of the higher-phase material will decrease
- Δm and Q are negative

Sample Latent Heat Values

TABLE 20.2 *Latent Heats of Fusion and Vaporization*

Substance	Melting Point (°C)	Latent Heat of Fusion (J/kg)	Boiling Point (°C)	Latent Heat of Vaporization (J/kg)
Helium	−269.65	5.23×10^3	−268.93	2.09×10^4
Oxygen	−218.79	1.38×10^4	−182.97	2.13×10^5
Nitrogen	−209.97	2.55×10^4	−195.81	2.01×10^5
Ethyl alcohol	−114	1.04×10^5	78	8.54×10^5
Water	0.00	3.33×10^5	100.00	2.26×10^6
Sulfur	119	3.81×10^4	444.60	3.26×10^5
Lead	327.3	2.45×10^4	1 750	8.70×10^5
Aluminum	660	3.97×10^5	2 450	1.14×10^7
Silver	960.80	8.82×10^4	2 193	2.33×10^6
Gold	1 063.00	6.44×10^4	2 660	1.58×10^6
Copper	1 083	1.34×10^5	1 187	5.06×10^6

Graph of Ice to Steam



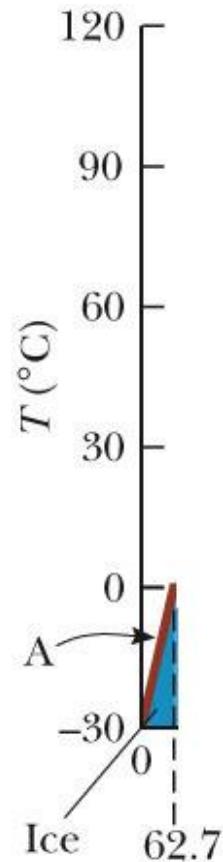
Warming Ice, Graph Part A

Start with one gram of ice at -30.0°C .

During phase A, the temperature of the ice changes from -30.0°C to 0°C .

Use $Q = m_i c_i \Delta T$

- In this case, 62.7 J of energy are added.



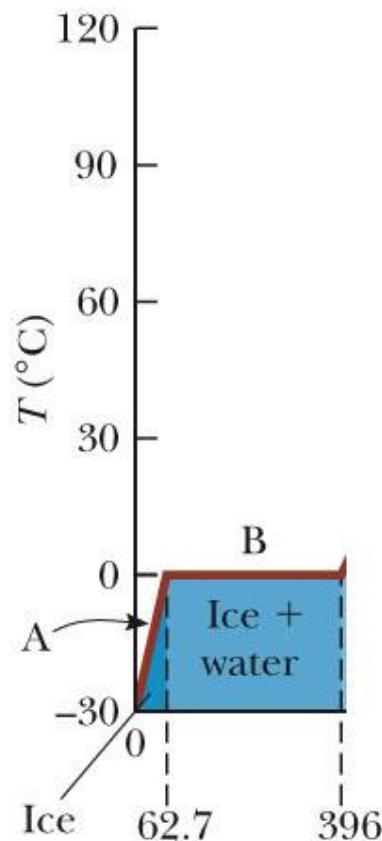
Melting Ice, Graph Part B

Once at 0°C, the phase change (melting) starts.

The temperature stays the same although energy is still being added.

Use $Q = L_f \Delta m_w = L_f m_i$

- The energy required is 333 J.
- On the graph, the values move from 62.7 J to 396 J.



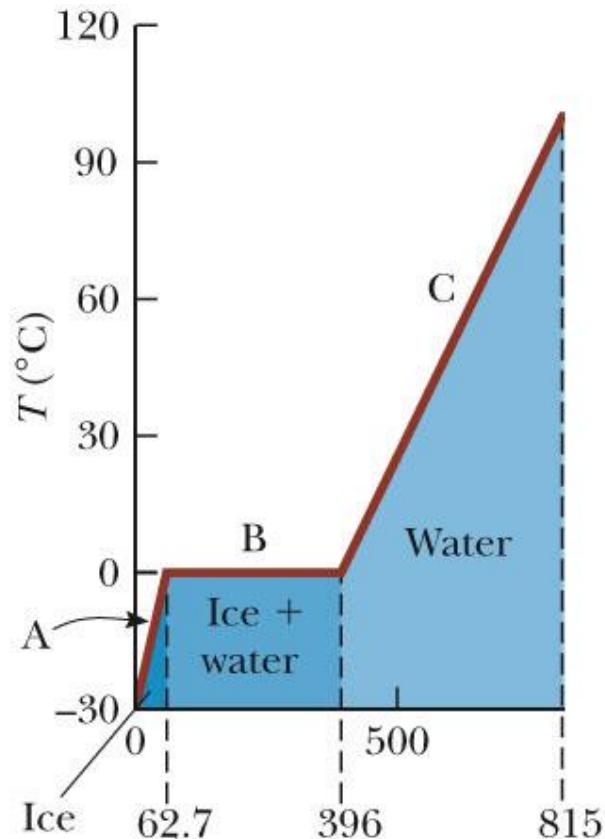
Warming Water, Graph Part C

Between 0°C and 100°C , the material is liquid and no phase changes take place.

Energy added increases the temperature.

Use $Q = m_w c_w \Delta T$

- 419 J are added
- The total is now 815 J



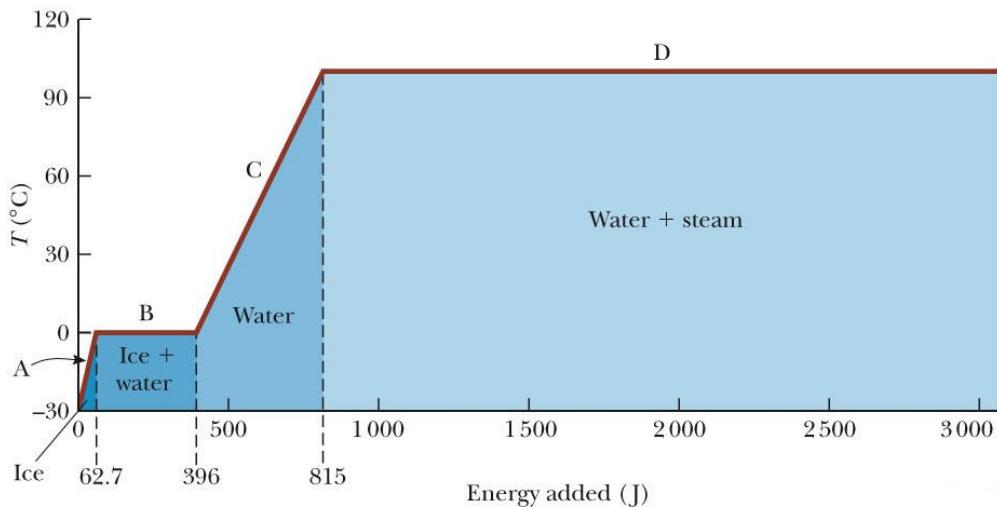
Boiling Water, Graph Part D

At 100°C, a phase change occurs (boiling).

Temperature does not change.

Use $Q = L_v \Delta m_s = L_v m_w$

- This requires 2260 J
- The total is now 3070 J



Heating Steam, Graph Part E

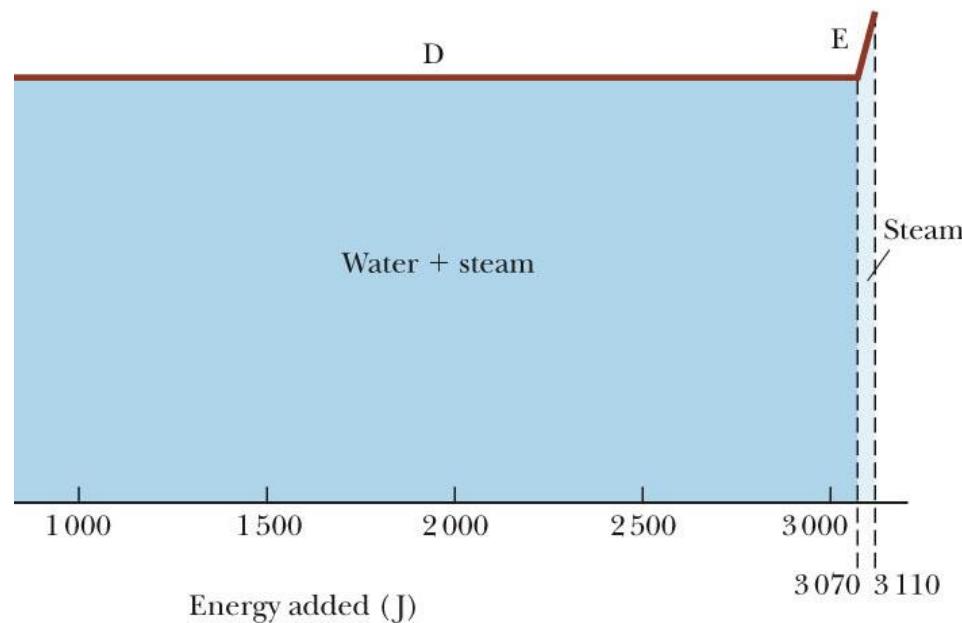
After all the water is converted to steam, the steam will heat up.

No phase change occurs.

The added energy goes to increasing the temperature.

Use $Q = m_s c_s \Delta T$

- In this case, 40.2 J are needed.
- The temperature is increasing from 100° C to 120° C .
- The total is now 3110 J.



Mechanisms of Energy Transfer In Thermal Processes

We want to know the rate at which energy is transferred.

There are various mechanisms responsible for the transfer:

- Conduction
- Convection
- Radiation

Conduction

The transfer can be viewed on an atomic scale.

- It is an exchange of kinetic energy between microscopic particles by collisions.
 - The microscopic particles can be atoms, molecules or free electrons.
- Less energetic particles gain energy during collisions with more energetic particles.

Rate of conduction depends upon the characteristics of the substance.

Conduction, cont.

In general, metals are good thermal conductors.

- They contain large numbers of electrons that are relatively free to move through the metal.
- They can transport energy from one region to another.

Poor conductors include asbestos, paper, and gases.

Conduction can occur only if there is a difference in temperature between two parts of the conducting medium.

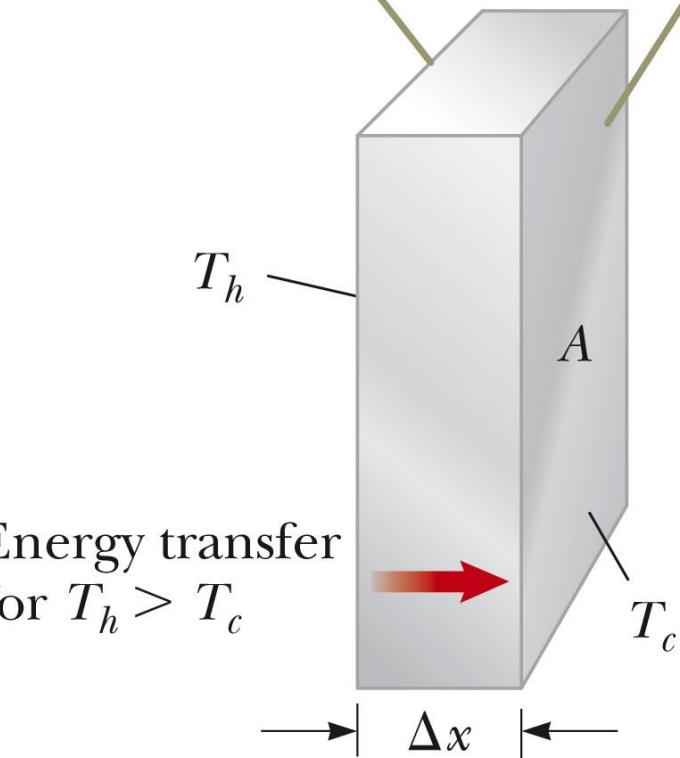
Conduction, equation

The slab at right allows energy to transfer from the region of higher temperature to the region of lower temperature.

The rate of transfer is given by:

$$P = \frac{Q}{\Delta t} = kA \left| \frac{dT}{dx} \right|$$

The opposite faces are at different temperatures where $T_h > T_c$.



Energy transfer
for $T_h > T_c$

Conduction, equation explanation

A is the cross-sectional area.

dT is the temperature difference.

dx is the thickness of the slab.

- Or the length of a rod

P is in Watts when Q is in Joules and t is in seconds.

k is the thermal conductivity of the material.

- Good conductors have high k values and good insulators have low k values.

Some Thermal Conductivities

TABLE 20.3

Thermal Conductivities

Substance	Thermal Conductivity (W/m · °C)
<i>Metals (at 25°C)</i>	
Aluminum	238
Copper	397
Gold	314
Iron	79.5
Lead	34.7
Silver	427
<i>Nonmetals (approximate values)</i>	
Asbestos	0.08
Concrete	0.8
Diamond	2 300
Glass	0.8
Ice	2
Rubber	0.2
Water	0.6
Wood	0.08
<i>Gases (at 20°C)</i>	
Air	0.023 4
Helium	0.138
Hydrogen	0.172
Nitrogen	0.023 4
Oxygen	0.023 8

Temperature Gradient

The quantity $|dT / dx|$ is called the **temperature gradient** of the material.

- It measures the rate at which temperature varies with position.

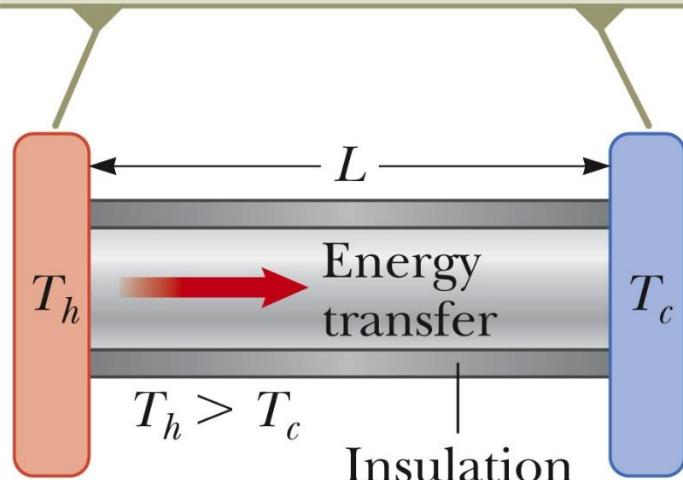
For a rod, the temperature gradient can be expressed as:

$$\left| \frac{dT}{dx} \right| = \frac{T_h - T_c}{L}$$

Using the temperature gradient for the rod, the rate of energy transfer becomes:

$$P = kA \left(\frac{T_h - T_c}{L} \right)$$

The opposite ends of the rod are in thermal contact with energy reservoirs at different temperatures.



Compound Slab

For a compound slab containing several materials of various thicknesses (L_1, L_2, \dots) and various thermal conductivities (k_1, k_2, \dots) the rate of energy transfer depends on the materials and the temperatures at the outer edges:

$$P = \frac{A(T_h - T_c)}{\sum_i (L_i/k_i)}$$

Home Insulation

Substances are rated by their R-values.

- $R = L / k$ and the rate becomes

$$P = \frac{A(T_h - T_c)}{\sum_i R_i}$$

- For multiple layers, the total R value is the sum of the R values of each layer.

Wind increases the energy loss by conduction in a home.

Insulation Values

TABLE 20.4 R-Values for Some Common Building Materials

Material	R-value ($\text{ft}^2 \cdot {}^\circ\text{F} \cdot \text{h/Btu}$)
Hardwood siding (1 in. thick)	0.91
Wood shingles (lapped)	0.87
Brick (4 in. thick)	4.00
Concrete block (filled cores)	1.93
Fiberglass insulation (3.5 in. thick)	10.90
Fiberglass insulation (6 in. thick)	18.80
Fiberglass board (1 in. thick)	4.35
Cellulose fiber (1 in. thick)	3.70
Flat glass (0.125 in. thick)	0.89
Insulating glass (0.25-in. space)	1.54
Air space (3.5 in. thick)	1.01
Stagnant air layer	0.17
Drywall (0.5 in. thick)	0.45
Sheathing (0.5 in. thick)	1.32

Convection

Energy transferred by the movement of a substance.

It is a form of matter transfer:

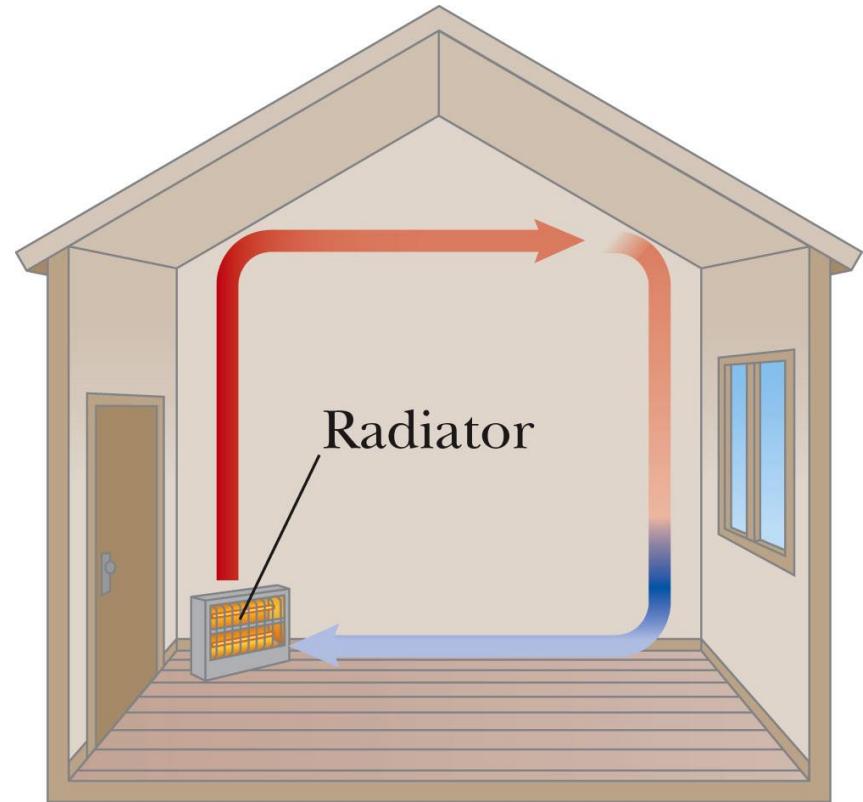
- When the movement results from differences in density, it is called *natural convection*.
- When the movement is forced by a fan or a pump, it is called *forced convection*.

Convection example

Air directly above the radiator is warmed and expands.

The density of the air decreases, and it rises.

A continuous air current is established.



Radiation

Radiation does not require physical contact.

All objects radiate energy continuously in the form of electromagnetic waves due to thermal vibrations of their molecules.

Rate of radiation is given by **Stefan's law**.

Stefan's Law

$$P = \sigma A e T^4$$

- P is the rate of energy transfer, in Watts.
- $\sigma = 5.6696 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4$
- A is the surface area of the object.
- e is a constant called the emissivity.
 - e varies from 0 to 1
 - The emissivity is also equal to the absorptivity.
- T is the temperature in Kelvins.

Ideal Absorbers

An *ideal absorber* is defined as an object that absorbs all of the energy incident on it.

- $e = 1$

This type of object is called a **black body**.

An ideal absorber is also an ideal radiator of energy.

Energy Absorption and Emission by Radiation

With its surroundings, the rate at which the object at temperature T with surroundings at T_o radiates is

- $P_{\text{net}} = \sigma A e (T^4 - T_o^4)$
- When an object is in equilibrium with its surroundings, it radiates and absorbs at the same rate.
 - Its temperature will not change.

State Variables

State variables describe the state of a system.

Variables may include:

- Pressure, temperature, volume, internal energy

The state of an isolated system can be specified only if the system is in thermal equilibrium internally.

- For a gas in a container, this means every part of the gas must be at the same pressure and temperature.

Transfer Variables

Transfer variables are zero unless a process occurs in which energy is transferred across the boundary of a system.

Transfer variables are not associated with any given state of the system, only with changes in the state.

- Heat and work are transfer variables.

Transfer variable can be positive or negative, depending on whether energy is entering or leaving the system.

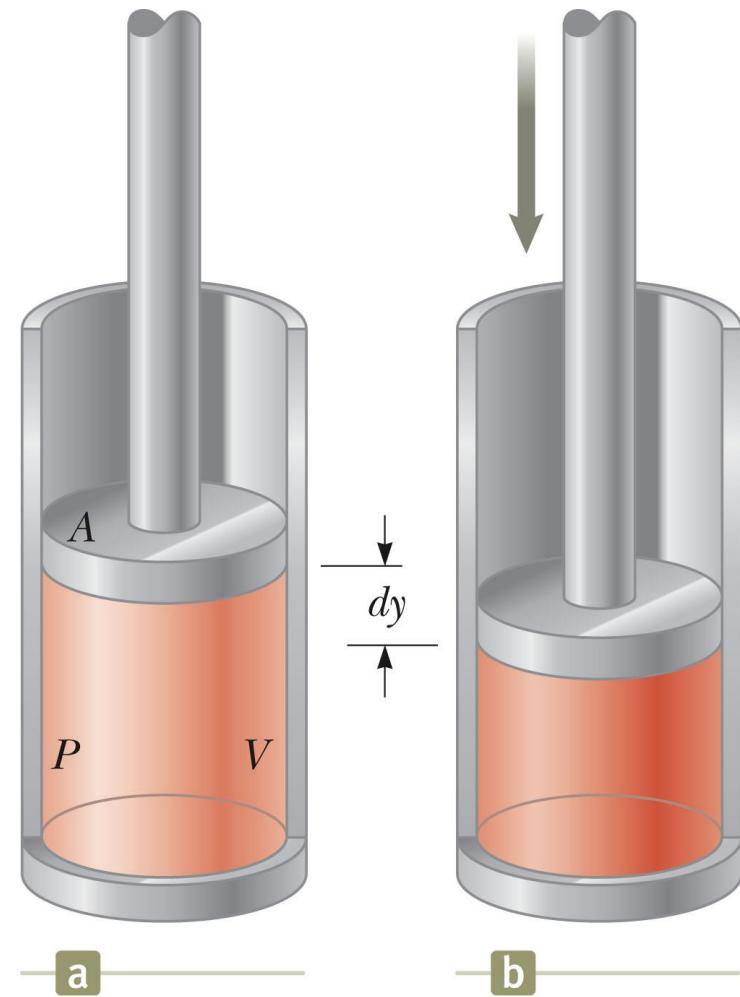
Work in Thermodynamics

Work can be done on a deformable system, such as a gas.

Consider a cylinder with a moveable piston.

A force is applied to slowly compress the gas.

- The compression is slow enough for all the system to remain essentially in thermal equilibrium.
- This is said to occur **quasi-statically**.



Work, 2

The piston is pushed downward by a force through a displacement of:

$$dW = \vec{F} \cdot d\vec{r} = -F\hat{j} \cdot dy\hat{j} = -Fdy = -PA dy$$

$A \cdot dy$ is the change in volume of the gas, dV .

Therefore, the work done on the gas is

$$dW = -P dV$$

Work, 3

Interpreting $dW = - P dV$

- If the gas is compressed, dV is negative and the work done on the gas is positive.
- If the gas expands, dV is positive and the work done on the gas is negative.
- If the volume remains constant, the work done is zero.

The total work done is:

$$W = - \int_{V_i}^{V_f} P dV$$

PV Diagrams

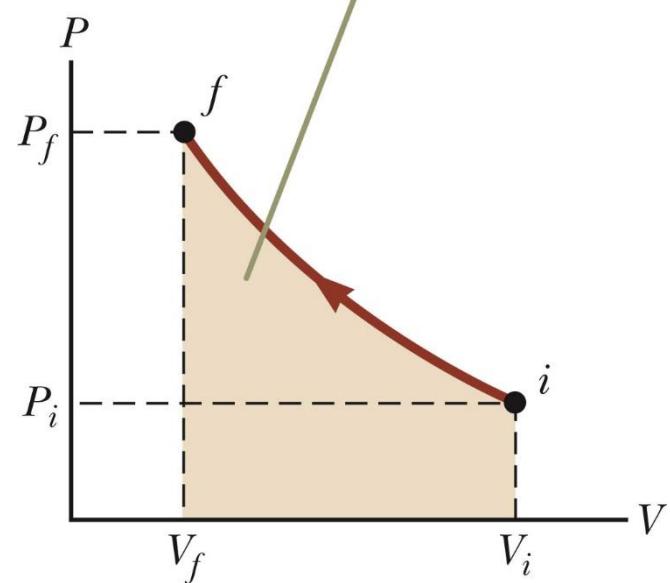
Used when the pressure and volume are known at each step of the process.

The state of the gas at each step can be plotted on a graph called a **PV diagram**.

- This allows us to visualize the process through which the gas is progressing.

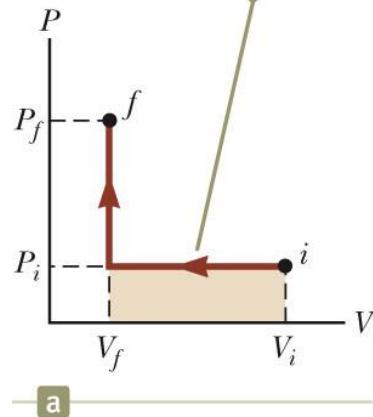
The curve is called the *path*.

The work done on a gas equals the negative of the area under the *PV* curve. The area is negative here because the volume is decreasing, resulting in positive work.



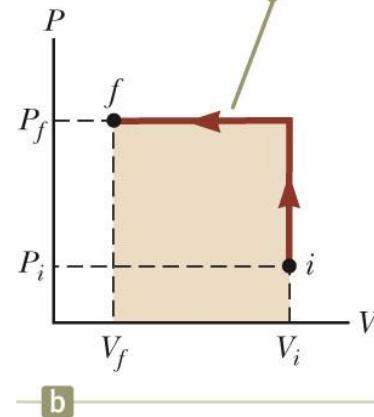
Work Done By Various Paths

A constant-pressure compression followed by a constant-volume process



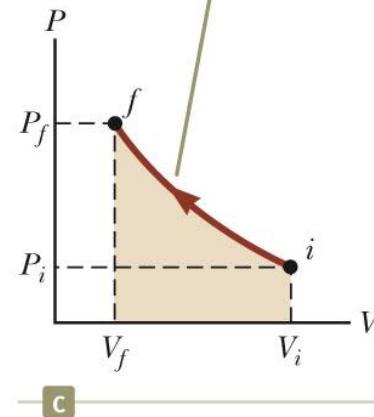
a

A constant-volume process followed by a constant-pressure compression



b

An arbitrary compression



c

Each of these processes has the same initial and final states.

The work done differs in each process.

The work done depends on the path.

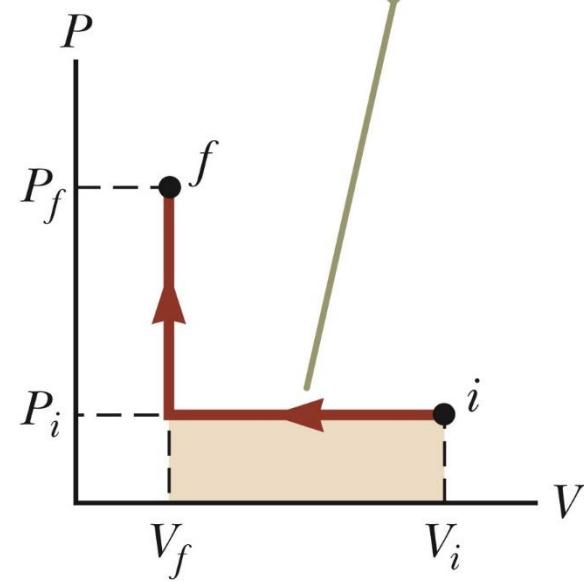
Work From a *PV* Diagram, Example 1

The volume of the gas is first reduced from V_i to V_f at constant pressure P_i .

Next, the pressure increases from P_i to P_f by heating at constant volume V_f .

$$W = -P_i(V_f - V_i)$$

A constant-pressure compression followed by a constant-volume process



a

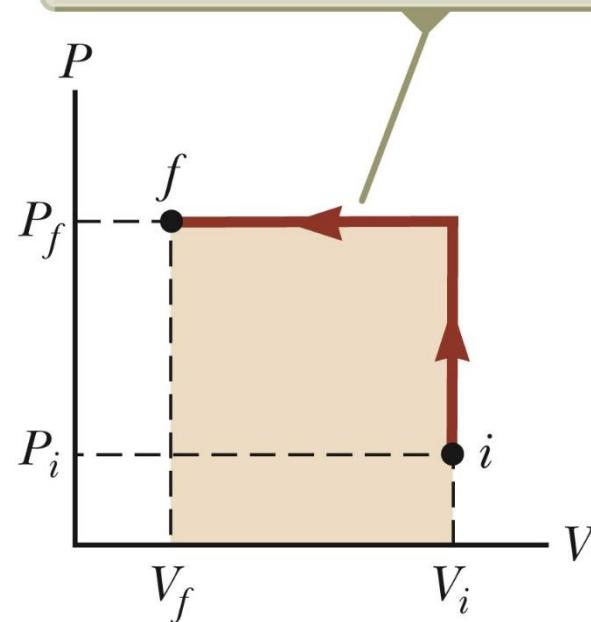
Work From a *PV* Diagram, Example 2

The pressure of the gas is increased from P_i to P_f at a constant volume.

The volume is decreased from V_i to V_f .

$$W = -P_f(V_f - V_i)$$

A constant-volume process followed by a constant-pressure compression



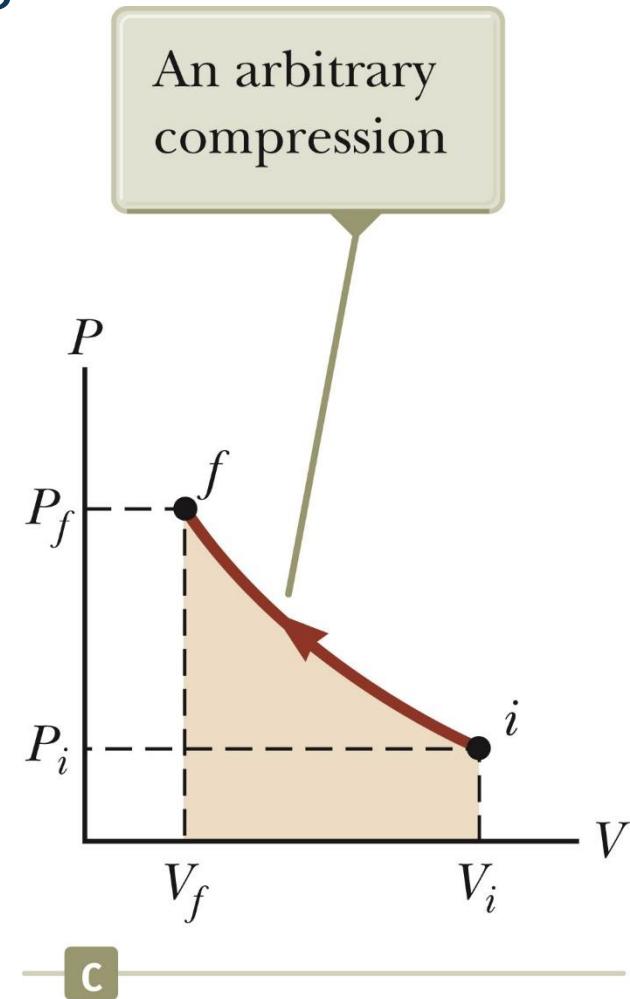
b

Work From a *PV* Diagram, Example 3

The pressure and the volume continually change.

The work is some intermediate value between $-P_f(V_f - V_i)$ and $-P_i(V_f - V_i)$.

To evaluate the actual amount of work, the function $P(V)$ must be known.



The First Law of Thermodynamics

The First Law of Thermodynamics is a special case of the Law of Conservation of Energy.

- It is a special cases when only the internal energy changes and the only energy transfers are by heat and work.

The First Law of Thermodynamics states that

$$\Delta E_{\text{int}} = Q + W$$

- All quantities must have the same units of measure of energy.

One consequence of the first law is that there must exist some quantity known as internal energy which is determined by the state of the system.

- The internal energy is therefore a state variable.

Isolated Systems

An isolated system is one that does not interact with its surroundings.

- No energy transfer by heat takes place.
- The work done on the system is zero.
- $Q = W = 0$, so $\Delta E_{\text{int}} = 0$

The internal energy of an isolated system remains constant.

Cyclic Processes

A cyclic process is one that starts and ends in the same state.

- This process would not be isolated.
- On a *PV* diagram, a cyclic process appears as a closed curve.

The internal energy must be zero since it is a state variable.

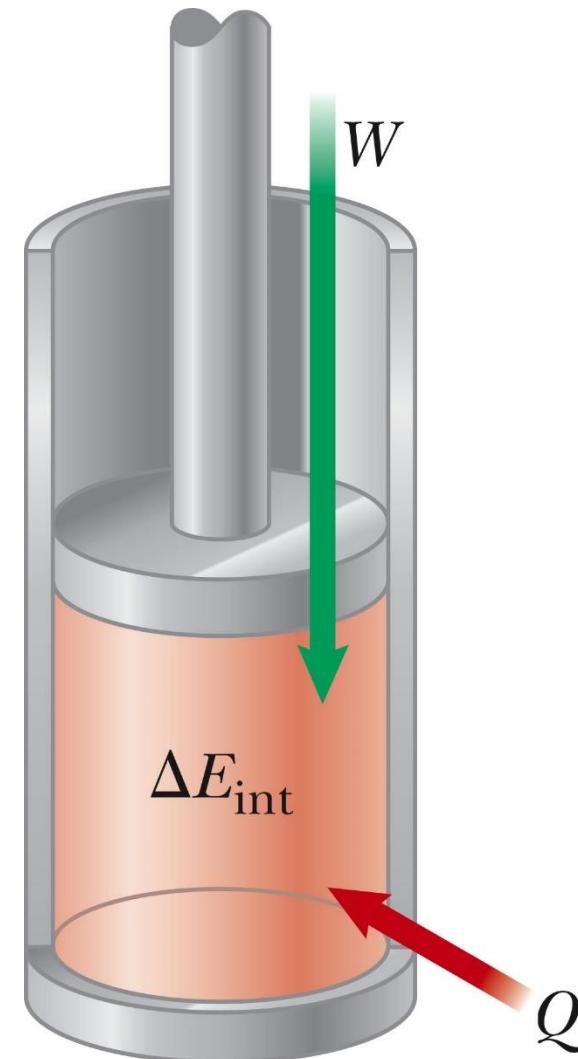
If $\Delta E_{\text{int}} = 0$, $Q = -W$

In a cyclic process, the net work done on the system per cycle equals the area enclosed by the path representing the process on a *PV* diagram.

Adiabatic Process

An adiabatic process is one during which no energy enters or leaves the system by heat.

- $Q = 0$
- This is achieved by:
 - Thermally insulating the walls of the system
 - Having the process proceed so quickly that no heat can be exchanged



Adiabatic Process, cont

Since $Q = 0$, $\Delta E_{\text{int}} = W$

If the gas is compressed adiabatically, W is positive so ΔE_{int} is positive and the temperature of the gas increases.

If the gas expands adiabatically, the temperature of the gas decreases.

Some important examples of adiabatic processes related to engineering are:

- The expansion of hot gases in an internal combustion engine
- The liquefaction of gases in a cooling system
- The compression stroke in a diesel engine

Adiabatic Free Expansion

The process is adiabatic because it takes place in an insulated container.

Because the gas expands into a vacuum, it does not apply a force on a piston and $W = 0$.

Since $Q = 0$ and $W = 0$, $\Delta E_{\text{int}} = 0$ and the initial and final states are the same

- No change in temperature is expected.

Isobaric Processes

An isobaric process is one that occurs at a constant pressure.

- May be accomplished by allowing the piston to move freely so that it is always in equilibrium between the net force from the gas pushing upward and the weight of the piston plus the force due to atmospheric pressure pushing downward

The values of the heat and the work are generally both nonzero.

The work done is $W = -P(V_f - V_i)$ where P is the constant pressure.

Isovolumetric Processes

An isovolumetric process is one in which there is no change in the volume.

- This may be accomplished by clamping the piston at a fixed position.

Since the volume does not change, $W = 0$.

From the first law, $\Delta E_{\text{int}} = Q$

If energy is added by heat to a system kept at constant volume, all of the transferred energy remains in the system as an increase in its internal energy.

Isothermal Process

An isothermal process is one that occurs at a constant temperature.

- This can be accomplished by putting the cylinder in contact with some constant-temperature reservoir.

Since there is no change in temperature, $\Delta E_{\text{int}} = 0$.

Therefore, $Q = -W$

Any energy that enters the system by heat must leave the system by work.

Isothermal Process, cont

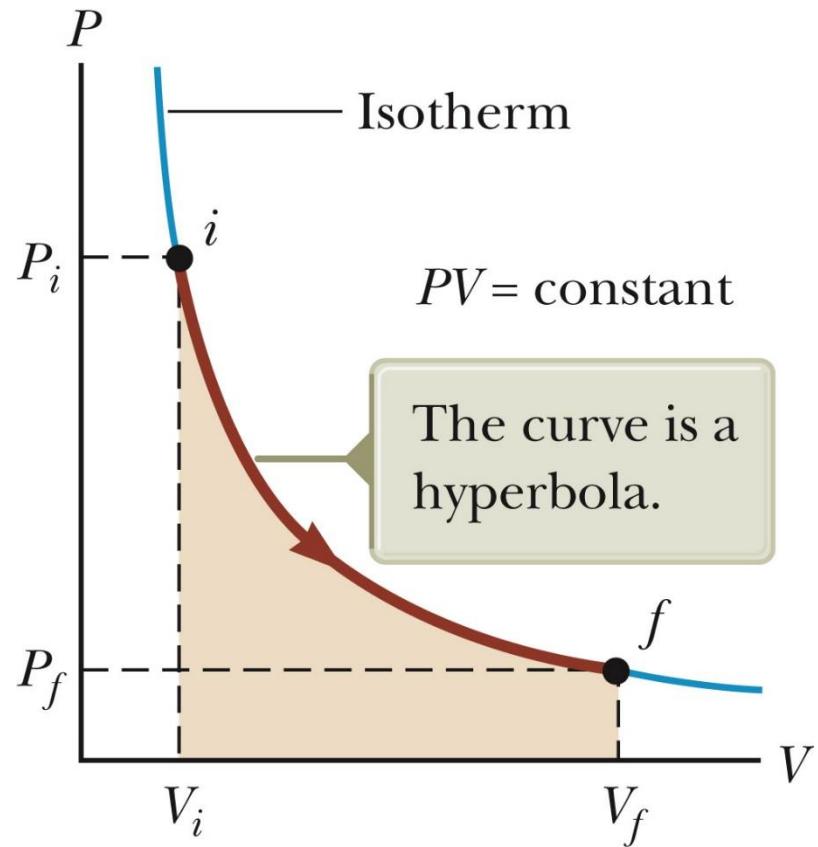
At right is a PV diagram of an isothermal expansion.

The curve is a hyperbola.

The equation of the curve is

$$P V = n R T = \text{constant}.$$

The curve is called an **isotherm**.



Isothermal Expansion, Details

Because it is an ideal gas and the process is quasi-static, the ideal gas law is valid for each point on the path.

$$W = - \int_{V_i}^{V_f} P dV = - \int_{V_i}^{V_f} \frac{nRT}{V} dV = -nRT \int_{V_i}^{V_f} \frac{dV}{V}$$

$$W = nRT \ln\left(\frac{V_f}{V_i}\right)$$

Numerically, the work equals the negative of the area under the PV diagram.

Because the gas expands, $V_f > V_i$ and the value of the work done on the gas is negative.

If the gas is compressed, $V_f < V_i$ and the value of the work done on the gas is positive.

Special Processes, Summary

Adiabatic

- No heat exchanged
- $Q = 0$ and $\Delta E_{int} = W$

Isobaric

- Constant pressure
- $W = P(V_f - V_i)$ and $\Delta E_{int} = Q + W$

Isothermal

- Constant temperature
- $\Delta E_{int} = 0$ and $Q = -W$

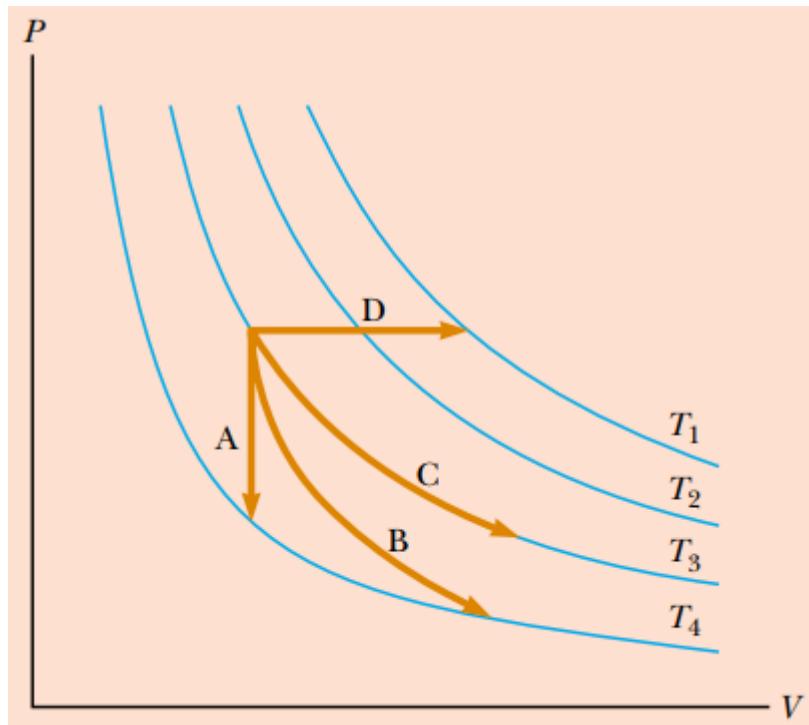
In the last three columns of the following table, fill in the boxes with +, -, or 0.

Situation	System	Q	W	ΔE_{int}
(a) Rapidly pumping up a bicycle tire	Air in the pump			
(b) Pan of room-temperature water sitting on a hot stove	Water in the pan			
(c) Air quickly leaking out of a balloon	Air originally in the balloon			

Situation	System	Q	W	ΔE_{int}
(a) Rapidly pumping up a bicycle tire	Air in the pump	0	+	+
(b) Pan of room-temperature water sitting on a hot stove	Water in the pan	+	0	+
(c) Air quickly leaking out of a balloon	Air originally in the balloon	0	-	-

- (a) Because the pumping is rapid, no energy enters or leaves the system by heat. Because $W > 0$ when work is done *on* the system, it is positive here. Thus, we see that $\Delta E_{\text{int}} = Q + W$ must be positive. The air in the pump is warmer.
- (b) There is no work done either on or by the system, but energy transfers into the water by heat from the hot burner, making both Q and ΔE_{int} positive.
- (c) Again no energy transfers into or out of the system by heat, but the air molecules escaping from the balloon do work on the surrounding air molecules as they push them out of the way. Thus W is negative and ΔE_{int} is negative. The decrease in internal energy is evidenced by the fact that the escaping air becomes cooler.

Characterize the paths in Figure as isobaric, isovolumetric, isothermal, or adiabatic. Note that $Q = 0$ for path B.



An Isothermal Expansion

A 1.0-mol sample of an ideal gas is kept at 0.0°C during an expansion from 3.0 L to 10.0 L.

- (A)** How much work is done on the gas during the expansion?
- (B)** How much energy transfer by heat occurs with the surroundings in this process?
- (C)** If the gas is returned to the original volume by means of an isobaric process, how much work is done on the gas?

$$\begin{aligned}
 W &= nRT \ln \left(\frac{V_i}{V_f} \right) \\
 &= (1.0 \text{ mol}) (8.31 \text{ J/mol}\cdot\text{K}) (273 \text{ K}) \ln \left(\frac{3.0 \text{ L}}{10.0 \text{ L}} \right) \\
 &= -2.7 \times 10^3 \text{ J}
 \end{aligned}$$

From the first law, we find that

$$\Delta E_{\text{int}} = Q + W$$

$$0 = Q + W$$

$$Q = -W = 2.7 \times 10^3 \text{ J}$$

$$\begin{aligned}
 W &= -P(V_f - V_i) = -\frac{nRT_i}{V_i} (V_f - V_i) \\
 &= -\frac{(1.0 \text{ mol}) (8.31 \text{ J/mol}\cdot\text{K}) (273 \text{ K})}{10.0 \times 10^{-3} \text{ m}^3} \\
 &\quad \times (3.0 \times 10^{-3} \text{ m}^3 - 10.0 \times 10^{-3} \text{ m}^3) \\
 &= 1.6 \times 10^3 \text{ J}
 \end{aligned}$$

Boiling Water

Suppose 1.00 g of water vaporizes isobarically at atmospheric pressure (1.013×10^5 Pa). Its volume in the liquid state is $V_i = V_{\text{liquid}} = 1.00 \text{ cm}^3$, and its volume in the vapor state is $V_f = V_{\text{vapor}} = 1\,671 \text{ cm}^3$. Find the work done in the expansion and the change in internal energy of the system. Ignore any mixing of the steam and the surrounding air—imagine that the steam simply pushes the surrounding air out of the way.

Solution Because the expansion takes place at constant pressure, the work done on the system (the vaporizing water) as it pushes away the surrounding air is, from Equation 20.11,

$$\begin{aligned}W &= -P(V_f - V_i) \\&= -(1.013 \times 10^5 \text{ Pa})(1.671 \times 10^{-6} \text{ m}^3 - 1.00 \times 10^{-6} \text{ m}^3) \\&= -169 \text{ J}\end{aligned}$$

To determine the change in internal energy, we must know the energy transfer Q needed to vaporize the water. Using Equation 20.6 and the latent heat of vaporization for water, we have

$$Q = mL_v = (1.00 \times 10^{-3} \text{ kg})(2.26 \times 10^6 \text{ J/kg}) = 2260 \text{ J}$$

Hence, from the first law, the change in internal energy is

$$\Delta E_{\text{int}} = Q + W = 2260 \text{ J} + (-169 \text{ J}) = 2090 \text{ J}$$

The positive value for ΔE_{int} indicates that the internal energy of the system increases. We see that most of the energy ($2090 \text{ J}/2260 \text{ J} = 93\%$) transferred to the liquid goes into increasing the internal energy of the system. The remaining 7% of the energy transferred leaves the system by work done by the steam on the surrounding atmosphere.

Heating a Solid

A 1.0-kg bar of copper is heated at atmospheric pressure. If its temperature increases from 20°C to 50°C,

- (A) what is the work done on the copper bar by the surrounding atmosphere?
- (B) What quantity of energy is transferred to the copper bar by heat?
- (C) What is the increase in internal energy of the copper bar?

Solution Because the process is isobaric, we can find the work done on the copper bar using Equation 20.11, $W = -P(V_f - V_i)$. We can calculate the change in volume of the copper bar using Equation 19.6. Using the average linear expansion coefficient for copper given in Table 19.1, and remembering that $\beta = 3\alpha$, we obtain

$$\begin{aligned}\Delta V &= \beta V_i \Delta T \\ &= [5.1 \times 10^{-5} (\text{ }^{\circ}\text{C})^{-1}] (50\text{ }^{\circ}\text{C} - 20\text{ }^{\circ}\text{C}) V_i = 1.5 \times 10^{-3} V_i\end{aligned}$$

The volume V_i is equal to m/ρ , and Table 14.1 indicates that the density of copper is $8.92 \times 10^3 \text{ kg/m}^3$. Hence,

$$\begin{aligned}\Delta V &= (1.5 \times 10^{-3}) \left(\frac{1.0 \text{ kg}}{8.92 \times 10^3 \text{ kg/m}^3} \right) \\ &= 1.7 \times 10^{-7} \text{ m}^3\end{aligned}$$

The work done on the copper bar is

$$\begin{aligned}W &= -P \Delta V = -(1.013 \times 10^5 \text{ N/m}^2) (1.7 \times 10^{-7} \text{ m}^3) \\ &= -1.7 \times 10^{-2} \text{ J}\end{aligned}$$

Because this work is negative, work is done *by* the copper bar on the atmosphere.

Solution Taking the specific heat of copper from Table 20.1 and using Equation 20.4, we find that the energy transferred by heat is

$$\begin{aligned} Q &= mc \Delta T = (1.0 \text{ kg})(387 \text{ J/kg}\cdot\text{°C})(30\text{°C}) \\ &= 1.2 \times 10^4 \text{ J} \end{aligned}$$

Solution From the first law of thermodynamics, we have

$$\begin{aligned} \Delta E_{\text{int}} &= Q + W = 1.2 \times 10^4 \text{ J} + (-1.7 \times 10^{-2} \text{ J}) \\ &= 1.2 \times 10^4 \text{ J} \end{aligned}$$

Consider Joule's apparatus described in Figure 20.1. The mass of each of the two blocks is 1.50 kg, and the insulated tank is filled with 200 g of water. What is the increase in the temperature of the water after the blocks fall through a distance of 3.00 m?

The container is thermally insulated, so no energy flows by heat:

$$Q = 0$$

and

$$\Delta E_{\text{int}} = Q + W_{\text{input}} = 0 + W_{\text{input}} = 2mgh$$

The work on the falling weights is equal to the work done on the water in the container by the rotating blades. This work results in an increase in internal energy of the water:

$$2mgh = \Delta E_{\text{int}} = m_{\text{water}}c\Delta T$$

$$\begin{aligned}\Delta T &= \frac{2mgh}{m_{\text{water}}c} = \frac{2 \times 1.50 \text{ kg}(9.80 \text{ m/s}^2)(3.00 \text{ m})}{0.200 \text{ kg}(4186 \text{ J/kg}\cdot^\circ\text{C})} = \frac{88.2 \text{ J}}{837 \text{ J/}^\circ\text{C}} \\ &= \boxed{0.105^\circ\text{C}}\end{aligned}$$

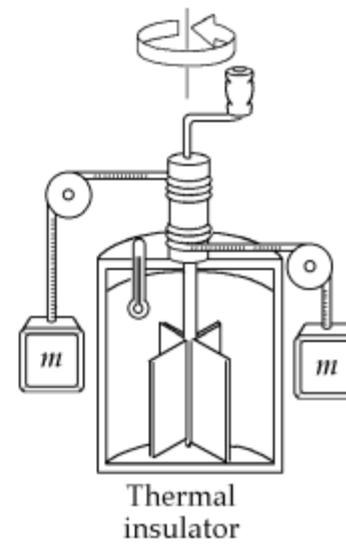


FIG. P20.2

The temperature of a silver bar rises by 10.0°C when it absorbs 1.23 kJ of energy by heat. The mass of the bar is 525 g. Determine the specific heat of silver.

$$\Delta Q = mc_{\text{silver}} \Delta T$$

$$1.23 \text{ kJ} = (0.525 \text{ kg})c_{\text{silver}}(10.0^{\circ}\text{C})$$

$$c_{\text{silver}} = [0.234 \text{ kJ/kg}\cdot{}^{\circ}\text{C}]$$

A 50.0-g sample of copper is at 25.0°C . If 1 200 J of energy is added to it by heat, what is the final temperature of the copper?

From $Q = mc\Delta T$

we find $\Delta T = \frac{Q}{mc} = \frac{1200 \text{ J}}{0.0500 \text{ kg}(387 \text{ J/kg}\cdot{}^{\circ}\text{C})} = 62.0^{\circ}\text{C}$

Thus, the final temperature is $[87.0^{\circ}\text{C}]$.

An aluminum cup of mass 200 g contains 800 g of water in thermal equilibrium at 80.0°C. The combination of cup and water is cooled uniformly so that the temperature decreases by 1.50°C per minute. At what rate is energy being removed by heat? Express your answer in watts.

Let us find the energy transferred in one minute.

$$Q = [m_{\text{cup}} c_{\text{cup}} + m_{\text{water}} c_{\text{water}}] \Delta T$$

$$Q = [(0.200 \text{ kg})(900 \text{ J/kg}\cdot^\circ\text{C}) + (0.800 \text{ kg})(4186 \text{ J/kg}\cdot^\circ\text{C})](-1.50^\circ\text{C}) = -5290 \text{ J}$$

If this much energy is removed from the system each minute, the rate of removal is

$$\mathcal{P} = \frac{|Q|}{\Delta t} = \frac{5290 \text{ J}}{60.0 \text{ s}} = 88.2 \text{ J/s} = \boxed{88.2 \text{ W}}$$

An aluminum calorimeter with a mass of 100 g contains 250 g of water. The calorimeter and water are in thermal equilibrium at 10.0°C. Two metallic blocks are placed into the water. One is a 50.0-g piece of copper at 80.0°C. The other block has a mass of 70.0 g and is originally at a temperature of 100°C. The entire system stabilizes at a final temperature of 20.0°C. (a) Determine the specific heat of the unknown sample. (b) Guess the material of the unknown, using the data in Table 20.1.

(a) $Q_{\text{cold}} = -Q_{\text{hot}}$

$$(m_w c_w + m_c c_c)(T_f - T_c) = -m_{\text{Cu}} c_{\text{Cu}}(T_f - T_{\text{Cu}}) - m_{\text{unk}} c_{\text{unk}}(T_f - T_{\text{unk}})$$

where *w* is for water, *c* the calorimeter, Cu the copper sample, and *unk* the unknown.

$$\begin{aligned} & [250 \text{ g}(1.00 \text{ cal/g}\cdot^{\circ}\text{C}) + 100 \text{ g}(0.215 \text{ cal/g}\cdot^{\circ}\text{C})](20.0 - 10.0)^{\circ}\text{C} \\ & = -(50.0 \text{ g})(0.0924 \text{ cal/g}\cdot^{\circ}\text{C})(20.0 - 80.0)^{\circ}\text{C} - (70.0 \text{ g})c_{\text{unk}}(20.0 - 100)^{\circ}\text{C} \\ & 2.44 \times 10^3 \text{ cal} = (5.60 \times 10^3 \text{ g}\cdot^{\circ}\text{C})c_{\text{unk}} \end{aligned}$$

or $c_{\text{unk}} = \boxed{0.435 \text{ cal/g}\cdot^{\circ}\text{C}}$.

(b) The material of the sample is beryllium.

A water heater is operated by solar power. If the solar collector has an area of 6.00 m^2 and the intensity delivered by sunlight is 550 W/m^2 , how long does it take to increase the temperature of 1.00 m^3 of water from 20.0°C to 60.0°C ?

The rate of collection of energy is $\mathcal{P} = 550 \text{ W/m}^2 (6.00 \text{ m}^2) = 3300 \text{ W}$. The amount of energy required to raise the temperature of 1000 kg of water by 40.0°C is:

$$Q = mc\Delta T = 1000 \text{ kg}(4186 \text{ J/kg}\cdot^\circ\text{C})(40.0^\circ\text{C}) = 1.67 \times 10^8 \text{ J}$$

Thus, $\mathcal{P}\Delta t = 1.67 \times 10^8 \text{ J}$

or $\Delta t = \frac{1.67 \times 10^8 \text{ J}}{3300 \text{ W}} = \boxed{50.7 \text{ ks}} = 14.1 \text{ h.}$



A sample of ideal gas is expanded to twice its original volume of 1.00 m^3 in a quasi-static process for which $P = \alpha V^2$, with $\alpha = 5.00 \text{ atm/m}^6$, as shown in Figure P20.23. How much work is done on the expanding gas?

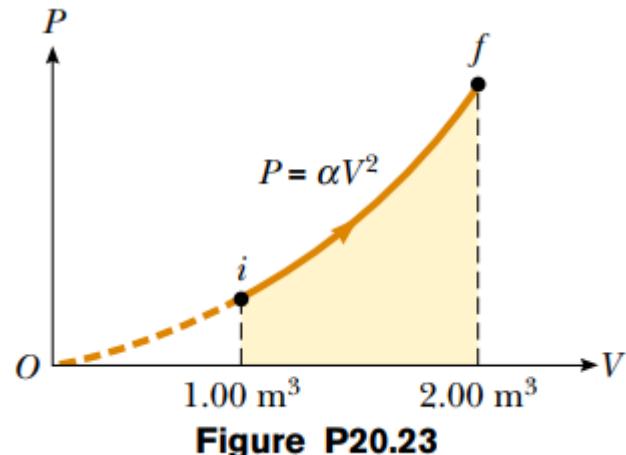


Figure P20.23

$$W_{if} = - \int_i^f P dV$$

The work done on the gas is the negative of the area under the curve $P = \alpha V^2$ between V_i and V_f .

$$W_{if} = - \int_i^f \alpha V^2 dV = -\frac{1}{3} \alpha (V_f^3 - V_i^3)$$

$$V_f = 2V_i = 2(1.00 \text{ m}^3) = 2.00 \text{ m}^3$$

$$W_{if} = -\frac{1}{3} [(5.00 \text{ atm/m}^6)(1.013 \times 10^5 \text{ Pa/atm})] [(2.00 \text{ m}^3)^3 + (1.00 \text{ m}^3)^3] = \boxed{-1.18 \text{ MJ}}$$

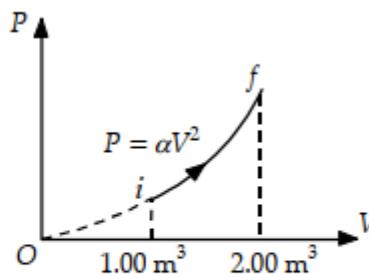


FIG. P20.23

(a) Determine the work done on a fluid that expands from i to f as indicated in Figure P20.24. (b) **What If?** How much work is performed on the fluid if it is compressed from f to i along the same path?

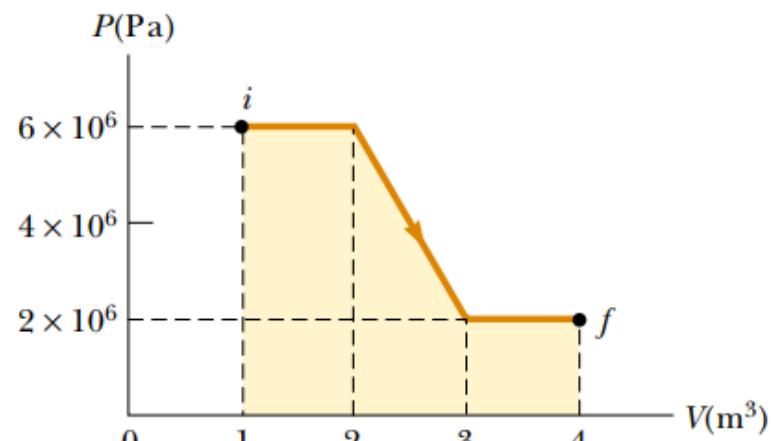


Figure P20.24

$$(a) \quad W = -\int P dV$$

$$W = -\left(6.00 \times 10^6 \text{ Pa}\right)(2.00 - 1.00) \text{ m}^3 + \\ -\left(4.00 \times 10^6 \text{ Pa}\right)(3.00 - 2.00) \text{ m}^3 + \\ -\left(2.00 \times 10^6 \text{ Pa}\right)(4.00 - 3.00) \text{ m}^3$$

$$W_{i \rightarrow f} = \boxed{-12.0 \text{ MJ}}$$

$$(b) \quad W_{f \rightarrow i} = \boxed{+12.0 \text{ MJ}}$$

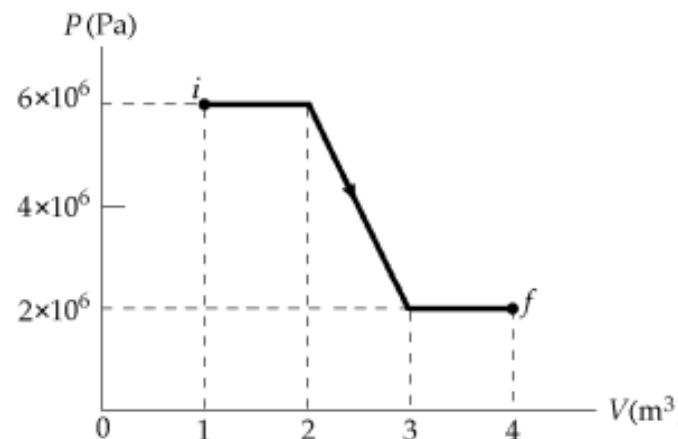


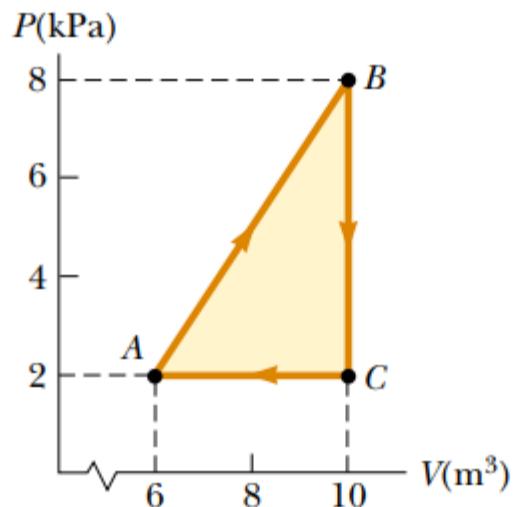
FIG. P20.24

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? (b) What is the change in its internal energy?

(a) $W = -P\Delta V = -(0.800 \text{ atm})(-7.00 \text{ L})\left(1.013 \times 10^5 \text{ Pa/atm}\right)\left(10^{-3} \text{ m}^3/\text{L}\right) = \boxed{+567 \text{ J}}$

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

A gas is taken through the cyclic process described in Figure P20.30. (a) Find the net energy transferred to the system by heat during one complete cycle. (b) **What If?** If the cycle is reversed—that is, the process follows the path $ACBA$ —what is the net energy input per cycle by heat?



(a)

$$Q = -W = \text{Area of triangle}$$

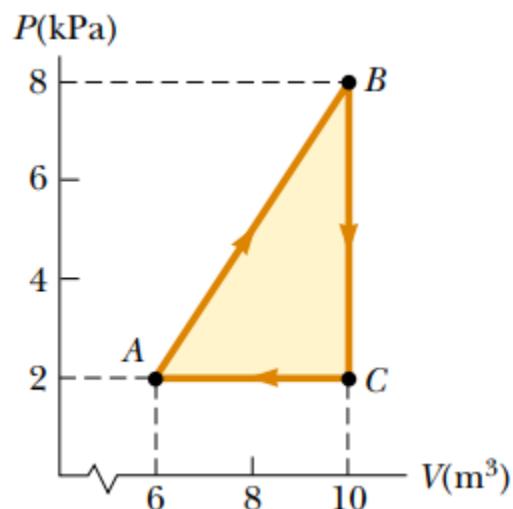
$$Q = \frac{1}{2} (4.00 \text{ m}^3)(6.00 \text{ kPa}) = \boxed{12.0 \text{ kJ}}$$

(b)

$$Q = -W = \boxed{-12.0 \text{ kJ}}$$

Figure P20.30 Problems 30 and 31.

Consider the cyclic process depicted in Figure P20.30. If Q is negative for the process BC and ΔE_{int} is negative for the process CA , what are the signs of Q , W , and ΔE_{int} that are associated with each process?



	Q	W	ΔE_{int}	
BC	-	0	-	$(Q = \Delta E_{\text{int}} \text{ since } W_{BC} = 0)$
CA	-	+	-	$(\Delta E_{\text{int}} < 0 \text{ and } W > 0, \text{ so } Q < 0)$
AB	+	-	+	$(W < 0, \Delta E_{\text{int}} > 0 \text{ since } \Delta E_{\text{int}} < 0 \text{ for } B \rightarrow C \rightarrow A; \text{ so } Q > 0)$

A sample of an ideal gas goes through the process shown in Figure P20.32. From A to B , the process is adiabatic; from B to C , it is isobaric with 100 kJ of energy entering the system by heat. From C to D , the process is isothermal; from D to A , it is isobaric with 150 kJ of energy leaving the system by heat. Determine the difference in internal energy $E_{\text{int}, B} - E_{\text{int}, A}$.

$$W_{BC} = -P_B(V_C - V_B) = -3.00 \text{ atm}(0.400 - 0.0900) \text{ m}^3 \\ = -94.2 \text{ kJ}$$

$$\Delta E_{\text{int}} = Q + W$$

$$E_{\text{int}, C} - E_{\text{int}, B} = (100 - 94.2) \text{ kJ} \\ E_{\text{int}, C} - E_{\text{int}, B} = 5.79 \text{ kJ}$$

Since T is constant,

$$E_{\text{int}, D} - E_{\text{int}, C} = 0$$

$$W_{DA} = -P_D(V_A - V_D) = -1.00 \text{ atm}(0.200 - 1.20) \text{ m}^3 \\ = +101 \text{ kJ}$$

$$E_{\text{int}, A} - E_{\text{int}, D} = -150 \text{ kJ} + (+101 \text{ kJ}) = -48.7 \text{ kJ}$$

$$\text{Now, } E_{\text{int}, B} - E_{\text{int}, A} = -[(E_{\text{int}, C} - E_{\text{int}, B}) + (E_{\text{int}, D} - E_{\text{int}, C}) + (E_{\text{int}, A} - E_{\text{int}, D})]$$

$$E_{\text{int}, B} - E_{\text{int}, A} = -[5.79 \text{ kJ} + 0 - 48.7 \text{ kJ}] = \boxed{42.9 \text{ kJ}}$$

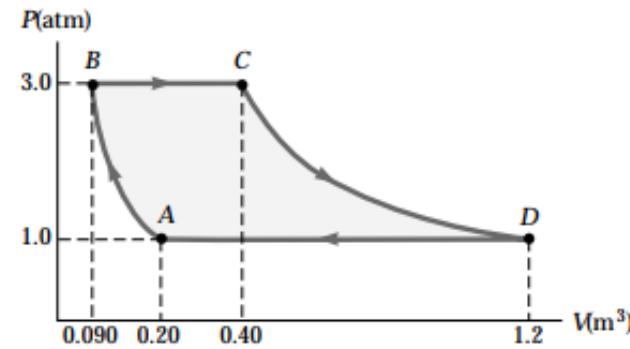
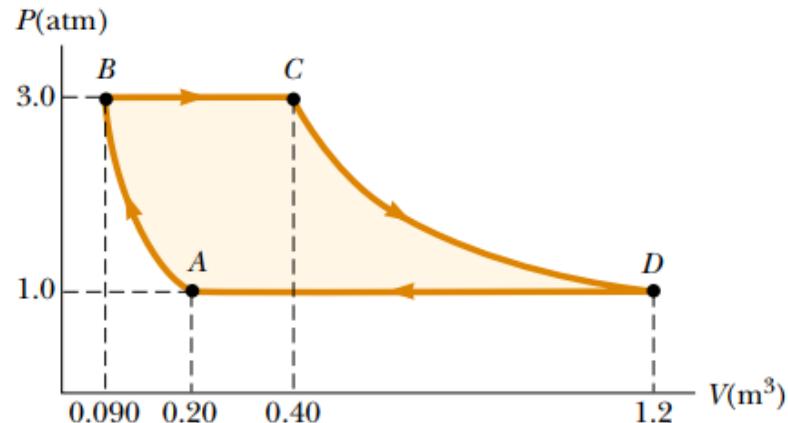


FIG. P20.32

A 1.00-kg block of aluminum is heated at atmospheric pressure so that its temperature increases from 22.0°C to 40.0°C. Find (a) the work done on the aluminum, (b) the energy added to it by heat, and (c) the change in its internal energy.

$$(a) \quad W = -P\Delta V = -P[3\alpha V\Delta T]$$

$$= -\left(1.013 \times 10^5 \text{ N/m}^2\right) \left[3\left(24.0 \times 10^{-6} \text{ }^\circ\text{C}^{-1}\right) \left(\frac{1.00 \text{ kg}}{2.70 \times 10^3 \text{ kg/m}^3}\right) (18.0 \text{ }^\circ\text{C}) \right]$$

$$W = \boxed{-48.6 \text{ mJ}}$$

$$(b) \quad Q = cm\Delta T = (900 \text{ J/kg}\cdot\text{ }^\circ\text{C})(1.00 \text{ kg})(18.0 \text{ }^\circ\text{C}) = \boxed{16.2 \text{ kJ}}$$

$$(c) \quad \Delta E_{\text{int}} = Q + W = 16.2 \text{ kJ} - 48.6 \text{ mJ} = \boxed{16.2 \text{ kJ}}$$

A 2.00-mol sample of helium gas initially at 300 K and 0.400 atm is compressed isothermally to 1.20 atm. Noting that the helium behaves as an ideal gas, find (a) the final volume of the gas, (b) the work done on the gas, and (c) the energy transferred by heat.

$$(a) \quad P_i V_i = P_f V_f = nRT = 2.00 \text{ mol}(8.314 \text{ J/K}\cdot\text{mol})(300 \text{ K}) = 4.99 \times 10^3 \text{ J}$$

$$V_i = \frac{nRT}{P_i} = \frac{4.99 \times 10^3 \text{ J}}{0.400 \text{ atm}}$$

$$V_f = \frac{nRT}{P_f} = \frac{4.99 \times 10^3 \text{ J}}{1.20 \text{ atm}} = \frac{1}{3} V_i = \boxed{0.0410 \text{ m}^3}$$

$$(b) \quad W = -\int PdV = -nRT \ln\left(\frac{V_f}{V_i}\right) = -(4.99 \times 10^3) \ln\left(\frac{1}{3}\right) = \boxed{+5.48 \text{ kJ}}$$

$$(c) \quad \Delta E_{\text{int}} = 0 = Q + W$$

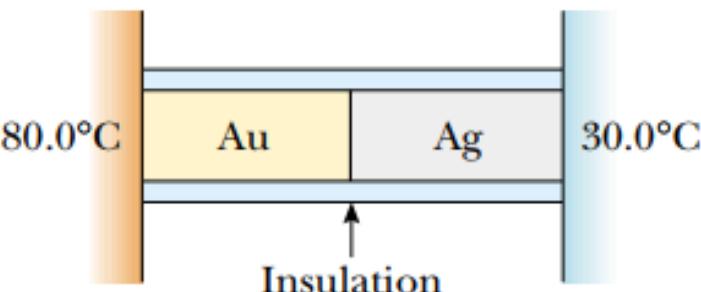
$$Q = \boxed{-5.48 \text{ kJ}}$$

A box with a total surface area of 1.20 m^2 and a wall thickness of 4.00 cm is made of an insulating material. A 10.0-W electric heater inside the box maintains the inside temperature at 15.0°C above the outside temperature. Find the thermal conductivity k of the insulating material.

$$\mathcal{P} = kA \frac{\Delta T}{L}$$

$$k = \frac{\mathcal{P}L}{A\Delta T} = \frac{10.0 \text{ W}(0.0400 \text{ m})}{1.20 \text{ m}^2(15.0^\circ\text{C})} = \boxed{2.22 \times 10^{-2} \text{ W/m}\cdot^\circ\text{C}}$$

A bar of gold is in thermal contact with a bar of silver of the same length and area (Fig. P20.43). One end of the compound bar is maintained at 80.0°C while the opposite end is at 30.0°C. When the energy transfer reaches steady state, what is the temperature at the junction?



In the steady state condition,

$$\mathcal{P}_{\text{Au}} = \mathcal{P}_{\text{Ag}}$$

so that

$$k_{\text{Au}} A_{\text{Au}} \left(\frac{\Delta T}{\Delta x} \right)_{\text{Au}} = k_{\text{Ag}} A_{\text{Ag}} \left(\frac{\Delta T}{\Delta x} \right)_{\text{Ag}}$$

In this case

$$A_{\text{Au}} = A_{\text{Ag}}$$

$$\Delta x_{\text{Au}} = \Delta x_{\text{Ag}}$$

$$\Delta T_{\text{Au}} = (80.0 - T)$$

and

$$\Delta T_{\text{Ag}} = (T - 30.0)$$

where T is the temperature of the junction.

Therefore,

$$k_{\text{Au}} (80.0 - T) = k_{\text{Ag}} (T - 30.0)$$

And

$$T = 51.2^{\circ}\text{C}$$

The surface of the Sun has a temperature of about 5 800 K. The radius of the Sun is 6.96×10^8 m. Calculate the total energy radiated by the Sun each second. Assume that the emissivity of the Sun is 0.965.

$$\mathcal{P} = \sigma A e T^4 = \left(5.6696 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4\right) \left[4\pi \left(6.96 \times 10^8 \text{ m}\right)^2\right] (0.965)(5800 \text{ K})^4$$

$$\mathcal{P} = \boxed{3.77 \times 10^{26} \text{ W}}$$

The tungsten filament of a certain 100-W light bulb radiates 2.00 W of light. (The other 98 W is carried away by convection and conduction.) The filament has a surface area of 0.250 mm^2 and an emissivity of 0.950. Find the filament's temperature. (The melting point of tungsten is 3 683 K.)

$$\mathcal{P} = \sigma A e T^4$$

$$2.00 \text{ W} = \left(5.67 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4\right) \left(0.250 \times 10^{-6} \text{ m}^2\right) (0.950) T^4$$

$$T = \left(1.49 \times 10^{14} \text{ K}^4\right)^{1/4} = \boxed{3.49 \times 10^3 \text{ K}}$$