

# Chapter 20

## The First Law of Thermodynamics



# 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

# 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.

In the US Customary system, the unit is a BTU (British Thermal Unit).

- One BTU is the amount of energy transfer necessary to raise the temperature of 1 lb of water from 63°F to 64°F.

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

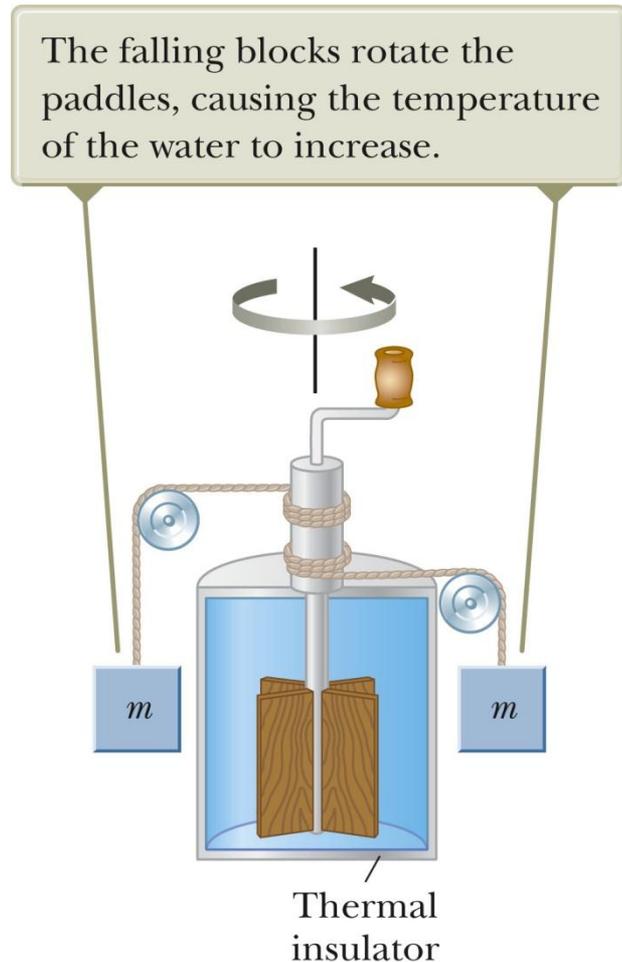


# Mechanical Equivalent of Heat

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.



## 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 more proper name would be the *equivalence between mechanical energy and internal energy*, but the historical name is well entrenched.

## 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  $1^\circ\text{C}$ .

If energy  $Q$  produces a change of temperature of  $\Delta T$ , 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  $\Delta T$ , 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>		<i>Other 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.

## 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  $\Delta T$  are positive
- Energy transfers into the system

If the temperature decreases:

- $Q$  and  $\Delta T$  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

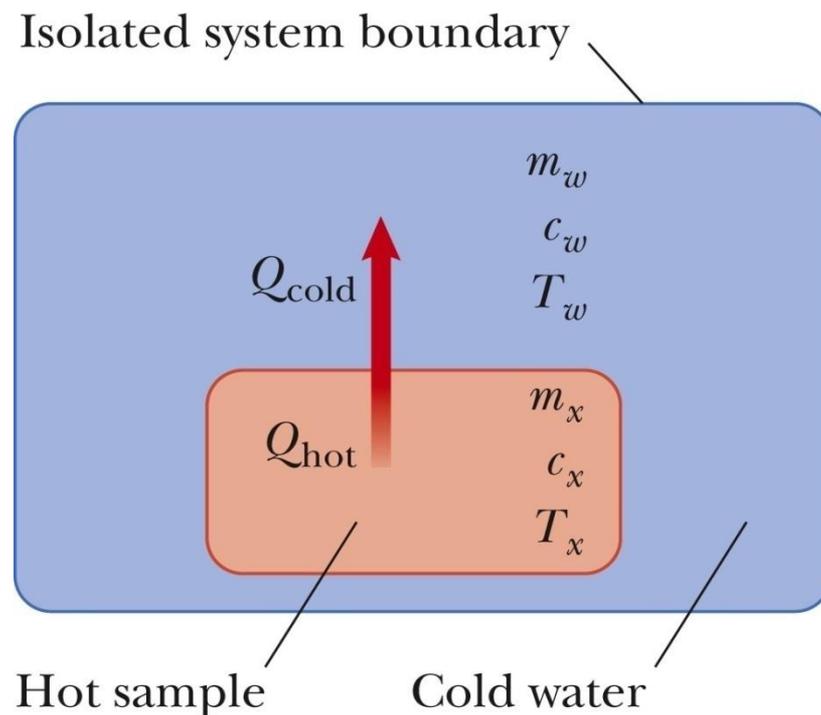
$T_f$  is the final temperature after the system comes to equilibrium.

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

An ingot of metal is heated and then dropped into a beaker of water. The equilibrium temperature is measured:

$$\begin{aligned}c_x &= \frac{m_w c_w (T_f - T_w)}{m_x (T_x - T_f)} \\ &= \frac{(0.400 \text{ kg})(4186 \text{ J/kg} \cdot ^\circ\text{C})(22.4^\circ\text{C} - 20.0^\circ\text{C})}{(0.0500 \text{ kg})(200.0^\circ\text{C} - 22.4^\circ\text{C})} \\ &= 453 \text{ J/kg} \cdot ^\circ\text{C}\end{aligned}$$

# 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  $Q = L \Delta m$ .

- $\Delta m$  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 is enters the system:

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

If energy is extracted from the system:

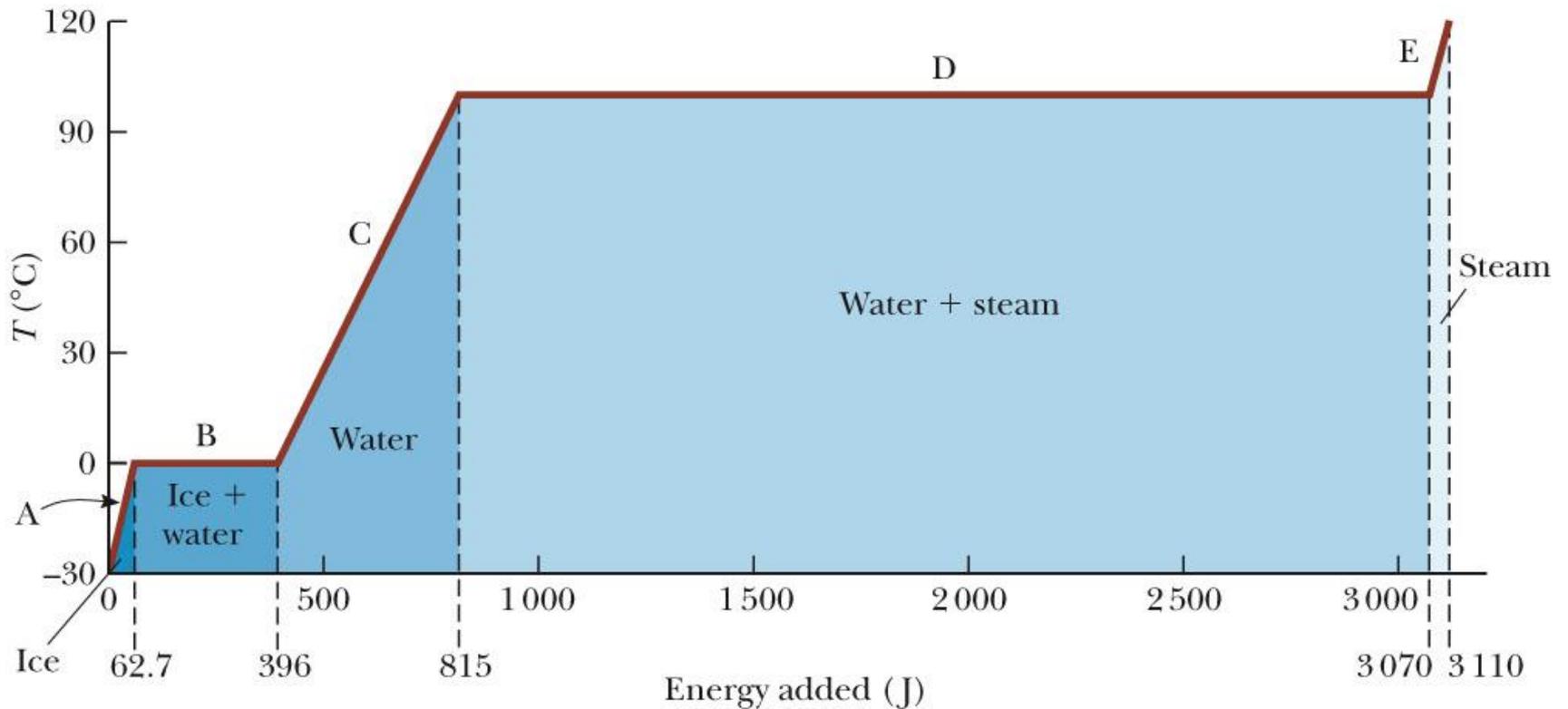
- This will result in freezing or condensation
- The amount of the higher-phase material will decrease
- $\Delta m$  and  $Q$  are negative

## 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	$5.23 \times 10^3$	-268.93	$2.09 \times 10^4$
Oxygen	-218.79	$1.38 \times 10^4$	-182.97	$2.13 \times 10^5$
Nitrogen	-209.97	$2.55 \times 10^4$	-195.81	$2.01 \times 10^5$
Ethyl alcohol	-114	$1.04 \times 10^5$	78	$8.54 \times 10^5$
Water	0.00	$3.33 \times 10^5$	100.00	$2.26 \times 10^6$
Sulfur	119	$3.81 \times 10^4$	444.60	$3.26 \times 10^5$
Lead	327.3	$2.45 \times 10^4$	1 750	$8.70 \times 10^5$
Aluminum	660	$3.97 \times 10^5$	2 450	$1.14 \times 10^7$
Silver	960.80	$8.82 \times 10^4$	2 193	$2.33 \times 10^6$
Gold	1 063.00	$6.44 \times 10^4$	2 660	$1.58 \times 10^6$
Copper	1 083	$1.34 \times 10^5$	1 187	$5.06 \times 10^6$

# Graph of Ice to Steam



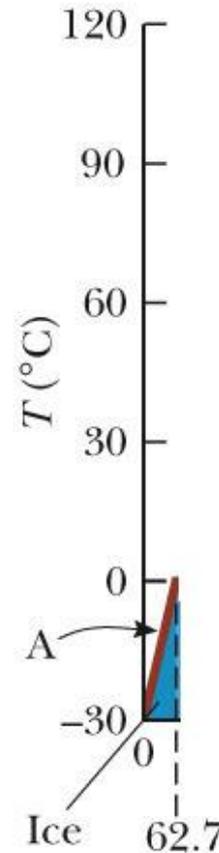
## Warming Ice, Graph Part A

Start with one gram of ice at  $-30.0^{\circ}\text{C}$ .

During phase A, the temperature of the ice changes from  $-30.0^{\circ}\text{C}$  to  $0^{\circ}\text{C}$ .

Use  $Q = m_i c_i \Delta T$

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



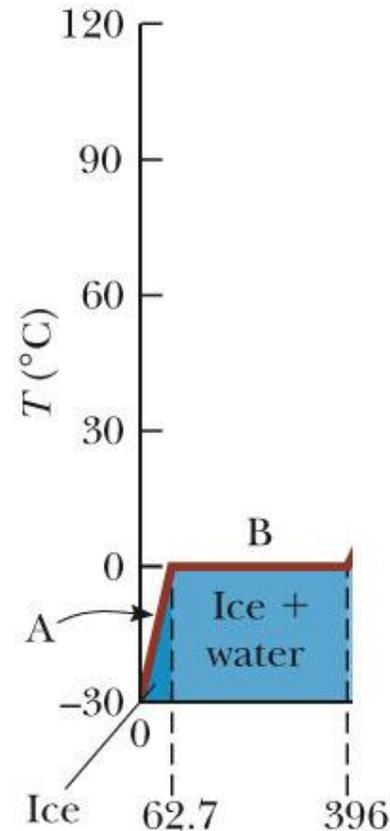
## Melting Ice, Graph Part B

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.



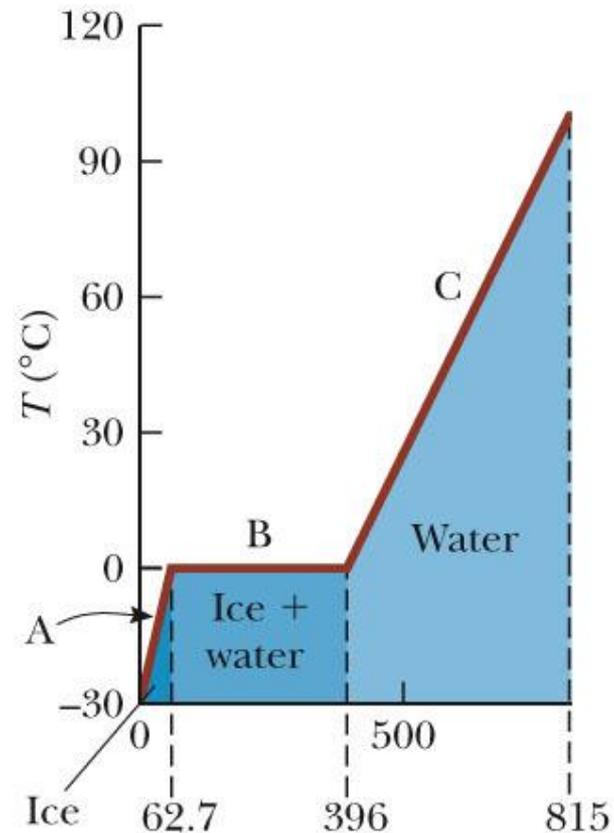
## Warming Water, Graph Part C

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



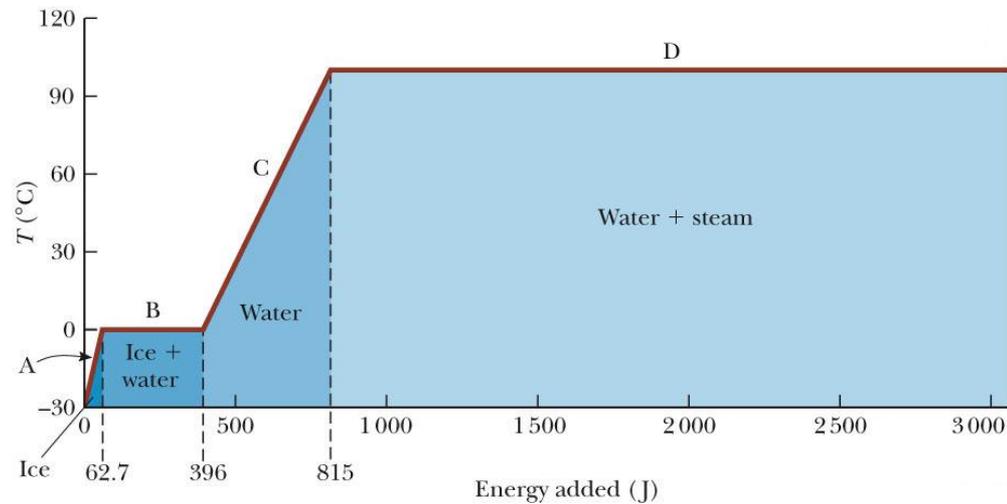
## Boiling Water, Graph Part D

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

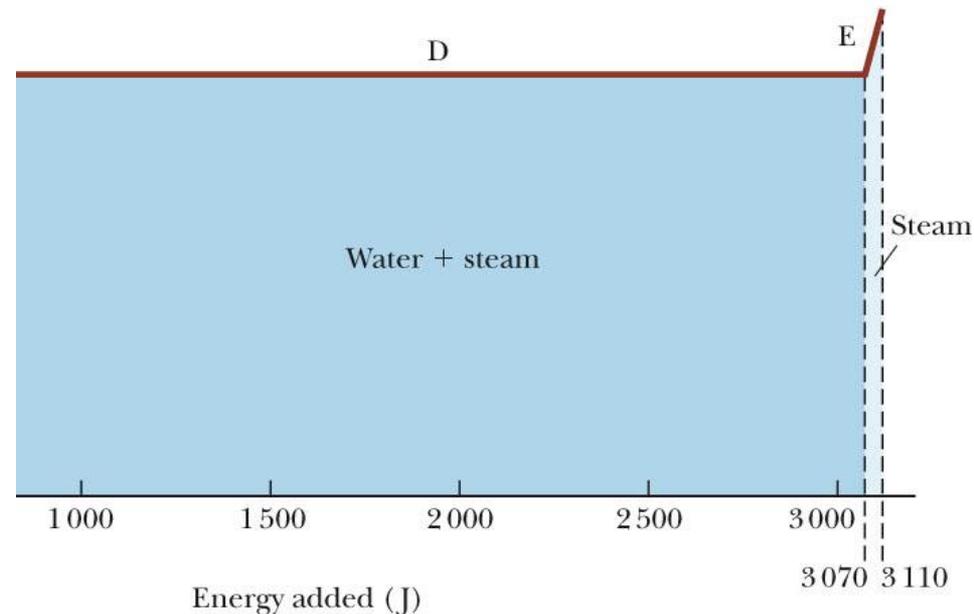
After all the water is converted to steam, the steam will heat up.

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.



# Supercooling

If liquid water is held perfectly still in a very clean container, it is possible for the temperature to drop below  $0^{\circ}\text{C}$  without freezing.

This phenomena is called **supercooling**.

It arises because the water requires a disturbance of some sort for the molecules to move apart and start forming the open ice crystal structures.

- This structure makes the density of ice less than that of water.

If the supercooled water is disturbed, it immediately freezes and the energy released returns the temperature to  $0^{\circ}\text{C}$ .

# Superheating

Water can rise to a temperature greater than 100° C without boiling.

This phenomena is called **superheating**.

The formation of a bubble of steam in the water requires nucleation site.

- This could be a scratch in the container or an impurity in the water.

When disturbed, the superheated water can become explosive.

- The bubbles will immediately form and hot water is forced upward and out of the container.

# 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.

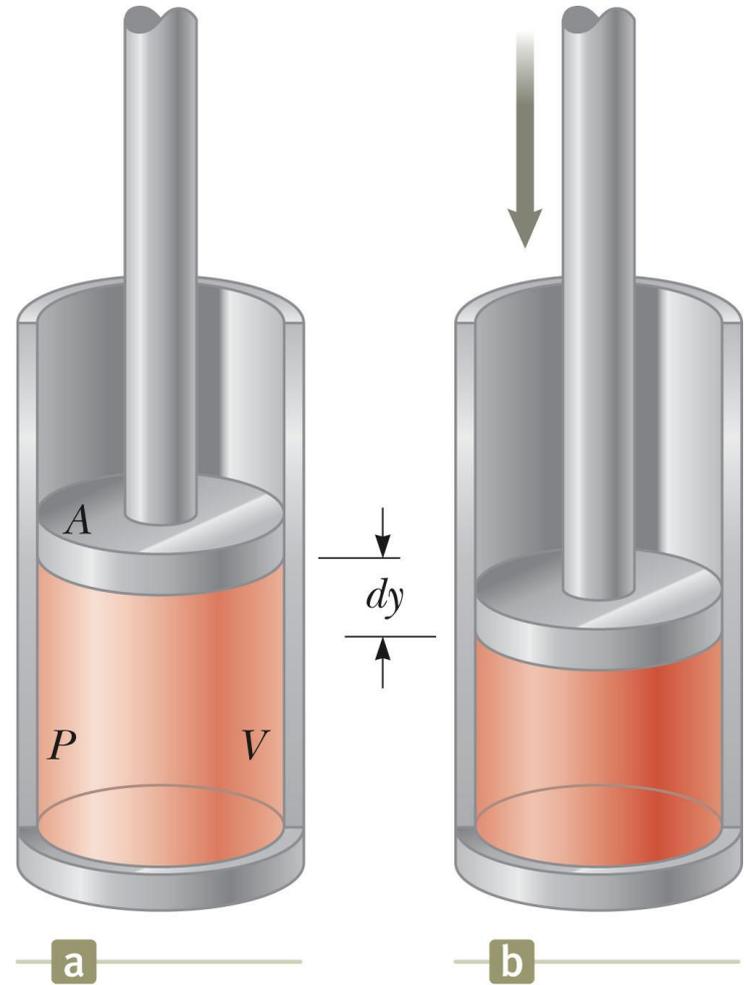
# Work in Thermodynamics

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.

- Force  $F = PA$
- 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:

$$dW = \vec{\mathbf{F}} \cdot d\vec{\mathbf{r}} = -F_{\text{air}} dy \mathbf{j} = -F dy = -PA dy$$

$A \cdot dy$  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:

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

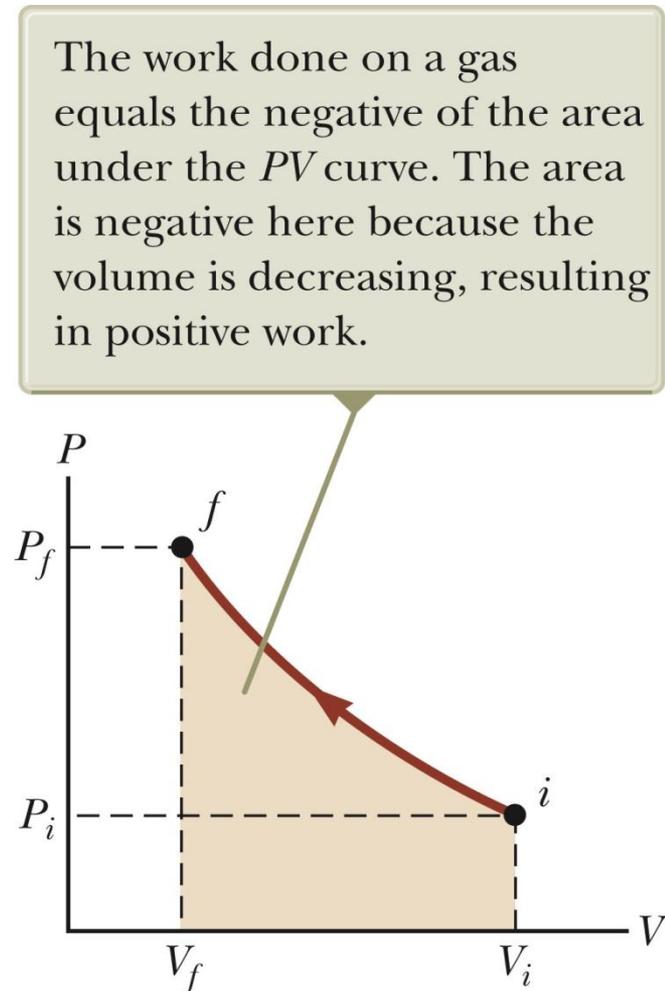
## PV Diagrams

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*.

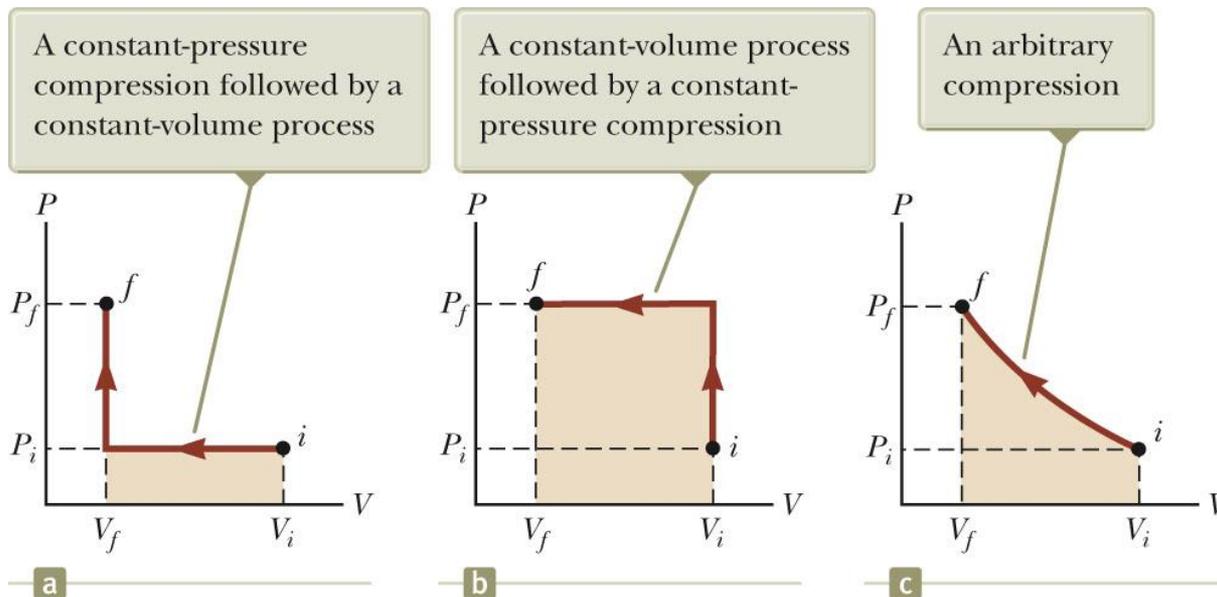


## PV Diagrams, cont

The work done on a gas in a quasi-static process that takes the gas from an initial state to a final state is the negative of the area under the curve on the PV diagram, evaluated between the initial and final states.

- This is true whether or not the pressure stays constant.
- The work done does depend on the path taken.

## Work Done By Various Paths



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

The work done differs in each process.

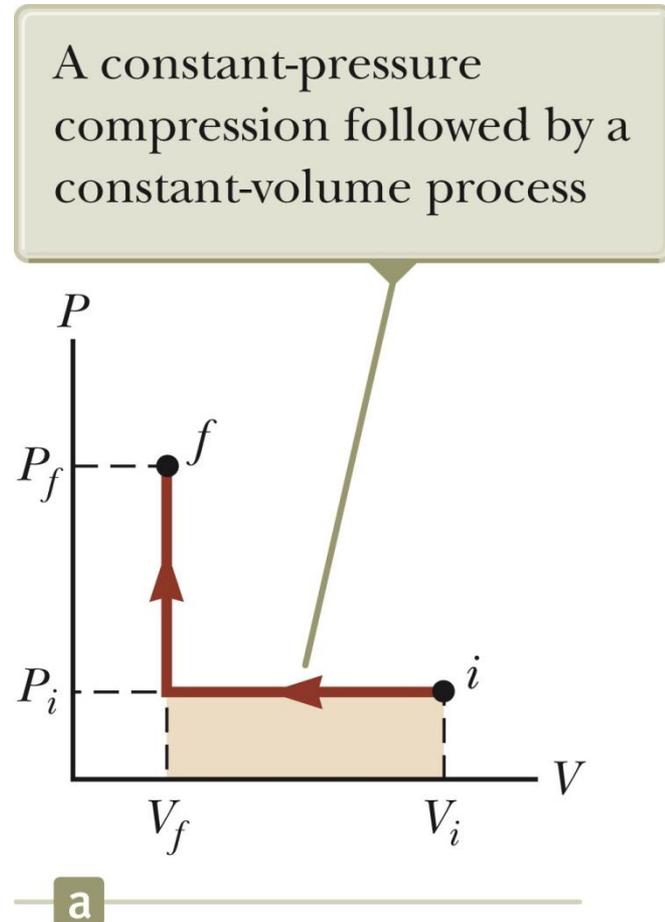
The work done depends on the path.

## 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  $V_f$ .

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



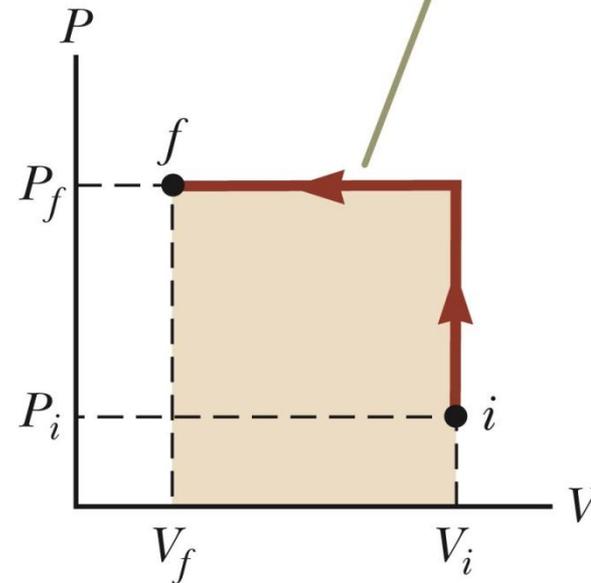
## Work From a $PV$ Diagram, Example 2

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



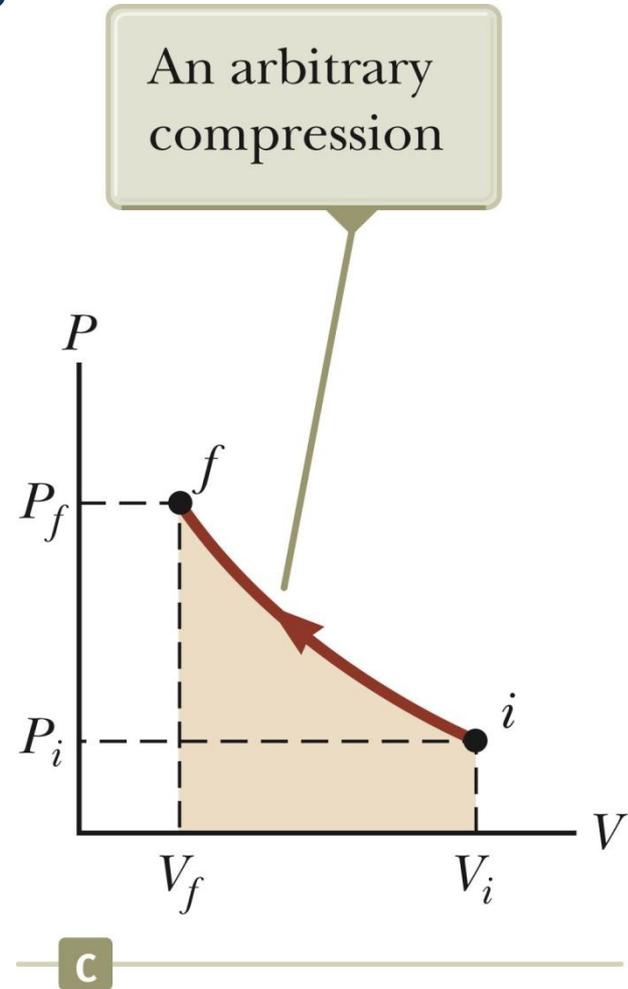
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## Work From a $PV$ Diagram, Example 3

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.

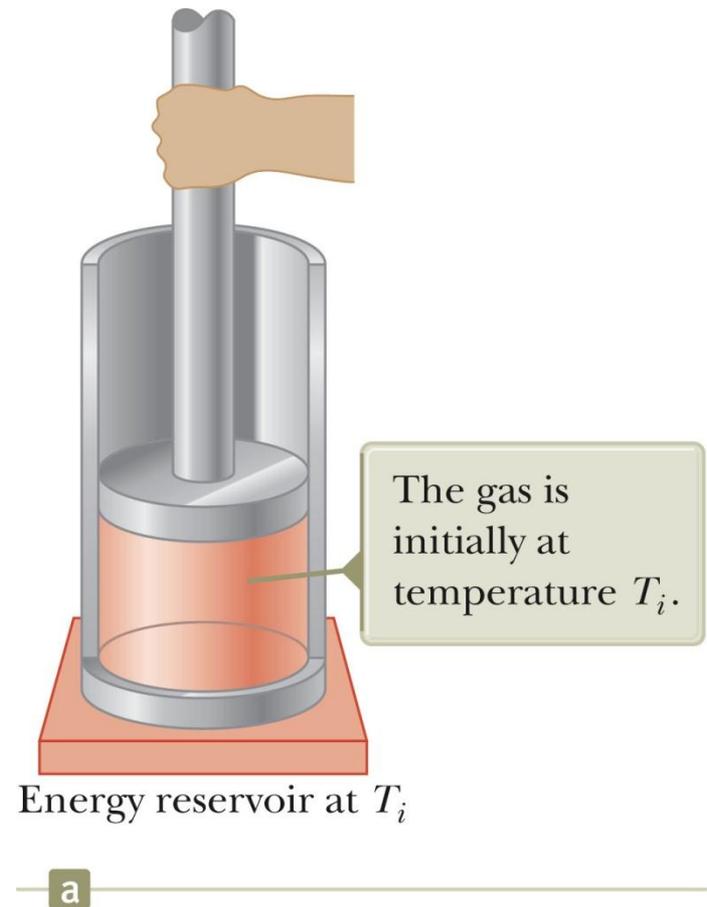


## Energy Transfer, 1

The energy transfer,  $Q$ , into or out of a system also depends on the process.

The **energy reservoir** is a source of energy that is considered to be so great that a finite transfer of energy does not change its temperature.

The piston is held at its internal position by an external agent.

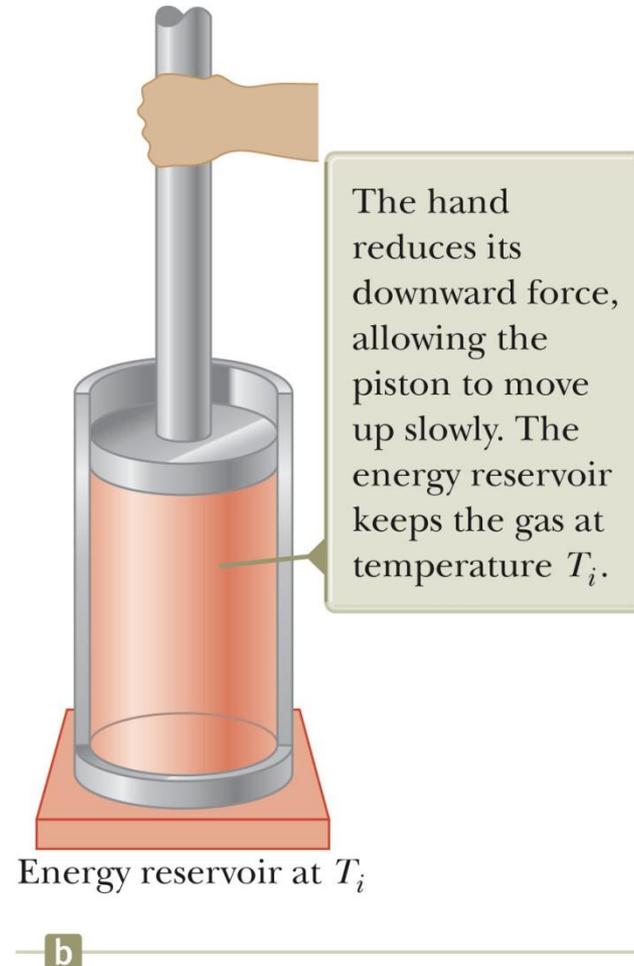


## Energy Transfer, 2

The external force is reduced.

The piston is moving upward and the gas is doing work on the piston.

During this expansion, just enough energy is transferred by heat from the reservoir to the gas to maintain a constant temperature.



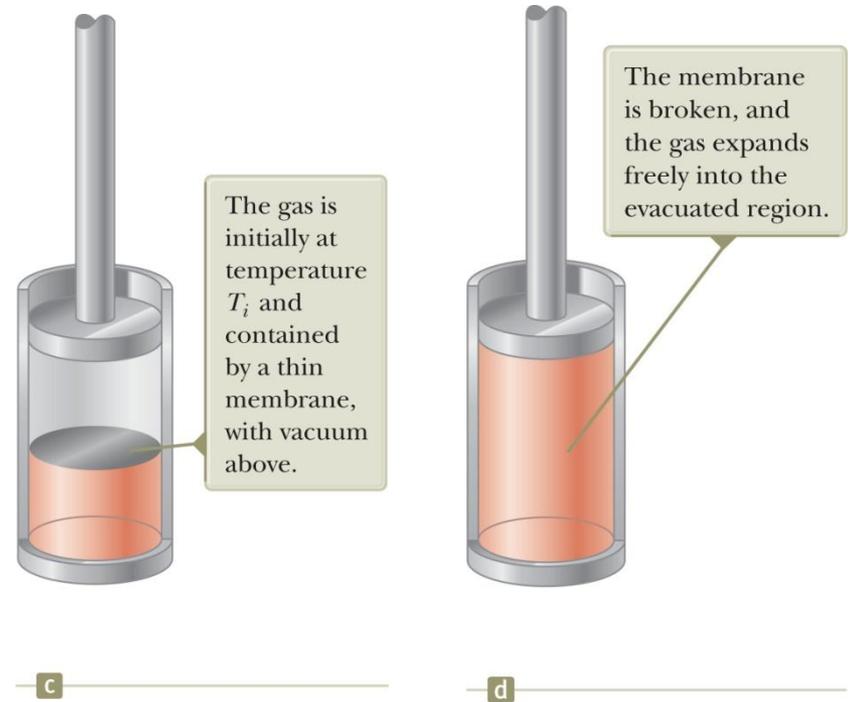
## Energy Transfer, Isolated System 1

The system is completely thermally insulated.

When the membrane is broken, the gas expands rapidly into the vacuum until it comprises the final volume.

The gas does no work because it does not apply a force.

No energy is transferred by heat through the insulating wall.



## Energy Transfer, Summary

Energy transfers by heat, like the work done, depend on the initial, final, and intermediate states of the system.

Both work and heat depend on the path taken.

Neither can be determined solely by the end points of a thermodynamic process.

# 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

Exchange energy with surroundings

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.

Energy  $Q$  added equals negative of work  $W$  done on the system

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

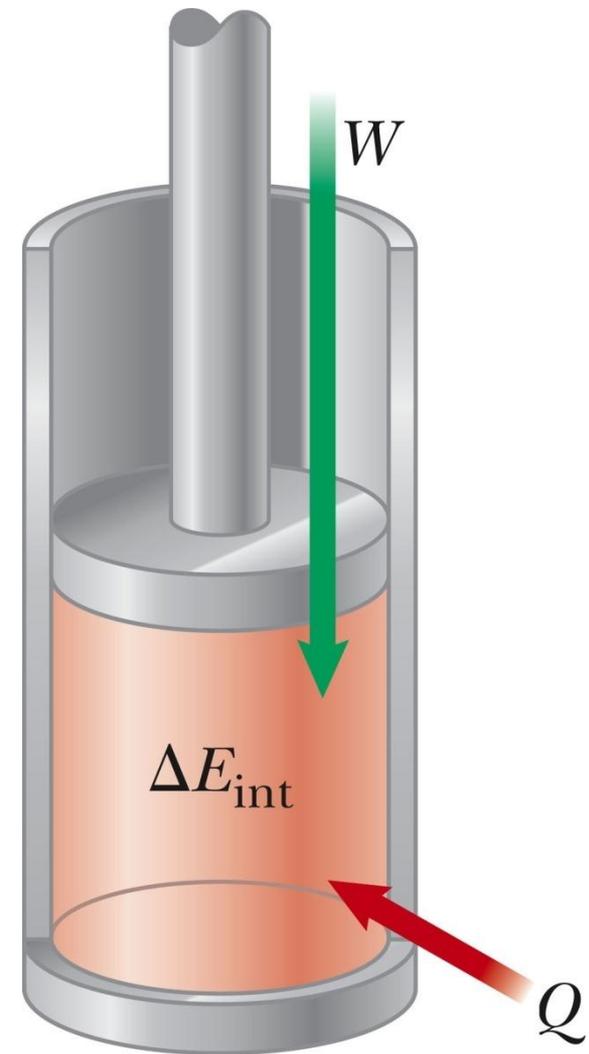
In a cyclic process, the net work done on the system per cycle equals the area enclosed by the path representing the process on a  $PV$  diagram.

# Application of First Law of Thermodynamics

## Adiabatic Processes

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  $\Delta E_{\text{int}}$  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

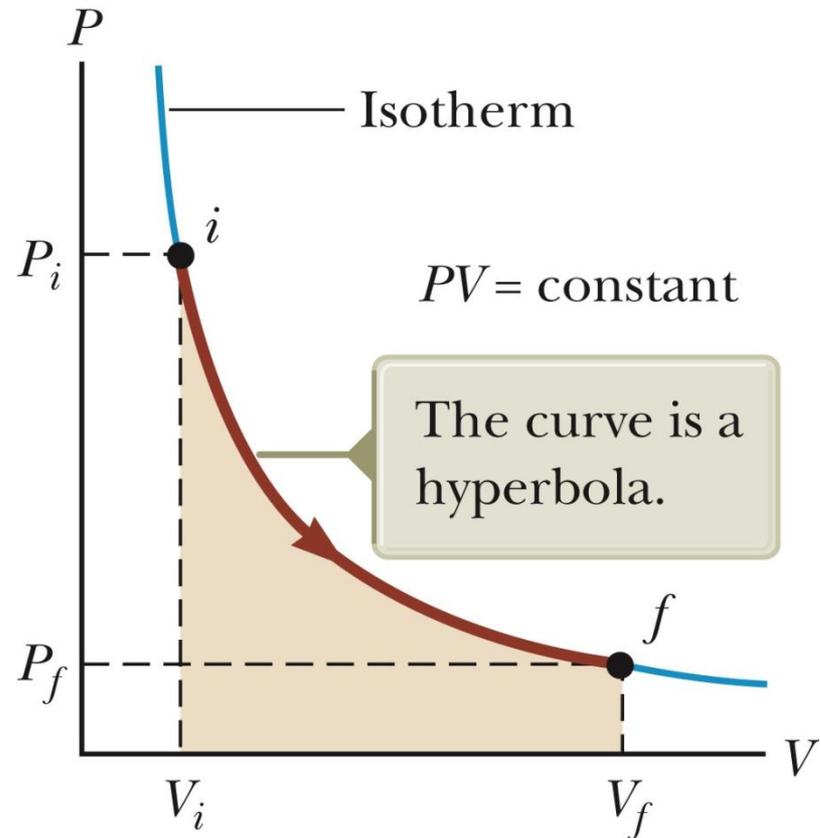
At right is a  $PV$  diagram of an isothermal expansion.

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.

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

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

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

## Isobaric

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

## Isothermal

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

# 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

Conduction /Thermal Conduction - The process of energy transfer by heat

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.

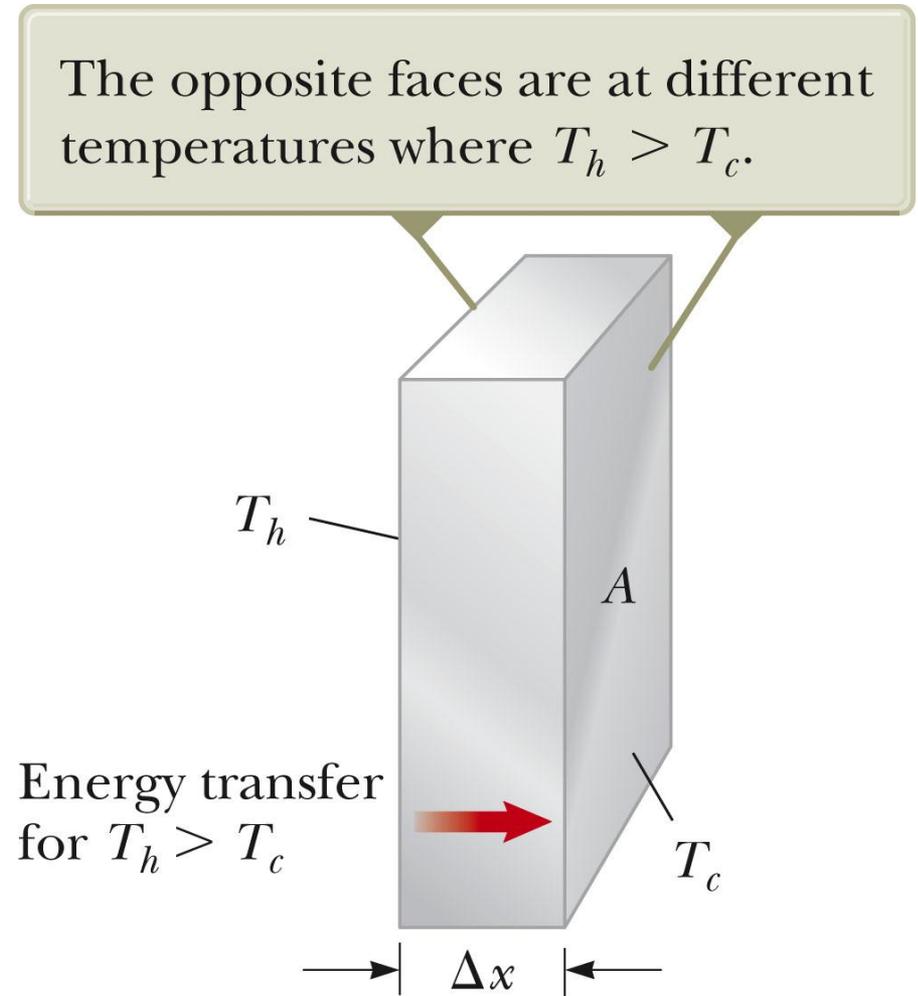
## Conduction, equation

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:

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

(Law of thermal conduction)



## 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*

<b>Substance</b>	<b>Thermal Conductivity (W/m · °C)</b>
<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.

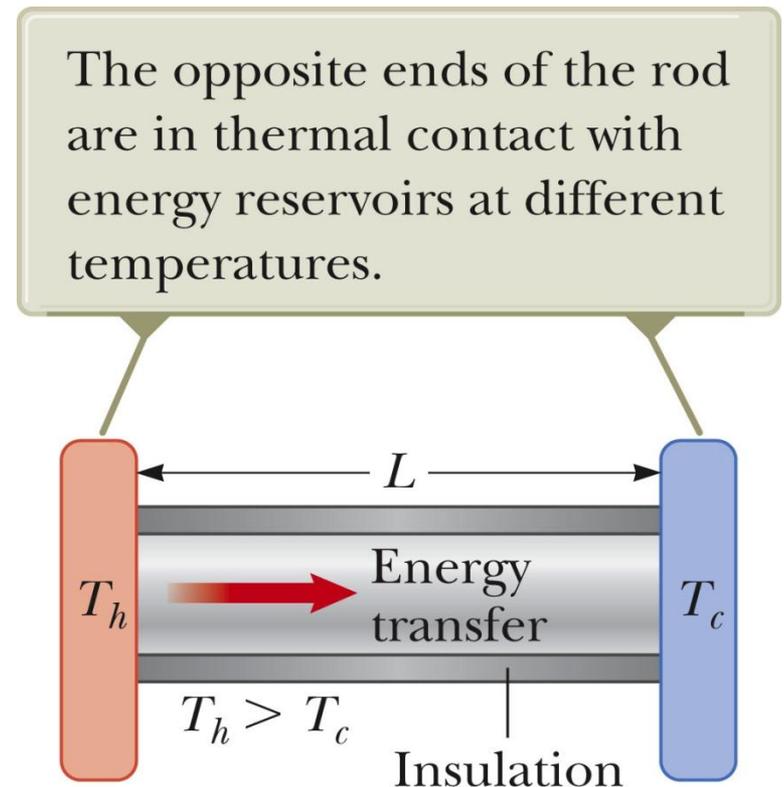
- It measures the rate at which temperature varies with position.

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:

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



## 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*

<b>Material</b>	<b>R-value (ft<sup>2</sup> · °F · h/Btu)</b>
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:

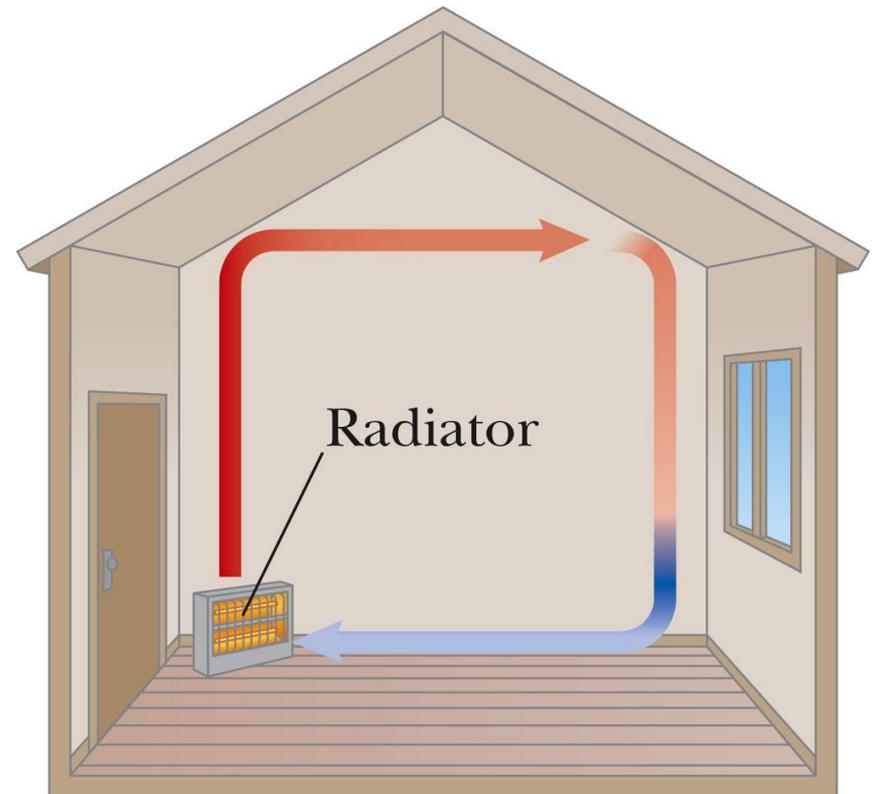
- When the movement results from differences in density, it is called *natural convection*.
- When the movement is forced by a fan or a pump, it is called *forced convection*.

## 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.

# The Dewar Flask

A Dewar flask is a container designed to minimize the energy losses by conduction, convection, and radiation.

- Invented by Sir James Dewar (1842 – 1923)

It is used to store either cold or hot liquids for long periods of time.

- A Thermos bottle is a common household equivalent of a Dewar flask.

## Dewar Flask, Details

The space between the walls is a vacuum to minimize energy transfer by conduction and convection.

The silvered surface minimizes energy transfers by radiation.

- Silver is a good reflector.

The size of the neck is reduced to further minimize energy losses.

