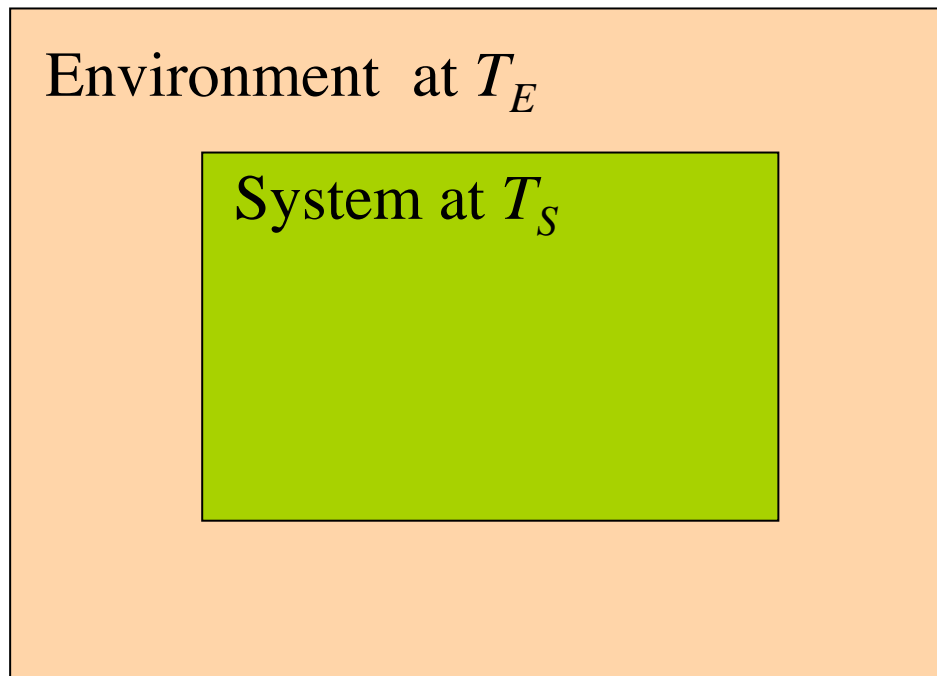


Temperature and Heat

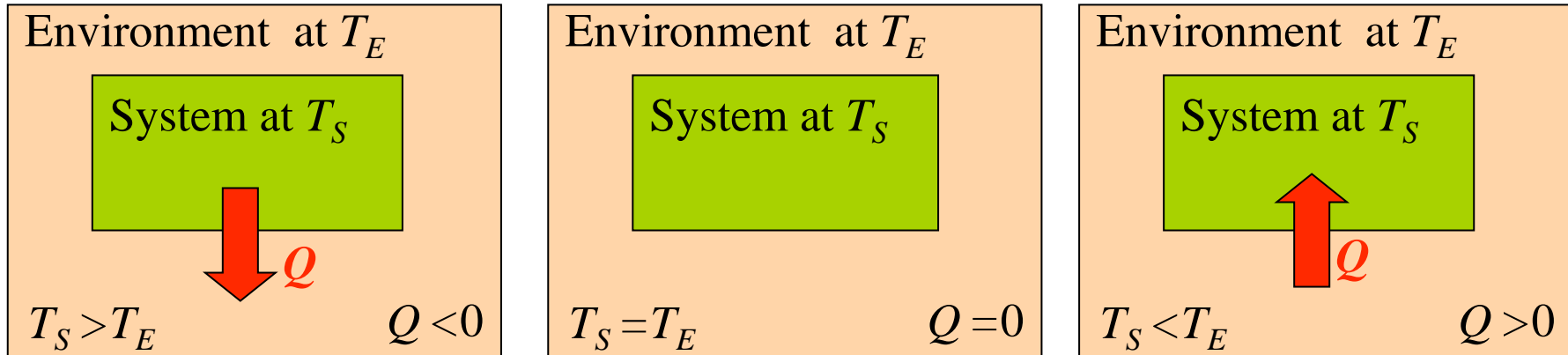
Terminology of Thermodynamics



Closed system
*No energy exchange
with the outside*

Place **system S**, at temperature T_S , within **environment E**, at temperature T_E . If $T_S \neq T_E$ temperatures will change until **thermal equilibrium** is reached ($T_S = T_E$)

Energy transferred by virtue of a temperature difference



Heat leaves S


Q negative ←

Sign convention →

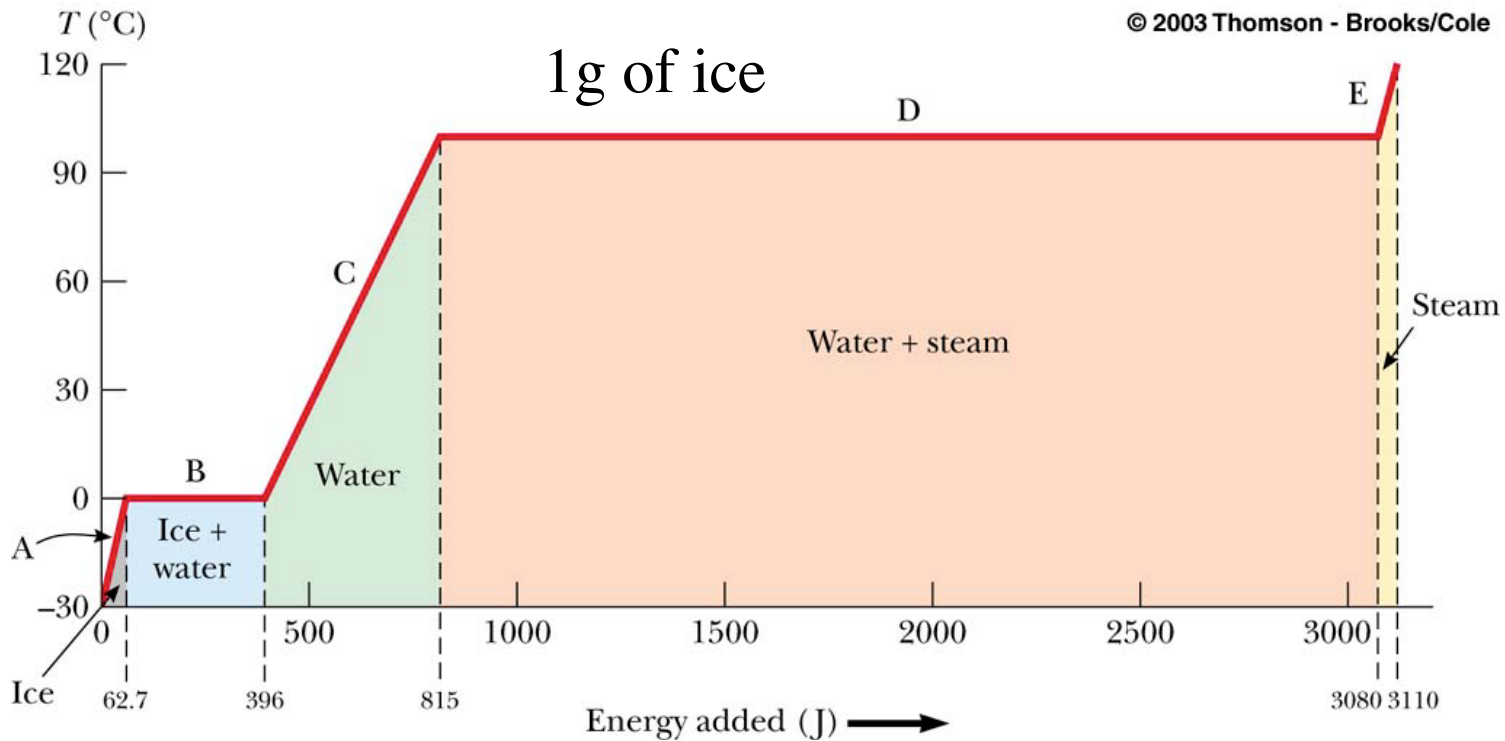
Heat enters S

Q positive

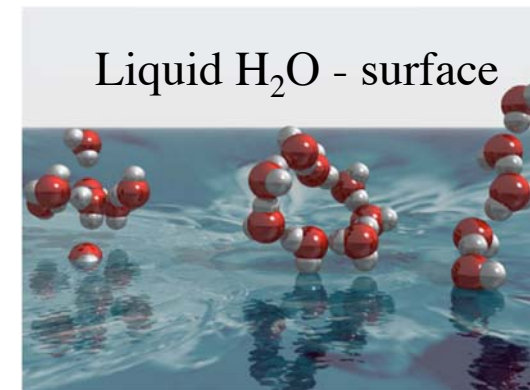
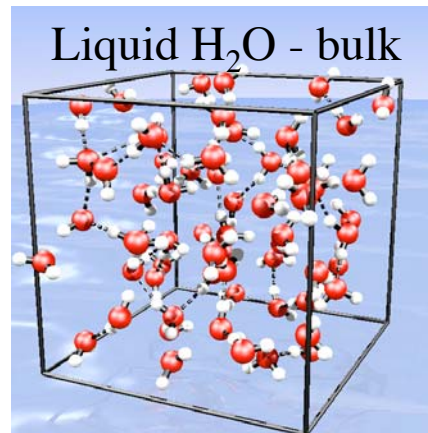
- **Heat** is defined as the transfer of energy across the boundary of a system due to a temperature difference between the system and its surroundings
- The term *heat* will also be used to represent the *amount* of energy transferred by this method

- 
- 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 (here 1 cal = 4.1855 J)
 - The “Calorie” used for food is actually 1 kilocalorie (here 1 Calorie = 4.186 kJ)
 - 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 (from the latin “libra” = “scale”) of water from 63°F to 64°F
 - The **international standard** (SI units) is to use **Joules**

Heat absorption by Solids and Liquids



- Heating results in either
- increase in temperature
 - phase change




Latent Heat

- If an amount of energy Q is required to change the phase of a sample of mass m , then

$$Q = L m$$

The diagram shows the equation $Q = L m$ enclosed in a red box. Three red arrows point from the units to the variables: an arrow from $[J]$ points to Q , an arrow from $[J\ g^{-1}]$ points to L , and an arrow from $[g]$ points to m .

- The quantity L is called the **latent heat** of the material
 - Latent means “hidden” (it is hidden in the chemical bonds of the atoms and molecules)
 - The value of L depends on the substance as well as the actual phase change
- The energy required to change the phase is $Q = \pm mL$

- 
- 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
 - The positive sign is used when the energy is transferred into the system
 - This will result in melting or boiling
 - The negative sign is used when energy is transferred out of the system
 - This will result in freezing or condensation



Sample Latent Heat Values

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
Nitrogen	-209.97	2.55×10^4	-195.81	2.01×10^5
Oxygen	-218.79	1.38×10^4	-182.97	2.13×10^5
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



Molecular View of Phase Changes

- Phase changes can be described in terms of the rearrangement of molecules (or atoms in an elemental substance)
- Liquid to Gas phase change
 - Molecules in a liquid are close together
 - The forces between them are stronger than those in a gas
 - Work must be done to separate the molecules
 - The latent heat of vaporization is the energy per unit mass needed to accomplish this separation



- Solid to Liquid phase change

- The addition of energy will cause the amplitude of the vibration of the molecules about their equilibrium position to increase
- At the melting point, the amplitude is great enough to break apart bonds between the molecules
- The molecules can move to new positions
- The molecules in the liquid are bound together less strongly than those of the solid
- The latent heat of fusion is the energy per unit mass required to go from the solid-type to the liquid-type bonds



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 (1K)
- If energy Q produces a change of temperature of ΔT , then


$$Q = C \Delta T$$


- **Specific heat**, c , is the heat capacity per unit mass

$$Q = cm\Delta T$$

$$C = cm$$

Units: Q in J, C in J K^{-1} , c in $\text{J K}^{-1} \text{g}^{-1}$ or $\text{J K}^{-1} \text{mol}^{-1}$

- 
- The specific heat is essentially a measure of how insensitive a substance is to the addition of energy
 - The greater the substance's specific heat, the more energy that must be added to cause a particular temperature change
 - 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

- 
- Technically, the specific heat varies with temperature
 - The corrected equation is $Q = m \int_{T_i}^{T_f} c(T) dT$
 - However, if the temperature intervals are not too large, the variation can be ignored and c can be treated as a constant

$$Q = m \int_{T_i}^{T_f} c(T) dT \approx mc \int_{T_i}^{T_f} dT = mc(T_f - T_i) = mc\Delta T$$

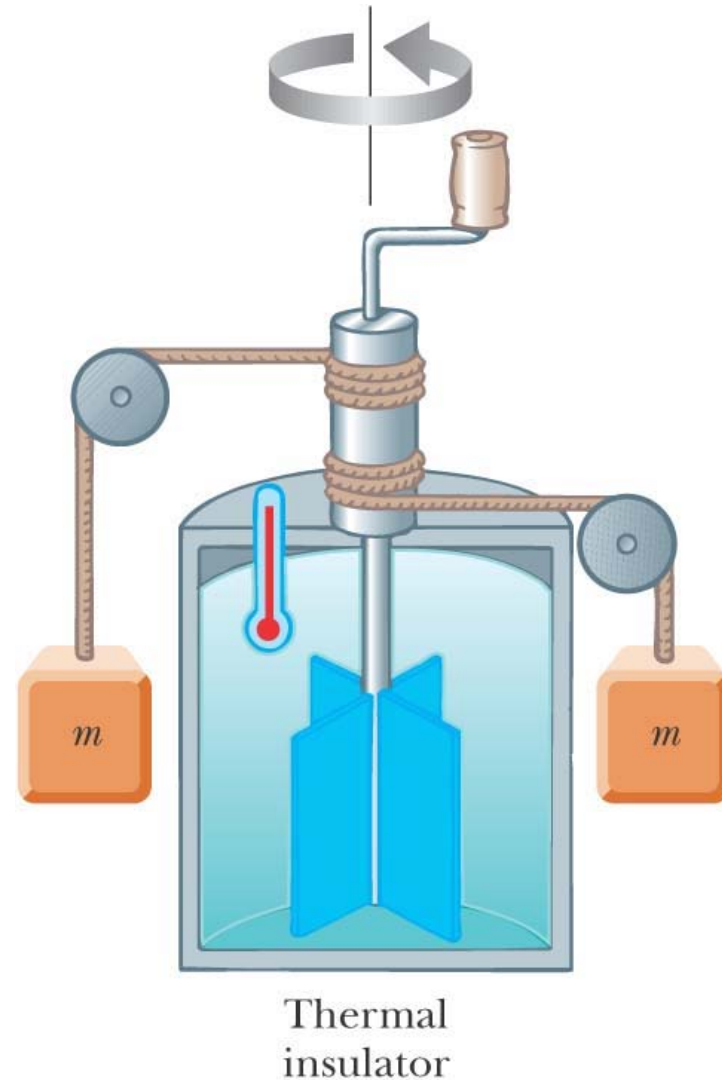
- Typically there is only about a 1% variation in c between 0° and 100°C
- Water has the highest specific heat of common materials
- This is responsible for many weather phenomena like
 - Moderate temperatures near large bodies of water
 - Global wind systems
 - Land and sea breezes


Some Specific Heat Values c

	J/kg·°C	cal/g·°C		J/kg·°C	cal/g·°C
<i>Elemental Solids</i>			<i>Other solids</i>		
Aluminium	900	0.215	Brass	380	0.092
Beryllium	1830	0.436	Glass	837	0.200
Cadmium	230	0.055	Ice (-5°C)	2090	0.50
Copper	387	0.092	Marble	860	0.21
Germanium	322	0.077	Wood	1700	0.41
Gold	129	0.030	<i>Liquids</i>		
Iron	448	0.107	Alcohol	2400	0.58
Lead	128	0.030	Mercury	140	0.033
Silicon	703	0.168	Water	4186	1.00
Silver	234	0.056	<i>Gas</i>		
			Water steam	2010	0.48

Mechanical Equivalent of Heat

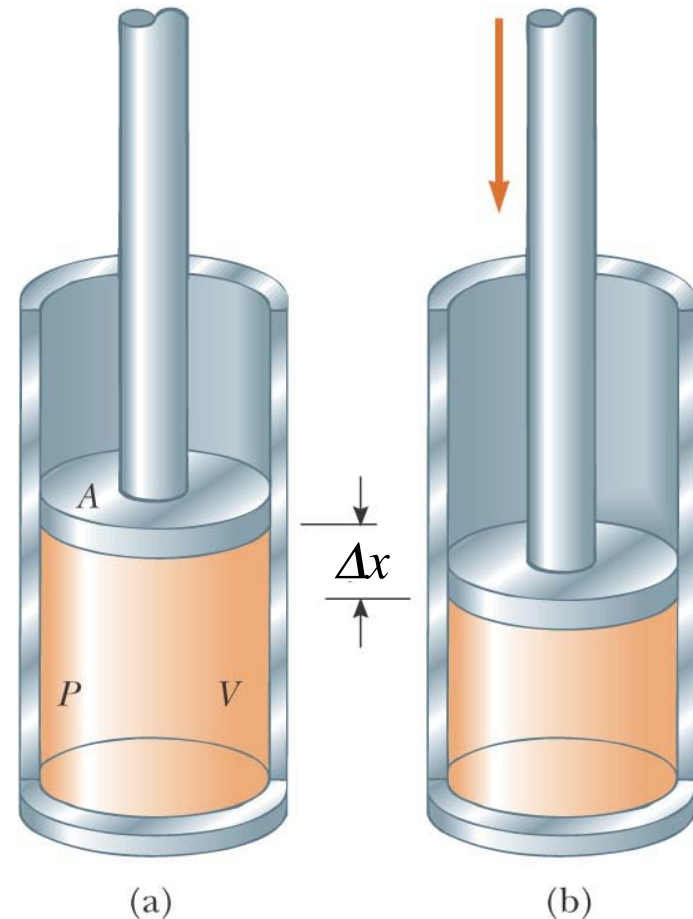
- Joule established the equivalence between mechanical energy and internal energy
- His experimental setup is shown on the right
- The loss in potential energy associated with the blocks equals the work done by the paddle wheel on the water



- 
- Joule found that it took approximately 4.18 J of mechanical energy to raise the water (1 g) by 1°C
 - Later, more precise, measurements determined the amount of mechanical energy needed to raise the temperature of 1g of water from 14.5°C to 15.5°C
 - 1 cal = 4.186 J
 - This is known as the **mechanical equivalent of heat**

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** (i.e. by slow dynamics)



© 2004 Thomson/Brooks Cole

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

$$\Delta W = -F\Delta x = -PA\Delta x = -P\Delta V \quad \text{or} \quad dW = -PdV$$

- Interpreting $dW = -P dV$ (*differential form*)

- If the gas is compressed, dV is negative and the work done on the gas is positive

(NB: $\Delta V = V_f - V_i$ and $V_f < V_i$)

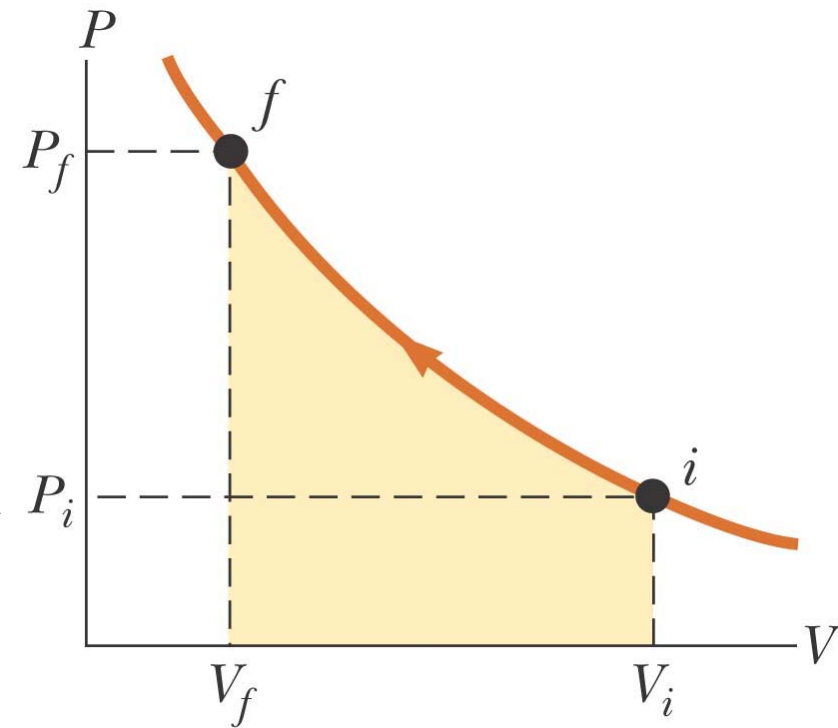
- If the gas expands, dV is positive and the work done on the gas is negative

- If the volume remains constant, the work done is zero

- The total work done is: $W = -\int_{V_i}^{V_f} PdV$ (*integral form*)

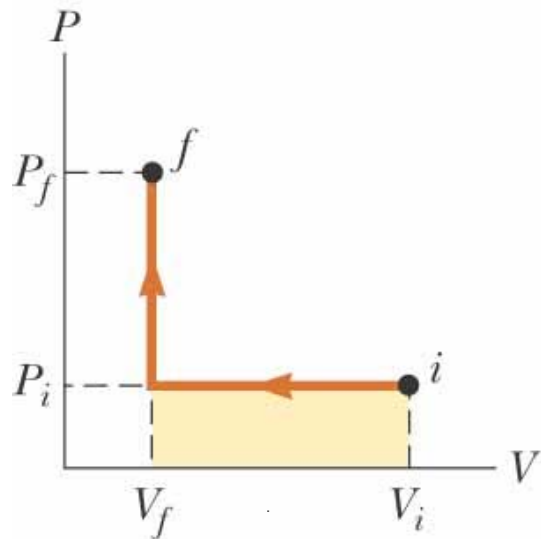
PV Diagrams for Gases

- 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*

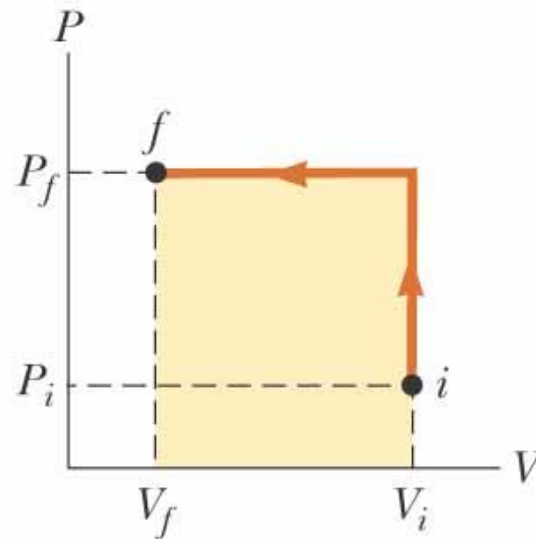


© 2004 Thomson/Brooks Cole

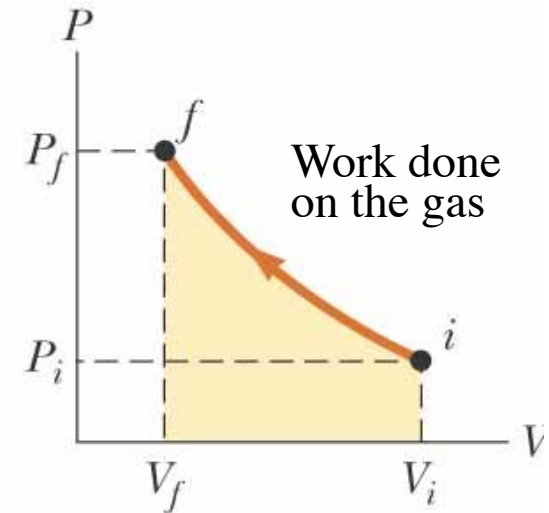
Work Done By Various Paths



(a)



(b)



(c)

© 2004 Thomson/Brooks Cole

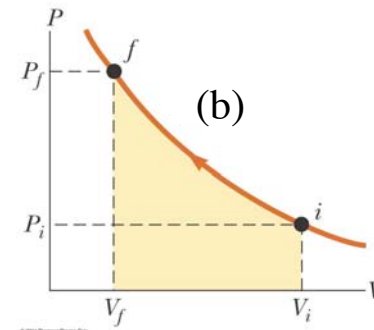
- Each of these processes has the same initial and final states
- The work done differs in each process
- The work done depends on the path

- 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

Path (a): $W_a = -P_i(V_f - V_i)$

Path (b): $W_b = -P_f(V_f - V_i)$

Path (c): $W_c = -\int_{V_i}^{V_f} P(V)dV$



$\longrightarrow W_b < W_c < W_a$

Example: An ingot of metal (50 g) is heated to 200 °C and then dropped into a beaker of 20 °C water (400 mL). The equilibrium temperature is measured. The temperature of water raises to 22.4 °C. What metal was used?

raising the temp. of water cooling the metal


$\Delta T = T_f - T_i$

$$Q_{cold} = -Q_{hot} \quad \Rightarrow \quad c_w m_w \Delta T_w = -c_M m_M \Delta T_M$$

$$c_M = -\frac{c_w m_w \Delta T_w}{m_M \Delta T_M} = \frac{c_w m_w (T_f - T_w)}{m_M (T_M - T_f)} = \frac{4186 \text{ J kg}^{-1} \text{ }^\circ\text{C}^{-1} \cdot 0.4 \text{ kg} (22.4^\circ\text{C} - 20.0^\circ\text{C})}{0.05 \text{ kg} (200^\circ\text{C} - 22.4^\circ\text{C})}$$

$$= 453 \text{ J kg}^{-1} \text{ }^\circ\text{C}^{-1} = 453 \text{ J kg}^{-1} \text{ K}^{-1}$$

Iron has $c = 448 \text{ J kg}^{-1} \text{ }^\circ\text{C}^{-1}$, which is close to our calculated value (within the experimental error)

- 
- One technique for measuring the 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
 - 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: $Q_{\text{cold}} = -Q_{\text{hot}}$
 - This technique is also used for measuring heat released in chemical reactions (e.g. combustions)