

Heat and Thermal Energy

Heat - an energy transfer that occurs because of a difference in temperature.

Internal Energy - the energy a substance has because of its temperature.

Energy may be transferred between two objects without heat flow.

Example: rubbing two coins together. Both internal energies are increased due to mechanical work but both remain in thermal equilibrium throughout.

Units of Heat

- Calorie - the amount of heat energy required to raise the temperature of 1 gram of H₂O from 14.5°C to 15.5°C. A *food calorie* is 10³ "physics" calories
- BTU - the amount of heat energy required to raise the temperature of 1 lb. of H₂O from 63°F to 64°F.
- A candy bar contains more energy (by a fair amount) than a stick of dynamite
- Joule - the S.I. unit of heat and work.

$$1 \text{ calorie} = 4.186 \text{ Joules} = 3.9 \times 10^{-3} \text{ Btu}$$

Example: A student eats a 1000 food calorie meal consisting of pizza, soda, and snacks. How many 100kg clean & jerks (225 lbs.) must they do to work off this meal just by lifting weights (neglecting metabolism)?

Work required is:

$$1000 \text{ food cal} \times 10^3 \text{ cal} = 1 \times 10^6 \text{ calories} = 4.186 \times 10^6 \text{ J}$$

The work done lifting weights is against gravity $\therefore w = mgh$. For n repetitions through some height h , $w = nmgh$. Assume $h = 2.0$ meters.

$$n = \frac{4.186 \times 10^6 \text{ J}}{(100\text{kg})(9.8\text{m} \cdot \text{s}^{-2})(2.0\text{m})} \approx 2136\text{reps!}$$

How far would you have to run uphill to burn off these calories (neglecting metabolism)?

4271 meters!

Heat Capacity and Specific Heat

- The *heat capacity* (C) of a substance is the amount of heat energy required to raise the temperature of a given mass of the substance by some amount. Heat capacity varies from substance to substance.
- The *specific heat* (C/m) of a substance is the heat capacity per unit mass of the substance.

The heat transferred between a substance and its surroundings not in thermal equilibrium may be expressed in terms of specific heat as (if the temperature range is small):

$$Q = mc\Delta T$$

Calorimetry - measuring specific heat

Conservation of energy requires that for objects in thermal contact in an isolated system any heat leaving a hot object must equal the heat entering the cold object.

Cooling a hot ingot:

One method for measuring the specific heat of a substance is to heat it to a known temperature, place it in an insulated vessel filled with water ($c = 4186 \text{ J/kg}\cdot^\circ\text{C}$) of known mass and temperature, and measure the temperature of the water after equilibrium (22.4°C) has been reached.

heat gained = -heat lost

$$m_w c_w (T_f - T_{wi}) = -m_x c_x (T_f - T_{xi}) \therefore c_x = \frac{m_w c_w (T_f - T_{wi})}{m_x (T_{xi} - T_f)}$$

ingot $m_x = 0.05\text{kg}$ $T_{xi} = 200^\circ\text{C}$ $c_x = ?$

water $m_w = 0.4\text{kg}$ $T_{wi} = 20^\circ\text{C}$ $c_w = 4186 \text{ J/kg}\cdot^\circ\text{C}$

$$m_w c_w (T_f - T_{wi}) = m_x c_x (T_{xi} - T_f) \therefore c_x = 453 \text{ J/kg}\cdot^\circ\text{C}$$

According to CRC tables, this closely matches iron @ $448 \text{ J/kg}\cdot^\circ\text{C}$.

Example: A silver bullet of mass 2 grams (0.002kg) impacts a well insulated wall at 200 m/s. How much does the temperature of the bullet increase upon impact?

$$w = \Delta KE = Q = mc\Delta t$$

$$\Delta KE = 40J = mc\Delta t \Rightarrow \Delta t = \frac{40J}{(2 \times 10^{-3} kg)(234 J/kg^{\circ}C)} = 85.5^{\circ}C$$

Note: Room temperature is normally between 20 and 25^o C. What does this imply about the temperature of the bullet after it comes to rest?

Latent Heat

- Normally the transfer of heat to a substance from its environment results in a change in temperature of the substance.
- When a substance undergoes a *change in phase* it is possible to add or extract quite a bit of heat to or from a substance without changing its temperature.

Example Phase Changes

- solid → liquid
- liquid → gas
- crystal A → crystal B

All phase changes involve a change in internal energy known as the *heat of transformation*. The energy required to change the phase of a mass m of a given substance is:

$$Q = mL$$

where L is the *latent heat* of the substance (J/kg).

Heat of fusion (L_f) is used when the phase change is from solid to liquid, and *heat of vaporization* (L_v) is used when the phase change is from liquid to vapor, a.k.a. heats of solidification and condensation. Heats of vaporization are normally much greater than heats of fusion.

Phase changes involve dramatic adjustment in *long range order*.

- Solids - each atom in a solid can translate, vibrate, or rotate about a fixed equilibrium position. Rotational states tend to be very low energy while vibrational states tend to be a little higher. As heat is added to a solid substance the vibrations of atoms become large enough in amplitude that the attractive forces between atoms are overcome so that they no longer occupy fixed positions.
- Liquids - less long range order. Atoms or molecules exist in small groups weakly bonded to each other.
- Gasses - No long range order.

More work is required to vaporize a substance than to melt it because the average distance between atoms/molecules is greater.

$$W = Fs \cos \theta, \text{ as } s \text{ increases so does } W \therefore Q.$$

Example: The heat required to convert 1 gram of ice at -30°C to steam at 120°C .

1. The heat required to change the temperature of the ice from $-30^{\circ}\text{C} \rightarrow 0^{\circ}\text{C}$, no phase change, is $Q = mc\Delta t$.

$$Q = (mc\Delta t)_{ice} = (10^{-3}\text{kg})(2090\text{ J/kg}^{\circ}\text{C})(30^{\circ}\text{C}) = 62.7\text{ J}$$

Total heat added = 62.7 Joules

2. When the ice reaches 0°C it begins to melt. The ice/water mixture remains at 0°C even though heat is being added because of the phase change. This continues until all of the ice melts. The heat required for the mixture to complete the phase change is $Q = mL_f$.

$$Q = mL_f = (10^{-3}\text{ kg})(3.33 \times 10^5\text{ J/kg}) = 333\text{ J}$$

Total heat added = 395.7 Joules

3. The heat required to change the temperature of the water from $0^{\circ}\text{C} \rightarrow 100^{\circ}\text{C}$, no phase change, is $Q = mc\Delta t$.

$$Q = (mc\Delta t)_{water} = (10^{-3}\text{kg})(4.19 \times 10^3\text{ J/kg}^{\circ}\text{C})(100^{\circ}\text{C}) = 419\text{ J}$$

Total heat added = 813.7 Joules

4. When the water reaches 100°C it begins to boil. The water/steam mixture remains at 100°C even though heat is being added because of the phase change. This continues until all of the water is converted to steam. The heat required for the mixture to complete the phase change is $Q = mL_v$.

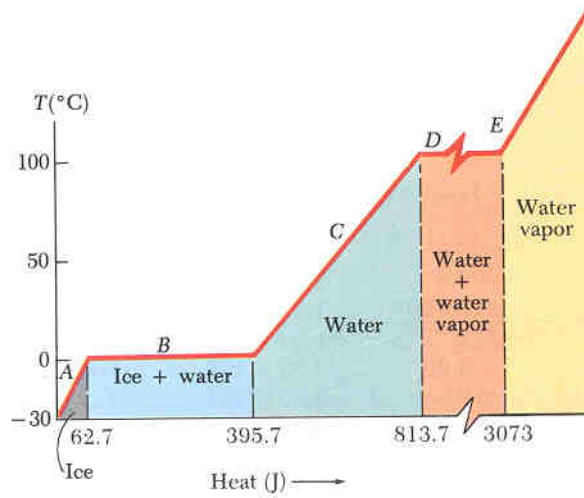
$$Q = mL_v = (10^{-3}\text{ kg})(2.26 \times 10^6\text{ J/kg}) = 2260\text{ J}$$

Total heat added = 3073.7 Joules

5. The heat required to change the temperature of the steam from $100^{\circ}\text{C} \rightarrow 120^{\circ}\text{C}$, no phase change, is $Q = mc\Delta t$.

$$Q = (mc\Delta t)_{\text{steam}} = (10^{-3}\text{kg})(2.01 \times 10^3 \text{ J/kg}^{\circ}\text{C})(20^{\circ}\text{C}) = 40.2 \text{ J}$$

Total heat added = 3113.9 Joules



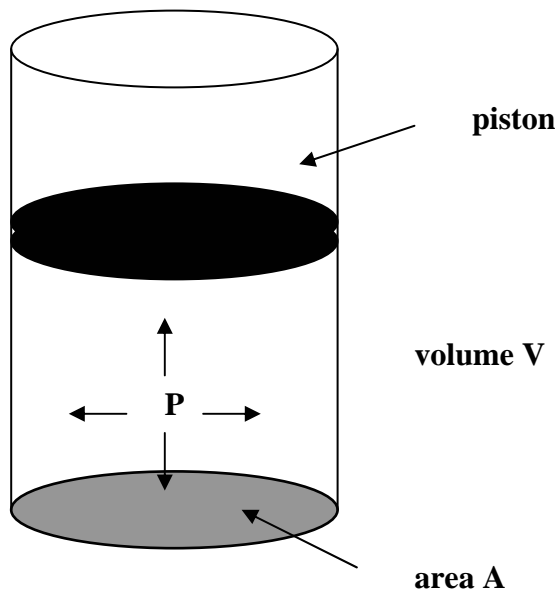
Work and Heat in Thermodynamic Processes

The macroscopic state of a system is defined by state variables:

1. pressure
 2. volume
 3. temperature
 4. internal energy
 5. entropy
- The number of variables needed to completely describe a system depends upon the nature of the system.
 - Homogeneous systems, e.g., molecular gasses such as N_2 , O_2 , etc., usually require only two variables.
 - The state must be the same everywhere within the system in order for a macroscopic state to be specified.

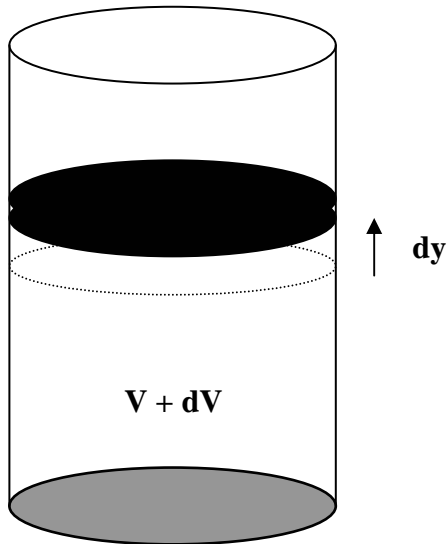
Work done by expansion of a gas

Consider a gas contained to a cylinder with a moveable piston.



- In equilibrium, the gas occupies volume V and exerts uniform pressure P on the walls of the cylinder and the piston.
- If the piston has a cross sectional area A , the force on it due to the gas pressure is: $F = PA$.

- Assume a *quasi-static* expansion, e.g., slow enough that the system remains in equilibrium at all times.



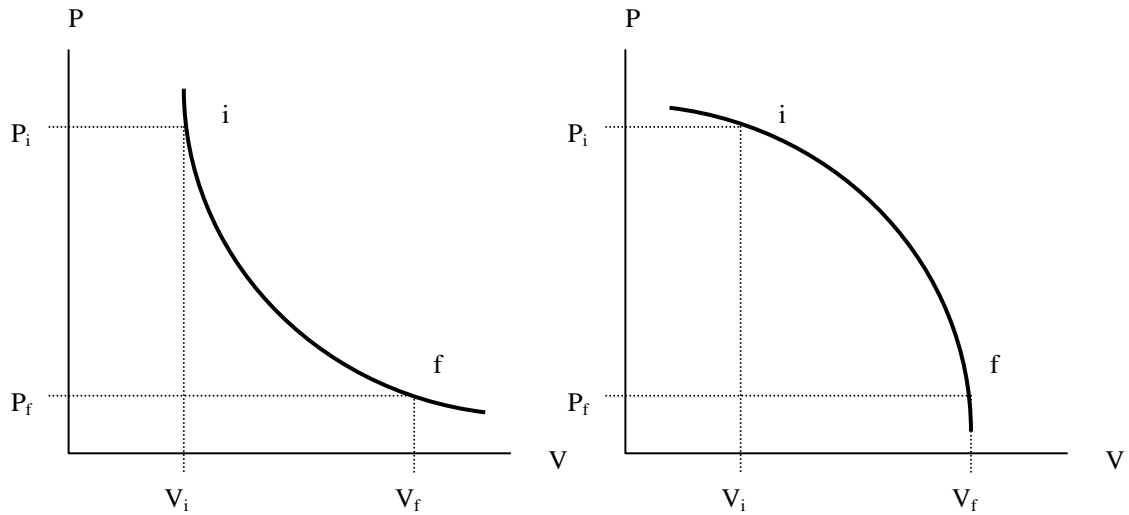
- As the piston moves a distance Δy : $\Delta W = F\Delta y = P\Delta y = P\Delta V \Rightarrow \Delta W = P\Delta V$
- If ΔV increases (the gas expands), the work done by the gas is positive (+), i.e., if the gas expands, the work done by the gas on the system is positive (+).
- If ΔV decreases (the gas is compressed), the work done by the gas is negative (-), i.e., if the gas is compressed work is done on the gas by the system.
- If the volume remains constant, $W = 0$.

Symbolically: $W = P\Delta V$

Definition: The work done in the expansion of an ideal gas from the initial state to the final state is equal to the area under the curve in a PV diagram.

Work and Path Taken by Expansion of an Ideal Gas

Consider our ideal gas expanding quasi-statically from state i to state f . Does work depend upon the path taken?

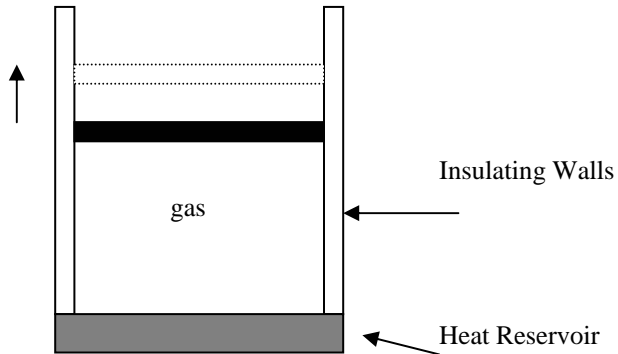


Yes, since the areas under the curves differ.

- Work depends upon shape of the PV curve.
- The work done by a system depends upon the process by which the system goes from the initial to the final state.
- *Heat transfer* depends upon intermediate states as well as initial and final states.

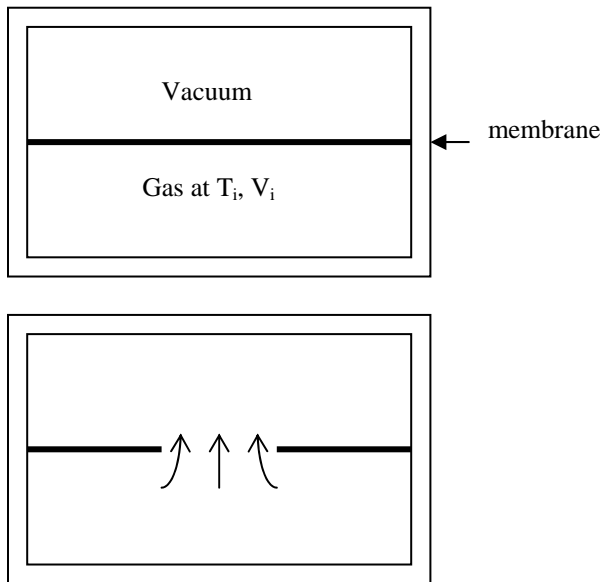
Consider two thermodynamic processes with the same initial and final states:

Isothermal - overall heat of the system remains constant



- As the gas expands due to the addition of heat from the reservoir the piston rises.
- Enough heat is transferred from the reservoir during the process to maintain the gas at a constant temperature.
- Perfectly insulating walls prevent heat loss.

Adiabatic - No heat transfer in or out of the system



- Adiabatic free expansion (no work is done).
- The gas expands without doing work.
- Walls are perfect insulators so no heat is transferred in or out of the system.

- In both the isothermal and adiabatic processes the initial and final states were the same but the intermediate states were different, i.e., the path between the initial and final states was different in each case.
- Since heat and work depend upon path neither process is *independently* conserved during a thermodynamic process.

The First Law of Thermodynamics

The First Law is a statement of energy conservation that takes into account changes in internal energy.

So far we have defined energy conservation in terms of conservation of mechanical energy, $E_i = E_f - W_{nc}$.

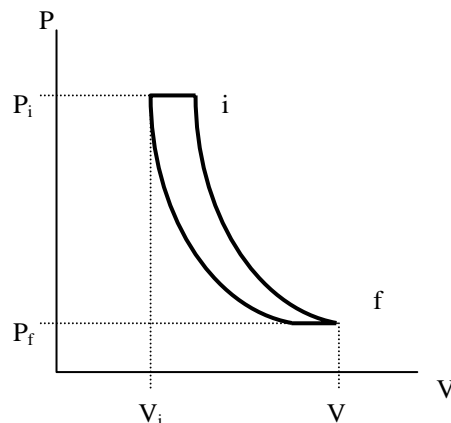
$$\Delta U = U_f - U_i = Q - W \quad (\text{The First Law of Thermodynamics})$$

- +Q - the system absorbs heat
- -Q - the system loses heat
- +W - the system does work on the environment
- -W - the environment does work on the system

Consider a thermodynamic process: $P_i V_i \rightarrow P_f V_f$. While Q depends upon the path taken, and W independently depends upon the path taken, the quantity $Q - W$ (the change in internal energy) is independent of path, i.e., depends only upon initial and final states.

Special Cases of the First Law

1. Isolated Systems - no interaction with the environment \Rightarrow no heat flow, no work done. $Q = W = 0$, $\Delta U = 0$, $U_i = U_f \therefore$ the internal energy of an isolated system remains constant.
2. Cyclic processes - originate and end in the same thermodynamic state. $Q = W$, $\Delta U = 0$. Net work done per cycle equals the area enclosed by a PV curve.

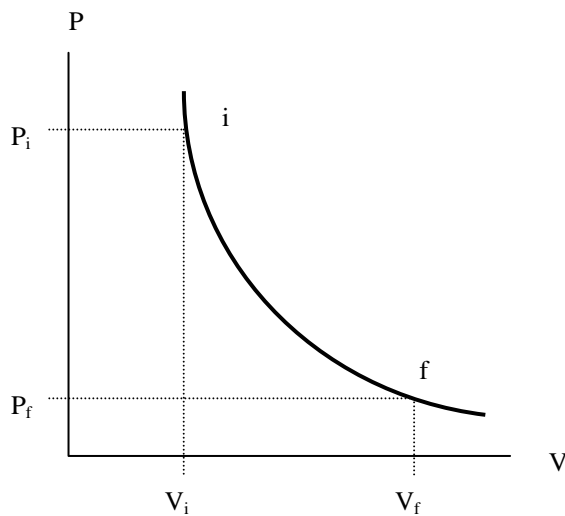


3. Processes in which $W = 0 \rightarrow$ internal energy increases or decreases as heat enters or leaves the system.

Applications of the First Law

Computing Work in Common Thermodynamic Processes

- **adiabatic** - a process for which no heat enters or leaves the system, i.e., $Q = 0$, $\Rightarrow \Delta U = -W$. An adiabatic process may be achieved by either placing the thermodynamic system in a highly insulated environment or by having the system evolve rapidly enough that heat transfer is inconsequential.
- **isobaric** - a process that occurs at constant pressure. Both Q and W are nonzero in such a process. $W = P(V_f - V_i)$.
- **isovolumetric** - a process that takes place at constant volume. $W = 0$, $\Rightarrow \Delta U = Q$. Heat added to a system in an isovolumetric process goes directly into increasing the internal energy of the system.
- **isothermal** - a process that occurs at a constant temperature. The product PV is constant (though P and V vary independently). The PV curve for an isothermal process is known as an *isotherm*. The internal energy of an ideal gas is a function of temperature only, hence $\Delta U = 0$ for an ideal gas in an isothermal process.



Consider the isothermal, quasi-static expansion of an ideal gas as shown in the PV diagram left.

The work done by the gas in expanding from state i to state f is the area under the PV curve from V_i to V_f .

Placing the gas in thermal contact with a heat reservoir at the same temperature will produce such an arrangement.

Since the gas is ideal the process quasi-static, $PV = nRT$ for each point in the process. Since the process is isothermal:

$$W = P\Delta V = nRT \ln \frac{V_f}{V_i}$$

If the gas expands isothermally, $V_f > V_i$ and the work done by the gas is positive.

Mechanisms of Heat Transfer

There are three mechanisms of heat transfer:

- Conduction
- Convection
- Radiation

Conduction occurs when heat is transferred directly via the exchange of kinetic energy between less energetic and more energetic atoms and molecules within a substance. Substances that readily allow large amplitude vibrations of their constituent atoms make good conductors of both heat and electricity. Conduction occurs only between materials in thermal contact at different temperatures or when there is a temperature gradient within a thermally conducting medium.

Conductive heat transfer rate for most materials is given by:

$$H = kA \frac{T_2 - T_1}{L}$$

where k is the thermal conductivity of the material, A is the surface area, L the thickness, and T_2 & T_1 the boundary temperatures.

When heat is conducted across several dissimilar materials the heat transfer rate is:

$$H = A \frac{(T_2 - T_1)}{\sum_i \frac{L_i}{k_i}}$$

A common method of computing heat transfer rates for building materials is:

$$H = A \frac{(T_2 - T_1)}{\sum_i R_i}$$

where L_i/k_i has been replaced with a so called R value.

Convection is a mechanism of heat transfer that involves movement of the heated substance. Convection is an important mechanism for heat transfer in fluids. Boiling water in a heated pot is an example of convection. Because most fluids are poor conductors, heat transfer in fluids would be very inefficient without convection.

Because convection involves the phenomena of mass transport, it is very difficult to quantify without the use of some fairly sophisticated differential equations.

Radiation is the transfer of heat via electromagnetic waves. All objects emit energy continuously in the form of e/m waves. At low temperatures this radiation is composed of longer wavelengths (infrared) that decrease as the temperature of the object is raised.

The rate at which an object emits radiant energy is proportional to the cube of its absolute temperature. This relationship is known as Stefan's law:

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

- P is the power radiated (watts)
- σ is a constant equal to $5.6696 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4$
- A is the surface area of the emissive object (m^2)
- e is a constant known as emissivity that varies between zero and unity depending upon the ability of the surface to absorb and emit radiant energy
- T is the temperature in K

- When an object is in equilibrium with its environment it radiates and absorbs energy at the same rate. Its temperature, therefore, remains constant. When an object is cooler than its environment it absorbs more energy than it radiates and gets hotter (and vice versa).

Examples of Heat Transfer

The Dewar flask

Designed to defeat all three mechanisms of heat transfer

Cumulus Clouds