Heat and Internal Energy

Common usage of the word **heat** is something more closely equivalent to what in physics is known as **internal energy**. Although these two types of energy are related through the **1st Law of Thermodynamics**, important differences require clarification.

Internal Energy U: Atomic or molecular kinetic and potential energies due to translations, rotations, vibrations, and interatomic or intermolecular interactions.

Heat Q: Energy transferred resulting from a **temperature gradient** between two thermodynamic systems or between a system and the surrounding environment.

If $+Q$ heat is transferred into a system, then the system internal energy U will increase. Similarly transferring $-Q$, i.e., removing heat will lower the internal energy. This is quantified in the 1st Law of Thermodynamics: $\Delta U = Q - W$

'Heating something up' usually references a ***resulting*** increase in temperature i.e., internal energy increase, and not the quantity of heat energy transferred. ***An effect versus a cause***.

Mechanical Equivalent of Heat

The connection between mechanical energy and heat energy was established by **J. Joule** in 1843. Examples of this connection are:

1) **An oscillating mass on a non-zero coefficient of friction surface can raise the temperature of the system above ambient temperature ➔ ΔT .**

2) **Heat engines such as the automobile use temperature differences to produce rotational mechanical energy $\tau\Delta\theta$ of crankshafts etc.**

Calorie and BTU

1) **1 Calorie = amount of heat that raises the temperature of 1 g of H_2O by 1 ^{0}C from 14.5 to 15.5 ^{0}C.**

2) **1 BTU = amount of heat that raises the temperature of 1 lb of H_2O by 1 ^{0}F from 63 to 64 ^{0}F.**

The conversion factor between heat energy and mechanical energy is called the **mechanical equivalent of heat: 1 Cal = 4.186 J**

Absorption of Heat and Specific Heat

How effectively the heat input into or the heat extracted from a substance produces a temperature change in that substance leads to the definition of **specific heat**:

$$Q = mC\Delta T$$

Q = Heat transferred to or from the substance

ΔT = Temperature change of mass m that results from that transfer

C = Specific Heat in units of J/kg ^{0}C.

If a substance has a **large** specific heat C , it will exhibit a **small** ΔT if compared to a material with low specific heat assuming Q was the same in both cases.

Aluminum foil vs. tin foil:

$C = 0.051 \,{}^{Cal}\!/_{g \cdot K}$ For Sn foil. $C = 0.215 \,{}^{Cal}\!/_{g \cdot K}$ For Al foil.

Calorimetry

Measuring heat energies (Calories) exchanged between systems of differing temperature.

An ideal **calorimeter** is a thermally isolated container that does not exchange any heat with the surrounding environment.

Starting with $T_{object} > T_{calorimeter}$, **Equilibrium** proceeds as:

$$Q_{Lost\_by\_object} = Q_{Gained\_by\_calorimeter}$$

$$M_{object} C_{object} \{T_{0,object} - T_f\} = M_{Cal} C_{Cal} \{T_f - T_{0,Cal}\}$$

Experimentally if the substance is an unknown, its specific heat is obtained:

$$C_{object} = M_{Cal} C_{Cal} \{T_f - T_{0,Cal}\} * \frac{1}{M_{object} \{T_{0,object} - T_f\}}$$

Phase Change

To change the matter **phase** (solid, liquid, gas) of a substance using heat, heat may be added or extracted from the substance at its melting or boiling point.

Example:

Adding heat increases system disorder or entropy. A transition from ice into water requires the addition of heat in an amount **proportional to the ice mass** sufficient to disorder the initially well-organized ice lattice.

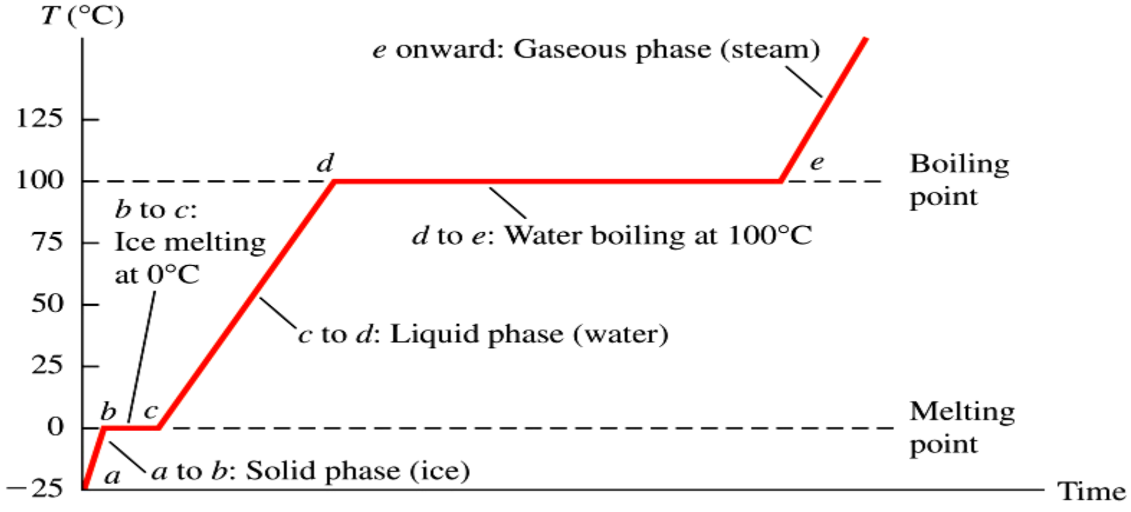
Latent Heat of Fusion

For **solid** ⇔ **liquid** transitions, the amount of heat added or extracted is:

$$Q = mL_F \qquad L_F \text{ is the } \textbf{latent heat of fusion in } {}^{J}\!/_{kg}$$

For **gas** ⇔ **liquid** transitions, the amount of heat added or extracted is:

$$Q = mL_V \qquad L_V \text{ is the } \textbf{latent heat of vaporization in } {}^{J}\!/_{kg}$$



If **both** a phase change and a substance temperature change are sought, then the total heat added or removed to effect the change will in general be:

$$Q = mL_F + mC\Delta T$$

Heat Transfer Mechanisms

1) **Heat may be transferred between two objects by placing the two objects in physical contact via the process of conduction.**

2) **Objects at finite temperatures radiate electromagnetic energy in a process called thermal radiation.**

3) **Convection requires relocation of a heated or cooled intermediate material to another physical location where the heat energy is then transferred.**

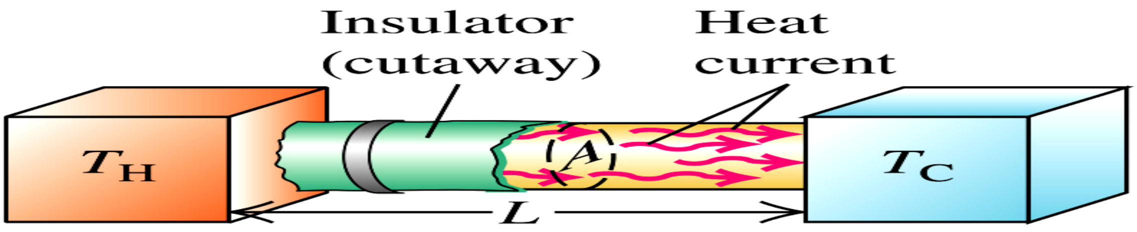
Conduction

For two objects in physical contact, molecular collisions between molecules in the higher temperature system with molecules of the lower temperature system leads directly to the transfer of energy between the two systems.

The rate at which any given substance transfers heat via conduction depends on the relative availability of molecules to transmit energy via the collision process.

Gases are relatively poor thermal conductors due to large intermolecular distances separating the gas molecules. Gases make good thermal insulators.

Metals have relatively large **thermal conductivities** resulting from the nearness of atoms in a metal and the large number of free electrons that may contribute to heat conduction.



The rate of conduction heat transfer is:

$$P = \frac{Q}{\Delta t} = \frac{kA\Delta T}{L}$$

$k =$ **Thermal Conductivity in** J/sm^0C

$\Delta T =$ **Temperature gradient** $T_H - T_L$

$A =$ **Cross-sectional area of material**

$L =$ **Material width or thickness over which the gradient is established.**

Thermal Resistance

$\frac{L}{k}$ **Is the " R " value. Higher R value ➔ better thermal insulator.**

For a multi-layered composite insulating wall, $P = \frac{Q}{\Delta t} = \frac{A \Delta T}{\sum_i \frac{L_i}{k_i}} = \frac{A \Delta T}{\sum_i R_i}$

Doubling wall thickness or insulating material thickness doubles the R-value and reduces the heat transfer rate by a factor of two.

Material	R-Value (ft^2 °F h/Btu)
Free stagnant air layer	**0.17**
Drywall (0.5 in. thick)	**0.45**
Wood shingles (lapped)	**0.87**
Flat glass (0.125 in thick)	**0.89**
Hardwood siding (1 in. thick)	**0.91**
Air space (3.5 in. thick)	**1.01**
Sheathing (0.5 in. thick)	**1.32**
Insulating glass (0.25 in space)	**1.54**
Concrete block (filled cores)	**1.93**
Brick (4 in. thick)	**4.00**
Fiberglass batting (6 in. thick)	**18.80**

Convection

From a quantitative aspect, convection is a bit more complex. Fluid flow effects and losses during the fluid relocation for example make analytic treatment difficult.

1) **At Home: furnaces and ovens heat air, and fans force the heated air throughout the home or oven volume. [Forced convection]**

2) **On Earth: ocean currents relocate massive volumes of heated / cooled ocean waters which either heat or cool the surroundings.**

3) **In Space: gases near the sun's core carry heat energy from the interior to the solar surface. Once at the surface, these gases radiate heat energy into space.**

Radiation

All objects radiate energy via thermal radiation at a rate according to **Stefan's Law**:

$$P = \frac{Q}{\Delta t} = \sigma A e T^4$$

$\sigma =$ **Stefan-Boltzmann Constant** $5.67x10^{-8}\, {}^{W}\!/_{m^2K^4}$

$T =$ **Temperature in Kelvin** $A =$ **Surface area of object.**

$e = emissivity$

$e = 1$ ➔ **Perfect radiator / absorber ➔ "Blackbody"**

$e = 0$ ➔ **No absorption / emission.**

Since an object can also absorb radiation from the environment while it radiates energy into the environment, the net heat transferred via radiation is:

$$P = \frac{Q}{\Delta t} = \sigma e A\{T_0^{\,4} - T^4\}$$

$\mathbf{T_0}$ is environment temperature and **T** is the object temperature. As thermal equilibrium with the environment is reached, the object absorbs and radiates heat at the same rate.