Energy, Work, and Heat

O **Energy** is the ability to do work or transfer heat.

O **Work** for a mechanical system is defined as force applied to an object times the distance through which it is moved:

$$w = fd$$

But force is mass times acceleration

$$f = ma$$

So,

$$w = mad$$

O **Work** is the transfer of energy from one body to another.

P Work is energy in the process of transfer.

P A unit of energy is that quantity transferred when a unit of work is done.

L In other words, energy and work have the same units.

Units of Work and Energy

$w = fd = mad$

Force Units

cgs system:

$$f = (\text{g})(\text{cm}/\text{sec}^2) = \text{g}\cdot\text{cm}/\text{sec}^2 = \text{g}\cdot\text{cm}\cdot\text{s}^{-2} \equiv \text{dyne}$$

SI units:

$$f = (\text{kg})(\text{m}/\text{sec}^2) = \text{kg}\cdot\text{m}\cdot\text{s}^{-2} \equiv \text{newton (N)}$$

Work and Energy Units

cgs system:

$$w = (\text{g}\cdot\text{cm}\cdot\text{s}^{-2})(\text{cm}) = \text{g}\cdot\text{cm}^2\cdot\text{s}^{-2} = (\text{dyne})(\text{cm}) \equiv \text{erg}$$

SI units:

$$w = (\text{kg}\cdot\text{m}\cdot\text{s}^{-2})(\text{m}) = \text{kg}\cdot\text{m}^2\cdot\text{s}^{-2} = (\text{newton})(\text{m}) \equiv \text{joule (J)}$$

Forms of Energy

- All forms of energy can be viewed in terms of equivalent amounts of **kinetic energy** (K) and **potential energy** (U).

- The total energy of a system may be defined as the sum of its kinetic and potential energies.

$$E_t = K + U$$

- Kinetic energy is energy of motion.

  - For a mechanical system

$$K = ½mv^2$$

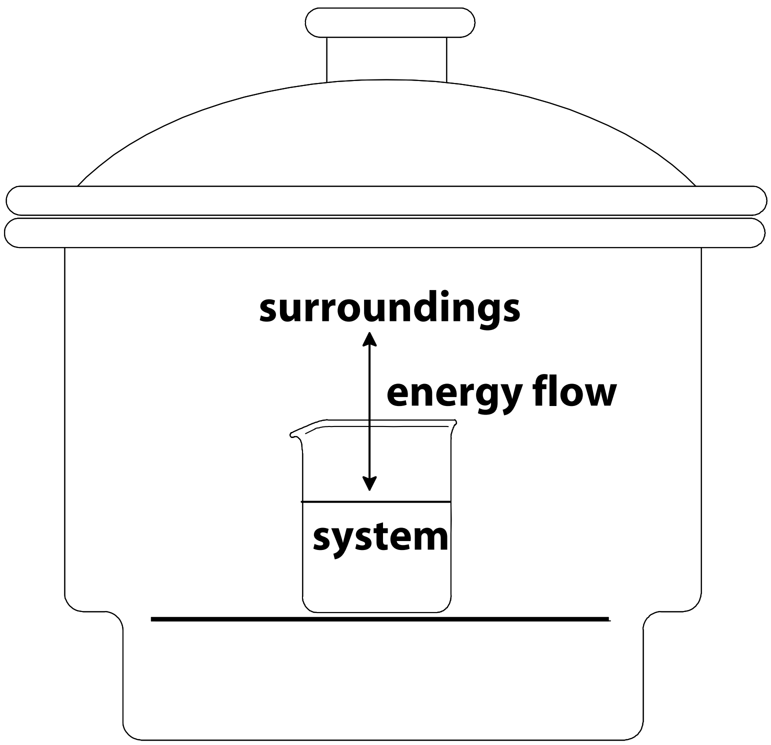
- Potential energy is energy of position or composition.

  - An object of mass m suspended a distance h above a surface has a potential energy due to position of

$$U_p = mgh$$

First Law of Thermodynamics

- Energy can be transferred from one object to another, and its forms can be interconverted, but energy can neither be created nor destroyed.



Transferring Heat

- We will be most interested in energy flow between system and surroundings that occurs through the transfer of heat (symbol q) during a chemical reaction or physical change.

- The system can give heat to the surroundings or receive heat from the surroundings.
  - The sense of the heat flow is indicated by the sign on q:
    $q < 0$ system gives heat to surroundings
    $q > 0$ surroundings give heat to system

Heat Flow Between Two Objects at Different Temperatures

U The sense of heat flow is always from the hotter object to the cooler object.

U Thermal equilibrium occurs when both are at the same temperature.

U The quantity of heat lost from the hotter object is equal to the heat gained by the cooler object.

Heat Transfer Between System and Surroundings

- As a chemical or physical change occurs, potential energy changes by gaining or losing heat, resulting in a temperature change.

- Heat transferred between the system and surroundings represents a **change in heat content of the system**, sometimes called the **heat of reaction**.

Endothermic vs. Exothermic

Endothermic

O If a system takes up heat from its surroundings, its heat content will be higher at the end of the process. Consequently, *q* will be positive

K Such a process is **endothermic**.

Examples of endothermic processes:

melting ice

$$\text{heat} + H_2O(s) \ 6 \ H_2O(l)$$

dissolving $NH_4NO_3(s)$

$$\text{heat} + NH_4NO_3(s) \ 6 \ NH_4^+(aq) + NO_3^-(aq)$$

Endothermic vs. Exothermic
Exothermic

O If the system gives off heat to its surroundings, the heat content at the end of the process will be less. Consequently, q will be negative.

K Such a process is **exothermic**.

Examples of exothermic processes:

freezing water

$$H_2O(l) \ 6 \ H_2O(s) + \text{heat}$$

dissolving $CaCl_2(s)$

$$CaCl_2(s) \ 6 \ Ca^{2+}(aq) + 2Cl^{-}(aq) + \text{heat}$$

Endothermic, Exothermic, and Direction

K If a process is endothermic in one direction it is exothermic in the opposite direction, and vice versa.

heat + $H_2O(s)$ 6 $H_2O(l)$ endothermic

$H_2O(l)$ 6 $H_2O(s)$ + heat exothermic

Units of Heat

- Formerly, heats of reaction were measured in units of calories, based on the following definition:

  One calorie is the heat needed to raise 1.000g of water from 14.50ºC to 15.50ºC.

- With the adoption of SI units, the calorie has been redefined in terms of the joule (J), which is the preferred unit today:

  $$1 \text{ calorie} = 4.184 \text{ J (exactly)}$$

- When large amounts of heat are transferred, the kilojoule (kJ) is used.

- The kilocalorie survives as the Calorie used in nutrition

  $$1 \text{ Calorie} = 1 \text{ kilocalorie} = 10^3 \text{ calories}$$