

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:

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

SI units:

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

Work and Energy Units

cgs system:

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

SI units:

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

Forms of Energy

- O All forms of energy can be viewed in terms of equivalent amounts of **kinetic energy** (K) and **potential energy** (U).
- O The total energy of a system may be defined as the sum of its kinetic and potential energies.

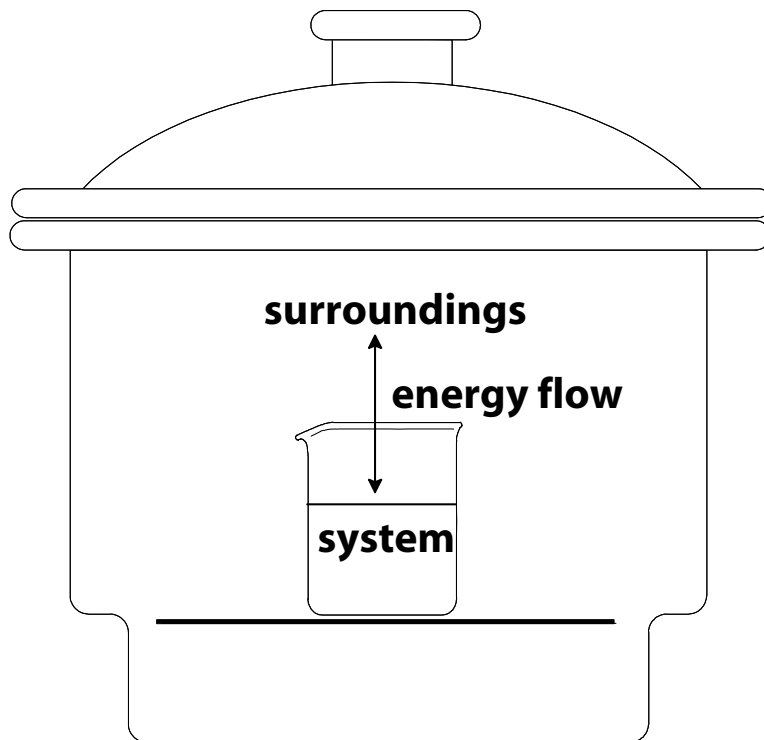
$$E_t = K + U$$

- O Kinetic energy is energy of motion.
 - P For a mechanical system
$$K = \frac{1}{2}mv^2$$
- O Potential energy is energy of position or composition.
 - P 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

- O 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.
- O The system can give heat to the surroundings or receive heat from the surroundings.
 - P 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

- O As a chemical or physical change occurs, potential energy changes by gaining or losing heat, resulting in a temperature change.
- O Heat transferred between the system and surroundings represents a **change in heat content of the system**, sometimes called the **heat of reaction**.

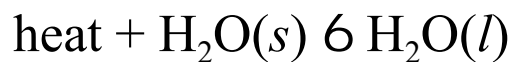
Endothermic vs. Exothermic

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



dissolving $\text{NH}_4\text{NO}_3(s)$



Endothermic vs. Exothermic

Exothermic

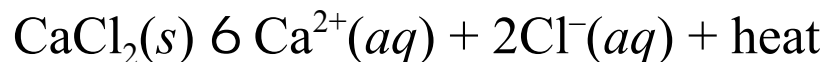
- O If the system gives off heat to its surroundings, the heat content at the end of the process will be less. Consequently, q will be negative.
- K Such a process is **exothermic**.

Examples of exothermic processes:

freezing water

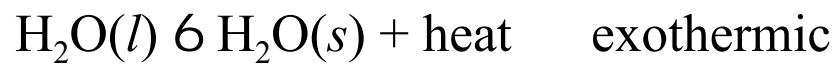


dissolving $\text{CaCl}_2(s)$



Endothermic, Exothermic, and Direction

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



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 14.50°C to 15.50°C.

- 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

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