THERMODYNAMICS

Courtesy of lab-initio.com

Definitions #1

Energy: The capacity to do work or produce heat

Potential Energy: Energy due to position or composition

Kinetic Energy: Energy due to the motion of the object

$$KE = \frac{1}{2}mv^2$$

Definitions #2

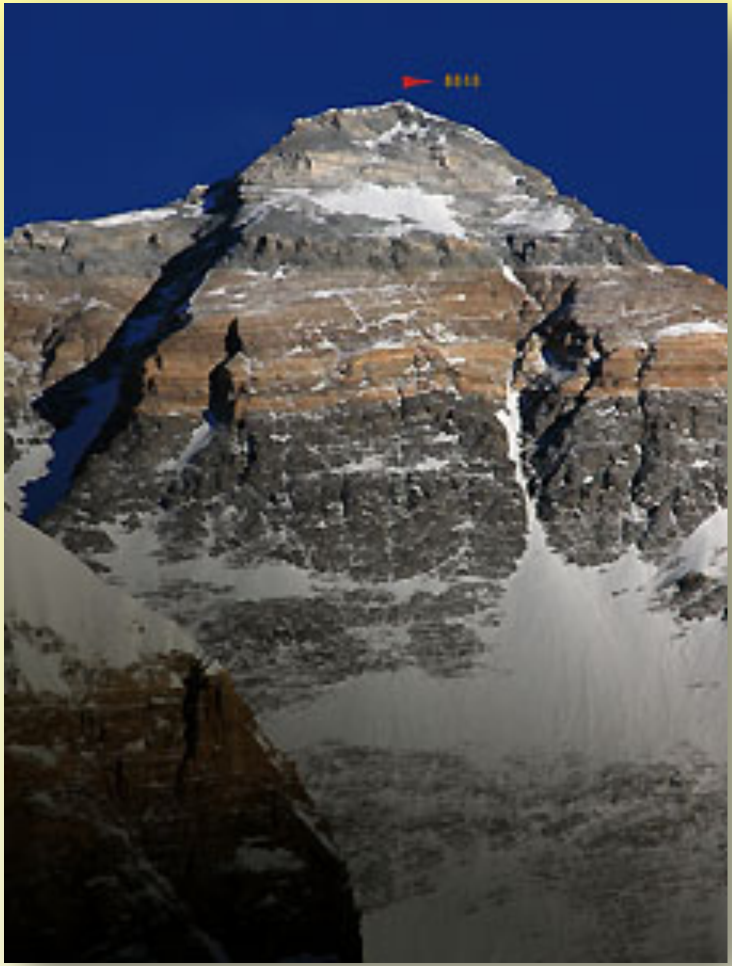
Law of Conservation of Energy: Energy can neither be created nor destroyed, but can be converted between forms

The First Law of Thermodynamics: The total energy content of the universe is constant

State Functions depend ONLY on the present state of the system

ENERGY IS A STATE FUNCTION

A person standing at the top of Mt. Everest has the same potential energy whether they got there by hiking up, or by falling down from a plane ☺



WORK IS NOT A STATE FUNCTION

WHY NOT???

$\Delta E = q + w$

ΔE = change in internal energy of a system

q = heat flowing into or out of the system

- $-q$ if energy is leaving to the surroundings
- $+q$ if energy is entering from the surroundings

w = work done by, or on, the system

- $-w$ if work is done by the system on the surroundings
- $+w$ if work is done on the system by the surroundings

Work, Pressure, and Volume

$$w = -P\Delta V$$

Expansion

$+\Delta V$ (increase)

↓

$-w$ results

↓

E_{system} decreases

↓

Work has been done by the system on the surroundings

Compression

$-\Delta V$ (decrease)

↓

$+w$ results

↓

E_{system} increases

↓

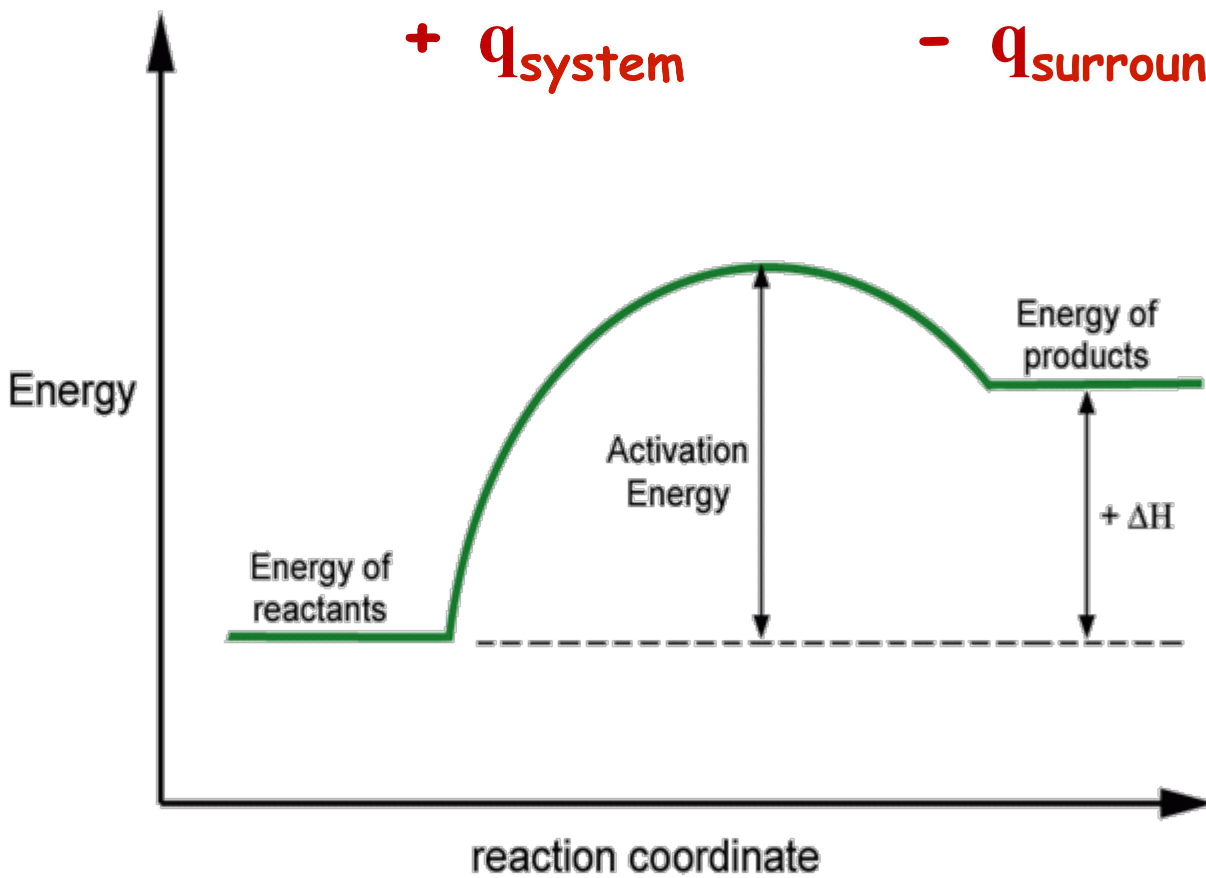
Work has been done on the system by the surroundings

Energy Change in Chemical Processes

Endothermic

Reactions in which energy flows *into* the system as the reaction proceeds.

$+\ q_{system}$ $-\ q_{surroundings}$

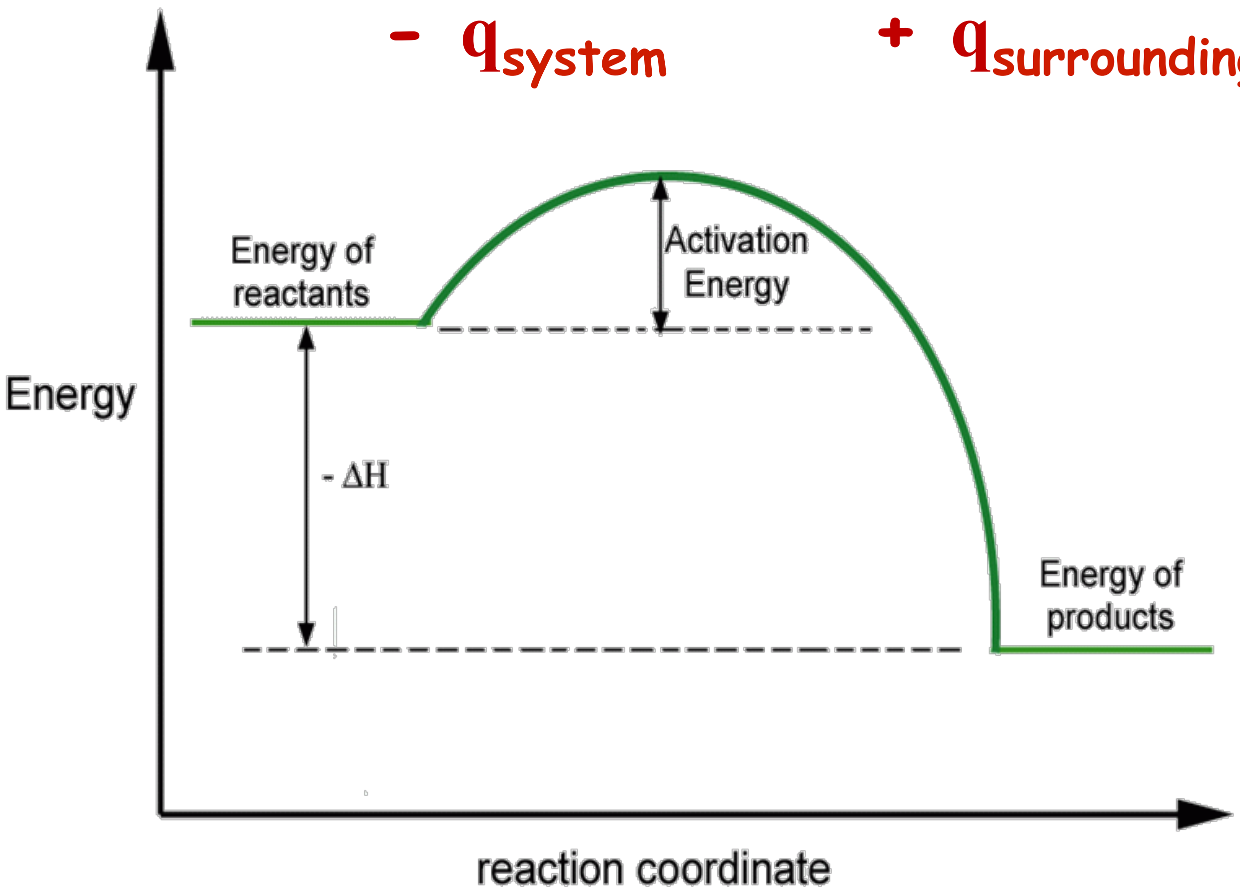


Energy Change in Chemical Processes

Exothermic:

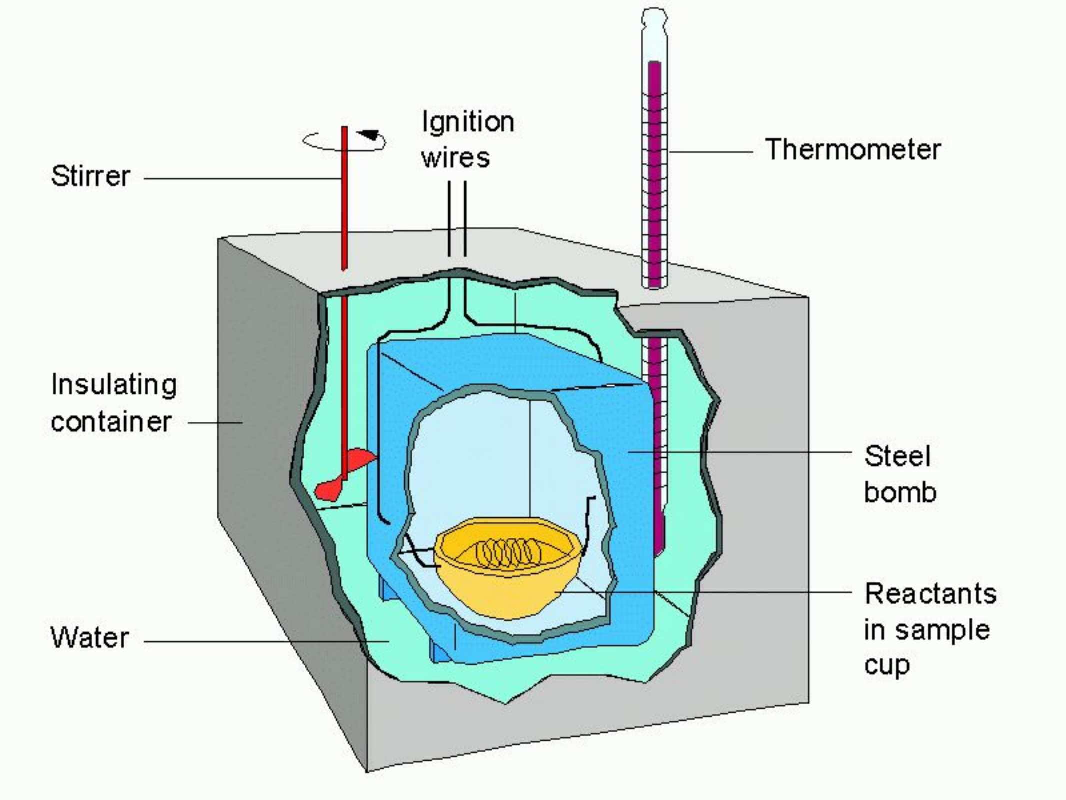
Reactions in which energy flows *out of* the system as the reaction proceeds.

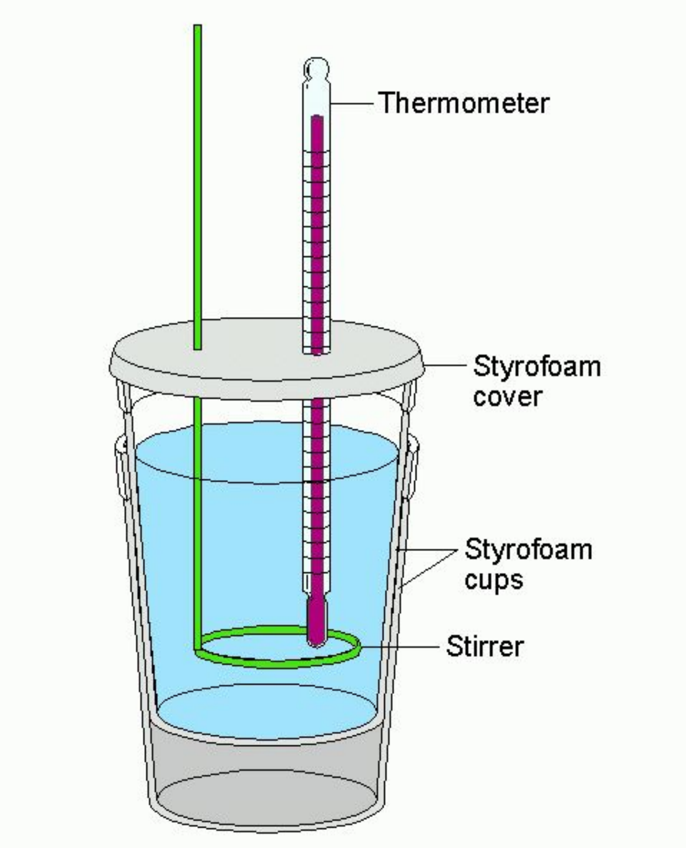
$- q_{system}$ $+ q_{surroundings}$



Calorimetry

The amount of heat absorbed or released during a physical or chemical change can be measured, usually by the change in temperature of a known quantity of water in a calorimeter.





Units for Measuring Heat

The **Joule** is the SI system unit for measuring heat:

$$1\,Joule = 1\,newton \cdot meter = \frac{1\,kg \cdot m^2}{s^2}$$

The **calorie** is the heat required to raise the temperature of 1 gram of water by 1 Celsius degree

$$1\,calorie = 4.18\,Joules$$

1 BTU is the heat required to raise the temperature of **1 pound of water by 1 °F**

Specific Heat

The amount of heat required to raise the temperature of one gram of substance by one degree Celsius.

Substance	Specific Heat (J/g·C)
Water (liquid)	4.18
Ethanol (liquid)	2.44
Water (solid)	2.06
Water (vapor)	1.87
Aluminum (solid)	0.897
Carbon (graphite,solid)	0.709
Iron (solid)	0.449
Copper (solid)	0.385
Mercury (liquid)	0.140
Lead (solid)	0.129
Gold (solid)	0.129

Calculations Involving Specific Heat

$$q = C \cdot m \cdot \Delta T$$

q = Heat lost or gained

C= Specific Heat Capacity

ΔT = Temperature change

Units

- **m - mass in kilograms**
- **ΔT - change the temperature in Kelvin**
- **C has units of J/kg K or kcal/kg K**
- **1 calorie = 4.184 Joules**

Entropy

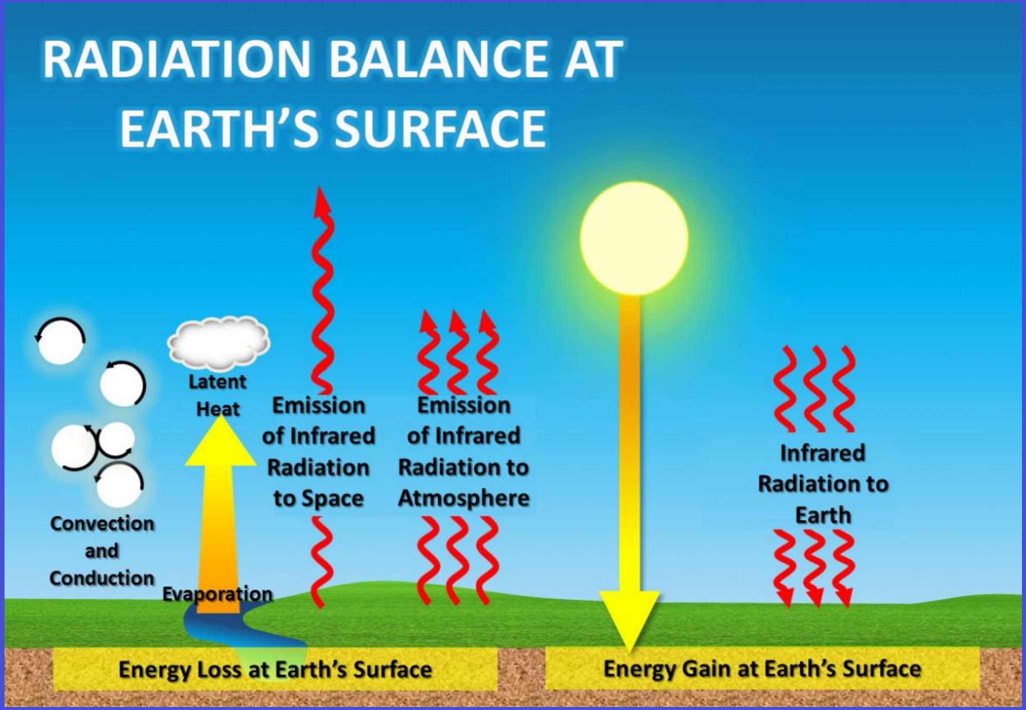
- Chaos, disorder
- When things increase in entopy they increase in disorder
- Temperature up = entropy up
- Temp down = entropy down

3 types of energy transfer

- Radiation
- Conduction
- Convection

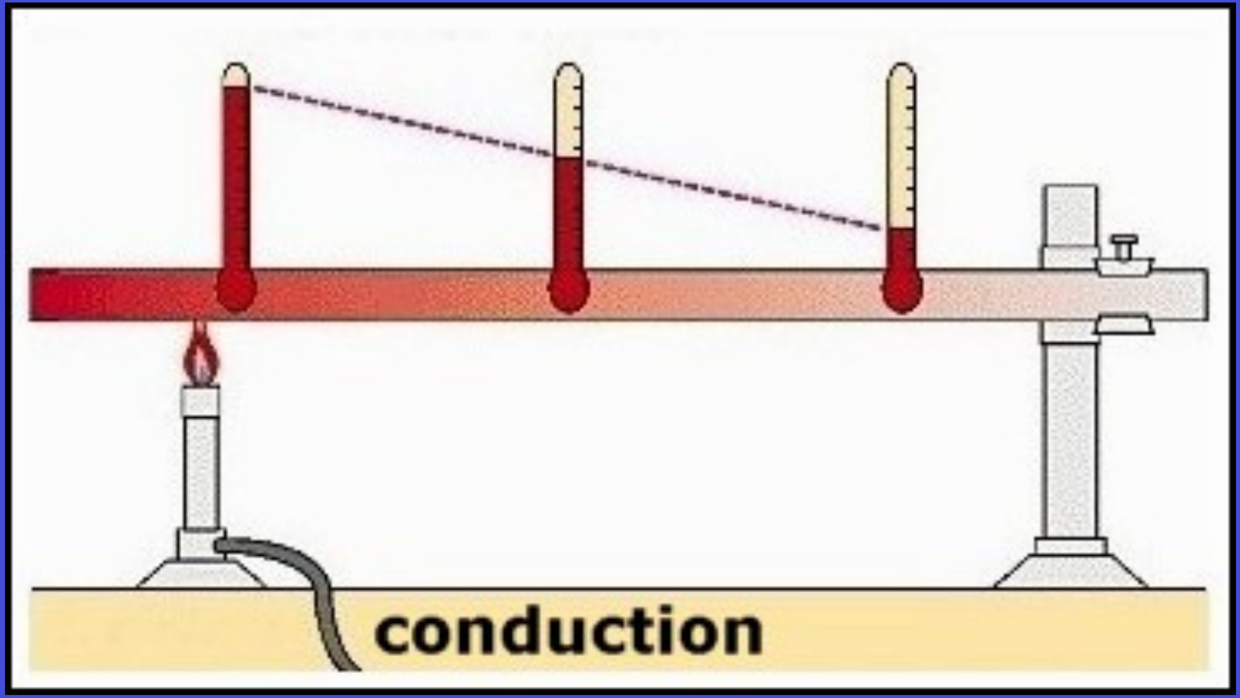
Radiation

- Comes from waves or electromagnetic sources
- Lights, IR heaters,



Conduction

- Transfer from touching objects.
- When you touch something hot and it feels hot.



Convection

- Energy transfer from density variations that cause currents.

This is why the upstairs of your often warmer. Circulation occurs in fluids that change density.