10/31/2018

Units of energy

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

m/s

Kg

$Kg\,m^2/s^2$

SI derived unit: 1 joule (J) = $1\ Kg\,m^2/s^2$

James Joule

-English scientist + brewer 1818 - 1889

$$\text{Power} = \frac{\text{Energy}}{\text{time}}$$

J

s

$\frac{J}{s}$ or W

└ watts (James Watt.)

light bulbs : 100 W = 100 J/s

1 J ≈ energy of a heartbeat

On av'g, in US, adult: 900 MJ/day.

900×10^6 J

9.00×10^8 J

The first law of thermodynamics

- total E of universe = constant

E_{UNIV} = const, $\Delta E_{UNIV} = 0$

└ change.

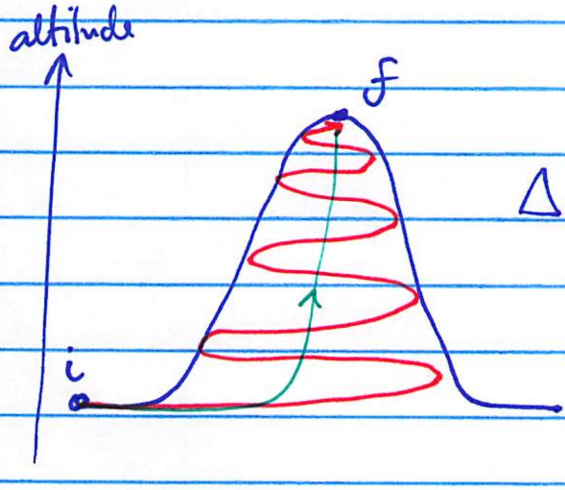
Internal energy, E = sum of all KE + PE of its particles

- it is a STATE FUNCTION
  - only depend on current state, not history

$\Delta E = E_{final} - E_{initial}$ = constant for any process/path

Analogy:

altitude
"
state fn



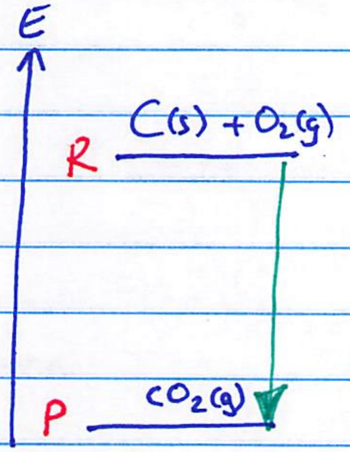
Δaltitude = same for red/green path!

For a chemical system:

initial: Reactants
final: Products

$$R \longrightarrow P$$
$$i \qquad f$$

ex: $C(s) + O_2(g) \longrightarrow CO_2(g)$

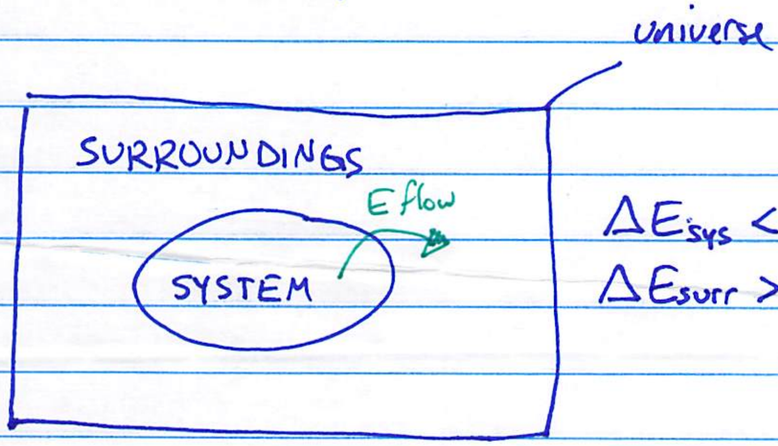


$\Delta E = E_f - E_i$

$\Delta E < 0$ (-ve)

Q: where does E go?

A:

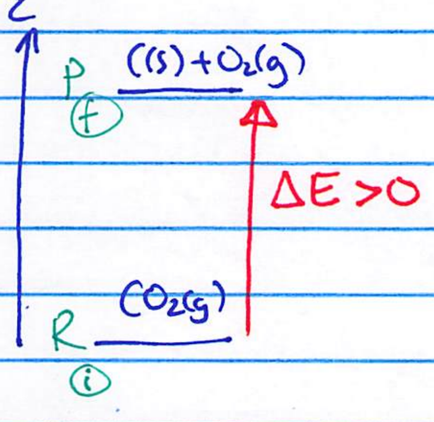


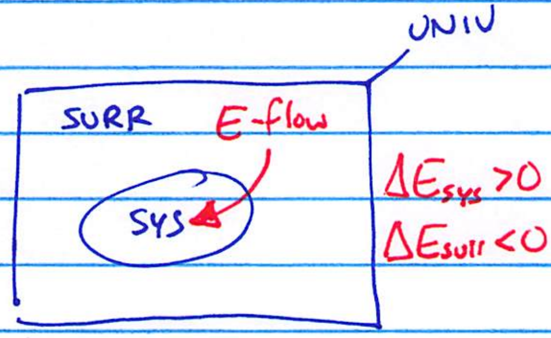
$\Delta E_{sys} < 0$

$\Delta E_{surr} > 0$

let's look @ reverse rxn:

$$CO_2(g) \longrightarrow C(s) + O_2(g)$$



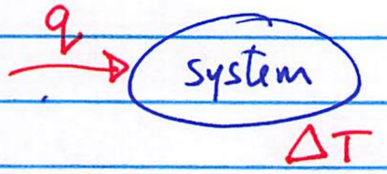


Read: pages 256 – 258 : heat + work

Although E can be transferred as work (force × distance)
we'll focus on heat in this course.
 └ q

Quantifying heat

(q) heat: flow of E due to T-difference



$q \propto \Delta T$

$q = C \times \Delta T$

 └ heat capacity

$q = C \times \Delta T \Rightarrow C = \dfrac{q}{\Delta T}$

1 mL H_2O

Q: Which one has the largest C?

takes much less q to change by 1°C

takes way more q to change by 1°C

J

$C = q/\Delta T$

J/°C

°C or K

small C

large C

Problem: C is extensive

Much better: use specific heat capacity: C_s

$$\underbrace{\frac{J}{g \cdot °C}}_{} \rightarrow C_s = \frac{C}{m}$$

J/°C

extensive

extensive

intensive

g

C_s = heat capacity per gram.

$q = C \times \Delta T$

$\cancel{q = m C_s}$ $\boxed{q = m \times C_s \times \Delta T}$