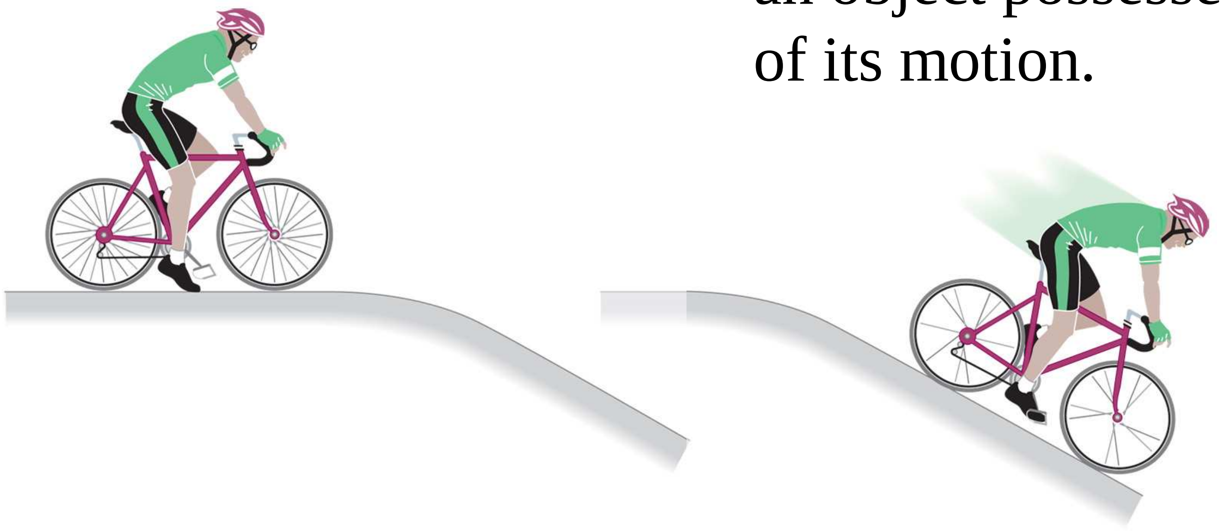
Chapter 5: Energy, Work, and Heat

Energy = The ability to do work or transfer heat.
Work = Energy used to cause an object that has mass to move.
Heat = Energy used to cause the temperature of an object to rise.

Potential Energy = Energy an object possesses by virtue of its position or chemical composition.

Kinetic Energy = Energy an object possesses by virtue of its motion.



Units of Energy

The SI unit of energy is the joule (J).

An older, non-SI unit is still in widespread use: The calorie (cal).

$$1 \text{ cal} = 4.184 \text{ J}$$

A ball weighing 0.5 kg is thrown at 40 m/s, what is its kinetic energy?

$$KE = \tfrac{1}{2}\, m\, v^2$$

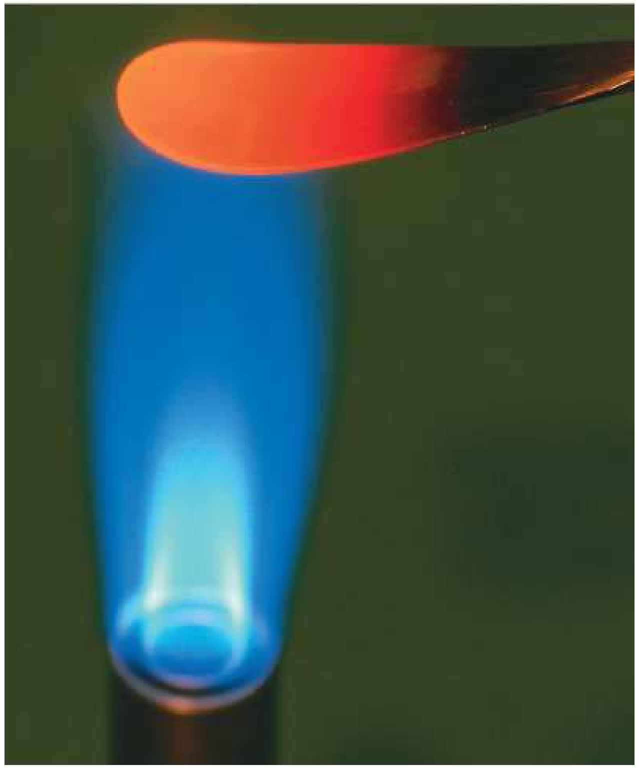
$$KE = \tfrac{1}{2}\,(0.5 \text{ kg})\,(40 \text{ m/s})^2$$
$$KE = 40 \text{ J}$$

Work and Heat

Work = Energy used to move an object over some distance.

Work = Force × distance

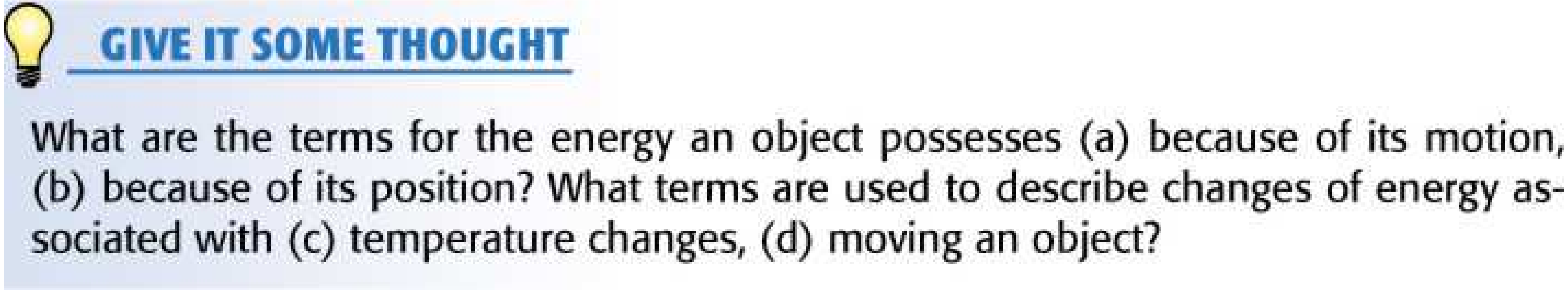
Work = w



Energy can also be transferred as heat.

Heat flows from warmer objects to cooler objects.

Heat = q



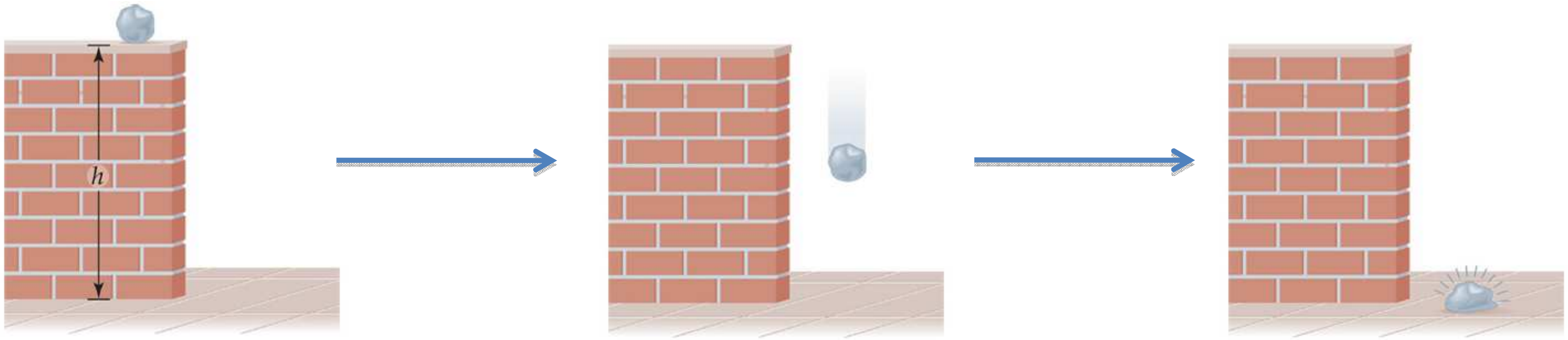
GIVE IT SOME THOUGHT

What are the terms for the energy an object possesses (a) because of its motion, (b) because of its position? What terms are used to describe changes of energy associated with (c) temperature changes, (d) moving an object?

A. heat
B. work
C. kinetic energy
D. potential energy

First Law of Thermodynamics

1st Law : Energy is neither created nor destroyed.



PE > 0
KE = 0

PE > 0
KE > 0

PE = 0
KE = 0

w > 0
q > 0

System & Surroundings

The system includes the molecules we want to study (here, the hydrogen and oxygen molecules).

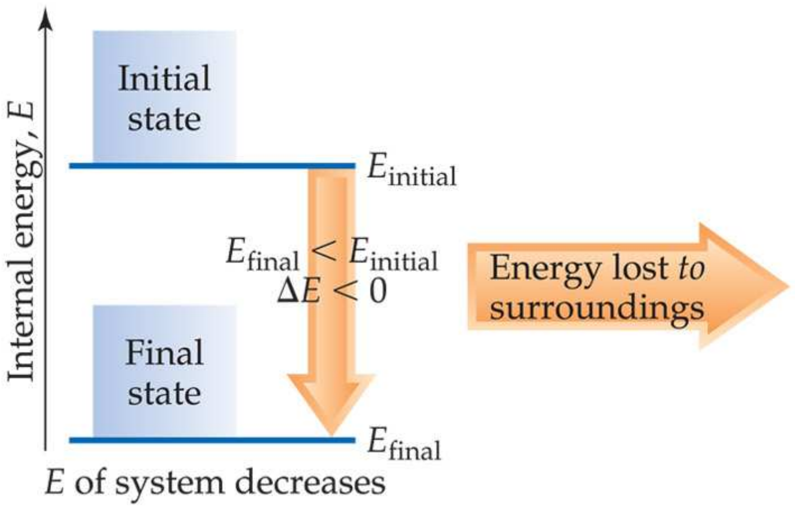
The surroundings are everything else (here, the cylinder and piston).

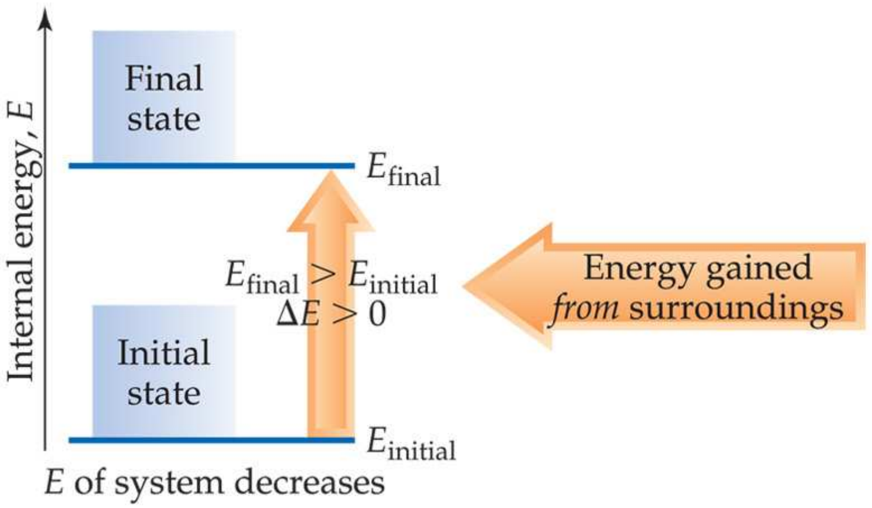
Internal Energy, E

The internal energy of a system is the sum of all kinetic and potential energies of all components of the system; we call it *E*.

By definition, the change in internal energy, ΔE, is the final energy of the system minus the initial energy of the system:

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

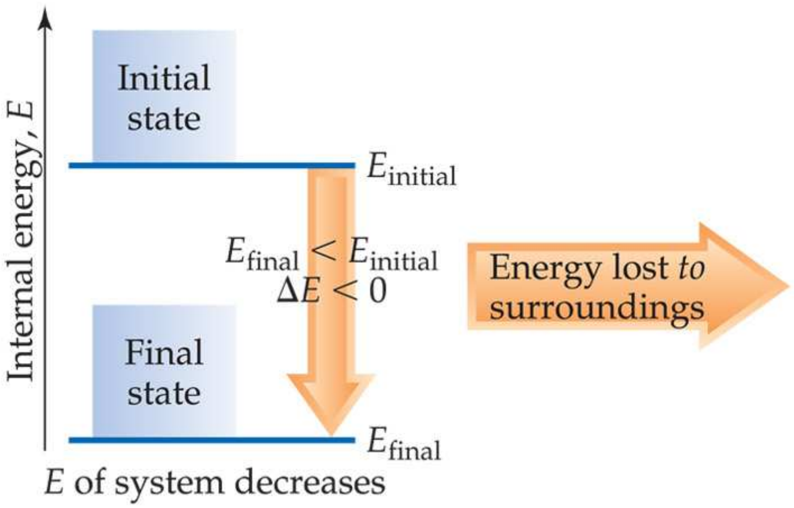


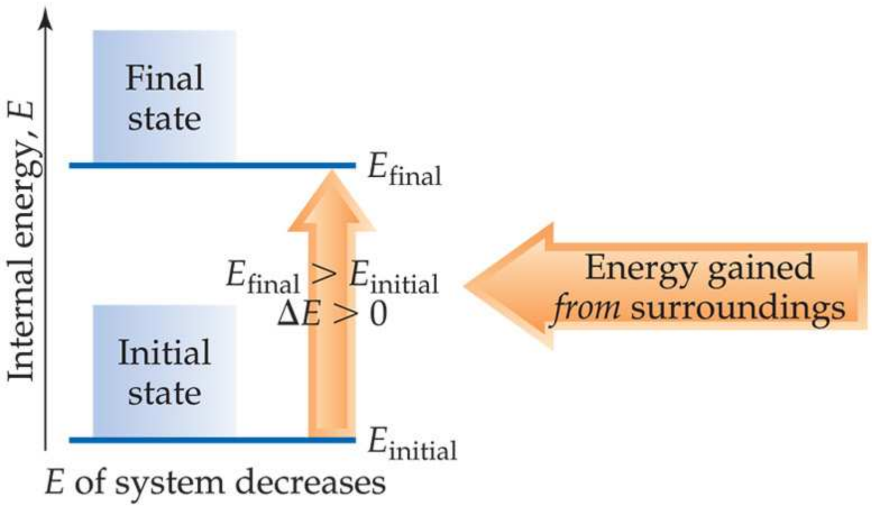


First Law of Thermodynamics

1^{st} Law : Energy is neither created nor destroyed.

The total energy of the universe is a constant; if the system loses energy, it must be gained by the surroundings, and vice versa.

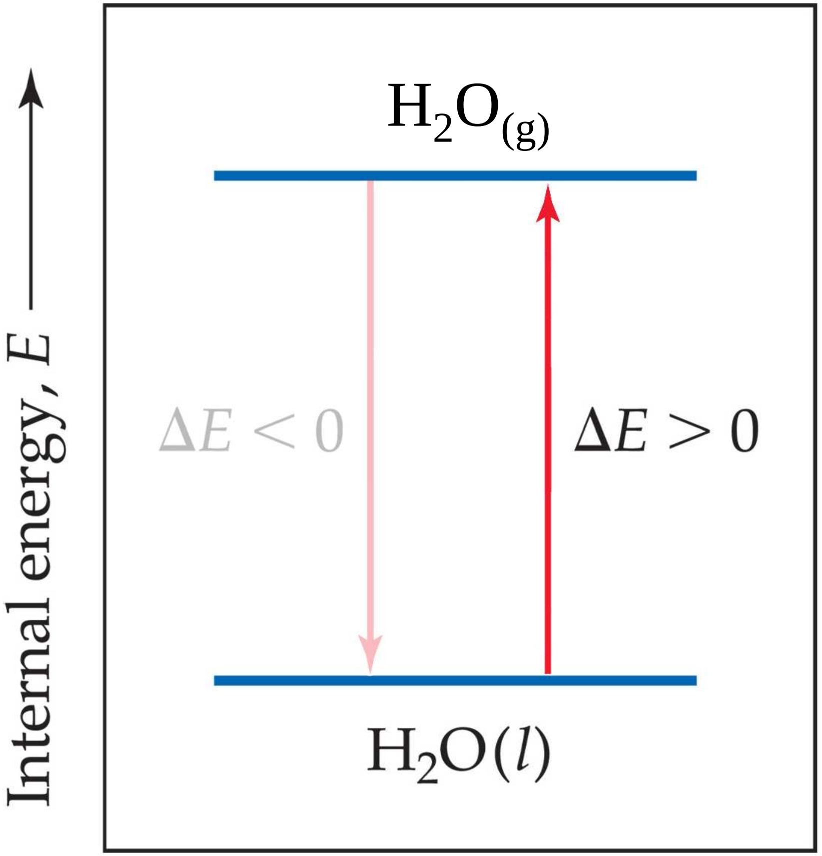




If a piece of metal at 85°C is place under water at 25°C, the final temperature of the water and metal is 30°C. Which of the following is true?

A. Heat lost by the metal > heat gained by water
B. Heat gained by water > heat lost by the metal
C. Heat lost by metal > heat lost by the water
D. Heat lost by the metal = heat gained by water

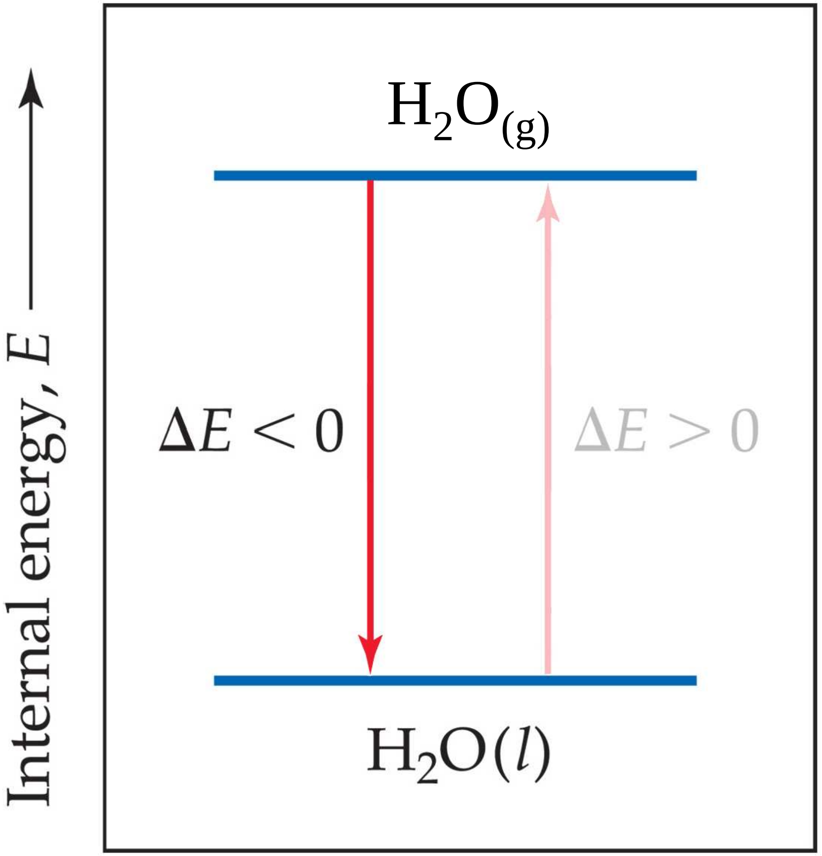
Change in Internal Energy



If $\Delta E > 0$, $E_{\text{final}} > E_{\text{initial}}$

Therefore, the system *absorbed* energy from the surroundings.

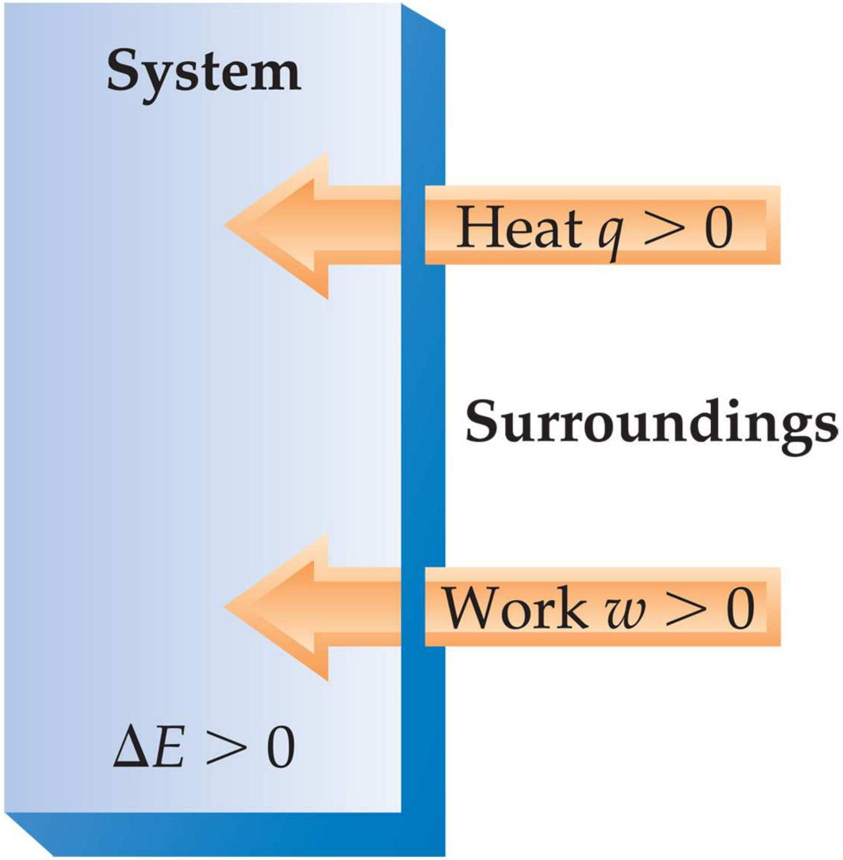
Change in Internal Energy



If $\Delta E < 0$, $E_{\text{final}} < E_{\text{initial}}$

Therefore, the system *released* energy to the surroundings.

Change in Internal Energy



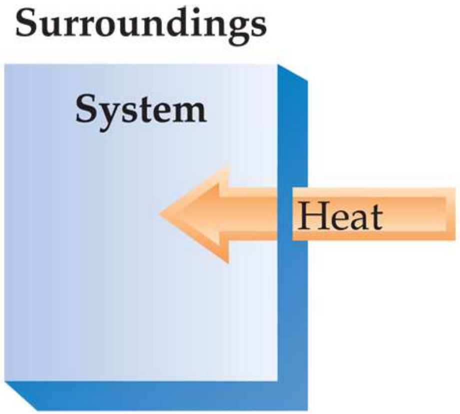
When energy is exchanged between the system and the surroundings, it is exchanged as either heat (*q*) or work (*w*).

$$\Delta E = q + w$$

For q	+ means system *gains* heat	− means system *loses* heat
For w	+ means work done *on* system	− means work done *by* system
For ΔE	+ means *net gain* of energy by system	− means *net loss* of energy by system

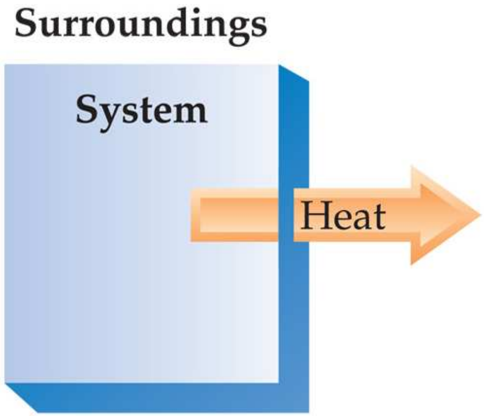
Endothermic & Exothermic

When heat is absorbed by the system from the surroundings, the process is endothermic.



$q > 0$
endothermic

When heat is released by the system to the surroundings, the process is exothermic.



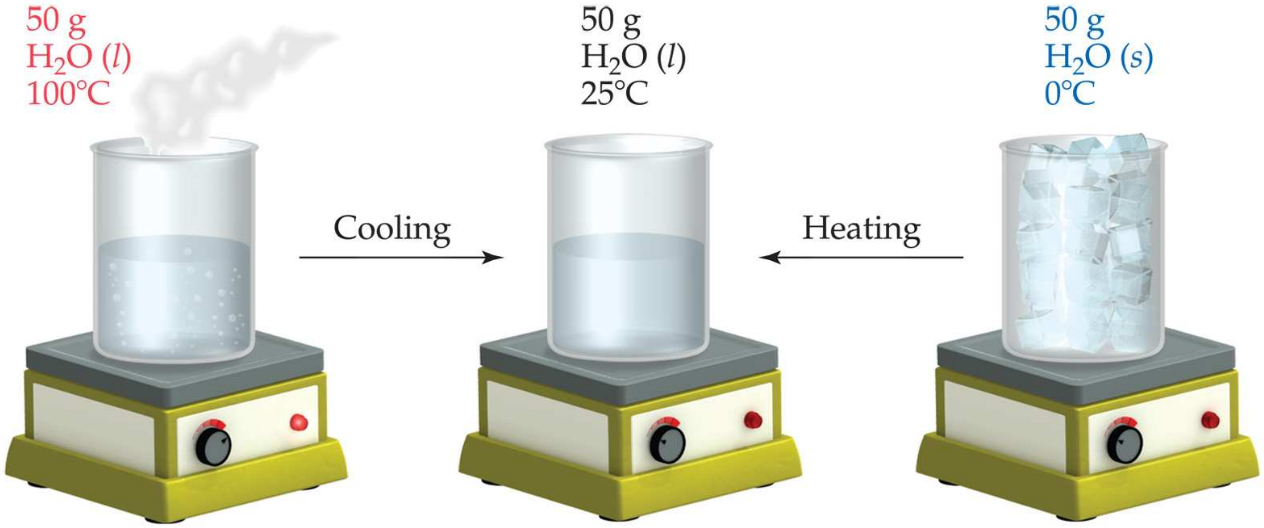
$q < 0$
exothermic

State Function

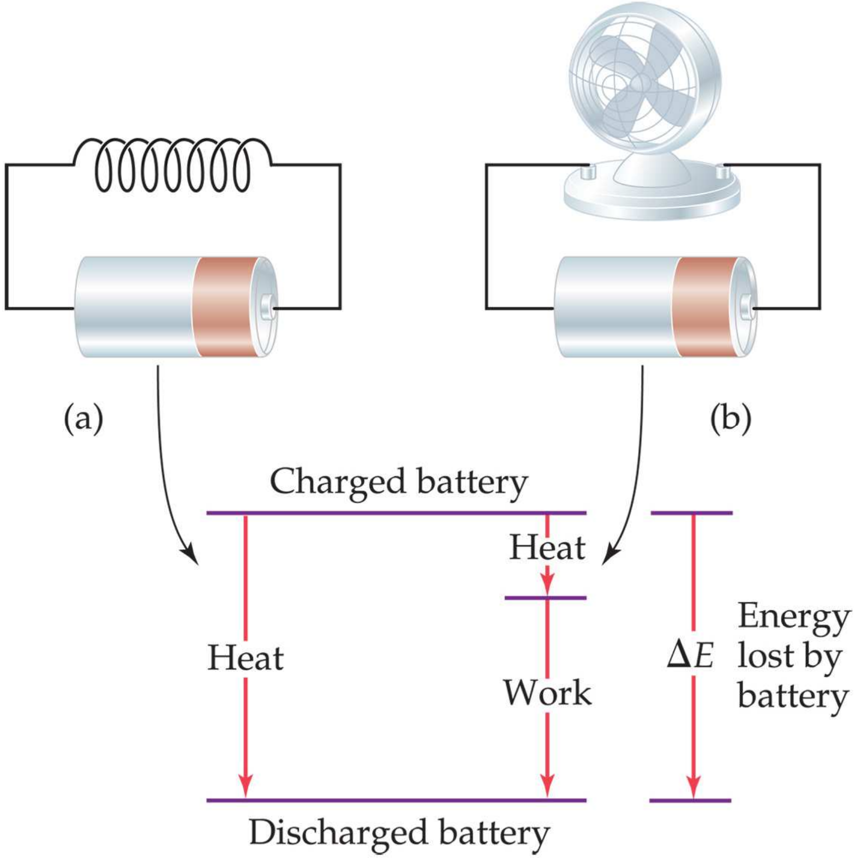
Internal Energy (E) depends only on the present state of the system, not on the path by which the system arrived at that state.

Internal energy (E) is a state function.

And so, ΔE depends only on E_{initial} and E_{final}.



Work and heat are not state functions



q and *w* are **not** state functions. Whether the battery is shorted out or is discharged by running the fan, its ΔE is the same.

But *q* and *w* are different in the two cases.

GIVE IT SOME THOUGHT

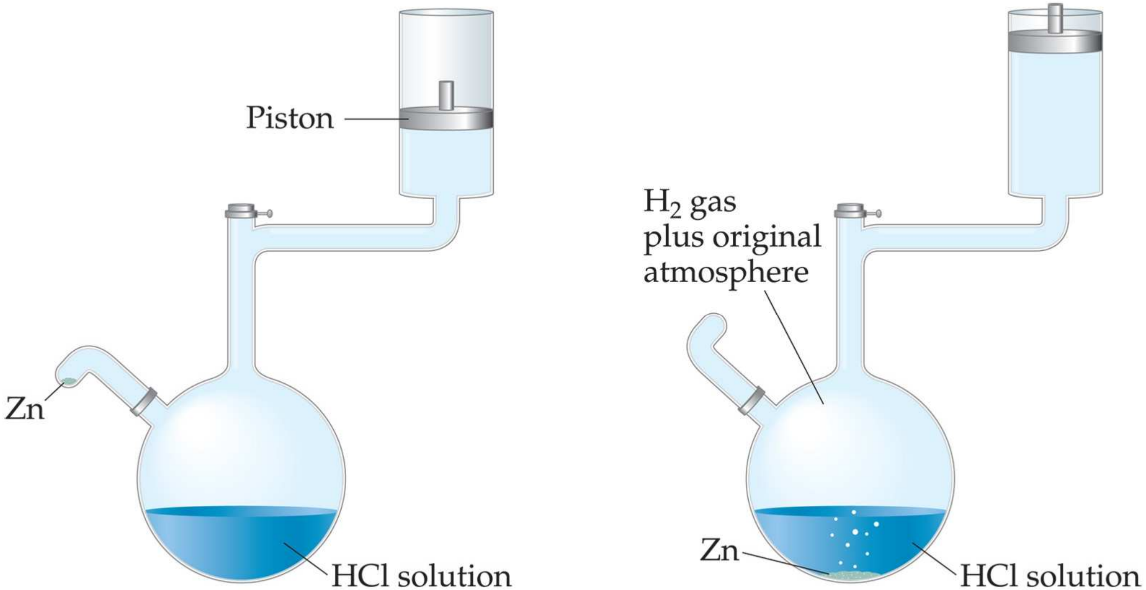
In what ways is the balance in your checkbook a state function?

A. It is determined by many factors.
B. It is a numerical value.
C. It is calculated from multiple transactions.
D. It doesn’t depend on how money got into your account or was taken out; it only depends on the net total of all transactions.

P-V Work

We can measure the work done by the gas if the reaction is done in a vessel that has been fitted with a piston and the pressure is constant.

$$w = -P\Delta V$$



Example: The piston is raised against atmospheric pressure (10^5 N/m^2) and the volume change is 1 L (10^{-3} m^3)

w = -PΔV

w = - (10^5 N/m^2)* (10^{-3} m$^{3)}$

w = -100 J

Enthalpy

If a process takes place at constant pressure and the only work done is this P-V work, then heat flow during the process is a measure the change in enthalpy (ΔH) of the system.

Enthalpy is the internal energy plus the product of pressure and volume:

When the system changes at constant pressure, the change in enthalpy, ΔH, is

$$\Delta H = \Delta(E + PV)$$

This can be written

$$\Delta H = \Delta E + P\Delta V$$

Enthalpy

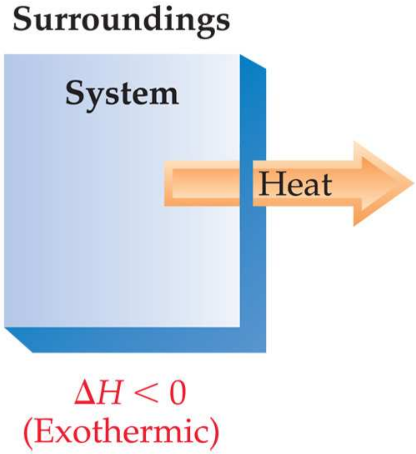
Since $\Delta E = q + w$ and $w = -P\Delta V$, we can substitute these into the enthalpy expression:

$$\Delta H = \Delta E + P\Delta V$$

$$\Delta H = (q+w) - w$$

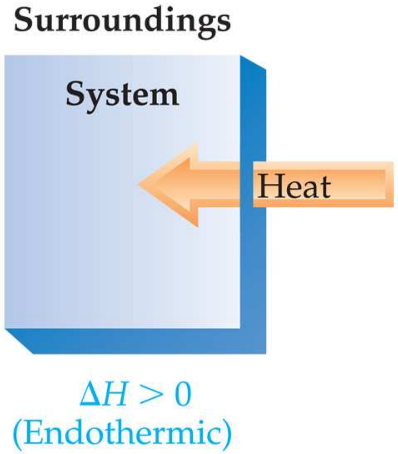
$$\Delta H = q$$

So, at constant pressure the change in enthalpy *is* the heat gained or lost.



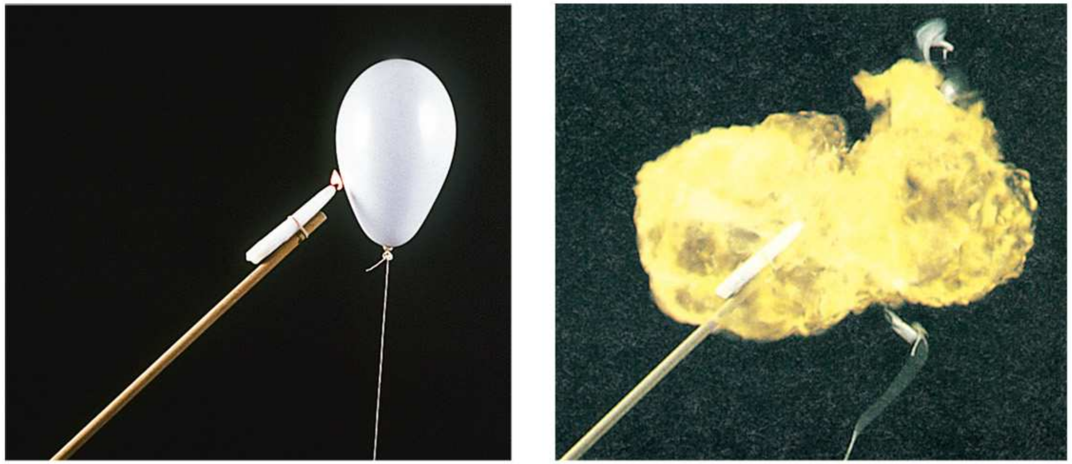
A process is endothermic when ΔH is positive.

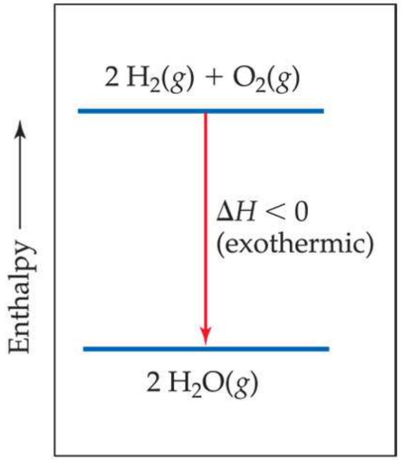
A process is exothermic when ΔH is negative.



Enthalpies of Reaction

This quantity, ΔH, is called the enthalpy of reaction, or the heat of reaction.

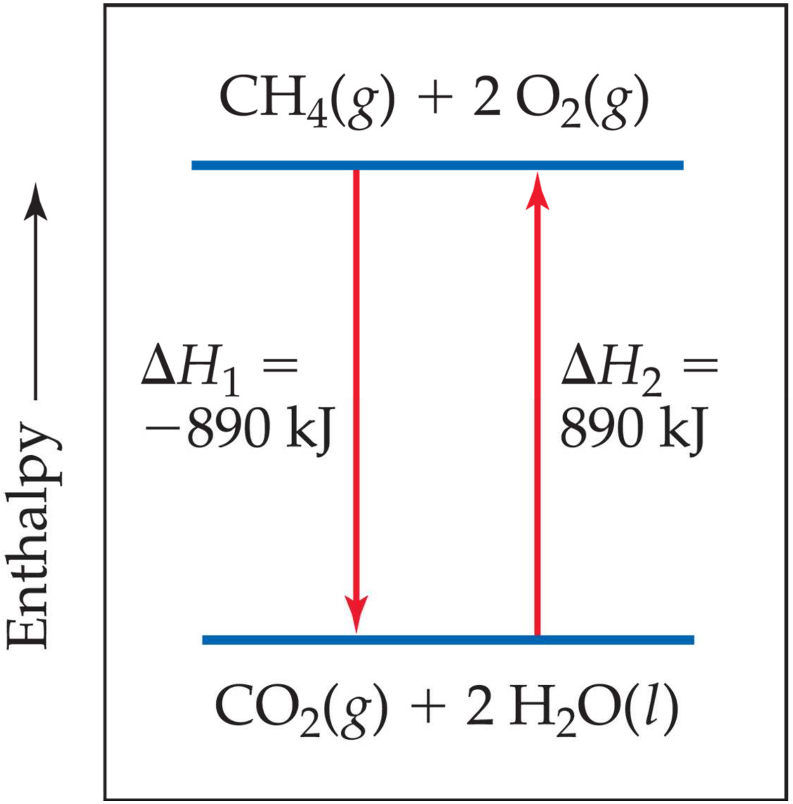




Enthalpies of Reaction

The *change* in enthalpy, ΔH, is the enthalpy of the products minus the enthalpy of the reactants:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

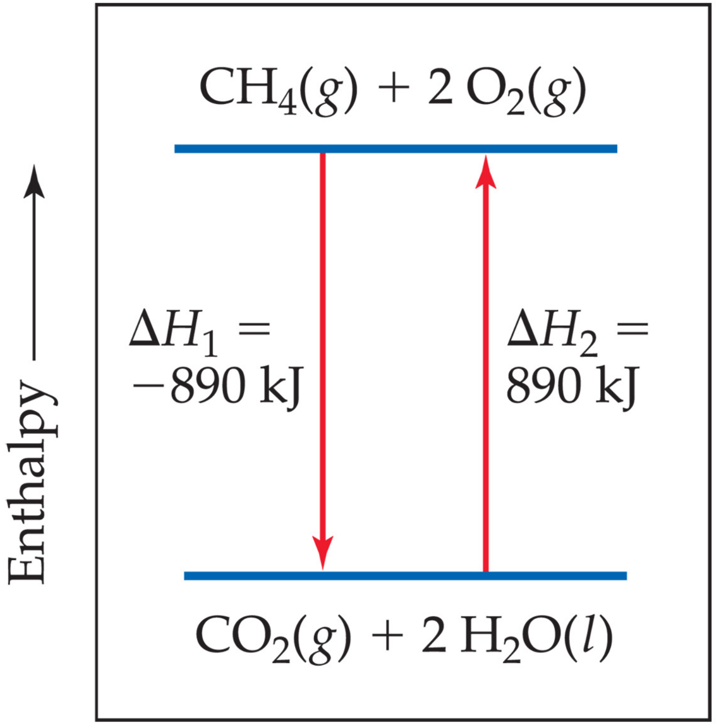


Facts About Enthalpy

- Enthalpy is an extensive property.
- ΔH for a reaction in the forward direction is equal in size, but opposite in sign, to ΔH for the reverse reaction.
- ΔH for a reaction depends on the state of the products and the state of the reactants.

Calorimetry

we measure ΔH through calorimetry, the measurement of heat flow.



Specific Heat and Heat Capacity

The amount of energy required to raise the temperature of a substance by 1 K (1°C) is its heat capacity.

We define specific heat capacity (or simply specific heat) as the amount of energy required to raise the temperature of 1 g of a substance by 1 K.

$$\text{Specific heat} = \frac{\text{heat transferred}}{\text{mass} \times \text{temperature change}}$$

$$s = \frac{q}{m \times \Delta T}$$

Which metal will undergo the greatest temperature change if an equal amount of heat is added to each?

A. Fe, $s = 0.45$ J/g K
B. Al, $s = 0.90$ J/g K
C. Cu, $s = 0.38$ J/g K
D. Pb, $s = 0.13$ J/g K

Hess’s Law

Hess’s law : If a reaction is carried out in a series of steps, Δ*H* for the overall reaction will be equal to the sum of the enthalpy changes for the individual steps.

Δ*H* is well known for many reactions, and the values are tabulated. Using these values of ΔH, ΔH for any can be calculated.

$$\boxed{C_{(s)} + \tfrac{1}{2}\, O_{2(g)} \rightarrow CO_{(g)} \quad \Delta H = }\,?$$

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} \qquad \Delta H = -393.5\ \text{kJ}$$

$$+ \qquad CO_{2(g)} \rightarrow CO_{(g)} + \tfrac{1}{2}\, O_{2(g)} \qquad \Delta H = 283.0\ \text{kJ}$$

$$C_{(s)} + O_{2(g)} + CO_{2(g)} \rightarrow CO_{(g)} + \tfrac{1}{2}\, O_{2(g)} + CO_{2(g\ )} \quad DH = -110.5\ \text{kJ}$$

Rules Regarding Hess's Law

1) When a chemical equation is multiplied by a constant, the ΔH is also multiplied.

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H = -393.5$ kJ

$2C_{(s)} + 2O_{2(g)} \rightarrow 2CO_{2(g)}$ $\Delta H = -787$ kJ

2) When a chemical equation is reversed, the sign of ΔH is changed.

$CO_{2(g)} \rightarrow CO_{(g)} + \frac{1}{2} O_{2(g)}$ $\Delta H = 283.0$ kJ

$CO_{(g)} + \frac{1}{2} O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H = -283.0$ kJ

3) Enthalpy depends on the state of the reactants and products

$H_2O_{(l)} \rightarrow H_2O_{(g)}$ $\Delta H = 88$ kJ

Calculate ΔH for $2\ C_{(s)} + O_{2(g)} + 4H_{2(g)} \rightarrow 2\ CH_3OH_{(g)}$

$2\ C_{(s)} + O_{2(g)} \rightarrow 2\ CO_{(g)}$ $\Delta H = –221$ kJ

$CO_{(g)} + 2\ H_{2(g)} \rightarrow CH_3OH_{(g)}$ $\Delta H = –91$ kJ

A. –120 kJ

B. +120 kJ

C –312 kJ

D –403 kJ

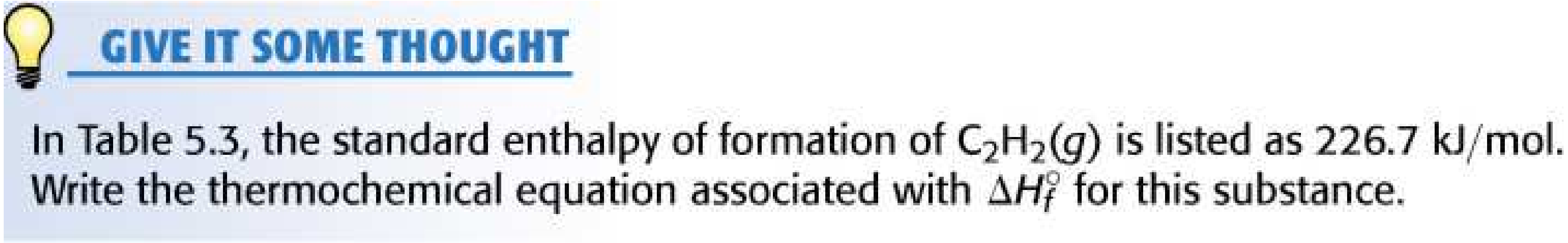
Enthalpy of Formation

Change in enthalpy when one mole of a compound is formed from its elements (in their standard state), ΔH_f

Standard state is atmospheric pressure and 25 C.

$$H_{2(g)} + \tfrac{1}{2}\, O_{2(g)} \rightarrow H_2O_{(l)} \quad \Delta H_f = -285.6 \text{ kJ/mol}$$

ΔH_f are tabulated and can be combined to calculate ΔH of any reaction



GIVE IT SOME THOUGHT

In Table 5.3, the standard enthalpy of formation of $C_2H_2(g)$ is listed as 226.7 kJ/mol. Write the thermochemical equation associated with ΔH_f° for this substance.

A. $C_{(s)} + H_{(g)} \rightarrow \frac{1}{2}\, C_2H_{2(g)}$

B. $C_{2(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$
C. $2C_{(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$

D. $2C_{(s)} + 2H_{(g)} \rightarrow C_2H_{2(g)}$

Given the following:

$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g) \quad \Delta H = –113\ kJ$

and ΔH_f for $NO_2(g)$ is 33.8 kJ,

what is ΔH_f for $NO(g)$?

- **A. –45.1 kJ**
- **B. +45.1 kJ**
- **C. +90.3 kJ**
- **D. –79.2 kJ**