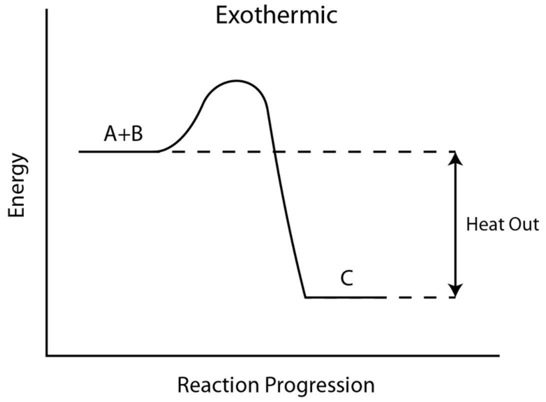
6.1 - Introduction to Thermochemistry



AP Big Idea #5:

The laws of thermodynamics describe the essential role of energy, and explain and predict the direction of changes in matter.

Energy

Energy is the ability to do work or transfer heat.

Energy used to cause an object that has mass to move is called work.

Energy used to cause the temperature of an object to rise is called heat.

Law of conservation of energy: Energy can be converted from one form to another but can be neither created nor destroyed.

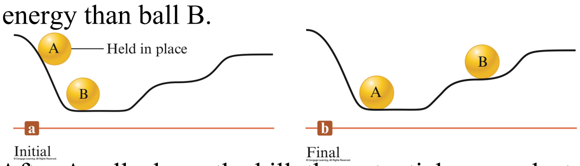
First Law of Thermodynamics: The total energy content of the universe is constant.

Potential vs. Kinetic Energy

Potential energy – energy due to position or composition.

Kinetic energy – energy due to motion of the object; depends on the mass of the object and its velocity.

In the initial position, ball A has a higher potential energy than ball B.



After A rolls down the hill, the potential energy lost by A has been converted to random motions of the components of the hill (frictional heating) and to the increase in the potential energy of B.

Kinetic Energy

Kinetic energy is energy an object possesses by virtue of its motion.

$KE = \frac{1}{2}mv^2$ AP Equation	m = mass (kg) v = velocity (m/s)

Units of Energy

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

$$1 J = 1\frac{kg \bullet m^2}{s^2}$$

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

1 cal = 4.184 J

Heat (Symbol = q)

Energy can also be transferred as heat.

Heat flows from warmer objects to cooler objects.

Heat is not a *thing*.

Heat is a way energy is transferred.

Energy

Heat involves the transfer of energy between two objects due to a temperature difference.

Work – force acting over a distance.

State Function: Property that does not depend in any way on the system's past or future (only depends on *present* state). Volume, pressure, number of particles, temperature, internal energy are state functions.

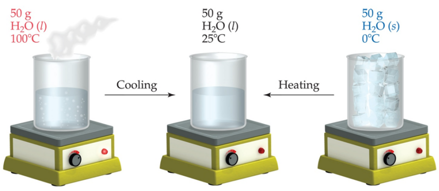
Work and heat are not state functions, because they are not intrinsic properties of a system. Instead, they are forms of energy in transit, and are only defined when there is a change in the system.

State Functions

Usually we have no way of knowing the internal energy of a system; finding that value is simply too complex a problem.

However, we do know that the internal energy of a system is independent of the path by which the system achieved that state.

In the system below, the water could have reached room temperature from either direction.



State Functions

Internal energy is a state function.

It depends only on the present state of the system, not on the path by which the system arrived at that state.

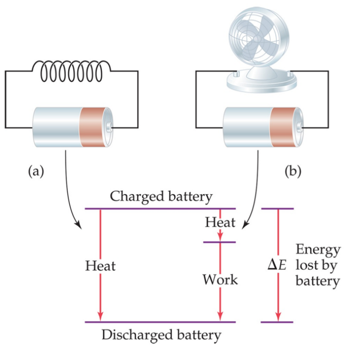
ΔE depends only on E_{initial} and E_{final}.

State Functions

However, q and w are *not* state functions.

In the diagram, 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.



Definitions: 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).

Chemical Energy

Endothermic Reaction: heat flow is into a system, absorbing energy from the surroundings.

Exothermic Reaction: energy flows out of the system.

Energy gained by the surroundings must be equal to the energy lost by the system.

Example 1. Is the freezing of water an endothermic or exothermic process? Explain.

2. Energy Examples

Classify each process as exothermic or endothermic. Explain. The system is underlined in each example.

Your hand gets cold when you touch ice.

The ice gets warmer when you touch it.

Water boils in a kettle being heated on a stove.

Water vapor condenses on a cold pipe.

Ice cream melts.

2. Energy Examples

Your hand gets cold when you touch ice. Exothermic - the hand loses heat.

The ice gets warmer when you touch it. Endothermic - absorbs energy.

Water boils in a kettle being heated on a stove. Both: endothermic as it is heated, exothermic as its molecules are boiled off.

Water vapor condenses on a cold pipe. Exothermic - cooling water molecules move slower until they condense.

Ice cream melts. Endothermic - absorbs energy, breaking bonds between water molecules.

3. System Examples

For each of the following, define a system and its surroundings and give the direction of energy transfer.

A. Methane is burning in a Bunsen burner in a laboratory.

B. Water drops, sitting on your skin after swimming, evaporate.

3. System Examples

A. Methane is burning in a Bunsen burner in a laboratory.
System - burner with gas. Surroundings - air around burner.
Energy flows from the burning gas into the surrounding air.

B. Water drops, sitting on your skin after swimming, evaporate.
System - skin with drops. Surroundings - air where evaporated water goes.
Energy flows through your skin to the water droplets, and the evaporating molecules escape into the atmosphere.

Internal Energy

Consider a system of molecules that can react with each other.

In such a system, molecules have kinetic energy by virtue of their motion.

They also have potential energy, in that they could react, absorbing or releasing heat in the process.

Internal Energy, *E*, of a system is the sum of the kinetic and potential energies of all the "particles" in the system.

1st Law of Thermodynamics, Again

The first Law of Thermodynamics looks at systems in terms of energetic quantities.

Since energy can't be created or destroyed, the 1st Law takes into account three parameters of energy within a system that amount to the total internal energy:

$\Delta E = q + w$ **1st Law of Thermodynamics**	ΔE = Internal Energy (J) q = heat (J) w = work (J)

Internal energy can be changed by a flow of heat, work, or both.

Internal Energy

Thermodynamic quantities consist of two parts:

Number gives the magnitude of the change.

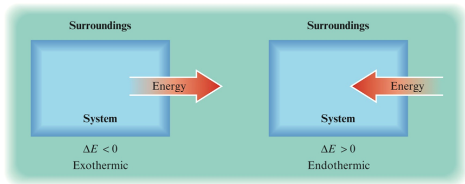
Sign indicates the direction of the flow, from the system's point of view:

Endothermic Process: q is positive.

Exothermic Process: q is negative.

System does work on surroundings: w is negative.

Surroundings do work on the system: w is positive.



Work

A common type of work associated with a chemical process involves work done by a gas (through expansion) or work done to a gas (by compression).

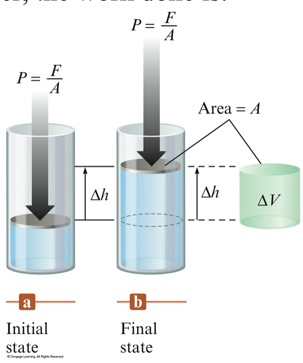
Imagine a cylinder with a gas that can expand or contract. In such a cylinder, the work done is:

Work = $P \times A \times \Delta h$

P is pressure, A is area

Δh is the piston

moving a distance.



Work

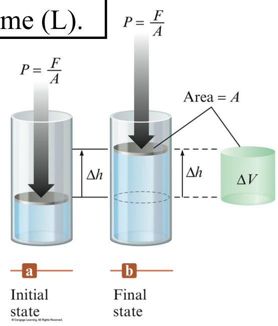
It is more practical to express work in terms of volume, and from the perspective of the gas (the system), if it pushing out on the piston (exerting energy on the surroundings:

Work = - $P\Delta V$	P = pressure (atm) ΔV = volume (L).

Some unit conversions:

Pressure: 1 atm = 760 torr

Energy: 1 L atm = 101.3 J



4. Work Example

How do the following amounts of work compare?

A gas expanding against a pressure of 2 atm from 1.0 L to 4.0 L.

A gas expanding against a pressure of 3 atm from 1.0 L to 3.0 L.

5. Work Energy

How much work is done by a system at constant pressure of 2.3 atm that changes in volume from 250 mL to 575 mL?

Work = - $P\Delta V$

= -2.3 atm x (0.575 L - 0.250 L) = -0.74 L atm

0.74 L atm x 101.3 J/L atm = -75 J

Introduction to Enthaply

The change in internal energy a system undergoes during a reaction is called enthaply, ΔH, and will be explored later in more detail.

For now, consider the reaction between hydrogen and oxygen:

$2H_2 + O_2$ --> $2\ H_2O$ $\Delta H = -572$ kJ.

The negative sign indicates a release of energy.

5. How much energy would be released if 0.5 moles of hydrogen reacted?

$\frac{0.5}{2} = \frac{1}{4} \cdot -572\text{kJ} = \underline{-143\text{kJ}}$

Homework

Read 6.1 & 6.2 in your textbook.

6.1 Problems in your Booklet
Due: Next Class.