Physics 23 Chapter 15

The First Law of Thermodynamics

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All of the gases discussed in this chapter are ideal monatomic gases. An ideal gas is one composed of particles (atoms) imagined to have zero size, and therefore are incapable of colliding with each other. The atoms are assumed to be incapable of bonding with each other. Such gases don't exist, but nearly so.

The total energy of the gas (sometimes called, "the internal energy") within the cylinder shown in the figure is the sum of all of the kinetic energies: As we saw in Chapter 14,

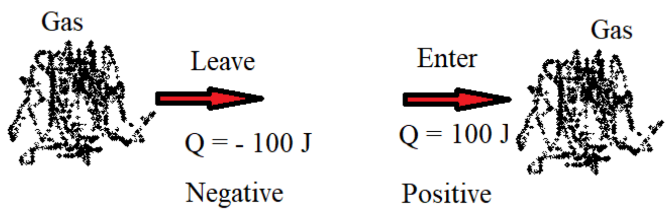
$$E = 3/2\ NkT \qquad E = 3/2\ nRT \qquad E = 3/2\ PV$$

In the first part of this chapter we consider thermodynamic processes whereby heat enters or leaves the gas while work is being done on, or by, the gas.

Later, we will look at thermodynamic processes involving more general systems, not ideal gases, but systems such as the human body, or machines.

The First Law of Thermodynamics

One way the internal energy of a gas can change is by thermal energy (heat) entering or leaving the gas. If a quantity Q of heat enters a gas, Q is positive; if heat leaves, Q is negative.



$$\Delta E = Q$$

Another way for the internal energy of the gas to change is for work to be done by the gas. We will let W be the work done <u>by</u> the gas. Recall that works W can be positive, negative, and even zero.

If gas is pushing upward on the piston as it rises, as happens in an expansion process, W is positive.

On the other hand, if the gas pushing upward on the piston as it falls, as happens during a compression of the gas, W is negative.

Either way, the internal energy changes according to the equation below:

$$\Delta E = -W$$

The equation below allows for heat to enter or leave the gas, and for the gas to do work:

$$\Delta E = Q - W$$

Heat Enters	Q Positive
Heat Leaves	Q Negative
Expansion	W Positive
Compression	W Negative

The equation above is the “First Law of Thermodynamics”

<u>Example A</u>:

A force pushing downward on the piston does 350 J of work *on* a gas. By Newton's Third Law, the work done *by* the gas is the opposite amount:

$W = -350$ J

At the same time, 400 J of heat energy leave the gas. What is the change in the internal energy of the gas?

$$\begin{aligned}\Delta E &= Q - W \\ &= -400 - (-350) \\ &= -50 \text{ J}\end{aligned}$$

<u>Example B</u>:

500 J of heat enters a gas, and while this is happening, 200 J of work is done *by* the gas.

$Q = 500$ J
$W = 200$ J

$$\begin{aligned}\Delta E &= Q - W \\ &= 500 - 200 \\ &= 300 \text{ J}\end{aligned}$$

Was the gas compressed, or expanded?

Answer: When gas does positive work, the gas pressure pushes upward on the piston, which also moves upward, leading to an expansion of the gas.

Example :

A container of gas is placed in a refrigerator and 200 J of heat leaves the gas, and while this is happening, -150 J of work is done by the gas.

(a) What is the change in the internal energy of the gas?

$Q = -200$ J
$W = -150$ J
$\Delta E = Q - W$
$= -200 - (-150)$
$= -50$ J

(b) Was the gas compressed, or expanded ?

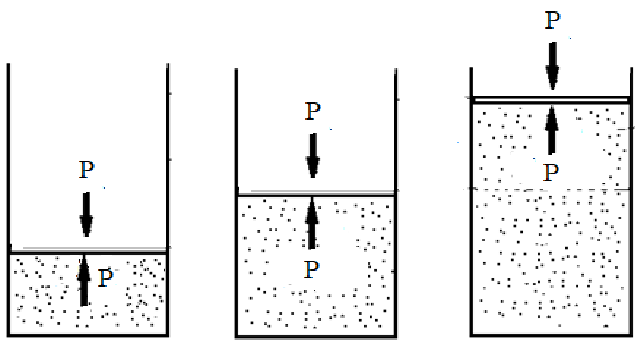
When negative work is done by a gas, it's because the piston is moving downward while the gas is pushing upward. Therefore, a compression occurred.

Isobaric Expansions and Compressions

Greek, "iso" same + "baros" weight

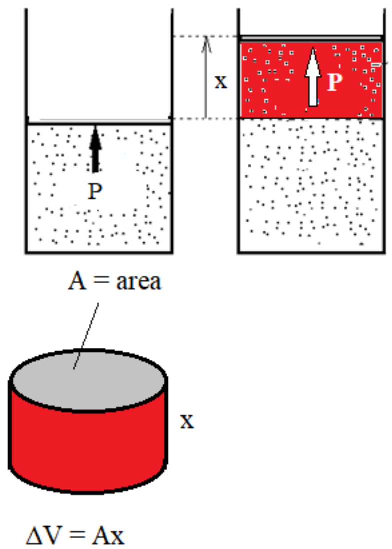
As a gas is slowly heated the gas expands as the piston moves upward. If we assume the piston's acceleration is virtually zero, the net force on it likewise is zero, which means that the internal gas pressure always matches whatever is the external pressure (often the constant atmospheric pressure). Constant pressure processes are called "*isobaric*."

The figures below show a cylinder of gas experiencing a constant-pressure expansion during heating. Note that the external pressure doesn't change, and neither does the internal pressure change: It's always equal to the external pressure, P.



Below we will calculate the work done by a gas during an isobaric expansion or contraction.

Calculating Work Done in Isobaric Expansions

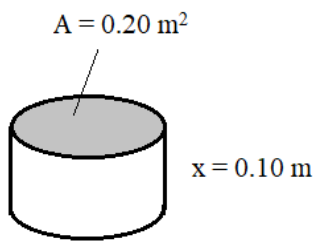
The additional volume created during the expansion shown at the right is the volume of the cylinder in the illustration: $\Delta V = Ax$ $W = Fx$ The force exerted on the underside of the piston by the gas pressure is $F = PA$. $W = (PA)x$ $= P(Ax)$ $= P\,\Delta V$ $W = P\,\Delta V$ This equation can equally well be used to deal with isobaric *compressions.*	

Example:

A cylinder of gas (not shown) has a base area A = 0.20 m^2. As the gas expands under a constant pressure of 116,000 Pa, the piston rises by 0.10 m. Assume 6500 J of heat enters the gas.

What is the change in the internal energy of the gas?

The amount of volume increase is illustrated below.



$$\Delta V = Ax$$
$$= (0.20)(0.10)$$
$$= 0.020\ m^3$$

$$P = 116{,}000\ Pa$$

$$W = P\Delta V$$
$$= (116{,}000)\ (0.020)$$
$$= 2320\ J$$

$$Q = 6500\ J$$

$$\Delta E = Q - W$$
$$= 6500 - 2320$$
$$= 4180\ J$$

Isochoric Processes

Greek, *"iso"* same + *"chora"* space

Constant volume processes are called "isochoric" processes. Such a process involving a gas in a cylinder with a piston could be achieved by locking the cylinder lid (piston) in place. The air inside may neither be expanded nor compressed ($\Delta V = 0$) so no work can be done by, or on, the gas:

$$W = 0$$

Example:

Six hundred joules of heat enter an ideal gas in an isochoric (constant volume) process. What is the change in the gas's internal energy?

$$Q = 600 \text{ J}$$
$$W = 0$$
$$\begin{aligned}\Delta E &= Q - W \\ &= 600 - 0 \\ &= 600 \text{ J}\end{aligned}$$

Adiabatic Processes

Greek, *"a"* not + *"diabatos"* passable

In "adiabatic" processes, gases are thermally insulated: Heat cannot pass through container walls:

$$Q = 0$$

<u>Example</u>:

In a certain adiabatic process, the internal energy of a gas contained in a cylinder decreases by 600 J.

What was the work done by the gas, and was the gas compressed, or expanded?

$$\Delta E = Q - W$$
$$-600 = 0 - W$$
$$W = 600\text{ J}$$

The work done by the gas is positive, which means the gas pushed upward on the piston and the air above it, which likewise moved upward. Therefore, the gas *expanded*.

Internal Energy Change

Recall the internal energy of an ideal monatomic gas is

$E = (3/2)\ nRT$, where T is the Kelvin temperature.

A change in the internal energy is represented as shown below:

$$\Delta E = (3/2)\ nR\Delta T$$

Example:

Two moles of an ideal gas at a pressure of 150,000 Pa undergo an isobaric (constant pressure) expansion in which the change in volume is $2.0 \times 10^{-3}\ m^3$ as 400 joules of heat enters the gas.

What is the change in temperature?

Solution:

$n = 2$ moles
$Q = 400$ J
$P = 150{,}000$ Pa
$\Delta V = 2.0 \times 10^{-3}\ m^3$

$$\begin{aligned} W &= P\Delta V \\ &= (150{,}000)\ (2.0 \times 10^{-3}) \\ &= 300\ J \end{aligned}$$

$$\begin{aligned} \Delta E &= Q - W \\ &= 400 - 300 \\ &= 100\ J \end{aligned}$$

$$(3/2)\ nR\ \Delta T = \Delta E$$
$$(3/2)\ (2)(8.31)\ \Delta T = 100$$

$$\Delta T = 4.01\ K^o$$

Example A:

As 600 J of heat exit a container consisting of three moles of a monatomic gas, the gas does -350 J of work. What is the change in the temperature of the gas?

$n = 3$ moles
$Q = -600$ J
$W = -350$ J

$\Delta E = Q - W$
$= -600 - (-350)$
$= -250$ J

$(3/2)\ nR\ \Delta T = \Delta E$
$(3/2)\ (3)(8.31)\Delta T = -250$

$\Delta T = -13.37\ K^o$

Example B:

Three thousand joules of heat is added to 20 grams of helium gas (A = 4.003) as the gas does 1600 J of work. What is the change in the temperature?

$n = 20 / 4.003$
$= 5.00$ moles
$Q = 3000$ J

$\Delta E = Q - W$
$= 3000 - 1600$
$= 1400$ J

$\Delta E = (3/2)\ n\ R\Delta T$
$1400 = (3/2)\ (5.00)\ (8.31)\ \Delta T$
$\Delta T = 22.46\ K^o$

Example:

2400 J of heat is added to three moles of an ideal gas at a temperature of 350 ^oK, and 1000 J of work is done by external forces in compressing the gas.

What is the final temperature of the gas?

The work done *by* the gas is the negative of the work done *on* the gas:

$W = -1000$ J

$Q = 2400$ J
$\Delta E = Q - W$
$\Delta E = 2400 - (-1000)$
$= 3400$ J

$(3/2)(3)(8.31)\ \Delta T = 3400$
$\Delta T = 90.9\ K^o$
$T = 350 + 90.9$
$= 440.9\ ^oK$

Isothermal Processes

Greek, *"iso"* same + *"therme"* heat
Processes in which the temperature doesn't change are called "isothermal" processes.

In an isothermal process involving an ideal gas, the change in internal energy is zero:

$\Delta T = 0$
$\Delta E = (3/2)\ nR\Delta T$
$= (3/2)\ nR\ (0)$
$= 0$

Example A:

In an isothermal ideal gas process, 400 J of heat enter the gas. How much work was done by the gas and was the gas compressed, or expanded?

$\Delta T = 0$
$\Delta E = (3/2)\, nR\Delta T$
$= (3/2)\, nR\, (0)$
$= 0$

$Q = 400$ J
$\Delta E = Q - W$
$0 = 400 - W$
$W = 400$ J

The gas does positive work, which means the piston pushed upward on the piston and air above it which likewise moved upward. Thus, the gas expanded.

Example B:

In an isothermal ideal gas process, 400 J of heat leave the gas. How much work was done by the gas and was the gas compressed, or expanded?

$\Delta T = 0$
$\Delta E = (3/2)\, nR\Delta T$
$= 0$
$Q = -400$ J

$\Delta E = Q - W$
$0 = -400 - W$
$W = -400$ J

The gas did negative work, which means the gas pushed upward on the piston that was moving *downward* ($\theta = 180^o$). Thus, the gas is compressed.

Chemical Energy

Chemical energy is defined as the energy which is stored in the bonds of chemical compounds (molecules and atoms). When certain chemicals undergo an “exothermic” reaction, heat is released.

An example of an exothermic reaction is the combustion of octane gasoline, which produces heat according to the equation below:

$C_8H_{18} + 25/2\ O_2 \rightarrow 9\ H_2O + 8\ CO_2$ + Heat

Without proof, we state that the heat released is 4.79×10^7 J per kilogram of octane.

Thermodynamic Efficiency

The “thermodynamic efficiency” of an agent (machine, animal, person) is defined to be the ratio of work it does, divided by the energy used.

Efficiency = Work Done/Energy Used

> Example:
>
> 100,000 J of internal chemical energy is used by a certain gasoline-burning engine that does 15,000 J of work.
>
> What is the engine’s efficiency?
>
> Efficiency = Work Done/Energy Used
> = 15,000/100,000
> = 0.15

Example:

A woman is doing work at the rate of 30 J/s while climbing stairs. Her efficiency is 20%.

(a) How much of her stored chemical energy will she use if she continues working at this rate for one hour?

Solution:

In one hour (3600 s), she does the following amount of work:

(30 J/s) (3600 s = 108,000 J

108,000 = 0.20 (Energy Used)
Energy Used = 540,000 J

540,000 J of energy is consumed, but only 108,000 J is used to do work. What happens to the other 432,000 J?

Answer: The remaining energy appears as heat that is released as the by-product of exothermic reactions that facilitate muscle contraction.

Example:

(a) How much of a person's internal energy will be used to do 40,000 joules of work with an efficiency of 5%?

0.05 = 40,000 Joules/Energy Used
Energy Used = 800,000 joules

(b) How much heat is created inside the person?

800,000 – 40,000 = 760,000 joules

(c) If the temperature of the person is to remain constant, 760,000 J of the heat will have to leave her body. Such heat loss is commonly achieved by

-- Evaporation of perspiration
--Air moving across skin takes heat away (convection).
--Direct contact with a cooler object (conduction).
--Radiation
--Respiration. Air breathed in is warmed, then exhaled.

Example:

When one gallon of gasoline is burned in a car engine, 1.20×10^8 J of the gasoline's internal energy is used to do work and create heat:

Suppose that during the same time 1.00×10^8 J of heat is created and lost to the environment:

How much work was done?

Chemical Energy Used = Heat Released + Work Done

1.20×10^8 J = 1.00×10^8 J + Work Done

Work Done = 0.20×10^8 J

Metabolic Rate

The average metabolic rate in a person or animal is the rate at which the body's internal chemical (food) energy is used by a person or animal, measured in joules per second, or watts.

What is the average metabolic rate of woman who metabolizes 10,500 kJ of food energy in one day, i.e., uses 10,500 kJ of his body's chemical internal energy in 24 hours?

$R = (1.05 \times 10^7 \text{ J}) / (24 \times 3600 \text{ s})$
$\quad = 122 \text{ W}$