Thermodynamics
Heat & Internal Energy
Heat Capacity

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Last time

- the ideal gas equation
- moles and molecules

Overview

- finish applying the ideal gas equation
- thermal energy
- introduced heat capacity

Ideal Gas Equation

The equation of state for an ideal gas:

$$PV = nRT$$

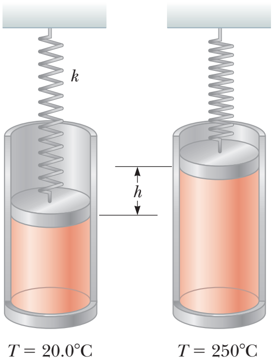
where

- P is pressure
- V is volume
- n is the number of moles (amount of gas)
- $R = 8.314$ J mol^{-1} K^{-1} is the universal gas constant
- T is temperature

The LHS and RHS of this equation both have units of Joules (energy).

Problem #74

A cylinder is closed by a piston connected to a spring of constant 2.00×10^3 N/m. With the spring relaxed, the cylinder is filled with 5.00 L of gas at a pressure of 1.00 atm and a temperature of 20.0°C.
(a) If the piston has a cross-sectional area of 0.0100 m^2 and negligible mass, how high will it rise when the temperature is raised to 250°C?



[1] Serway & Jewett, page 588.

Problem #74

(a) Find h.

$$PV = nRT$$

We have enough info to know the number of moles, n, or we can work around that because the amount of gas does not change as it is heated.

$$\frac{P_f V_f}{T_f} = \frac{P_i V_i}{T_i}$$

Also, $kh + P_0 A = P_f A$

and $P_i = P_0$ and $V_f = V_i + Ah$.

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$$\begin{aligned} P_f V_f &= P_i V_i \frac{T_f}{T_i} \\ \left(\frac{kh}{A} + P_0\right)(V_i + Ah) &= P_0 V_i \frac{T_f}{T_i} \\ k h^2 + \left(\frac{kV_i}{A} + AP_0\right) h + P_0 V_i \left(1 - \frac{T_f}{T_i}\right) &= 0 \end{aligned}$$

Problem #74

(a) Find h.
Solving quadratic:

$$k h^2 + \left(\frac{kV_i}{A} + AP_0\right) h + P_0 V_i \left(1 - \frac{T_f}{T_i}\right) = 0$$

Remember: $T_f = 250 + 273$ K, $T_i = 20 + 273$ K

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positive solution:

$$\underline{h = 0.169 \text{ m}}$$

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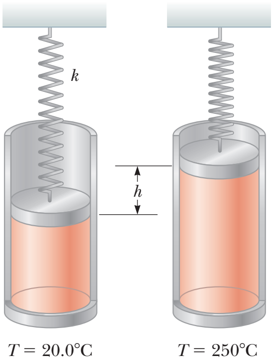
A cylinder is closed by a piston connected to a spring of constant 2.00×10^3 N/m. With the spring relaxed, the cylinder is filled with 5.00 L of gas at a pressure of 1.00 atm and a temperature of 20.0°C.
(b) What is the pressure of the gas at 250°C?



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$$P_f = 1.35 \times 10^5 \text{ Pa}$$

Heat and Energy

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Benjamin Thompson (Count Rumford, or simply "Rumford") was an American Loyalist soldier during the American revolution and moved to Europe after the war.

While studying canon manufacture in Munich, he noticed that the process of boring canons (drilling out the barrel) produced an incredible amount of heat, especially if the drill bit was dull.

Heat and Energy

This fluid model could not explain why *friction* of the drill bit on the canon would produce enough heat to keep water boiling, basically for as long as the drilling continued.

Wearing away the canon metal was not producing the heating (a dull bit wears away the metal more slowly), the friction was.

Heat and Energy

James Prescott Joule realized that this meant there was an equivalence between work and heat: both were kinds of energy transfers.

He did many experiments to quantify this relationship.

About 4,180 Joules of mechanical energy are needed to increase the temperature of 1 kg of water by 1°C. (More on this later.)

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About 4,180 Joules of mechanical energy are needed to increase the temperature of 1 kg of water by 1°C. (More on this later.)

We conclude that there is another type of energy: a hot object has more energy than a similar cold object.

Thermal Energy and Internal Energy

thermal energy[1]

The energy that an object has as a result of its temperature.

Internal energy, E_{int} or U

The energy that a system has as a result of its temperature and all other molecular motions, effects, and configurations, when viewed from a reference frame at rest with respect to the center of mass of the system.

Internal energy can be thought of as a combination of kinetic and potential energies of microscopic particles, but it is different from mechanical energy, because it cannot be directly converted into work.

[1] This definition is not universal!

Internal energy vs. Heat

Internal energy: U is the symbol most commonly used for internal energy, but it should not be confused with potential energy! They are different: potential energy can be directly converted to work, internal energy cannot.

The textbook uses E_{int} for internal energy.

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Heat, Q

Energy that is transferred into or out of a system in thermal contact with its environment because of a temperature difference between the system and environment.

Heat and internal energy are not the same thing. Heat changes the internal energy of the system.

Bond Energy

bond energy

The energy that an object has as a result of the configuration of its constituent particles at a microscopic level.

$$\text{internal energy} = \text{thermal energy} + \text{bond energy}$$

Intuitively,

bond energy is an intermolecular potential energy of all of the atoms or molecules due to how they are bonded, and

thermal energy is the the kinetic energy of the random motion of the atoms or molecules.

Units of Internal Energy and Heat: Calories

The units of both internal energy and heat are Joules, J.

However, heat was not always understood to be an amount of energy, so other units have been defined for it, and are still sometimes used.

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The "calories" listed on food labels are sometimes called "Calories" (capital C) because they are in fact kilocalories.

1 Calorie = 1 kilocalorie = the heat required to raise the temperature of 1 kilogram of water by 1 degree Celsius.

1 calorie = 4.18 Joules.

First Law of Thermodynamics

Where this is headed:

1st Law

The change in the internal energy of a system is equal to the sum of the heat added to the system and the work done on the system.

$$\Delta E_{\text{int}} = W + Q$$

This is just the conservation of energy written in a different way!

It takes into account that heat is energy.

Heat Capacity

Before we look more closely at the first law, let's look at the effect of adding heat to a substance that is not near a phase change.

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It requires less energy to raise the temperature of some objects compared to others.

Obviously, a small amount of water requires less heat to raise its temperature by 1 degree than a large amount of water.

But even two objects of the same mass may require different amounts of heat to change their temperature by 1 degree if they are made of different materials.

Different materials have different **heat capacities**.

Heat Capacity

Heat Capacity, C

of a sample of substance is the quantity of heat required to change the temperature **of that sample** by 1 degree C (or K).

$$Q = C\,\Delta T$$

where ΔT is the change in temperature and Q is the heat.

$$Q \propto \Delta T$$

and C is the constant of proportionality.

Specific Heat Capacity

However, it is usually more useful to compare one kind of substance to another for a given mass (*eg.* 1 kg).

Specific Heat Capacity[2], c

of a substance is the quantity of heat required to change the temperature of a unit mass of that substance by 1 degree C (or K).

$$Q = cm\Delta T$$

m is the mass of the object.

For example, water has a specific heat capacity $c = 4186$ J kg^{-1} K^{-1}.

$$C = cm$$

[2] Here, this refers to a process where temperature changes at constant pressure. In solids and liquids this distinction is not too significant.

Specific Heat Capacity

Different materials have different heat capacities.

- for Lead, $c = 129$ J kg^{-1} K^{-1}
- for Hydrogen, $c = 14300$ J kg^{-1} K^{-1}

Hydrogen gas's heat capacity is phenomenally high. (Its molar mass is small.)

Most substances are in the range 500 - 2000 J kg^{-1} K^{-1}.

This means that water also has quite a high heat capacity (4186 J kg^{-1} K^{-1}). This has an effect on Earth's weather and climate, since oceans make most of Earth's surface.

Summary

- applying the ideal gas equation
- thermal energy
- introduced heat capacity

Homework Serway & Jewett:

- Look at examples 20.1 – 20.3.
- new: Ch 20, onward from page 615. (OQs: 3, 5, 7; CQs: 3, 11); Probs: 1, (3, 9, 13, 63, 69)