

PHYS430 - Thermal Physics

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Chapter 1

Energy in Thermal Physics

1.1 Thermal Equilibrium

- After two objects have been in contact long enough, we say that they are in **thermal equilibrium**.
- The time required for a system to come to thermal equilibrium is called the **relaxation time**.
- **Temperature** is a measure of the tendency of an object to spontaneously give up energy to its surroundings.
- The flow of energy is from the object with a higher temperature to the lower one.
- For low-density gas at constant pressure, the volume should go to *zero* at approximately -273°C . which defines the **absolute zero**, in the **absolute temperature scale**, in K (kelvin).

1.2 The Ideal Gas

$$PV = nRT; \quad R = 8.31 \text{ J/mol K} \quad (1.1)$$

- A **mole** of molecules is Avogadro's number of them, 6.02×10^{23} .
- Number of molecules is $N = n \times N_A$
- Ideal gas law becomes $PV = NkT$, where k is Boltzmann's constant.
- The average transnational kinetic energy is $\bar{K}_{\text{trans}} = \frac{3}{2}kT$, where $kT = \frac{1}{40}\text{eV}$

1.3 Equipartition of Energy

Equipartition theorem At a temperature T , the average energy of any quadratic degree of freedom is $\frac{1}{2}kT$. For a system of N molecules, each with f degree of freedom, and there are no other (non-quadratic) temperature-dependent forms of energy, then its **total thermal energy** is

$$U = Nf\frac{1}{2}kT \quad (1.2)$$

Note, This is the *average* total thermal energy, but for large N , fluctuations become negligible.

1.4 Heat and Work

- Total amount of energy in the universe never changes, **Conservation of energy**
- **Heat** any spontaneous flow of energy from one object to another, caused by difference in temperature.
- **Work**, in thermodynamics, is any other transfer of energy into or out of a system.
- Work and heat refer to energy *in transit*.
- The total energy in a system is determined, but not the work nor the heat, it's meaningless.
- We ask about how much heat *entered* a system and how much work *was done on* a system.
- $\Delta U = Q + W$ is just a statement of the law of conservation of energy, but it's still called **first law of thermodynamics**.
- Processes of heat transfer: Conduction, Convection, and Radiation.

1.5 Compression Work

- From classical mechanics work is $W = \vec{F} \cdot d\vec{r}$
- Consider compressing gas with a piston of area A a distance Δx , the change in volume is $\Delta V = -A\Delta x$
- Volume change should be quasistatic, meaning very slow so that the pressure defined is uniform. then $W = P\Delta V$, but $\Delta x = -\Delta V/A$; minus since the volume decreases.
- $W = -P\Delta V$ - quasistatic.
- If P is not constant,

$$W = - \int_{V_i}^{V_f} P(V) dV \quad (1.3)$$

- **isothermal compression** is slow that the temperature doesn't raise.
- **adiabatic compression** is so fast that no heat escapes from the gas.
- Isothermal process

- the change will be along an **isotherm** line, with $P = NkT/V$.
- The work done is

$$W = - \int_{V_i}^{V_f} P(V) dV = NkT \ln \frac{V_i}{V_f} \quad (1.4)$$

- The heat enters the system, from the first law, is

$$Q = \Delta U - W = \underbrace{\Delta \left(\frac{1}{2} N f k T \right)}_0 - W = NkT \ln \frac{V_f}{V_i} \quad (1.5)$$

- adiabatic process

- In the PV plot the change is from one isotherm to another.
- There should be no transfer of heat so

$$\Delta U = Q + W = W \quad (1.6)$$

- If it's *ideal* gas, U is proportional to T , so the temperature increases.
- By the equipartition theorem $U = \frac{f}{2}NkT$, so $dU = \frac{f}{2}Nk dT$, then from (1.3)

$$\frac{f}{2}Nk dT = -P dV \quad (1.7)$$

Using the ideal gas law for P and integrate

$$\frac{f}{2} \ln \frac{T_f}{T_i} = -\frac{V_f}{V_i} \quad \text{or} \quad V_f T_f^{f/2} = V_i T_i^{f/2} = \text{const.} \quad (1.8)$$

- Using the ideal gas law to eliminate T , $V^\gamma P = \text{const.}$, γ is the **adiabatic exponent**.

1.6 Heat Capacities

- **Heat capacity** of an object is the amount of heat needed to raise its temperature, per degree change

$$C = \frac{Q}{\Delta T} \quad (1.9)$$

- The more matter you have the larger the heat capacity, by factoring out the mass m we get **specific heat**

$$c \equiv \frac{C}{m} \quad (1.10)$$

- Note (1.10) is ambiguous, plug in the first law

$$C \frac{Q}{\Delta T} = \frac{\Delta U - W}{\Delta T} \quad (1.11)$$

Even if the energy of an object is a well-defined function of its temperature alone, the work W done on the object is not; it depends on the process path on PV plot.

- From (1.11) The **heat capacity at constant volume**, denoted C_V

$$C_V = \left(\frac{\Delta U}{\Delta T} \right)_V = \left(\frac{\partial U}{\partial T} \right)_V \quad (1.12)$$

- From (1.11) and (1.3) the **heat capacity at constant pressure**, denoted C_P

$$C_P = \left(\frac{\Delta U - (-P\Delta V)}{\Delta T} \right)_P = \left(\frac{\partial U}{\partial T} \right)_P + P \left(\frac{\partial V}{\partial T} \right)_P \quad (1.13)$$

for solid last term is almost negligible.

- At a **phase transformation**, you add heat in a system without increasing its temperature; such as melting of boiling water. Then $C = \frac{Q}{\Delta T} = \infty$
- The amount of heat needed to do this phase transformation is called **latent heat** L , and the **specific latent heat** is

$$l \equiv \frac{L}{m} = \frac{Q}{m} \quad (1.14)$$

It's ambiguous, but we assume the pressure is constant, and no other work done.

- Adding PV onto the energy gives the **enthalpy**

$$H = U + PV \quad (1.15)$$

it's the *total* energy you would need to create the system out of nothing.

Chapter 2

The Second Law

2.1 Two-State system