Thermal physics.

The temperature of a body is a measure of <u>the mean of the energy</u> of the translational, vibrational and rotational motions of its molecules or atoms. More precisely defined quantity is the **thermodynamic temperature** T, often called as **absolute temperature**, for two reasons: 1) it does not depend on the properties of a particular material; 2) it refers to an absolute zero according to the properties of the *ideal gas*. The SI unit of the absolute temperature is **kelvin** (1 K).

The Kelvin scale of the temperature selects the **absolute zero** ($T = 0$ K) as the theoretically lowest temperature. At this point, the particle constituents of matter have minimal motion and cannot become colder. In the classical physics, the thermal motion of particles is completely stopped at $T = 0$ K. According to the quantum mechanics, matter at absolute zero is in its ground state, the state of lowest possible energy (but not zero!). The second reference point of the Kelvin scale is the *triple point* of water at 273.16 K.

Heat is some kind of energy. When heat is transferred from one body to another, we speak about the <u>amount of tranferred heat</u> Q. **Heat transfer** between two bodies can occur through: 1) <u>direct transfer</u> due to mutual collisions of atoms and molecules, 2) <u>convection</u> , when hotter gas with lower density raises up and colder and heavier gas comes down due to gravitational forces, 3) electromagnetic <u>thermal radiation</u> (through the field) that is the most efficient way.

Heat capacity of the body C shows how much heat should be transferred to the body to increase its temperature by 1 K. $C = Q / \Delta T$. The unit of heat capacity in SI is 1 J/ K. **Molar heat capacity** of the substance is the heat capacity of one mole of this substance.

Reduced heat capacity c shows how much heat should be transferred to the **unit mass** of substance to increase its temperature by one degree $c = Q / (m \; \Delta T)$. The SI unit is 1 J / (kg K). Therefore, the transferred heat is $Q = c \; m \; \Delta T$ and heat capacity of the body is $C = c \; m$.

Work done by the gas due to expansion can be expressed in the form $W = p \; \Delta V$. Differentially small work $dW = p \; dV$. According to the common agreement used in physics, the system makes positive work when it expands and negative when it is being pressed. In this case the positive work is done by outer forces.

Thermodynamics (TD) is studying the heat phenomena on a macroscopic level ignoring the microstructure of the bodies. Main attention is paid to the <u>conditions, when the heat can be transferred from one body to another</u>. The bodies are in **thermodynamical (TD) equilibrium** when heat is not transferred between bodies. When two bodies are in TD equilibrium they have the same temperature.

Internal energy U of some amount of substance is the sum of the energies (kinetic and potential) of all particles of the body with respect to each other. The energy of the <u>ideal gas</u> is proportional to its temperature: For example, the internal energy of one mole of ideal gas $U = N_A \; E_{mk} = N_A \; (i/2) \; k \; T = (i/2) \; R \; T$, where i is *number of degrees of freedom* of gas molecules. Here, N_A is the Avogadro constant $N_A = 6.02 \cdot 10^{23}$ mol^{-1}.

Universal gas constant $R = 8{,}31$ J/(K $\cdot$ mol) describes the work that is done by one mol of ideal gas when it expands isobarically so that its temperature increases by 1 degree (1 K).

1st principle of thermodynamics is a version of the <u>law of conservation of energy</u>, adapted for thermodynamic systems. It states that the total energy of an isolated system is constant; energy can be transformed from one form to another, but cannot be created or destroyed. The first principle is often formulated in following way: the change ΔU of the internal energy U is equal to the amount of heat ΔQ supplied to the system, minus the amount of work $W = p \; \Delta V$ done by the system: $\Delta U = \Delta Q - p\Delta V$ or $dQ = dU + pdV$ in differential form.

Entropy S is a state parameter of the TD system that describes irreversible dispersion of energy during heat processes. Differential change of entropy is expressed as $dS = dQ/T$. The unit of entropy is 1 J/K. <u>Entropy describes the disorder of the system</u>. Because $dQ = T \; dS$, then every time the system receives some heat, the disorder of the system increases due to increase of chaotic motion of system particles.

2nd principle of thermodynamics: <u>the heat is never completely convertible to work</u>. It is impossible to create a periodically working heat engine that will not cause changes in surrounding bodies. The total entropy of an isolated system always increases over time, or remains constant in ideal cases where the system is in a steady state or undergoing a reversible process.

Thermal expansion is the change of the linear dimensions l or volume V (in most cases) with the change of temperature T (proportionally). It is expressed as: $\Delta l = \alpha \; l \Delta T$ (or $\Delta V = \beta \; V \Delta T$), where α is linear thermal expansion coefficient, β - volume/cubic thermal expansion coefficient, ΔT – change of the temperature.

Coefficient of thermal expansion α *(linear)* shows how great is the relative elongation of the body composed of a specified substance due to the increase of the temperature by 1 degree (1 K). $\alpha = \Delta l / (l\, \Delta T)$. Volume coefficient of thermal expansion β shows how great is the relative change of volume of the body composed of a specified substance due to the increase of the temperature by 1 degree (1 K). $\beta = \Delta V / (V\, \Delta T)$. The unit of both thermal expansion coefficients is the reciprocial kelvin (1 K^{-1}).

Ideal gas is simplified model of real gas which does not take into account: 1) particle <u>dimensions</u> and 2) <u>forces</u> with which particles affect each other.

Isothermal process is a heat process during which the <u>temperature</u> of the gas and thus its <u>internal energy</u> do not change. According the 1st principle $dQ = dU + pdV$, hence in the isothermal process the expansion of the gas ($pdV > 0$) is possible only with the help of external heat ($dQ > 0$). If $T = $ const, then $pV = $ const.

Isobaric process is the heat process during which gas <u>pressure</u> does not change. If $p = $ const, then $V/T = $ const. External heat dQ allows to do work and increases internal energy.

Isochoric process is the heat process during which gas <u>volume</u> does not change. If $V = $ const, then $p/T = $ const. External heat dQ is used only for increase of internal energy (there is no expansion, $dV = 0$).

Adiabatic process is the heat process during which there is no heat exchange with external environment. That means that the <u>entropy</u> does not change ($S = $ const). When there is no external heat, then the work ($p\, dV > 0$) is possible only due to reduction of internal energy ($dU < 0$).

Number of the degrees of freedom i is the number of independent coordinates that describe the motion of the particles of the system. The independent coordinate is coordinate that cannot be expressed through other coordinates. Single-atomic molecule can have only 3 translational coordinates. 2-atomic (N_2, O_2) or 3-atomic linear (CO_2) molecules have 3 translational and 2 rotational degrees of freedom ($i = 5$). Non-linear (2- or 3-dimensional) molecules (H_2O, CH_4) have 3 translational and 3 rotational degrees of freedom (6 in total).

Equation of state combines gas state parameters: pressure p, volume V and temperature T in the form of: $pV = zRT$, where z is number of moles of the gas (gas mass divided by molar mass) and R - <u>universal gas constant</u>.

Basic equation of the kinetic theory of the gas states that gas pressure depends on gas molecule **concentration** $n = N / V$ (number of molecules per unit volume) and **mean kinetic energy** of these molecules E_{mk}: $p = 2/3 n E_{mk}$. From this, $E_{mk} = 3/2\, kT$ and $p = n\, kT$, where k is Boltzmann' constant. In general, $E_{mk} = (i/2)kT$, where i is number of degrees of freedom of gas molecules.

Root mean squared (RMS) speed of the gas molecule (obtained during averaging of squares of velocities and taking square root of resulting average value) is $v_r = (3kT/m_0)^{1/2} = (3RT/M)^{1/2}$, where m_0 is the mass of single molecule and M is the molar mass.

Boltzmann constant k is the ratio of universal gas constant R and Avogadro constant (gas constant per ideal gas molecule) $k = R / N_A$. $k = 1,38 \cdot 10^{-23}$ J/K.

Triple point on a phase diagram describes a set of conditions when all three states of matter (solid, liquid and gas) are in equilibrium.

Saturated vapour is the state of the substance when only liquid and gaseous phases of the substance are in equilibrium (vaporization and condensation "compensate" each other). All substances have a unique defined <u>saturated vapour pressure</u> dependence on <u>temperature</u>.

Liquid is boiling, when its saturated vapour pressure becomes (due to increase of temperature) equal to the pressure of external environment. It can be observed as emerging of bubbles within the whole volume of liquid.

Latent heat (melting, vaporization or other heat) is amount of heat required for unit mass of substance to change its phase. Latent heat SI-unit is 1 J/kg. During change to a more energetic state, the latent heat is absorbed from the outside. During change to the less energetic state, the latent heat is dissipated to the environment.

Surface tension force is a mutual tractive force of the molecules of the liquid on its surface. Due to the effect of surface tension force, a liquid tends to take the form with the minimal surface area (sphere).

Coefficient of surface tension α shows what is the force required to rip the body from a liquid surface along the contact borderline acting as a 'zipper' $\alpha = F_p / l$ (SI unit is 1 N/m).

Heat engine is a device that converts <u>heat to the work</u>. Heat engine takes heat Q_1 from hotter body (heater), converts it to the work W and gives the remaining heat Q_2 to the colder body (cooler). The efficiency factor of heat engine is $\eta = W / Q_1 = (Q_1 - Q_2) / Q_1$ and its maximum value is $\eta_{max} = (T_1 - T_2) / T_1$, where T_1 and T_2 are corresponding temperatures of heater and cooler.