PRACTICAL 8

ENTROPY OF AN ISOLATED SYSTEM

O b j e c t i v e : study of the entropy increment in irreversible adiabatic processes.

E q u i p m e n t a n d a c c e s s o r i e s : calorimeter, heater, thermometer (0.1 K division), tank with water, measuring glass (100–150 ml), flask (500 ml), samples (steel, aluminum, brass), ice-cubes, balance scale, set of balances.

INTRODUCTION

One of the fundamental laws of nature, *the second law of thermodynamics*, determines the direction of the processes occurring in an isolated macrosystem. Using the concept of entropy, the second law of thermodynamics can be formulated as follows: *the entropy of an isolated system does not decrease whatever processes are occurring in it:*

$$\Delta S \geq 0, \qquad (1)$$

with ΔS being the change of the system's entropy. Moreover, the equality take place for reversible processes, and the inequality takes place for irreversible ones. Hence, ΔS can be used as a measure of the reversibility of a process occurring in an isolated system: the smaller ΔS the closer the process under study to a reversible one.

In this practical, change in the entropy of an isolated system is measured for an irreversible process of heat exchange, and for a reversible process of ice melting.

In the first experiment, heat transfer occurs when different bodies, heated to the same temperature T, are lowered into water, which is in a calorimeter at a temperature T_w. The presence of an external calorimeter cup makes the system practically heat-insulated. In the proposed experiment, three agents participate in the heat exchange:

1. Sample cylinder (referred to as “sample”) with mass m_s, specific heat c_s and heated up to temperature T_s (which is practically the temperature of the boiling water steam).

2. Aluminum can with mass m_c, specific heat c_c and temperature T_w.

3. Water in the calorimeter’s can with mass m_w, specific heat c_w and temperature T_w.

After the heat exchange, equilibrium is achieved at temperature T_0. For each of the agents, change of the entropy can be calculated:

$$\Delta S = \int_1^2 \frac{dQ}{T}.$$

For solid bodies and liquids, $dQ = cmdT$, hence:

change of the sample’s entropy: $\Delta S_s = c_s m_s \ln\frac{T_0}{T_s}$

change of the can’s entropy: $\Delta S_c = c_c m_c \ln\frac{T_0}{T_w}$;

change of the water’s entropy: $\Delta S_w = c_w m_w \ln\frac{T_0}{T_w}$.

According to the additivity of the entropy, the change in entropy of the entire system is:

$$\Delta S = (m_c c_c + m_w c_w)\, ln\frac{T_0}{T_w} + m_s c_s ln\frac{T_0}{T_s}. \qquad (2)$$

In the second experiment, a piece of ice (1 or 2 ice-cubes whose temperature is T_{ice}) is put into the calorimeter filled in with water having temperature T_w. An equilibrium temperature T_0 is set after the ice is warmed up (the change of its entropy is $c_{ice}m_{ice}ln(T_m/T_{ice})$) and melted at the melting temperature T_m (the change of its entropy is $\lambda_{ice}m_{ice}/T_m$, λ_{ice} – specific heat of ice melting). Melt water of mass m_{ice} is heated up to temperature T_0, and the calorimeter and water initially contained in it (having common temperature T_w) cool down to same equilibrium temperature T_0. The total change in the entropy of the system by the time thermal equilibrium is established can be calculated using the formula:

$$\Delta S = m_{ice}c_{ice}\, ln\frac{T_m}{T_{ice}} + \frac{\lambda_{ice}m_{ice}}{T_m} + m_{ice}c_w ln\frac{T_0}{T_m} + (c_w m_w + c_c m_c) ln\frac{T_0}{T_w} \qquad (3)$$

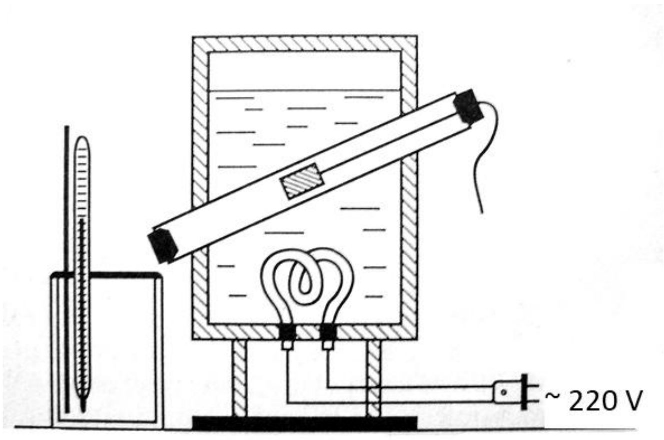

Fig. 1. Experimental setup

EXPERIMENTAL SETUP

As shown in Fig. 1, the sample has the form of a cylinder to which a thread is tied up. The sample is placed in the center of the tube which is closed at both ends with rubber stoppers (the thread tied to the sample is pressed by the top stopper, and its end hangs out). The tube is immersed in the boiling water that is heated by an electric coil. The calorimeter is an aluminum can (measure its mass) set into a heat-insulating jacket with a lid which has an opening for a thermometer. In this practical, can is used to catch the sample.

Attention! Remove the thermometer before grasping the sample!

If you first remove the lower stopper and then the upper stopper, the sample slips into the calorimeter's can along the tube.

The specific heat of ice melting and the specific heats of water, ice and aluminum, from which the calorimeter's can is made, are given at the Appendix to this description.

MEASUREMENT AND DATA PROCESSING

Task 0. Preparing the setup

Measure masses of the samples and cans with the help of the balance scales.

Using the flask prepare ~400 ml of water and put it away from the heater. It will be used to fill in the calorimeter's can.

Check that the water level in the vessel with the heater is not lower than the upper end of the tube immersed in it. Turn on the heater and place the first sample in the tube.

While the water and the sample in the heater vessel are heated up for at least 10 minutes (from the moment the boiling of water sets on), investigate the ice-melting process using ice-cubes and the calorimeter.

Task 1. Study of the entropy change during warming up and melting the ice-cubes

Using the measuring glass, pour 100 ml of water at room temperature from the flask into the calorimeter’s can, measure temperature of the water with a thermometer.

Take two pieces of ice from the freezer compartment of the refrigerator, weigh the ice using the balance scales (as quickly as possible) and place it into the calorimeter’s can filled in with water. Observe a decrease in water temperature in the calorimeter. Measure the equilibrium temperature t_0 established after a while. Determine the initial temperature of the ice t_{ice} by reading the thermometer in the freezer compartment of the refrigerator.

After completing the measurements, pour out the water from the calorimeter glass.

Write down the **heat balance equation** (the algebraic sum of heat received or given by the bodies in an isolated system is zero) and determine the theoretical value of the equilibrium temperature $t_{0\ \mathrm{theor}}$ for this system. Compare it to the measured value.

Compute the change of the entropy using (3). Fill in Table 1.

Table 1. Ice-melting experimental results

m_{ice} = m_c = m_w =	t_w, °C (T_w, K)	T_{ice}, °C (T_{ice}, K)	$t_{0\ exp}$, °C ($T_{0\ exp}$, K)	$t_{0\ theor}$, °C ($T_{0\ theor}$, K)	ΔS , J/K

Task 2. Study of the entropy change during the heat transfer

Pour 100 ml of water at room temperature from the flask into the calorimeter's can, use the measuring glass.

Place the calorimeter away from the heater. Measure the water temperature t_w **before** immersing the sample into the water. **Remove the thermometer!**

Allow ~ 10 min (after the boiling sets on) for the sample to heat up, remove the lower stopper and then the upper stopper from the tube, let the sample slide into the calorimeter's can. Quickly close the calorimeter, place it away from the heater and observe the water temperature rise. Measure the temperature $t_{0\ exp}$ after heat exchange is completed.

Pour the water out of the calorimeter and let it cool down. To accelerate cooling, water at room temperature may be used. Repeat measurements with the other samples (the mass of water in all experiments should be the same).

For each experiment, use the **heat balance equation** to calculate the temperature $t_{0\ theor}$ established in the system after the heat exchange is completed. The temperature of the heated up sample is assumed to be $t_s = 95\ °C$, i.e. 5 degrees below the boiling temperature of water. Compare the calculated values of the equilibrium temperature with the experimental ones. Using formula (2), find the change in the entropy of the system.

The results of all measurements and calculations write down into Table 2.

Table 2. Experimental results

sample	m, g	t_w, °C (T_w, K)	$t_{0\ exp}$, °C ($T_{0\ exp}$, K)	$t_{0\ theor}$, °C ($T_{0\ theor}$, K)	ΔS, J/K
1 aluminum 2 brass 3 steel					

QUESTIONS AND EXERSIZES

1. Formulate the thermodynamic definition of entropy. In which cases the entropy of the system increases; decreases? Is it possible to change the entropy of the system without telling it the heat?

2. What is the statistical meaning of entropy?

3. In what state - liquid or crystalline (at the same temperature) - does the body have more entropy and why?

4. Give the different formulations of the second law of thermodynamics.

5. Prove that the prohibition of Clausius (the statement about the impossibility of processes whose only result is the transfer of heat from a cold body to a hot body) is equivalent to the statement that it is impossible to reduce entropy in an isolated system.

6. What is the ideal heat engine? What is its maximum efficiency?

7. In a heat-insulated aluminum vessel weighing 120 g at a temperature of 20 ° C pour 210 g of water having a temperature of 100 ° C. How will the entropy of the system change as a result of heat exchange?

8. Find the entropy change for 30 g of ice as it turns to steam. The initial temperature of the ice is t_1 = - 40 ° C, the temperature of the vapor t_2 = 100 ° C.

Appendix

Reference data

Specific heat

Aluminum: $c_a = 879\ J/(kg \cdot K)$

Brass: $c_b = 385\ J/(kg \cdot K)$

Steel: $c_s = 461\ J/(kg \cdot K)$

Ice: $c_{ice} = 2.1 \cdot 10^3\ \ J/(kg \cdot K)$

Water: $c_w = 4.18 \cdot 10^3\ \ J/(kg \cdot K)$

Specific heat of ice melting $\lambda_{ice} = 333 \cdot 10^3\ J/kg$.