AE405 LAB-4 Heat capacity of metals

Fig. 1: Experimental set-up for determining the heat capacity of metals.

PRINCIPLE AND TASK

Heated specimens are placed in a calorimeter filled with water at low temperature. The heat capacity of the specimen is determined from the rise in the temperature of the water.

Equipment

Calorimeter, 500 ml 04401.00 1

Metal bodies, set of 3 04406.00 4

Stainless steel pot, 3000 ml 05934.00 1

Butane burner, Labogaz 206 type 32178.00 1

Butane cartridge 47535.00 1

Aneroid barometer 03097.00 1

Stopwatch, digital, 1/100 sec. 03071.01 1

Thermometer, -10...+ 50 °C 38033.00 1

Portable balance JR300 48891.00 1

Fish line, l = 100 m 02090.00 1

Triangle w. pipeclay, l = 60 mm 33278.00 1

Tripod, ring dia. 140 mm, h = 240 mm 33302.00 1

Glass beaker, short, 250 ml 36013.00 1

Glass beaker, short, 600 ml 36015.00 1

Glass beads, d = 6 mm, 850 pcs. 36756.25 1

Objective

1. To determine the heat capacity of the calorimeter by filling it with hot water and determining the rise in temperature.

2. To determine the specific heat capacity of aluminium, iron and brass.

Introduction

Different substances require different quantities of heat to produce a given temperature change. For example, about three and one-half times as much heat is needed to raise the temperature of 1 kg of iron through a given temperature interval ΔT as is needed to raise the temperature of 1 kg of lead by the same amount. This material behavior is characterized quantitatively by specific heat, which is ***the amount of heat necessary to raise the temperature of 1 gram of the substance 1 degree Celsius***. Thus, iron has a greater specific heat than lead The specific heat of a material is specific or characteristic for that material. As can be seen from the definition, the specific heat of a given material can be determined by adding a known amount of heat of a known mass of material and noting the corresponding temperature change. The purpose of this experiment is to determine the specific heats of some common metals by calorimetry methods.

Theory

The change in temperature ΔT of a substance is proportional to the amount of heat ΔQ added (or removed) from it:

$$\Delta Q \propto \Delta T, \tag{1}$$

In equation form, we may write

$$\Delta Q = C\,\Delta T. \tag{2}$$

Where the constant of proportionality C is called the heat capacity of the substance. However the amount of heat required to change the temperature of an object is also proportional to the mass of the object. Hence, it is convenient to define specific heat capacity c (or simply specific heat):

$$c = \frac{C}{m}$$

Which is the heat capacity per unit mass of a substance. Thus, equation (2) becomes:

$$\Delta Q = c\, m\, \Delta T \qquad \text{or} \qquad c = \frac{\Delta Q}{m \Delta T}. \tag{3}$$

And ***the specific heat is then the amount of heat (in calories) required to change the temperature of*** $1\ g$ ***of a substance*** $1c^{\circ}$***.*** The ***calorie unit of heat is then the amount of heat required to raise the temperature of*** $1\ g$ ***of water*** $1\ c^{\circ}$. By definition, then, water has a specific heat of $1\ cal/g\ c^{\circ}$.

$$c = \frac{\Delta Q}{m \Delta T} = \frac{1cal}{(1g)\,(1c^{\circ})} = 1cal/g\ c^{\circ}$$

The heat capacity of a calorimeter and the specific heat of a substance can be determined experimentally by measuring the temperature change of a given mass of substance produced by a quantity of heat. This is done indirectly by a calorimeter procedure known as ***the method of mixtures***. If several substances at various temperatures are brought together, the hotter substances will lose heat and the colder substances will gain heat until all the substances reach a common equilibrium temperature. If the system is insulated so that no heat is lost to the surroundings, then by the conservation of energy, the heat lost is equal to the heat gained.

In part ***(I)*** of this experiment an empty calorimeter at room temperature is filled with a known mass of hot water, then the mixture temperature is measured. So we may write

$$heat\ gained = heat\ lost$$

$$\Delta Q_{calorimeter} = \Delta Q_{water}$$

or

$$C \Delta T_{calorimeter} = m \times c \times \Delta T_{water} \tag{4}$$

In part ***(II)*** hot metal is added to water in a calorimeter cup and the mixture is stirred until the system is in thermal equilibrium. The calorimeter isolates the system from losing heat. In mathematical form, we may write:

$$heat\ lost = heat\ gained$$

$$\Delta Q_{metal} = \Delta Q_{calorimeter} + \Delta Q_{water}$$

or

$$m_m \times c_m \times (T_h - T_f) = (C_{cal} + m_w \times c_w) \times (T_f - T_c) \tag{5}$$

Where T_h is the initial temperature of the hot metal, T_c is the initial temperature of the colder water and calorimeter cup and stirrer, and T_f is the final intermediate equilibrium temperature of the system.