Student experiment (10th - 13th grade)	**The calorimetry of ethanol**	🕐 **Time:** max. 30 min.

Safety:

safety glasses

Instruments:

- an Erlenmeyer flask insulated with Styrofoam
- a porcelain bowl
- a lighter
- a scissors lift
- retort stand material
- a scale
- a thermometer

Chemicals:

- ethanol (H: 225; P: 210)
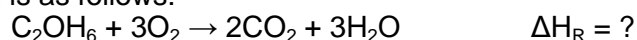- water

Experiment:

- Weigh the Erlenmeyer flask and record its mass.
- Put 500ml of water into the flask and note the mass again. Fasten the flask above table surface using retort stand clamps, etc.
- Measure the water to confirm that its temperature is constant. Do not begin the experiment until no temperature change is readable.
- Weigh ethanol (roughly 1g) and place it in a porcelain dish. Work quickly. Otherwise part of the ethanol will evaporate and skew the experimental results.
- Place the porcelain dish directly under the flask using the scissors lift and light the ethanol. Measure the water temperature as soon as the ethanol has completely burnt away. Calculate the temperature difference between the initial and final values.

Observations: The water temperature has risen.

Results:

Heat is given off during the combustion of ethanol and warms the water. The equation is as follows:

$$C_2OH_6 + 3O_2 \rightarrow 2CO_2 + 3H_2O \qquad \Delta H_R = ?$$

The heat of reaction, ΔH_R, is based on 46g of ethanol (1mol). First, however, we will calculate the overall heat, Q, released. This can be achieved using the heat capacity, c_P, of water:

$$Q = c_P \text{ (water) } * \text{ mass (water) } * \Delta T$$

In order to warm 1g (1g = 1ml) of water by 1 Kelvin (1°C) under STP conditions, 4.19 Joules of energy must be added. Ignore the heat capacity of the calorimeter.

Example: The test showed that 1g of ethanol warmed the 500g of water by 12.8 Kelvin (°C):

$$Q = 4.19 \text{ J/(g } * \text{ K) } * 500g * 12.8K = 26.816 \text{ kJ}$$

So burning 1g of ethanol releases 26.816 Joules of heat energy. We now can link the reaction energy to the mol amount of substance available for combustion. In 1g of ethanol there are 0.0217mol of ethanol total (1g divided by the formula mass of ethanol at 46g). The reaction warmth of 1mol of ethanol is therefore:

$$\Delta H_R = -26.816 \text{ kJ/0.0217mol} = -1237 \text{ kJ/mol}$$

The specific heat capacity of the calorimeter we ignored for the first step above. But the calorimeter is also warmed during the process, not just the water inside. If we do take the heat capacity of the calorimeter into account (e.g. 200 J/g * K), we get an expanded and more accurate formula:

$$Q = 4.19 \text{ J/(g } * \text{ K) } * 500g * \Delta K + 200 \text{ J/(g } * \text{ K) } * \Delta K$$

We can do this by previously calculating the heat capacity of the calorimeter by observation. For example, burning a substance with a known amount of energy released and seeing how much the calorimeter temperature changes.

SALiS
Student Active Learning in Science

Student Active Learning in Science – SALiS – Low-Cost –Experiments
Recorded by: S. Markic, N. Poppe, I. Eilks; University of Bremen
Translated by: N. Giles, University of Bremen

2