Because calorimetry is used to measure the heat of a reaction, it is a crucial part of thermodynamics. In order to measure the heat of a reaction, the reaction must be isolated so that no heat is lost to the environment. This is achieved by use of a calorimeter, which insulates the reaction to better contain heat. Coffee cups are often used as a quick and easy to make calorimeter for constant pressure. More sophisticated bomb calorimeters are built for use at constant volumes.

**Introduction**

Using calorimetry, it is possible to calculate the amount of heat gained and lost by each part of the system. The amount of heat lost or gained by any part of the reaction is related to the heat capacity of the substance. Heat capacity is the amount of energy, \( q \), needed to raise the temperature 1°C or 1 K. To calculate the amount of heat released or absorbed by a reaction in terms of \( \Delta T \), the equation for object's total heat capacity (below) is used:

\[
q = C \times \Delta T \tag{1}
\]

where

- \( q \) is the amount of heat gained or lost,
- \( C \) is the object's total heat capacity and
- \( \Delta T \) is the final temperature – initial temperature

The total heat capacity of an object (or sample) can be decomposed into one of two subsequent relationships. Either based on mass: \( C = m \times c_{sp} \) or based number of moles of atoms or molecule in the object: \( C = n \times c_p \)

with

- \( m \) is mass of the object in unites of g
- \( c_{sp} \) is the specific heat of the object in unite of J/g °C
- \( n \) is number of moles of molecules in the object in unites of moles
- \( c_p \) is the molar heat capacity in unites of J/mole °C (specifically under constant pressure; it would be \( c_v \) under constant volume)

Specific heat capacity takes into account the mass of a substance and the energy needed to raise its temperature by \( \Delta T \), as shown by combining equations 1 and 2a:

\[
q = m \times c_{sp} \times \Delta T \tag{3}
\]

Theoretically, the \( q_{products} \) should be equivalent to \( q_{reactants} \) because the amount of heat gained by one substance should be equal to the amount lost by another. However, the calorimeter also absorbs some of the heat from the reaction. Because of this, the heat capacity of the calorimeter is also necessary for calculations and results in the following equation:

\[
q_{reactants} + q_{products} + q_{calorimeter} = 0 \tag{4}
\]
The above equation can also be written as:

\[ q_{\text{cal}} = -(q_{\text{lead}} + q_{\text{water}}) \tag{5} \]

Example 1

If 150 g of lead at 100°C were placed in a calorimeter with 50 g of water at 28.8°C and the resulting temperature of the mixture was 22°C, what are the values of \( q_{\text{lead}} \), \( q_{\text{water}} \) and \( q_{\text{cal}} \)? The specific heat of water is 4.184 J/g °C and the specific heat of lead is 0.128 J/g °C.

**SOLUTION**

For lead, we know that:

\[ m = 150 \text{ g}, \]
\[ T_i = 100^\circ \text{C}, \]
\[ T_f = 28.8^\circ \text{C}, \]
\[ c_{sp} = 0.128 \text{ J/g °C} \]

For water:

\[ m = 50 \text{ g}, \]
\[ T_i = 22^\circ \text{C}, \]
\[ T_f = 28.8^\circ \text{C}, \]
\[ c_{sp} = 4.184 \text{ J/g °C} \]

\[ q = m \cdot c_{sp} \cdot \Delta T \]

\[ q_{\text{lead}} = (0.128 \text{ J/g °C}) \cdot (150 \text{ g}) \cdot (28.8^\circ \text{C} - 100^\circ \text{C}) = -1.37 \times 10^3 \text{ J} \]

\[ q_{\text{water}} = (4.184 \text{ J/g °C}) \cdot (50 \text{ g}) \cdot (28.8^\circ \text{C} - 22^\circ \text{C}) = 1.42 \times 10^3 \text{ J} \]

\[ q_{\text{cal}} = -(q_{\text{lead}} + q_{\text{water}}) \]

\[ q_{\text{cal}} = -(1.42 \times 10^3 + -1.37 \times 10^3) = -50.0 \text{ J} \]

In a **Coffee-Cup Calorimeter**, reactants are mixed in a solution, usually water, and measure the temperature change. Since styrofoam acts as a very effective insulator, little heat transfer takes place between the cup and the surrounding air. Everything within the cup, including the cup itself, is considered an isolated system. The heat of reaction is clearly defined as “the quantity of heat that would be exchanged with the surroundings in restoring the calorimeter to its initial temperature.” But the calorimeter cannot be restored back to its original conditions, which is why we take the negative quantity of heat producing the temperature change in the calorimeter.

\[ q_{\text{reaction}} + q_{\text{calorimeter}} = 0 \]

Which means,
\(q_{\text{reaction}} = - q_{\text{calorimeter}} \)

Temperature can be in Kelvin or Celsius, since the change in temperature, \(\Delta T = T_{\text{final}} - T_{\text{initial}}\), of either scale is identical.

Example 2

When 20 g of Al (s) at 98 degrees Celsius is placed in 50 g of \(\text{H}_2\text{O}\), the final temperature is 26.5 degrees Celsius. What was the initial temperature of \(\text{H}_2\text{O}\)? Given: \(C_{\text{spAl}} = 0.903 \text{ J/g}^\circ\text{C}\) and \(c_{\text{sp}{\text{H}_2\text{O}(l)}} = 4.18 \text{ J/g}^\circ\text{C}\)

**SOLUTION**

\[(q_{\text{H}_2\text{O}} = -q_{\text{Al}})\]

\[(m_{\text{H}_2\text{O}}) \cdot c_{\text{sp}{\text{H}_2\text{O}(l)}} \cdot (\Delta T) = (m_{\text{Al}}) \cdot (C_{\text{spAl}}) \cdot (\Delta T)\]

\[(50 \text{ g}) (4.18 \text{ J/g}^\circ\text{C}) (26.5^\circ\text{C} - T_{\text{iH}_2\text{O}}) = -(20 \text{ g}) (0.903 \text{ J/g}^\circ\text{C}) (26.5 - 98^\circ\text{C})\]

Using algebra, solve for \(T_{\text{iH}_2\text{O}}\):

\[T_{\text{iH}_2\text{O}} = 20^\circ\text{C}\]

Example 3

A 28.2 gram sample of copper \((c_{\text{sp}} (\text{Cu}) = 0.685 \text{ J/g}^\circ\text{C})\) is placed in a coffee cup calorimeter containing 100 grams of water that just stopped boiling. After some time the temperature of the water becomes constant at 92.3 °C. Assuming the atmospheric pressure is 1 atmosphere, calculate the initial temperature of the Copper block. Assume no heat is lost to the surroundings. \(c_{\text{p}{\text{H}_2\text{O}(l)}} = 4.18 \text{ J/g}^\circ\text{C}\)

**SOLUTION**

\[(q_{\text{H}_2\text{O}} = -q_{\text{Cu}})\]

\[(m_{\text{H}_2\text{O}}) \cdot c_{\text{sp}{\text{H}_2\text{O}(l)}} \cdot (\Delta T) = m_{\text{Cu}} \cdot c_{\text{sp}{\text{Cu}}} \cdot (T_{\text{iCu}} - 92.3^\circ\text{C})\]

\[(100 \text{ g}) (4.18 \text{ J/g}^\circ\text{C})(92.3^\circ\text{C} - 7.7^\circ\text{C}) = -(28.2 \text{ g}) (0.685 \text{ J/g}^\circ\text{C})(92.3^\circ\text{C} - T_{\text{iCu}})\]

Solve for \(T_{\text{iCu}}\):

\[-3210 \text{ J} = -19.3 \text{ J}^\circ\text{C} (92.3^\circ\text{C} - T_{\text{iCu}})\]

\[\frac{-3210}{-19.3} = 92.3^\circ\text{C} - T_{\text{iCu}}\]

\[T_{\text{iCu}} = 166^\circ\text{C} - 92.3^\circ\text{C} = -T_{\text{iCu}}\]
The lead is rather cold, which is necessary since it has a 6-fold lower specific heat than water.

Notice the difference in temperature units given by the two $c_p$ values. The $c_p$ of Cu = 0.685 J/g K is given in Kelvin and the $c_p$ of $H_2O_{(l)}$ = 4.18 J/g °C is given in Celsius. There is no need for conversions because when $c_p$ values are given, the change in temperature will remain constant. No matter what the units are, the change will still occur and be exactly the same as would naturally.

References


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