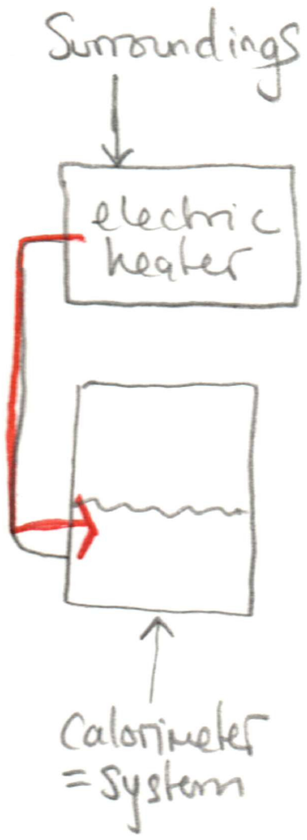
Calorimetry

→ red arrows indicate the direction of heat flow

1. A calorimeter was calibrated with an electric heater, which supplied 22.5 kJ of energy to the calorimeter and increased the temperature of the calorimeter and its water from 22.45°C to 23.97°C. What is the heat capacity of the calorimeter?



- $q_{cal} = C_{cal} \Delta T$

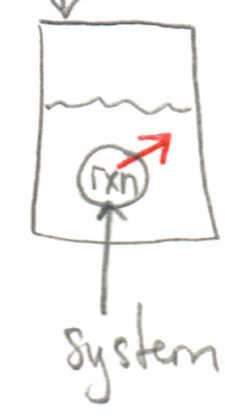
where q_{cal} is the heat absorbed by the calorimeter
C_{cal} is the specific heat of " "
ΔT is the change of temperature of " "

- $q_{cal} = +22.5$ kJ b/c heat was absorbed

$\Delta T = T_f - T_i = 23.97 - 22.45 = 1.52°C$

- So, $C_{cal} = \frac{q_{cal}}{\Delta T} = \frac{22.5 kJ}{1.52°C} = 14.8 \frac{kJ}{°C}$

2. When 0.113 g of benzene, C_6H_6, burns in excess oxygen in a calibrated constant-pressure calorimeter with a heat capacity of 551 J/°C, the temperature of the calorimeter rises by 8.60°C. Write the thermochemical equation and calculate the reaction enthalpy for $2\ C_6H_6(l) + 12\ O_2(g) \rightarrow 12\ CO_2(g) + 6\ H_2O(l)$.



- $q_{rxn} = -q_{released} = -q_{cal} = -C_{cal} \Delta T = -551 \frac{J}{°C} \times 8.60°C = -4738.6 J = -4.74 kJ$

- the ΔH calculated above was for 0.113g of benzene which is equal to $\frac{0.113 g}{78 g/mol} = 0.00145$ mol.

b/c MW(benzene) = (12×6)+(1×6) = 78g/mol

- ΔH_{rxn} or enthalpy of reaction is ΔH for 1 mol reaction. Since this is a combustion reaction where $\Delta H_{combustion}$ is the enthalpy of combustion. $\Delta H_{rxn} = \Delta H_{combustion}$
$\Delta H_{combustion}$ is reported for 1 mole of fuel, in this case C_6H_6
So the chemical equation will be $C_6H_{6(l)} + 6O_{2(g)} \rightarrow 6CO_{2(g)} + 3H_2O_{(l)}$.
ΔH_{rxn} for this equation is $-4.74 kJ \times \frac{1 mol}{0.00145 mol} = -3269 kJ$

b/c we want the ΔH for the combustion of 1 mol of benzene NOT 0.00145 moles.

The 2 underlined pieces of information make up the thermochemical equation

3. A reaction known to release 2.00 kJ of heat takes place in a calorimeter containing 0.200 L of solution and the temperature rose by 4.46°C. When 100 mL of nitric acid and 100 mL of sodium hydroxide were mixed in the same calorimeter, the temperature rose by 2.01°C. What is the heat output for the neutralization reaction? The chemical equation is $HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$.

1 We first have to find C_{cal} following the same steps as in problem #1 using the information about the rxn w/ the known amount of heat released. $C_{cal} = \frac{q_{cal}}{\Delta T} = \frac{2.00kJ}{4.46°C} = \frac{0.448kJ}{°C} = 488\frac{J}{°C}$

2 $q_{neutralization\ rxn} = -q_{cal} = -C_{cal}\Delta T = -(C_{s,H_2O} \times m_{H_2O})\Delta T$ b/c $C_{cal} = C_{H_2O} + C_{hardware}$

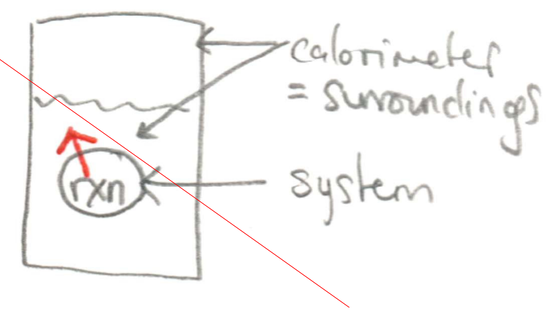
∴ we assume $C_{hardware} = 0$

∴ $C_{H_2O} = C_{s,H_2O} \times m_{H_2O}$

$= -\left(4.184\frac{J}{gK}\right)\left((100mL + 100mL) \times 1\frac{g}{mL}\right)(2.01°C)$

$= -1681.97\ J$

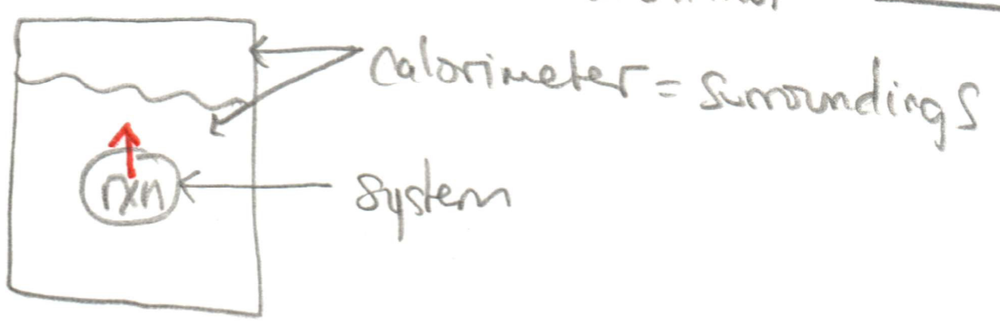
$= -1.68\ kJ$



q_{rxn} is negative. So, this rxn releases heat.

4. When a solution of 1.691 g of silver nitrate is mixed with an excess of sodium chloride in a calorimeter of heat capacity 216 J/(°C), the temperature rises 3.03°C. What is the reaction enthalpy for $NaCl(aq) + AgNO_3(aq) \rightarrow NaNO_3(aq) + AgCl(s)$?

- for this rxn $q = -q_{cal} = -C_{cal}\Delta T = -\left(216\frac{J}{°C}\right)(3.03°C) = 654.48\ J$
  $= 0.654kJ$
- this q is for 1.691 g of silver nitrate which is equal to $\frac{1.691g}{153.9\ g/mol} = 0.011\ mol$ b/c $MW(AgNO_3) = 107.9 + 14 + (16 \times 2) = 153.9\frac{g}{mol}$
- However, ΔH_{rxn} (reaction enthalpy) is for 1 rxn mol which according to the chemical equation above is for 1 mol of $AgNO_3$.
- So $\Delta H_{rxn} = .654kJ \times \frac{1\ mol}{0.011\ mol} = 59.5\ kJ$



*** The next two problems deal with combustion reactions. When reporting the molar heat or enthalpy of combustion for a certain fuel, one assumes the value of the heat of combustion or enthalpy of combustion is per 1 mole of the fuel. ***

5. If we set up a bomb calorimetry experiment to determine the molar enthalpy of combustion of ethane (C_2H_6) using 1 L of water as our heat sink, 2.805 g of ethane, and measure an initial and final temperature of 25.20°C and 58.92°C, respectively, what will be the experimentally determined molar enthalpy of combustion of ethane? The chemical equation is $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$? Assume the density of the water is 1.00 g/mL. Assume the calorimeter itself absorbs no heat. The specific heat capacity of water is 4.184 J/(g*K).

- fuel = C_2H_6
- 2.805 g ethane is $\frac{2.805\ g}{((12\times2)+(1\times6))\ g/mol} = 0.0935\ mol$
- $q_{rxn} = -q_{cal} = -C_{cal}\Delta T = -(C_{s,H_2O} \times M_{H_2O})\Delta T$ b/c we assume $C_{hardware} = 0$

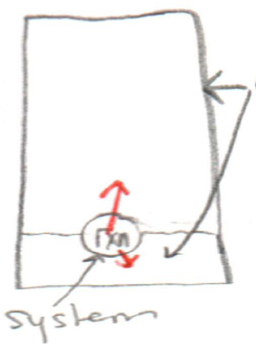
$$= -\left(4.184\ \frac{J}{g\,K}\right)\left(1\ L \times \frac{1000\ mL}{1\ L} \times \frac{1\ g}{mL}\right)(58.92 - 25.20)$$

$$= -141084\ J = -141.08\ kJ$$

So, 141.08 kJ of heat is released by the combustion of 0.0935 moles of ethane.

- $\Delta H_{comb} = -141.08\ kJ \times \frac{1\ mol}{0.0935\ mol} = \boxed{-1508.9\ kJ}$ since ΔH_{comb} is for 1 mole of fuel

This picture is the same for #5 & #6



6. 1.14 g of octane (C_8H_{18}) is combusted in a bomb calorimeter surrounded by 1L of water. The initial and final temperatures of the water are 25°C and 33°C respectively. The heat capacity of the calorimeter hardware (all of the calorimeter except for the water) is 456 J/(°C). The chemical equation is $2\ C_8H_{18} + 25\ O_2 = 16\ CO_2 + 18\ H_2O$. Determine the molar enthalpy of combustion of octane.

- fuel = C_8H_{18}
- 1.14 g octane is $\frac{1.14\ g}{(12\times8)+(1\times18)\ g/mol} = 0.01\ mol$
- $q_{rxn} = -q_{cal} = -C_{cal}\Delta T = -(C_{H_2O} + C_{hardware})\Delta T = -(C_{s,H_2O} \times M_{H_2O})(C_{hardware})\Delta T$

$$= -\left[\left(4.184\ \frac{J}{g\,K}\right)(1000\ g) + \left(456\ \frac{J}{°C}\right)\right](33 - 25\ °C)$$

$$= -(4640\ J)(8) = 37120\ J = -37.12\ kJ$$

So, -37.12 kJ of heat is released by the combustion of 0.01 moles of octane.

- $\Delta H_{combustion} = (-37.12\ kJ)\left(\frac{1\ mol}{0.01\ mol}\right) = \boxed{-3712\ kJ}$, b/c $\Delta H_{combustion}$ is reported for 1 mole of fuel