CHEM1109 Worksheet 1 – Answers to Critical Thinking Questions

The worksheets are available in the tutorials and form an integral part of the learning outcomes and experience for this unit.

Model 1: Calorimetry

1. When $\Delta T$ is negative: when the temperature lowers.
2. $C = c / M$ or $c = C \times M$ where $M$ is the molar mass.
3. 10. K (2 significant figures).
4. No. The temperature difference is the same in both units.
5. 420 J
6. Heating up water by the same amount as olive oil requires more energy.
7. It would take 0.31 J to heat up if pure. The necklace is not pure.

Model 2: Energy

1. (a) $\Delta H < 0$ (i.e. negative)    (b) $\Delta H > 0$ (i.e. positive)

2. (a) exothermic         (b) endothermic

   “Thermic” = caused by heat.

3. Colder.
4. Negative.
5. Stronger in the products than in the reactants.
• A mass of 1.250 g of benzoic acid (C\textsubscript{7}H\textsubscript{6}O\textsubscript{2}) underwent combustion in a bomb calorimeter. If the heat capacity of the calorimeter was 10.134 kJ K\textsuperscript{–1} and the heat of combustion of benzoic acid is –3226 kJ mol\textsuperscript{–1}, what is the change in internal energy during this reaction?

The molar mass of benzoic acid is:

\[
(7 \times 12.01 \text{ (C)} + 6 \times 1.008 \text{ (H)} + 2 \times 16.00 \text{ (O)}) \text{ g mol}^{-1} = 122.1 \text{ g mol}^{-1}
\]

A mass of 1.250 g therefore corresponds to:

\[
\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{1.250 \text{ g}}{122.1 \text{ g mol}^{-1}} = 0.0102 \text{ mol}
\]

As 3226 kJ are released per mole, the change in internal change for this amount is:

\[
\Delta U = (-3226 \text{ kJ mol}^{-1}) \times (0.0102 \text{ mol}) = –33.02 \text{ kJ}
\]

Answer: –33.02 kJ

Calculate the temperature change that should have occurred in the apparatus.

In a constant volume apparatus like a calorimeter, the change in internal energy is equal to the heat change, \(q\textsubscript{V}\). Using \(q = C\textsubscript{p}\Delta T\), the temperature change is:

\[
\Delta T = \frac{(33.02 \text{ kJ})}{(10.134 \text{ kJ K}^{-1})} = 3.528 \text{ K}
\]

As the combustion reaction evolves heat, the temperature increases.

Answer: +3.528 K

• The specific heat capacity of water is 4.18 J g\textsuperscript{–1} K\textsuperscript{–1} and the specific heat capacity of copper is 0.39 J g\textsuperscript{–1} K\textsuperscript{–1}. If the same amount of energy were applied to a 1.0 mol sample of each substance, both initially at 25 °C, which substance would get hotter? Show all working.

Using \(q = C \times m \times \Delta T\), the temperature change for a substance of mass \(m\) and specific heat capacity \(C\) when an amount of heat equal to \(q\) is supplied is given by:

\[
\Delta T = \frac{q}{C \times m}
\]

The atomic mass of copper is 63.55. Hence, the temperature change for 1.0 mol of copper is

\[
\Delta T \text{ (copper)} = \frac{q}{(0.39 \times 63.55)} = \frac{q}{24.8} \text{ °C}
\]

The molar mass of H\textsubscript{2}O is (2 \times 1.008 \text{ (H)}) + 16.00 \text{ (O)} = 18.016. Hence, the temperature change for 1.0 mol of water is

\[
\Delta T \text{ (water)} = \frac{q}{(4.18 \times 18.016)} = \frac{q}{75.3} \text{ °C}
\]

Hence,

\[
\Delta T \text{ (copper)} > \Delta T \text{ (water)}
\]

Answer: copper
A 150.0 g block of iron metal is cooled by placing it in an insulated container with a 50.0 g block of ice at 0.0 °C. The ice melts, and when the system comes to equilibrium the temperature of the water is 78.0 °C. What was the original temperature (in °C) of the iron?

Data: The specific heat capacity of liquid water is 4.184 J K$^{-1}$ g$^{-1}$.
The specific heat capacity of solid iron is 0.450 J K$^{-1}$ g$^{-1}$.
The molar enthalpy of fusion of ice (water) is 6.007 kJ mol$^{-1}$.

The heat from the iron is used to melt the ice and to warm the water from 0.0 °C to 78.0 °C.

The molar mass of H$_2$O is (2 × 1.008 (H) + 16.00 (O)) g mol$^{-1}$ = 18.02 g mol$^{-1}$.
Hence 50.0 g of ice corresponds to:

number of moles = mass / molar mass = (50.0 g) / (18.02 g mol$^{-1}$) = 2.775 mol.

Hence the heat used to melt ice is:

\[ q_1 = 6.007 \text{ kJ mol}^{-1} \times 2.775 \text{ mol} = 16.67 \text{ kJ} = 16670 \text{ J} \]

The heat used to warm 50.0 g water by 78.0 °C is:

\[ q_2 = m \times C \times \Delta T = (50.0 \text{ g}) \times (4.184 \text{ J K}^{-1} \text{ g}^{-1}) \times (78.0 \text{ K}) = 16320 \text{ J} \]

Overall, the heat transferred from the iron is:

\[ q = q_1 + q_2 = 16670 \text{ J} + 16320 \text{ J} = 32990 \text{ J} \]

This heat is lost from 150.0 g of iron leading to it cooling by \( \Delta T \):

\[ q = m \times C \times \Delta T = (150.0 \text{ g}) \times (0.450 \text{ J K}^{-1} \text{ g}^{-1}) \times \Delta T = 32990 \text{ J} \]

\[ \Delta T = 489 \text{ K} = 489 \text{ °C} \]

As the final temperature of the iron is 78.0 °C, its original temperature was (78.0 + 489) °C = 567 °C.

Answer: 567 °C