CHEM1612 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 ΔT is negative: i.e., when the temperature lowers.

2. $C = c \times M$ or $c = C / M$ where M is the molar mass.

3. $\Delta T = 10.$ K (2 significant figures).

4. No. The temperature *difference* is the same in both units.

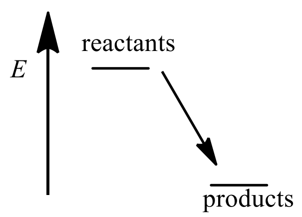
5. 420 J

6. Heating up water by the same amount as olive oil requires more energy.

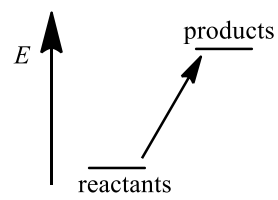
7. It would take 0.31 J to heat up if pure. The necklace is not pure.

Model 2: Energy

1. (a)



(b)



2. (a) exothermic (b) endothermic

“Thermic” = caused by heat.

3. Colder.

4. Negative.

5. Stronger in the products than in the reactants.

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

Marks 4

- A mass of 1.250 g of benzoic acid ($C_7H_6O_2$) underwent combustion in a bomb calorimeter. If the heat capacity of the calorimeter was 10.134 kJ K^{-1} and the heat of combustion of benzoic acid is –3226 kJ mol^{-1}, what is the change in internal energy during this reaction?

The molar mass of benzoic acid is:

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

A mass of 1.250 g therefore corresponds to:

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

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

$$\Delta U = (-3226\ kJ\ mol^{-1}) \times (0.0102\ mol) = -33.02\ kJ$$

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_V. Using $q = C_p\Delta T$, the temperature change is:

$$\Delta T = (33.02\ kJ) / (10.134\ kJ\ K^{-1}) = 3.258\ K$$

As the combustion reaction evolves heat, the temperature *increases*.

Answer: **+3.258 K**

Marks 2

- The specific heat capacity of water is 4.18 J g^{-1} K^{-1} and the specific heat capacity of copper is 0.39 J g^{-1} K^{-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 × m × ΔT, 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

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

The molar mass of H_2O is (2 × 1.008 (H)) + 16.00 (O) = 18.016. Hence, the temperature change for 1.0 mol of water is

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

Hence,

$$\Delta T\ (\text{copper}) > \Delta T\ (\text{water})$$

Answer: **copper**

Marks 4

- 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_2O 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 kJ mol^{-1} × 2.775 mol = 16.67 kJ = 16670 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 g) × (4.184 J K^{-1} g^{-1}) × (78.0 K) = 16320 J**

Overall, the heat transferred from the iron is:

$q = q_1 + q_2$ **= 16670 J + 16320 J = 32990 J**

This heat is lost from 150.0 g of iron leading to it cooling by ΔT:

$q = m \times C \times \Delta T$ **= (150.0 g) × (0.450 J K^{-1} g^{-1}) × ΔT = 32990 J**

ΔT **= 489 K = 489 °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**

Key to success: practice further by completing this week's tutorial homework

Key to even greater success: practice even further by completing this week's suggested exam questions