# Questions

# 7.2 Homework // Name: ________________________

## Multiple Choice

### Question 1:

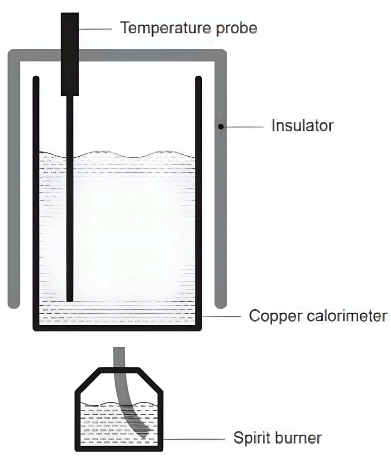
The energy from burning 0.250 g of ethanol causes the temperature of 150 $cm^3$ of water to rise by 10.5°C. What is the enthalpy of combustion of ethanol, in kJ $mol^{-1}$?

Specific heat capacity of water: 4.18 J $g^{-1}$ $K^{-1}$

A. $1.21 \times 10^6$
B. $1.21 \times 10^3$
C. $3.28 \times 10^7$
D. $3.28 \times 10^4$

### Question 2:

The enthalpy of combustion of a fuel was determined using the calorimeter shown. The final result was lower than the literature value.




Which factors could have contributed to this error?

I. Not all heat from the combustion was transferred to the calorimeter.
II. Incomplete combustion occurred.
III. The temperature probe touched the bottom of the calorimeter.

A. I and II only
B. I and III only
C. II and III only
D. I, II, and III

## Question 3:

A student obtained the following data to calculate q, using $q = mc\Delta T$.

$m = 20.2\ g \pm 0.2\ g$
$\Delta T = 10°C \pm 0.1°C$
$c = 4.18\ J\ g^{-1}\ K^{-1}$

What is the percentage uncertainty in the calculated value of q?

A. 0.2
B. 1.2
C. 11
D. 14

# Short Answer

## Question 4:

(a) Calculate the enthalpy of neutralization based on the following data. [4]

| | |
|---|---|
| Initial temperature of solutions / °C | 24.5 |
| Concentration of KOH (aq) / mol $dm^{-3}$ | 0.950 |
| Concentration of HCl (aq) / mol $dm^{-3}$ | 1.050 |
| Volume of HCl (aq) / $cm^3$ | 50.00 |
| Volume of KOH (aq) / $cm^3$ | 50.00 |
| Final temperature of mixture / °C | 30.3 |

(b) State the assumptions you have made in your calculation. [1]

## Question 5:

Alkanes undergo combustion and substitution.

Determine the molar enthalpy of combustion of an alkane if $8.75 \times 10^{-4}$ mols are burned, raising the temperature of 20.0 g of water by 57.3°C. [2]

Key

# Multiple Choice

## Question 1:

B

## Question 2:

A

## Question 3:

C

# Short Answer

## Question 4:

(a) $\Delta T = 30.3 - 24.5 = 5.8$ K ✔

$q = m(H_2O) \times c(H_2O) \times \Delta T(H_2O)$
$= 100.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times 5.8 \text{ K}$
$= 2424.4$ J ✔

KOH is the limiting reagent
$n(KOH) = (50 \times 0.950) / 1000 = 0.0475$ mol ✔

$\Delta H = -2424.4 / 0.0475 = -51040 \text{ J mol}^{-1}$
$= -51.04 \text{ kJ mol}^{-1}$ ✔

(b) Assumptions: no heat loss, c(solution) = c(water), m(solution) = $m(H_2O)$, density($H_2O$) = 1.00 ✔

## Question 5:

$q = mc\Delta T = 20.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ C}^{-1} \times 57.3 \text{ C} = 4790$ J ✔
$\Delta H = (4790/1000) / (8.75 \times 10^{-4} \text{ mol}) = -5470 \text{ kJ mol}^{-1}$ ✔

Award [2] for correct final answer.
Accept answers in the range -5470 to -5480 kJ $mol^{-1}$

# Optional Practice

# Multiple Choice

## Question 1:

In an experiment to determine the enthalpy of combustion of an alcohol, the mass of a burner plus its contents, and the temperature of a known mass of water in a calorimeter are measured before and after the experiment. What are the expected results?

| **Mass of burner and contents** | **Reading on thermometer** |
|---|---|
| A. Decreases | Increases |
| B. Decreases | Stays the same |
| C. Increases | Increases |
| D. Increases | Stays the same |

## Question 2:

A copper calorimeter was used to determine the enthalpy of combustion of butan-1-ol. The experimental value obtained was -2100 ± 200 kJ $mol^{-1}$, and the data booklet value is -2676 kJ $mol^{-1}$.

Which of the following accounts for the difference between the two values?

I. random measurement errors
II. incomplete combustion
III. heat loss to the surroundings

A. I and II only
B. I and III only
C. II and III only
D. I, II, and III

## Question 3:

The thermochemical equation for the combustion of propane is:

$C_3H_8\,(g) + 5O_2\,(g) \rightarrow 3CO_2\,(g) + 4H_2O\,(l)$ $\Delta H = -2219\ kJ\ mol^{-1}$

Identify the true statements.

I. If 1.00 g of propane is burnt completely, 2219 kJ of energy are produced.
II. The reactants have more potential energy than the products.
III. The reaction is exothermic.

A. I and II only
B. I and III only
C. II and III only
D. I, II and III

## Question 4:

20.0 $cm^3$ of 2.0 mol $dm^{-3}$ $HNO_3$ (aq) is mixed with 40 $cm^3$ of 1.0 mol $dm^{-3}$ KOH (aq) at 25.0°C. The temperature of the resulting solution increases by 9°C.

Predict the temperature change when 5.0 $cm^3$ of 2.0 mol $dm^{-3}$ $HNO_3$ (aq) is mixed with 10.0 $cm^3$ of 1.0 mol $dm^{-3}$ KOH (aq) at the same temperature.

A. 4.5°C
B. 9°C
C. 18°C
D. 27°C

# Short Answer

## Question 1:

A piece of brass is held in the flame of a Bunsen burner for several minutes. The brass is then quickly transferred into an aluminum calorimeter which contains 200.0 g of water.

Determine the temperature of the Bunsen flame from the following data.

| | |
|---|---|
| m(water) / ± 0.01 g | 200.00 |
| m(brass) / ± 0.01 g | 212.10 |
| m(aluminum calorimeter) / ± 0.01 g | 80.00 |
| c(brass) / J $g^{-1}$ $K^{-1}$ | 0.400 |
| c(aluminum) / J $g^{-1}$ $K^{-1}$ | 0.900 |
| Initial temperature of water / ± 0.1°C | 24.5 |
| Final temperature of water / ± 0.1°C | 77.5 |

## Question 2:

1.10 g of glucose was completely burnt and the heat produced increased the temperature of the water in a copper calorimeter from 25.85°C to 36.50°C.

(a) Calculate the enthalpy of combustion of glucose from the data below.

| | |
|---|---|
| Mass of water / g | 200.00 |
| Specific heat capacity of water / J $g^{-1}$ $K^{-1}$ | 4.18 |
| Mass of copper / g | 120.00 |
| Specific heat capacity of copper / J $g^{-1}$ $K^{-1}$ | 0.385 |

(b) Draw an enthalpy level diagram to represent this reaction.

## Question 3:

A student added 5.35 g of ammonium chloride to 100.0 $cm^3$ of water. The initial temperature of the water was 25.55°C but it decreased to 21.79°C. Calculate the enthalpy change that would occur when 1 mol of the solute is added to 1.000 $dm^3$ of water. [3]

## Question 4:

Calculate the enthalpy change, in kJ $mol^{-1}$, when 1.00 mol is burnt in each case (the molar volume of a gas at STP is 22.7 $dm^3$ $mol^{-1}$):

(a) 105,000 J is given out when 0.100 mol of A is burnt.
(b) 84,000 J is given out when 0.042 mol of B is burnt.
(c) 11,000 J is given out when 0.500 g of $CH_3OH$ is burnt.
(d) 6,000 J is given out when 0.150 g of $C_6H_6$ is burnt.
(e) 13,000 J is given out when 200 $cm^3$ (measured at STP) of $C_2H_6$ (g) is burnt.

## Question 5:

(a) When 1.20 g of hexane ($C_6H_{14}$) is burnt, the temperature of 250.0 g of water is raised by 56.0°C. Calculate the enthalpy change when one mole of hexane is burnt.
(b) When 0.870 g of pentan-1-ol ($C_5H_{11}OH$) is burnt, the temperature of 180.0 g of water is raised by 38.0°C. Calculate the enthalpy change when one mole of pentan-1-ol is burnt.
(c) When 0.521 g of benzene ($C_6H_6$) is burnt, the temperature of 320.0 g of water is raised by 12.2°C. Calculate the enthalpy change when one mole of benzene is burnt.
(d) When 2.00 kg of octane ($C_8H_{18}$) is burnt, the temperature of 500 kg of water is raised by 46.0°C. Calculate the enthalpy change when one mole of octane is burnt.

## Question 6:

A student conducts an experiment to determine the enthalpy change of combustion of propan-1-ol ($C_3H_7OH$)

(a) Determine the enthalpy change of combustion of propan-1-ol using the student's experimental data:
Mass of water = 200.0 g
Initial temperature of water = 18.2°C
Maximum temperature of water = 38.6°C
Initial mass of spirit burner = 185.51 g
Final mass of spirit burner = 184.56 g
(b) The actual value of the enthalpy change when one mole of propan-1-ol is burnt is -2010 kJ $mol^{-1}$. Account for any differences between this value and the one calculated from the experimental data in part (a).

# Optional Practice Key

## Multiple Choice

### Question 1:

A

### Question 2:

C

### Question 3:

C

### Question 4:

B

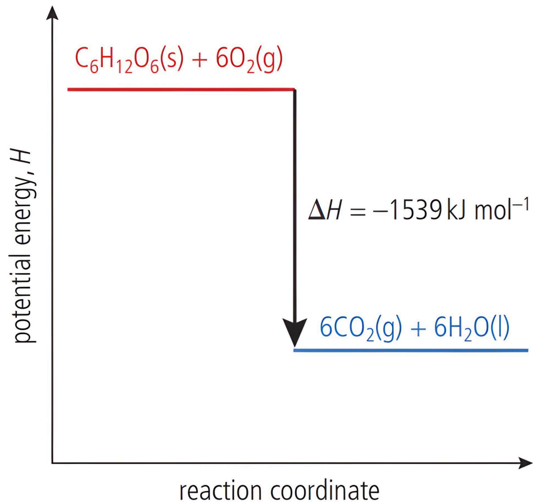
## Short Answer

### Question 1:

q = mcΔT
Temperature change (water and aluminum) = 77.5 - 24.5 = 53.0 ± 0.2 K
Energy lost by brass = energy gained by water and calorimeter
Energy gained by water and calorimeter = (200.00 x 4.18 x 53.0) + (80.00 x 0.900 x 53.0) J = 44,308 + 3,816 = 48,124 J
Energy lost by brass = 48,124 J
212.10 x 0.400 x ($T_{Bunsen}$ - 77.5) = 48,124 J
$T_{Bunsen}$ - 77.5 = 48,124 / (212.10 x 0.400) = 567.2°C
Temperature of brass in Bunsen flame = 567.2 + 77.5 = 644.7°C

### Question 2:

(a) ΔT = 36.50 - 25.85 - 10.65°C
q = mcΔT = [m($H_2O$) x c($H_2O$) x ΔT($H_2O$)] + [m(Cu) x c(Cu) x ΔT(Cu)] = (200.00 g x 4.18 J $g^{-1}$ $K^{-1}$ x 10.65 K) + (120.00 g x 0.385 J $g^{-1}$ $K^{-1}$ x 10.65 K) = 8903.4 + 492.0 J = 9395.4 J
n($C_6H_{12}O_6$) = 1.10 g / 180.18 g $mol^{-1}$ = 6.11 x $10^{-3}$ mol
$\Delta H_c$ = 9395.4 J / 6.11 x $10^{-3}$ mol = -1539 x $10^3$ J $mol^{-1}$ = -1539 kJ $mol^{-1}$




(b)

## Question 3:

$\Delta H_{reaction} = -\Delta H_{water}$
$q = mc\Delta T$
$\Delta H_{water} = 100.0\ g \times 4.18\ J\ g^{-1}\ K^{-1} \times (21.79 - 25.55)\ K = -1571.68\ J$ ✔

$n(NH_4Cl) = 5.35\ g / 53.50\ g\ mol^{-1} = 0.100\ mol$ ✔

$\Delta H = -\Delta H_{water} / n(NH_4Cl) = -(-1571.68\ J) / 0.100\ mol = +15716.8\ J\ mol^{-1} = +15.7\ kJ\ mol^{-1}$ ✔

## Question 4:

$\Delta H = -q/n$
Divide by 1,000 to convert to kJ

(a) -105,000 / (0.100 x 1000) = -1,050 kJ $mol^{-1}$
(b) -84,000 / (0.042 x 1000) = -2,000 kJ $mol^{-1}$
(c) Convert mass to amount in mol by dividing by molar mass.
-11,000 / ((0.500 / 32.05) x 1000) = -705 kJ $mol^{-1}$
(d) -6,000 / ((0.150 / 78.12) x 1000) = -3,120 kJ $mol^{-1}$
(e) Convert volume to $dm^3$ before dividing by the molar volume (22.7 $dm^3\ mol^{-1}$) to give the amount in mol.
-13,000 / ((0.200 / 22.7) x 1000) = -1,480 kJ $mol^{-1}$

## Question 5:

(a) q = 250.0 x 4.18 x 56.0 = 58,520 J
n = 1.20 / 86.20 = 0.0139 mol
$\Delta H$ = - 58,520 / 0.0139 = -4200000 J $mol^{-1}$ or -4200 kJ $mol^{-1}$
(b) q = 180.0 x 4.18 x 38.0 = 28,591 J
n = 0.870 / 88.17 = $9.867 \times 10^{-3}$ mol

$\Delta H = -28{,}591 / (9.867 \times 10^{-3}) = -2900000 J\ mol^{-1}$ or $-2900\ kJ\ mol^{-1}$

(c) $q = 320.0 \times 4.18 \times 12.2 = 16{,}329$ J
$n = 0.521 / 78.12 = 6.669 \times 10^{-3}$ mol
$\Delta H = -\ 16.329 / (6.669 \times 10^{-3}\ mol) = -2450000\ J\ mol^{-1}$ or $-2450\ kJ\ mol^{-1}$

(d) $q = 500{,}000 \times 4.18 \times 46.0 = 96{,}140{,}000$ J
$n = 2{,}000 / 114.26 = 17.50$ mol
$\Delta H = -5490000\ J\ mol^{-1}$ or $-5490\ kJ\ mol^{-1}$

## Question 6:

(a) $q = 200.0 \times 4.18 \times (38.6 - 18.2) = 17{,}054$ J
$n = (185.51 - 184.56) / 60.11 = 1.580 \times 10^{-2}$ mol
$\Delta H = -17{,}054 / (1.580 \times 10^{-2}) = -1080000\ J\ mol^{-1}$ or $-1080\ kJ\ mol^{-1}$

(b) Heat energy loss to the surroundings; incomplete combustion; (other, more minor, factors include evaporation of water and/or propan-1-ol)