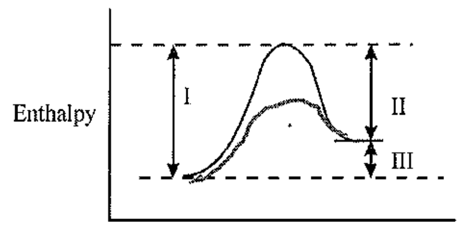
5.1 measuring energy changes key

1. Which statements about exothermic reactions are correct?
   - I. They have negative ΔH values.
   - II. The products have a lower enthalpy than the reactants.
   - III. The products are more energetically stable than the reactants.

   a. I and II only
   b. I and III only
   c. II and III only
   d. I, II and III

2. Which of the quantities in the enthalpy level diagram below is (are) affected by the use of a catalyst?
   a. I only
   b. III only
   c. I and II only
   d. II and III only



3. Which statements are correct for an endothermic reaction?
   - I. The system absorbs heat.
   - II. The enthalpy change is positive.
   - III. The bond enthalpy total for the reactants is greater than for the products.

   a. I and II only
   b. I and III only
   c. II and III only
   d. I, II and II

s → l ↑ disorder

4. When a sample of NH_4SCN is mixed with solid $Ba(OH)_2$*$8H_2O$ in a glass beaker, the mixture changes to a liquid and the temperature drops sufficiently to freeze the beaker to the table. Which statement is true about the reaction?

   Δ to surroundings

   a. The process is endothermic and ΔH is –
   b. The process is endothermic and ΔH is +
   c. The process is exothermic and ΔH is –
   d. The process is exothermic and ΔH is +

5. Which one of the following statements is *true* of all exothermic reactions?
   a. They produce gases
   b. They give out heat
   c. They occur quickly
   d. They involve combustion

6. If 500J of heat is added to 100.0g samples of each of the substances below, which will have the largest temperature increase?
   a. **Gold – 0.126 $Jg^{-1}K^{-1}$**
   b. Silver – 0.237 $Jg^{-1}K^{-1}$
   c. Copper – 0.385 $Jg^{-1}K^{-1}$
   d. Water – 4.18 $Jg^{-1}K^{-1}$

Smallest value, largest ↑, most easily affected

7. The specific heat of metallic mercury is 0.138 $Jg^{-1}°C^{-1}$. If 100.0 J of heat is added to a 100.0g sample of mercury at 25.0°C, what is the final temperature of mercury?

$100.0 = (100.0)(0.138)(x - 25.0)$

$100.0 = 13.8(x - 25)$

$100.0 = 13.8x - 345$

$445.0 = 13.8x$

$32.2°C = x$

8. The temperature of a 2.0g sample of aluminum increases from 25°C to 30°C. How many joules of heat energy were added? (specific heat of Al = 0.90 $Jg^{-1}K^{-1}$)
   a. 0.36
   b. 2.3
   c. **9.0**
   d. 11

9. A sample of a metal is heated. Which of the following are needed to calculate the heat absorbed by the sample?
   I. The mass of the sample – m
   II. The density of the sample – not needed but could be used to find vol or mass
   III. The specific heat capacity of the sample – c
   a. I and II only
   b. **I and III only**
   c. II and III only
   d. I, II and III

10. What is the energy change (in kJ) when the temperature of 20 g of water increases by 10°C?
   a. $20 \times 10 \times 4.18$
   b. $20 \times 283 \times 4.18$
   c. $\frac{20 \times 10 \times 4.18}{1000}$ ← ÷ 1000 to convert to KJ
   d. $\frac{20 \times 283 \times 4.18}{1000}$

11. In aqueous solution, potassium hydroxide and hydrochloric acid react as follows.

$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

The data below is from an experiment to determine the enthalpy change of this reaction.

50.0 cm^3 of a 0.500 mol dm^{-3} solution of KOH was mixed rapidly in a glass beaker with 50.0 cm^3 of a 0.500 mol dm^{-3} solution of HCl.

Initial temperature of each solution = 19.6°C
Final temperature of the mixture = 23.1°C

(a) State, with a reason, whether the reaction is exothermic or endothermic.

Exothermic, heat is released into water

(b) Explain why the solutions were mixed rapidly.

Minimize heat loss, quick heat release

(c) Calculate the enthalpy change of this reaction in kJ mol^{-1}. Assume that the specific heat capacity of the solution is the same as that of water.

$\Delta H = -(mc\Delta T)$

$\Delta H = -(100)(4.18)(23.1-19.6)$

$= -1463J \rightarrow \frac{1.463kJ}{0.025 mols}$

$\boxed{= -58.5 kJ mol^{-1}}$

Conc. = $0.5 = \frac{n}{0.05}$

$= 0.025$ mols

(d) Identify the **major** source of error in the experimental procedure described above. Explain how it could be minimized.

Heat loss to surroundings → insulate apparatus, cover with lid

(e) The experiment was repeated but with an HCl concentration of 0.510 mol dm^{-3} instead of 0.500 mol dm^{-3}. State and explain what the temperature change would be.

3.5°C (same) → amount of base is the same (becomes the limiting reactant) – 1:1 ratio

12. The mass of the burner and its content is measured before and after the experiment. The thermometer is read before and after the experiment. What are the expected results?

a. **Mass: decreases; Thermometer: increases**
b. Mass: decreases; Thermometer: stays the same
c. Mass: increases; Thermometer: increases
d. Mass: increases; Thermometer: stays the same

13. The heat released from the combustion of 0.0500g of white phosphorous increases the temperature of 150.00g of water from 25.0°C to 31.5°C. Calculate a value for the enthalpy change of combustion of phosphorous. Discuss possible sources of error in the experiment.

$m = \frac{0.05g}{30.97g} = 0.0016$ mols P

$\Delta H_{rxn} = -(150.00)(4.18)(31.5-25.0)$

$\Delta H_{rxn} = -4075.5J$

$\frac{-4.0755 kJ}{0.0016 mols}$

(sig fig) $\boxed{-2506 kJmol^{-1}}$

* heat loss
* incomplete combustion

14. When 8.00g of ammonium nitrate completely dissolved in 100 cm^3 of water, the temperature fell from 19.0°C to 14.5°C. Calculate the enthalpy change of solution.

$m: \frac{8.00g\ NH_4NO_3}{80.04g} = 0.099$ mols

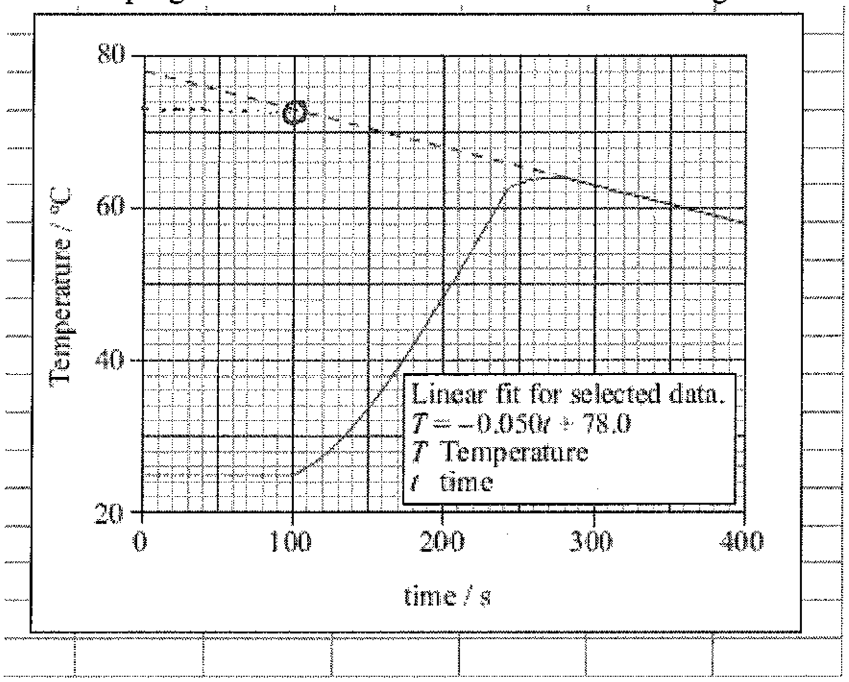
$\Delta H_{rxn} = -(100)(4.18)(14.5-19.0)$

$= 1881.9J \rightarrow \frac{1.8819 kJ}{0.099 mols} \rightarrow \boxed{+18.8\ kJmol^{-1}}$

15. The data below are from an experiment to measure the enthalpy change for the reaction of aqueous copper(II) sulfate, $CuSO_4(aq)$ and zinc, Zn(s).

$$Cu^{2+}(aq) + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq)$$

50.0 cm^3 of 1.00 mol dm^{-3} copper(II) sulfate solution was placed in a polystyrene cup and zinc powder was added after 100 seconds. The temperature-time data was taken from a data-logging software program. The table shows the initial 23 readings.



A straight line has been drawn through some of the data points. The equation for this line is given by the data logging software as

$T = -0.050t + 78.0$ where T is the Temperature at time t.

(a) The heat produced by the reaction can be calculated from the temperature change, ΔT, using the expression below.

Heat change = Volume of $CuSO_4(aq)$ × Specific heat capacity of H_2O × ΔT

Describe two assumptions made in using this expression to calculate heat change.

All heat is transferred to water/copper sulfate solution/no heat loss
Specificity heat capacity of Zn is negligible
Density of water/solution = 1.0
Heat capacity of the cup is negligible
Specific heat capacity of solution = specific heat capacity of water
Temperature is uniform throughout solution

(2)

(b) (i) Use the data presented by the data logging software to deduce the temperature change, ΔT, which would have occurred if the reaction had taken place instantaneously with no heat loss.

T_{final} = 72-74°C

ΔT = 73 – 25 = 48°C

(2)

(ii) State the assumption made in part (b)(i).

Temperature decreases at uniform rate (when above room temperature)
(1)

(iii) Calculate the heat, in kJ, produced during the reaction using the expression given in part (a).

H = (0.050)(4.18)(48) = 10.0 kJ

(1)

(c) The colour of the solution changed from blue to colourless. Deduce the amount, in moles, of zinc which reacted in the polystyrene cup.

$n_{Zn} = n_{CuSO_4}$ $\quad 1.00 = \frac{n}{0.05}$ $\quad n = 0.050\,mol$

(1)

(d) Calculate the enthalpy change, in kJ mol^{-1}, for this reaction.

$-\frac{10.0\ kJ}{0.05\,mol} = -201\ kJ\ mol^{-1}$

(1)