

5.2 Calculation of Enthalpy Changes

During a chemical reaction or physical change heat energy is either transferred into the surroundings from the chemicals or absorbed from the surrounding.

The amount of heat energy (Q) transferred depends on three factors:

1. The capacity of the substance to absorb heat – called the specific heat capacity (c).
2. The amount (moles) of the substances reacting
3. The change in temperature

The amount of heat transferred is proportional to each of these three factors. $T \propto$ moles and $Q \propto$ moles. So if the number of moles increases the temperature increases and the amount of heat energy transferred will be greater.

<http://www.chm.davidson.edu/ChemistryApplets/calorimetry/HeatOfSolutionOfAmmoniumNitrate.html> This simulation shows the effect of the mass of ammonium nitrate on the temperature change when it is dissolved in water.

To calculate the energy / enthalpy / heat energy / heat absorbed or released the following formula is used:

$$Q = cm\Delta T$$

heat added specific heat mass change in temperature

Where:

Mass of the substance (m)

The substance experiencing the temperature change. It is measured in grams (g).

- In an aqueous solution, m represents the total mass of aqueous solution. We make the assumption that 100 ml of solution = 100 g.
- In reactions involving a solid reactant which is often in excess, we assume m is the mass of the aqueous reactant. 100 ml of solution = 100 g
- For combustion reactions, m refers to the mass of water that the heat energy is transferred into. 100 ml water = 100 g water.
- For metals, m represents the mass of the metal.

Specific heat capacity (c)

Each substance absorbs heat to different extents. The specific heat is defined as the amount of heat (per unit mass of the substance) energy required to raise 1 gram of the substance by 1°C. It is measured in $\text{J g}^{-1}\text{K}^{-1}$ (= $\text{kJ kg}^{-1}\text{K}^{-1}$) The larger the specific heat capacity the more heat energy 1 g of the substance needs to absorb to cause a 1°C temperature rise.

Water has a high heat capacity ($4.18 \text{ J g}^{-1} \text{ K}^{-1}$ / $\text{kJ kg}^{-1}\text{K}^{-1}$) compared to other substances like air (1.00

$\text{J g}^{-1} \text{K}^{-1}$). This is an important physical property of water and means that water can absorb (and release) large amounts of heat energy with relatively small temperature changes. In the summer for example the air temperature can be 35°C but the water in a swimming pool is only 23°C . This is because when the water absorbs heat energy from the sun, its higher heat capacity means it will change the temperature of the water less than the air.

In most calculations involving aqueous solutions we make the assumption that the specific heat capacity of the solution is equal to the specific heat capacity of water. Remember an assumption is something you assume to be true but you are not sure. A lot of assumptions are made in science because they simplify models.

Energy / enthalpy / heat energy (q/Q)

Measured in Joules (J) or kilojoules (kJ). If the reaction is exothermic Q is given a negative sign and if endothermic a positive sign. Express in KJ unless J is specifically asked for.

Temperature

Measured in Kelvin (K). $0^\circ\text{C} = 273 \text{ K}$. The change in temperature, ΔT has the same numerical value on both the Celsius and Kelvin scales, so no conversion is needed.

Unit Conversions

The units given for c must be the same as those used for Q, m and ΔT . E.g. if $c = 4.18 \text{ J g}^{-1} \text{K}^{-1}$ then q must be in Joules, m in grams and ΔT in K. Make conversions where necessary. The specific heat capacity of water is found on page 2 of the Chemistry data booklet.

$$\text{g} \div 1000 = \text{kg}$$

$$\text{kg} \times 1000 = \text{g}$$

$$\text{J} \div 1000 = \text{kJ}$$

$$\text{kJ} \times 1000 = \text{J}$$

$$0^\circ\text{C} + 273 = 273 \text{ K (Kelvin)}$$

$$273 \text{ K} - 273 = 0^\circ\text{C}$$

Worked example:

If you drink a cold glass of water (250 g) at 0.0°C , how much heat is transferred to the water as it warms to 37.0°C . The specific heat capacity of water is $4.18 \text{ J g}^{-1} \text{K}^{-1}$.

Answer:

$$m = 250 \text{ g of water (3SF, measured value } \pm 1\text{g)}$$

$$c = 4.18 \text{ J g}^{-1} \text{K}^{-1}$$

$$\Delta T = 37.0 - 0.0 = 37.0^\circ\text{C} (37.0 \text{ K})$$

$$Q = m c \Delta T$$

$$= (250 \text{ g}) (4.18 \text{ J g}^{-1} \text{K}^{-1}) (37.0 \text{ K})$$

$$= 3.87 \times 10^4 \text{ J}$$

$$= 38.7 \text{ kJ (3SF)}$$

Enthalpy change, ΔH

The enthalpy change is a measure of the difference in the heat energy of the products and reactants. It is measured in kJmol^{-1} (unless Jmol^{-1} is specifically asked for).

$$\text{enthalpy change, } \Delta H = \frac{m \times c \times \Delta T}{n \text{ (limiting reagent)}}$$

or

$$\Delta H = \frac{Q}{n}$$

Where n is the actual number of moles of the limiting reagent in the chemical reaction:

$$n = \frac{m}{M} \quad \text{or} \quad c = \frac{n}{v}$$

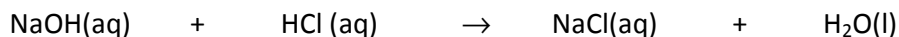
Remember to determine the limiting reagent

1. Find the actual number of moles of each reactant.
2. Find the number of moles needed (required) of each reactant from the moles ratio in the equation.
3. If the moles of reactant you have $<$ moles needed then this is the limiting reactant

Problems

1. When 1.00g of sodium hydroxide is dissolved in 100.0 g of water in an insulated container the temperature rises from 20.00°C to 22.66°C . Calculate the enthalpy change for the solution. State any assumptions made. NOTE: Dissolving reactions due not have limiting reagents.

2. When 100.0 ml of 1.00 mol dm⁻³ NaOH is added to 100.0 ml of 1.00 mol dm⁻³ HCl in an insulated container the temperature rises from 21.0°C to 34.6°C.



Calculate the enthalpy change for the neutralization reaction. State any assumptions made.

The enthalpy change can also be determined using the mole ratios in the equation. For example in the following reaction:

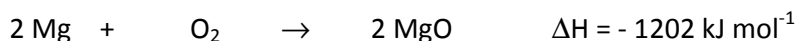


The reaction says:

2 moles Mg reacts with 1 mole O₂ to produce 2 mole of MgO 1202 kJ mol⁻¹ of energy is released or ΔH = -1202 kJ mol⁻¹

Using mole ratio's:

1 mole Mg reacts with ½ mole O₂ to produce 1mole of MgO -1202 ÷ 2 = 601 kJ mol⁻¹ of energy is released or ΔH = - 601 kJ mol⁻¹



mole ratios: 2 1 2 ΔH = -1202 kJ mol⁻¹

or 1 ½ 1 ΔH = -1202 ÷ 2 = 601 kJ mol⁻¹

Standard conditions

The change in enthalpy, ΔH varies according to the conditions of the experiment and is affected by temperature, pressure and concentration. All literature values for enthalpy changes are determined under standard conditions.

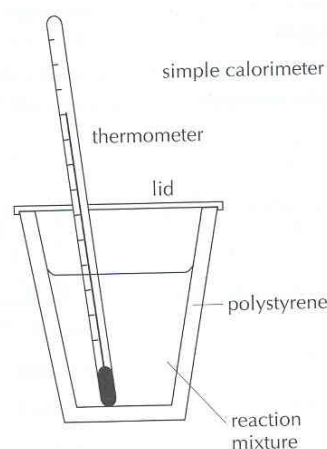
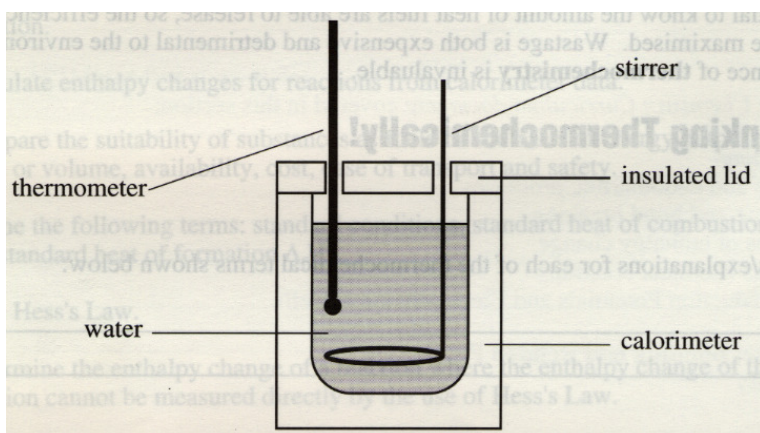
Standard conditions are:

- 1 atm or 1.01×10^5 Pa or 101 kPa
- 1 mole of gas
- Temperature of 298K or 25°C.

Measuring heat transfer (enthalpy changes)

Many enthalpy changes can be measured simply in the laboratory. If we measure the temperature change of the water and we know its mass and specific heat capacity we can work out how much energy was transferred to or from the water during the chemical reaction. We can find the change in enthalpy if we know the mass of the reactant/s. A calorimeter can be used to get more accurate values for enthalpy changes. A calorimeter container usually is well-insulated with styrofoam (polystyrene) and has a lid to minimize heat loss. If it is not insulated properly the reaction can either gain or lose heat energy to the surroundings.

A calorimeter



Procedure for finding enthalpy change of a reaction in an aqueous solutions

1. Take one reactant of known volume and concentration. This will be the limiting reagent.
2. Place in the calorimeter or some other insulated container like a styrofoam cup.
3. Record the initial temperature in °C.
4. Take an excess (or equal) mole of the other reactant of known volume and concentration. Record its initial temperature in °C.
5. Quickly add the excess reactant to the calorimeter, cover and stir rapidly to minimize heat loss.
6. Record the temperature change over time.
7. Determine the maximum temperature in °C.
8. Calculate enthalpy change, ΔH using the number of moles of limiting reagent and the equation $\Delta H = (m \times c \times \Delta T) \div n_{\text{limiting reagent}}$

Procedure for finding the enthalpy change of a solid or liquid fuel

1. Take a calorimeter containing a known volume of water.
2. Record the temperature of the water.
3. Take a known mass of fuel.

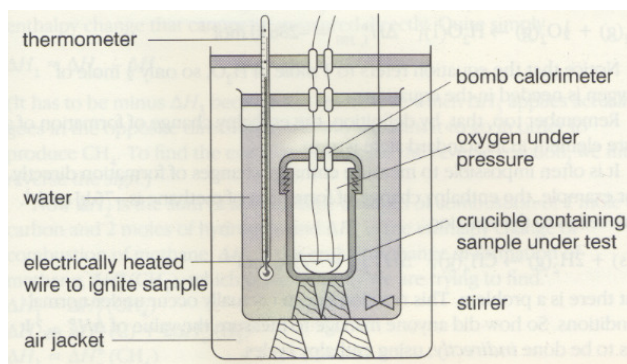
- Burn the fuel under the calorimeter minimizing heat loss to the surroundings. The energy released will be used to heat the water. The water in the calorimeter must be stirred to make sure that the heat is distributed evenly.
- At the end of the reaction record the change in mass of the fuel and the change in temperature.
- Calculate enthalpy change, ΔH using the number of moles of fuel and the equation $\Delta H = (m \times c \times \Delta T) \div n_{\text{limiting reagent}}$

Sources of systematic error and how they can be minimized

Systematic errors will cause the experimental enthalpy change, ΔH to be less than the accepted value.

- The reaction is not stirred or stirred too slowly. This will increase the time taken for the reaction and will increase the amount of heat lost. Minimize by stirring rapidly when the second reactant is added and extrapolate the graph of temp vs. time to determine the temperature change for an instantaneous reaction.
- The reaction container is not insulated adequately resulting the reaction can either gain or lose heat energy to the surroundings. This will make the change in temperature less than accepted and decrease the enthalpy change. Minimize by ensuring the reaction container is insulated correctly with styrofoam, and has an insulating lid. Alternatively a temperature versus time graph can be plotted and the instantaneous change in temperature determined from extrapolation.
- The transfer of heat to the temperature probe and metal reactants is ignored.
- The reaction is not carried out under standard conditions of 1 atm (1.01×10^5 Pa), concentration of 1 mol l^{-1} and a temperature of 298K or 25°C. All literature values are determined under standard conditions.
- Assumption that the specific heat capacity of the solutions is equal to that of water.

A bomb calorimeter is used for making more accurate measurements that are closer to the accepted value in combustion reactions (not for aqueous solutions). The fuel to be burned is electrically ignited and burns in oxygen. Energy is transferred to the surroundings, water, whose temperature rise is measured.

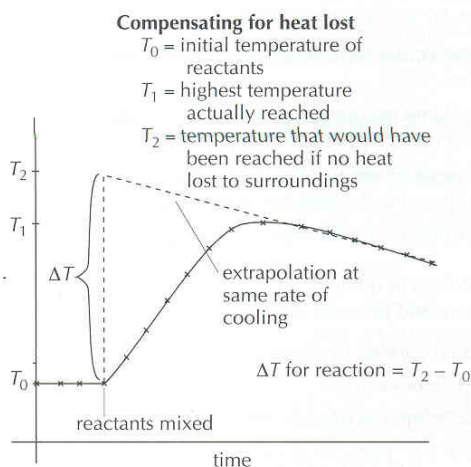


Experiments involving the calculation of enthalpy changes

CALORIMETRY

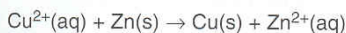
The enthalpy change for a reaction can be measured experimentally by using a calorimeter. In a simple calorimeter all the heat evolved in an exothermic reaction is used to raise the temperature of a known mass of water. For endothermic reactions the heat transferred from the water to the reaction can be calculated by measuring the lowering of temperature of a known mass of water.

To compensate for heat lost by the water in exothermic reactions to the surroundings as the reaction proceeds a plot of temperature against time can be drawn. By extrapolating the graph, the temperature rise that would have taken place had the reaction been instantaneous can be calculated.



50.0 cm³ of 0.200 mol dm⁻³ copper(II) sulfate solution was placed in a polystyrene cup. After two minutes 1.20 g of powdered zinc was added. The temperature was taken every 30 seconds and the following graph obtained. Calculate the enthalpy change for the reaction taking place.

Step 1. Write the equation for the reaction



Step 2. Determine the limiting reagent

$$\text{Amount of Cu}^{2+}(\text{aq}) = 50.0/1000 \times 0.200 = 0.0100 \text{ mol}$$

$$\text{Amount of Zn}(\text{s}) = 1.20/65.37 = 0.0184 \text{ mol}$$

$\therefore \text{Cu}^{2+}(\text{aq})$ is the limiting reagent

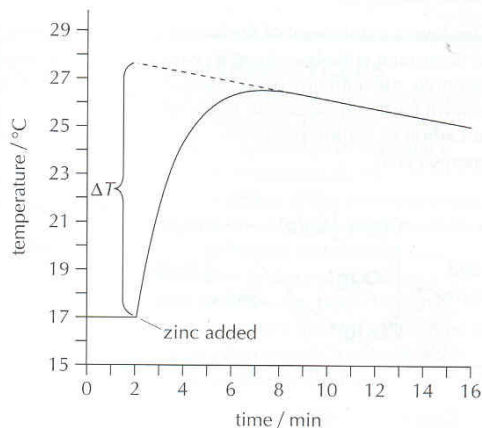
Step 3. Extrapolate the graph (*already done*) to compensate for heat loss and determine ΔT

$$\Delta T = 10.4 \text{ }^\circ\text{C}$$

Step 4. Calculate the heat evolved in the experiment for 0.0100 mol of reactants

$$\text{Heat evolved} = 50.0/1000 \times 4.18 \times 10.4 \text{ }^\circ\text{C} = 2.17 \text{ kJ}$$

Step 5. Express this as the enthalpy change for the reaction $\Delta H = -2.17 \times \frac{1}{0.0100} = -217 \text{ kJ mol}^{-1}$



Questions

To receive full credit, the method used and the steps involved at arriving at your final answer must be shown clearly. For numerical calculations please pay proper attention to significant figures.

1. Aluminum has a specific heat of $0.902 \text{ Jg}^{-1}\text{K}^{-1}$. How much heat is lost when a piece of aluminum with a mass of 23.984 g cools from a temperature of 415.0°C to a temperature of 22.0°C ?
2. The temperature of a sample of water increases by 69.5°C when 24 500 J are applied. The specific heat of liquid water is $4.18 \text{ Jg}^{-1}\text{K}^{-1}$. What is the mass of the sample of water?
3. 850 joules of heat are applied to a 250 g sample of liquid water with an initial temperature of 13.0°C . Find a) the change in temperature and b) the final temperature.
4. When 34 700 J of heat are applied to a 350 g sample of an unknown material the temperature rises from 22.0°C to 173.0°C . Calculate the specific heat of this material?
5. Calculate how much energy in kJ is needed to warm 250.0 ml water from 25.0°C to 95.0°C in order to make a cup of coffee? (250 ml = 250 g)
6. Determine the specific heat capacity of methanol if $1.80 \times 10^2 \text{ J}$ of heat is added to 70 g of liquid methanol and its temperature increases from 18.5°C to 30.5°C .
7. (N00/S) When 50.0 cm^3 of 1 mol dm^{-3} HCl is mixed rapidly with 50.0 cm^3 of 1 mol dm^{-3} NaOH, the temperature of the resulting solution increases by 6°C . Determine the temperature change when:
 - a) 100 cm^3 of each of these solutions is mixed
 - b) 2 mol dm^{-3} of each solution is mixed
8. When 8.00 g of ammonium nitrate (NH_4NO_3) was dissolved in 100.0 ml water, the temperature fell from 19.0°C to 14.5°C .
 - a) Calculate the enthalpy (Q) of the solution in J and kJ.
 - b) The experiment was repeated but with 10.0 g of NH_4NO_3 . State and explain if there would be a change in temperature.
9. An experiment was carried out to determine the amount of energy released in a "low fat" chocolate bar with a mass of 50.0 g. The complete combustion of a 10.0 g sample of the bar raised the temperature of 500.0g of water from 19.0°C to 86.5°C . Calculate the energy value (kJ) of the chocolate bar.
10. 500.0g of water were heated on a gas camping stove from 20.0°C to 100.0°C . If the pot was made of aluminum (specific heat capacity $875 \text{ kJ kg}^{-1} \text{ K}^{-1}$) and it weighed 100.0 g, calculate the heat energy (q) required in kJ to heat:
 - a) The pot
 - b) The water
 - c) The gas camping stove contained butane (C_4H_{10}). 14.5g of butane was used to boil the water. How many moles of butane was this?

11. Calculate the change in the enthalpy, ΔH of combustion of ethanol in Jmol^{-1} when 0.920g of ethanol was burned under a calorimeter containing 400.0 ml of water. The temperature rose by $10.0\text{ }^{\circ}\text{C}$



12. For the reaction: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \leftrightarrow 2\text{NH}_3(\text{g}) \quad \Delta H = -92.0 \text{ kJ mol}^{-1}$
- Determine the heat energy released when 2 moles of nitrogen reacts.
 - Determine the energy released when 0.50 mole of ammonia is formed.
 - Calculate the number of moles of nitrogen that react to produce 1840 kJ mol^{-1} of energy?
 - Determine the heat energy released when 10.0 g of hydrogen burns.
13. In cooking "Crepe Suzette", a delicious type of French pancake one tablespoon (2.50g) of brandy (containing ethanol) is poured over the pancakes and then it is ignited.
- The molar mass of the ethanol is 46.08 gmol^{-1} . Calculate the number of moles of ethanol, $\text{C}_2\text{H}_5\text{OH}$ in the tablespoon?
 - The enthalpy change for the reaction is -1350 kJmol^{-1} . Calculate the heat released when the brandy on the Crepe Suzette burns?
14. (N01/H/2) 10.0 g of NaOH is added to 100.0 g of water at 23.2°C in a glass beaker. The solution is stirred and the maximum temperature of 44.6°C is reached.
- Calculate the heat produced by the reaction. (Assume the specific heat capacity of the solution is $4.20 \text{ Jg}^{-1}\text{ }^{\circ}\text{C}^{-1}$). Is the reaction exothermic or endothermic [4]
 - Calculate the enthalpy change for dissolving the solid NaOH in water in kJmol^{-1} . [2]
 - Write an equation for the dissolving process. [1]
 - The value given in the literature under similar conditions is -42.7 kJmol^{-1} . Suggest a reason why the calculated value of the enthalpy change of the solution is different from the literature value and propose an improvement in the procedure to obtain a more accurate result. [2]
15. (M00/H/2) Describe an experiment that could be conducted in the school laboratory to determine the value of ΔH (in kJmol^{-1}) for the neutralization reaction between HCl and NaOH. Write an equation for the reaction and show the calculation that would need to be carried out to obtain the value for ΔH . [7]
16. (M05/S/2) In a neutralization reaction 50.0 cm^3 of a 0.500 moldm^{-3} solution of potassium hydroxide is mixed rapidly in a glass beaker with 50.0 cm^3 of a 0.500 moldm^{-3} solution of hydrochloric acid.
- Initial temperature of each solution = 19.6°C
Final temperature of the mixture = 23.1°C
- Write an equation for the reaction. [1]
 - State with a reason whether the reaction was exothermic or endothermic. [1]

- c) Explain why the solutions were mixed rapidly. [1]
- d) Calculate the enthalpy change in kJmol^{-1} . Assume that the specific heat capacity of the solution is the same as that of water. [1]
- e) If the experiment was repeated with an HCl concentration of 1 mol dm^{-3} state and explain what the temperature change would be.

17. (N98/H/2) The reaction between a strong base, sodium hydroxide and a strong acid, sulfuric acid was studied by measuring the heat produced during neutralization. Volumes of 1.25 mol dm^{-3} of the base ranging from 10.0 to 50.0 cm^3 were mixed with the required volumes of 1.25 mol dm^{-3} solution of the acid to produce a total volume of 60.0 cm^3 . The changes in temperature were measured and recorded in the table below.

Volume base (cm^3)	Temperature ($^{\circ}\text{C}$)
10.0	2.2
20.0	4.6
30.0	7.0
40.0	7.0
50.0	3.5

- a) Write an equation to represent the neutralization reaction. [1]
- b) Draw a graph of the results. [5]
- c) From the graph determine the greatest temperature increase. [2]
- d) Calculate the number of moles of base and the number of moles of acid in (c). [4]
- e) Determine the limiting reagent. [1]
- f) Calculate the energy released, Q in joules by the reaction when the maximum temperature was obtained on the graph. [2]
- g) Calculate the enthalpy of neutralization, ΔH [2]
- h) The accepted value for the enthalpy of neutralization is -55.5 kJmol^{-1} . Outline two sources of error in the experimental procedure that could result in a calculated value of ΔH that is smaller than the accepted value and suggest a way that **one** of these could be minimized. [3]

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Problems Answers

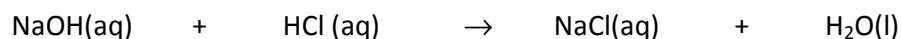
1. **When 1.00g of sodium hydroxide is dissolved in 100.0 g of water in an insulated container the temperature rises from 20.00°C to 22.66°C. Calculate the enthalpy change for the solution. State any assumptions made.**

$$\begin{aligned}\Delta H &= \frac{m \times c \times \Delta T}{n \text{ (limiting reagent)}} ; (1) & n(\text{NaOH}) &= m \div M = 1.00 \text{ g} \div 40.00 \text{ gmol}^{-1} = 0.0250 \text{ mol} ; (1) \\ & & & \\ &= \frac{100\text{g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times (22.66 - 20.00 \text{ K})}{0.0250 \text{ mol}} \\ &= 1111.88 \div 0.0250 \text{ Jmol}^{-1} \\ &= 44475 \text{ Jmol}^{-1} \\ &= -44.5 \text{ kJmol}^{-1} \text{ (3SF)}\end{aligned}$$

Assumptions

1. Specific heat capacity of the sodium hydroxide solution = specific heat capacity of water
2. 100 g water + 1.00g NaOH = 100g solution

2. Calculate ΔH when 100.0 ml of 1.00 mol dm⁻³ NaOH is added to 100.0 ml of 1.00 mol dm⁻³ HCl in an insulated container the temperature rises from 21.0°C to 34.6°C.



$$\text{Actual } n(\text{NaOH}) = c \times v = 1.00 \times 100.0 \div 1000 = 0.10 \text{ mol}$$

$$\text{Actual } n(\text{HCl}) = c \times v = 1.00 \times 100.0 \div 1000 = 0.10 \text{ mol}$$

	NaOH	:	HCl
Mole ratio from equation	1	:	1

Let χ = moles NaOH needed	χ	:	0.10
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by cross multiplying: $\chi = 0.10 \text{ mol NaOH}$ is needed to react with 0.10 mol HCl
so both get used up

Since both reactants react in a 1:1 ratio and the number of moles of each is the same both are used up in the reaction and there is no limiting reagent.

$$\Delta H = \frac{m \times c \times \Delta T}{n}; (1) \quad n = c \times v = 1.00 \text{ mol dm}^{-3} \times (100.0 \text{ ml} \div 1000); (1)$$

$$= \frac{200.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ K}^{-1} \times (34.6 - 21.0 \text{ K})}{0.100 \text{ mol}}; (1)$$

$$= -114 \text{ kJ mol}^{-1} \text{ (3SF)}; (1)$$

Assumptions

1. Specific heat capacity of the solution = specific heat capacity of water
2. Total mass of reaction = 200 ml
3. 200 ml solution = 200g solution

5.2 Questions

SELECTED ANSWERS

$$\begin{aligned} 5. \quad Q &= m \times c \times \Delta T \\ &= 250.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1} \times (95.0 - 25.0 \text{ }^\circ\text{C}) \\ &= -73150 \text{ J} \\ &= -73.2 \text{ kJ} \quad (3\text{SF}) \end{aligned}$$

$$\begin{aligned} 6. \quad c &= Q \div (m \times \Delta T) \\ c &= 180 \text{ J} \div (70 \text{ g} \times 12.0 \text{ }^\circ\text{C}) \\ c &= 0.20 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1} \quad (2\text{SF}) \end{aligned}$$

7. The temperature released or absorbed is proportional to the number of moles of the substance

- Even though the volume has doubled the number of moles of acid and alkali has remained the same so the temperature change = 6 °C
- When the concentration is doubled, the moles doubles, and so the temperature change will double 12 °C

8.

$$\begin{aligned} \text{a) } Q &= m \times c \times \Delta T \\ &= 100. \text{g} \times 4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1} \times (19.0 - 14.5 \text{ }^\circ\text{C}) \\ &\quad \text{assume } m \text{ is mass of aqueous reactant} \\ &= +1881 \text{ J} \\ &= +1.88 \text{ kJ} \quad (3\text{SF}) \\ &\quad \text{endothermic sign is positive} \end{aligned}$$

b) Since the enthalpy change is proportional to the number of moles, there would be a greater temperature change. By increasing the moles of ammonium nitrate more heat energy will be absorbed by the ammonium nitrate.

$$\begin{aligned} 9. \quad Q &= m \times c \times \Delta T \\ &= 500.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1} \times (86.5 - 19.0 \text{ }^\circ\text{C}) \\ &= -141075 \text{ J} \\ &\quad \text{heat released, exothermic, sign is negative} \\ &= -141 \text{ kJ for a } 10.0 \text{ g sample of chocolate} \\ &= -141 \times 5 = -705 \text{ kJ (3 SF) for a } 50.0 \text{ g chocolate bar} \end{aligned}$$

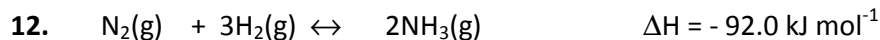
10.

$$\begin{aligned} \text{a) } Q_{(\text{pot})} &= m \times c \times \Delta T \\ &= 100.0 \div 1000 \text{ kg} \times 875 \text{ kJ kg}^{-1} \text{ } ^\circ\text{C}^{-1} \times (100.0 - 20.0^\circ\text{C}) \\ &= -7.00 \times 10^3 \text{ kJ (3SF)} \\ &\text{heat released, exothermic, sign is negative} \end{aligned}$$

$$\begin{aligned} \text{b) } Q_{(\text{water})} &= m \times c \times \Delta T \\ &= 500.0 \text{ g} \times 4.18 \text{ J g}^{-1} \text{ } ^\circ\text{C}^{-1} \times (80.0^\circ\text{C}) \\ &= -167200 \text{ J} \quad \text{heat released, exothermic, sign is negative} \\ &= -167 \text{ kJ} \end{aligned}$$

$$\begin{aligned} \text{c) } n_{(\text{butane})} &= m \div M \\ &= 14.5 \text{ g} \div 58.14 \text{ gmol}^{-1} \\ &= 0.249 \text{ mol (3SF)} \end{aligned}$$

-1 (SF) once - 1 (U) once



Set up ratios and cross multiply

$$\begin{aligned} \text{a) The reaction says:} \\ 1 \text{ mole N}_2 &= -92.0 \text{ kJ mol}^{-1} \\ 2 \text{ mol N}_2 &= -92.0 \times 2 = -184 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \text{b) } 2 \text{ mole NH}_3 &= -92.0 \text{ kJ mol}^{-1} \\ 0.50 \text{ mole NH}_3 &= -92.0 \times \frac{0.50}{2} = -23 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned} \text{c) } 1 \text{ mole N}_2 &= -92.0 \text{ kJ mol}^{-1} \\ x \text{ mol N}_2 &= -1840 \text{ kJ mol}^{-1} \quad \text{cross multiply} \\ 20.0 \text{ mol N}_2 &= -1840 \text{ kJ mol}^{-1} \end{aligned}$$

$$\begin{aligned}
 \text{d) } n(\text{H}_2) &= m \div M \\
 &= 10.0 \text{ g} \div 2.02 \text{ gmol}^{-1} \\
 &= 4.95 \text{ mol}
 \end{aligned}$$

$$3 \text{ mole H}_2 = -92.0 \text{ kJ mol}^{-1}$$

$$4.95 \text{ mol N}_2 = x \text{ kJ mol}^{-1}$$

cross multiply

$$4.95 \text{ mol N}_2 = \frac{92.0 \times 4.95}{3} = -152 \text{ kJ mol}^{-1}$$

13.

$$\begin{aligned}
 \text{a) } n_{(\text{ethanol})} &= m \div M \\
 &= 2.50 \text{ g} \div 46.08 \text{ gmol}^{-1} \\
 &= 0.0542 \text{ mol (3 SF)}
 \end{aligned}$$

$$\text{b) } \Delta H = (Q \div n)$$

$$Q = \Delta H \times n$$

$$= 1350 \text{ kJmol}^{-1} \times 0.0542 \text{ mol}$$

$$= -73.2 \text{ kJ (3SF)}$$

(or 73.2 kJ energy released)

14.

$$\begin{aligned}
 \text{a) } Q &= m \times c \times \Delta T \\
 &= 100.0 \text{ g} \times 4.20 \text{ Jg}^{-1}\text{°C}^{-1} \times 21.4\text{°C} \\
 &= 8988 \text{ J of energy released} \\
 &= -8.98 \times 10^3 \text{ J (3SF) Reaction is Exothermic}
 \end{aligned}$$

$$\text{b) } M(\text{NaOH}) = 40.00 \text{ gmol}^{-1};$$

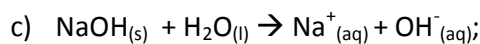
$$\Delta H = (Q \div n)$$

$$= 8988 \text{ J} \div (10.0 \text{ g} \div 40.00 \text{ gmol}^{-1});$$

$$= (-35952 \text{ Jmol}^{-1})$$

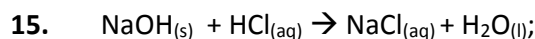
$$= -36.0 \text{ kJmol}^{-1} \quad (3\text{SF});$$

(Need minus sign for the mark; and units are not given)



d) Reason for the difference - To prevent heat loss to the surrounding (thus lowering expected value of ΔH);

Improvement - Use a Styrofoam cup **or** insulate (and cover) a beaker; Stir to increase the rate of the reaction



a) Take the dilute HCl of known volume and concentration. This will be the limiting reagent (or equal moles).

b) Place in the calorimeter or some other insulated container like a styrofoam cup.

c) Record the initial temperature in °C of the HCl.

d) Take an excess (or equal) mole of the aqueous NaOH of known volume and concentration. Record its initial temperature in °C.

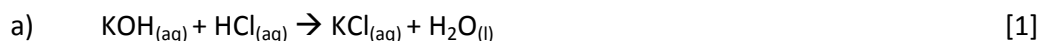
e) Quickly add the NaOH to the calorimeter, cover and stir rapidly to minimize heat loss.

f) Record the temperature change over time.

g) Determine the maximum temperature in °C. *[any 5 = 5 points]*

Calculate enthalpy change, ΔH using the number of moles of limiting reagent and the equation $\Delta H = (m \times c \times \Delta T) \div n_{\text{NaOH}}$; [1]

16.



b) Exothermic because the temperature increases / heat is released [1]

c) To make heat loss as small as possible / so that all the heat is given out rapidly.

(do not accept to produce a faster reaction) [1]

d) $Q = m \times c \times \Delta T$
 $= 100.0 \text{ g} \times 4.18 \text{ Jg}^{-1}\text{°C}^{-1} \times 3.5\text{°C}$ [1]

$= 1463 \text{ J}$

$= -1.46 \text{ kJ}$

or 1.46 kJ of energy is released

$n(\text{KOH}) \text{ used} = c \times v$ (no excess or limiting reagent)

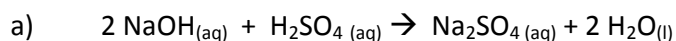
$= 0.500 \times 50.0$

$= 0.0250 \text{ mol}$

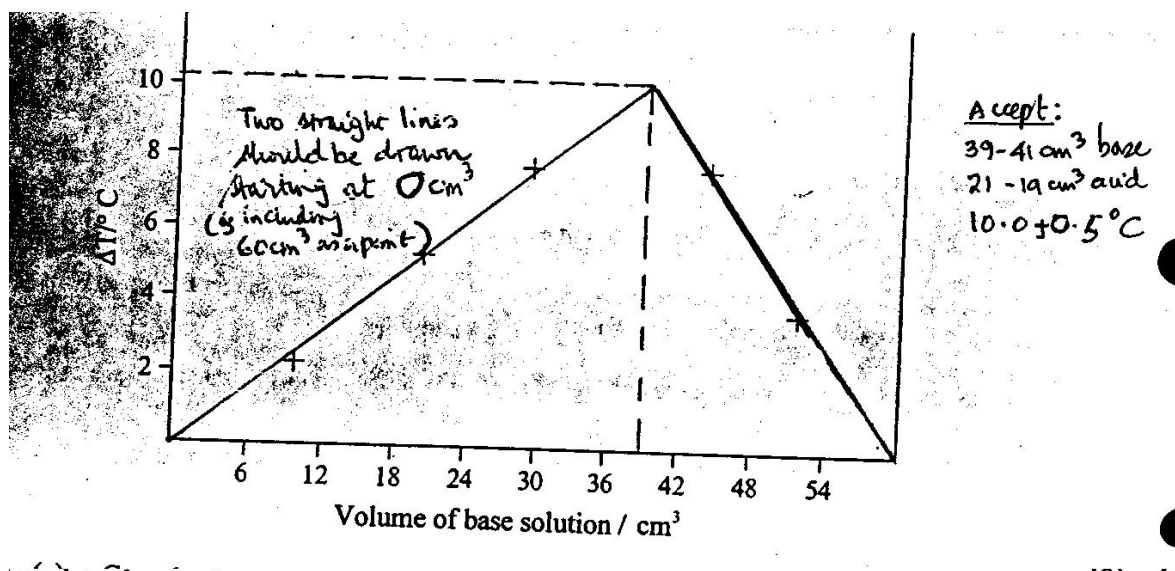
$$\begin{aligned} \Delta H &= (Q \div n) \\ &= 1.46 \text{ kJ} \div 0.025 \text{ mol} ; \\ &= -58.5 \text{ kJmol}^{-1} \text{ (3SF) ; (minus needed for the mark)} \end{aligned}$$

- e) Since the number of moles of each reactant is the same and both are used up in the reaction (there is no limiting reagent) doubling the concentration of HCl would not change the temperature because it would become the excess reagent and KOH the limiting reagent.

17.



b)



labeled x axis with variable and unit;

descriptive title;

correctly plotted points and straight line through points;

suitable scale / good shape / good size ;

[5]

- c) extrapolate each best fit straight line to get maximum temperature change until both lines intersect;

[1]

allow 10.0 to 10.2 °C ;

[1]

- d) volume of NaOH = 39 – 41 cm³

[1]

$$\text{moles NaOH} = c \times v = 1.25 \times (39 \div 1000) = \underline{0.04875 \text{ mol}} ; \quad [1]$$

$$\text{mole H}_2\text{SO}_4 = c \times v = 1.25 \times [(60 - 39) \div 1000] = \underline{0.02625 \text{ mol}} ; \quad [1]$$

$$\text{volume H}_2\text{SO}_4 = n \div c = 0.0265 \div 1.25 = 0.020 \text{ dm}^3 = 20. \text{ cm}^3 \quad [1]$$

n(H₂SO₄)required

	NaOH	:	H ₂ SO ₄
Mole ratio from equation	2	:	1
Let χ = moles HCl required	0.04875 mol	:	χ mol

by cross multiplying:

$$2 \chi = 0.04875 \text{ mol}$$

$$\chi = 0.024 \text{ mol H}_2\text{SO}_4 \text{ required}$$

n(NaOH) required

	NaOH	:	H ₂ SO ₄
Mole ratio from equation	2	:	1
Let χ = moles KOH required	χ	:	0.02625 mol

by cross multiplying:

$$\chi = 0.02625 \times 2 \text{ mol}$$

$$= 0.0525 \text{ mol KOH is required}$$

If the: actual moles of reactant < moles needed then this is the limiting reactant

0.026 moles H₂SO₄ available > 0.024 moles H₂SO₄ required to react with all of the NaOH it is the *excess reactant*.

0.049 moles NaOH available < 0.0525 moles **NaOH is the limiting reactant.** [1]

e) $Q = m \times c \times \Delta T$

$$= 60.0 \text{ g} \times 4.18 \text{ Jg}^{-1}\text{°C}^{-1} \times (10.2 - 2.2) \text{ °C}; \quad [1]$$

$$= 2558 \text{ J (2.56 kJ) of energy released (accept - 2558 J)} \quad [1]$$

f) $\Delta H = (Q \div n)$

$$= 2.56 \text{ kJ} \div \text{mol}_{\text{limiting reagent}}; \quad [1]$$

$$= 2.56 \text{ J} \div 0.04875 \text{ mol}$$

$$= -52 \text{ kJmol}^{-1} / -52 \text{ kJmol}^{-1}; \text{ (2SF)} \quad [1]$$

g) % Difference = $(55.5 - 52) \div 55.5 \times 100 = 6.0 \%$; [1]

source of error is heat loss to the surroundings; [1]

ensure reaction vessel is well insulated and has a lid (Styrofoam); [1]

(the reaction takes place in a glass beaker without a lid. You need to insulate the beaker with Styrofoam and use a lid)