

# Section 1

# Energy Changes

## **Energetics of Chemical and Physical Processes**

States of matter.

Absolute temperature as a measure of the average kinetic energy of particles.

The six thermochemical symbols  $\Delta_{\text{fus}}H^\circ$ ,  $\Delta_{\text{vap}}H^\circ$ ,  $\Delta_{\text{sub}}H^\circ$ ,  $\Delta_{\text{f}}H^\circ$ ,  $\Delta_{\text{c}}H^\circ$  and  $\Delta_{\text{atm}}H^\circ$  and their balanced equations.

Relationship between endothermic and exothermic reactions and bond making and bond breaking processes.

Calculations of enthalpy change using  $\Delta_{\text{f}}H^\circ$ ,  $\Delta_{\text{c}}H^\circ$ ,  $\Delta_{\text{r}}H^\circ$  or bond energies.

Hess's Law

Calculations of enthalpy change using Hess's Law

Activation energy and catalysed reactions

# The Particle Nature of Matter

Chemists divide matter into three states based on their observed properties. All matter can exist in three different states but there are some substances that do not exist in all three states under standard pressures and temperatures.

## Temperature and Kinetic Energy

### Difference between temperature and heat

Temperature of a substance (on the Kelvin scale) is a measure the average Kinetic Energy of its particles.  $E_k \propto T$

Heat is a measure of how much energy a system has in all together. For example, a small beaker of boiling water has a higher temperature than a large beaker of cold water. The large beaker contains many more water molecules and although each one might have less energy than a molecule in the boiling water, the total energy in the large beaker could be higher.

### Absolute zero

The point at which all particle motion stops, i.e. zero kinetic energy, is defined as 0 K (–273°C). Note that there is no degree sign (°) on Kelvin temperature. The average kinetic energy of the particles is directly proportional to the absolute temperature in Kelvin (K) not to the temperature in degree Celsius (°C).

### The Kelvin scale

The size of a degree in Kelvin is the same as in the Celsius scale but the zero point is different. To change between Celsius and Kelvin:

$$\text{Kelvin} = \text{degrees Celsius} + 273$$

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

	Kelvin	Celsius
Water Boils	373K	100°C
Water Freezes	273K	0°C
Absolute Temperature	0K	-273°C

## The Kinetic Theory of Matter

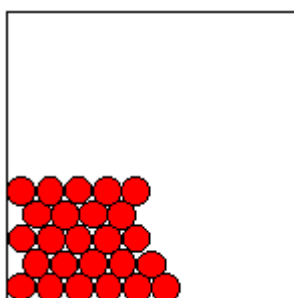
- All matter is composed of tiny particles (atoms, molecules or ions).
- Different substances have different types of particles of different sizes.
- The particles are in constant motion.
- The higher the temperature the faster the particles move and the more kinetic energy they have. Temperature is proportional to Kinetic Energy
- Heavier particles move slower than lighter particles at the same temperature.
- The arrangement of particles in different states (solid, liquid or gas) is different.
- Attractions (bonds) exist only between particles in the solid and liquid states.
- Melting and boiling points are determined by the strength of the forces attracting the particles together.

## States of Matter

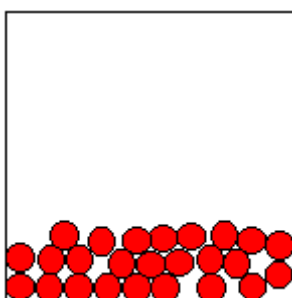
The three states of matter are **solid**, **liquid** and **gas**.

State	Particles	Energy	Bonding	Particle Movement
Solid	Close together	Not enough energy to break any attractions	All attractions between particles still exist	Vibrate about a fixed position
Liquid	Close together	Enough energy to break some attractions	Some attractions have been broken	Move slowly and randomly
Gas	Far apart	Enough energy to break all attractions	Negligible attractions between particles	Move rapidly and randomly

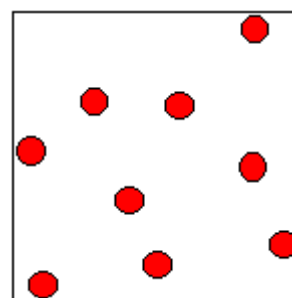
**Solid**



**Liquid**

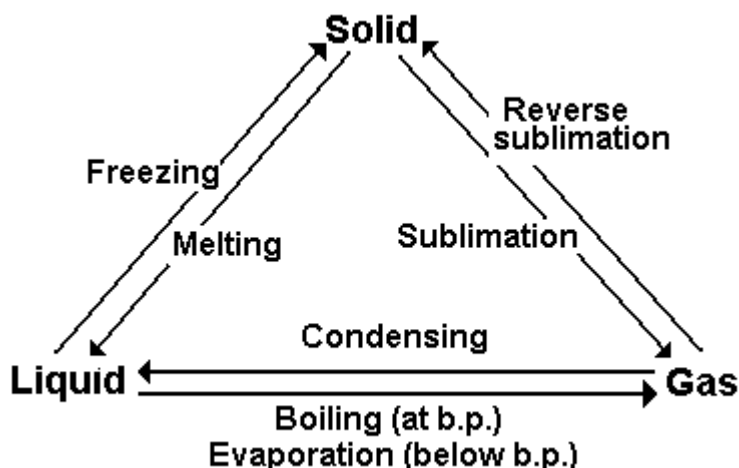


**Gas**



## State Changes

The kinetic theory explains how a substance changes from one state of matter to another.



### Melting

Added energy allows forces of attraction to be overcome and bonds broken between particles. The regular structure breaks down and the particles can move around as a liquid.

### Freezing

When a liquid is cooled enough it can freeze to form a solid. The particles have slowed down and organised themselves into a regular pattern (lattice).

### Boiling

Particles are given enough energy to overcome the forces of attraction between themselves and the other particles in the liquid and they escape forming a gas.

### Evaporation

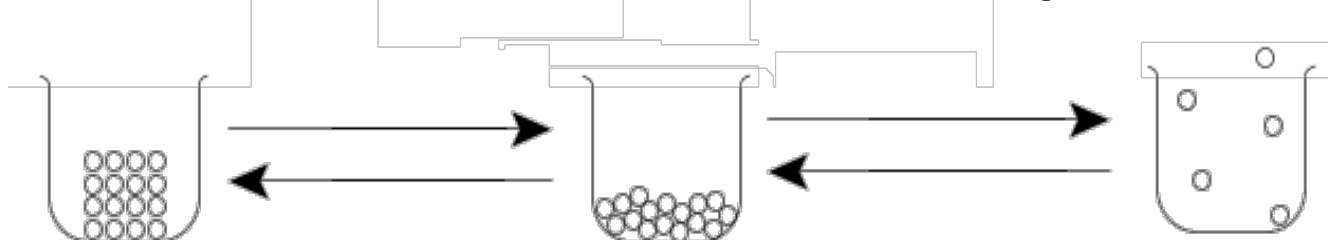
A liquid left standing in an open container evaporates slowly, even though its temperature is considerably lower than its boiling point. This happens because particles have a range of kinetic energy ( $E_k$ ) and because at any instant some will have enough energy to overcome the forces holding them in the liquid and escape from the surface.

### Condensing

The average kinetic energy of particles decreases and they move closer together. The attractive forces between the particles now cause the gas to condense into a liquid.

### Sublimation

When heated the solid particles spread out rapidly changing state directly from a solid to a gas. When cooled the gas particles change directly back to a solid. Sublimation means no liquid state.



## Questions 1

# Heats of Reaction

## Exothermic

A chemical reaction that releases heat energy as the reaction proceeds.

## Endothermic

A chemical reaction that absorbs heat energy as the reaction proceeds.

## Enthalpy (heat content)

The energy contained in substances in the form of either kinetic or potential energy.

Heat of reaction is measured as the **enthalpy change** ( $\Delta H$ ), the change of kinetic and potential energies of products and reactants during a chemical reaction assuming that the law of energy conservation is obeyed and that no other energy gains or losses take place.

$$\Delta H = H_P - H_R$$

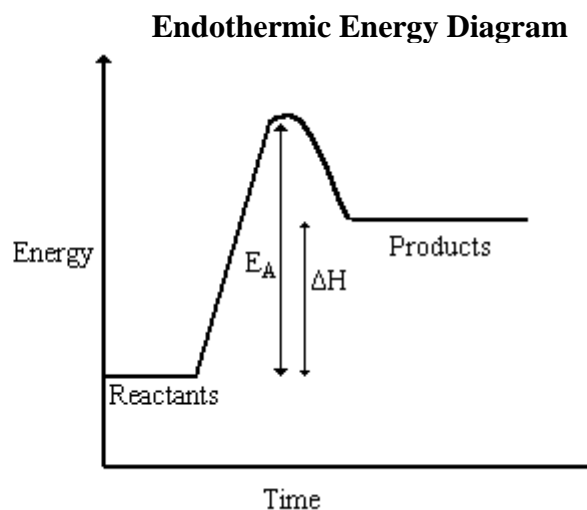
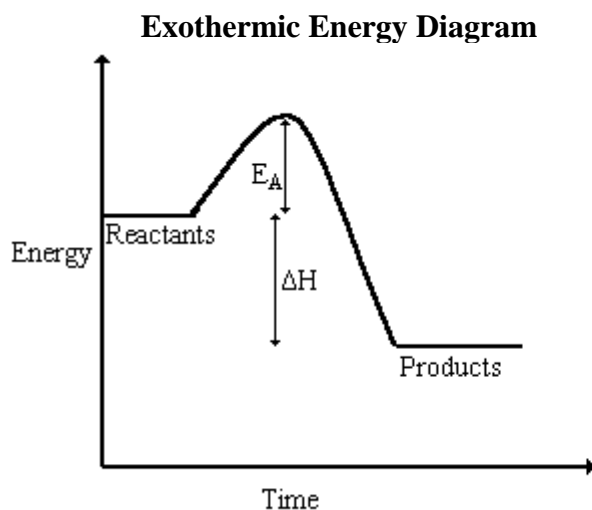
$H_P$  = Enthalpy or heat content of system in final state (Products).

$H_R$  = Enthalpy or heat content of system in initial state (Reactants).

$\Delta H$  = Enthalpy when initial and final temperatures and pressures are the same.

For exothermic reactions  $\Delta H < 0$  ( $H_P < H_R$ )

For endothermic reactions  $\Delta H > 0$  ( $H_P > H_R$ )



{  $E_A$  = Activation energy;  $\Delta H$  = Enthalpy change }

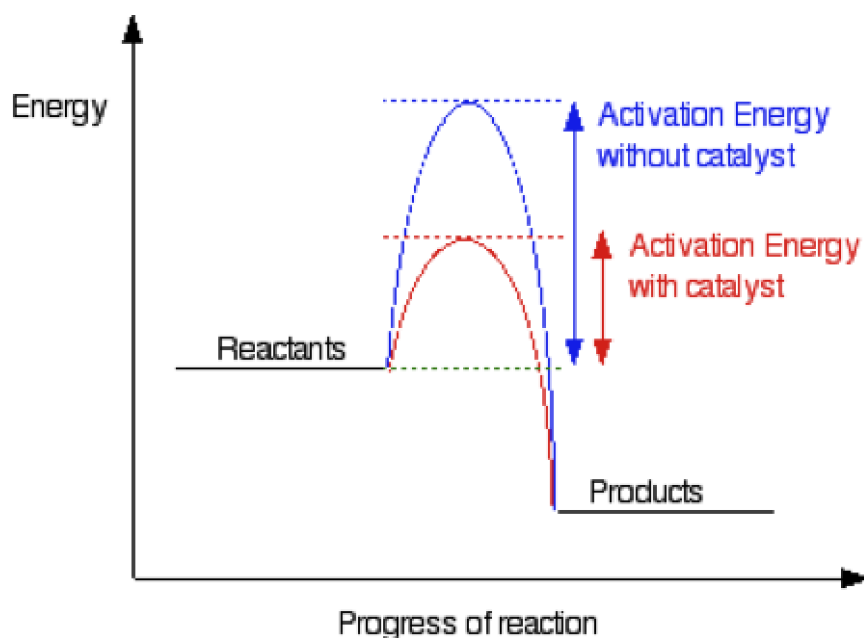
Activation Energy is the minimum energy required for a chemical reaction to occur.

## (a) The Effect of Heat on a Reaction

To speed up the reaction, you need to increase the average kinetic energy of the particles by increasing the temperature so that a greater proportion of the particles have energies equal to or greater than the activation energy of the reaction. Therefore, more successful collisions will occur so the reaction is faster.

## (b) The effect of a catalyst on a reaction

A catalyst is a substance that increases the rate of a chemical reaction without itself being used up. Catalysts do join reactions by providing an alternative reaction pathway with a lower activation energy. When the reaction has finished, the mass of the catalyst is exactly the same as that at the beginning. Showing this on an energy profile:



Adding a catalyst affects neither the energy level of products nor the  $\Delta H$  of the reaction, only shortens the time taken to complete the reaction.

# Thermochemical Symbols and Equations

When writing the thermochemical equation for a reaction or process, the physical state of each reactant or product must be recorded, using the notation shown below:

**(s): solid    (l): liquid    (g): gas    (aq): aqueous solution**

It is important to include physical states because state changes involve enthalpy changes.

The temperature and pressure at which the reaction takes place are also important. Data is recorded for a set of standard conditions, 101.3 kPa (or 1 atm) of pressure and usually 25°C (or 298 K) for temperature. Enthalpy changes measured under these conditions are indicated by the superscript °.

## Questions 3

### Enthalpy Changes for Changes of State

Enthalpy changes are associated with changes of state. Melting. Boiling and subliming are always endothermic.

#### (1) Heat of Fusion (melting)

The standard enthalpy of fusion ( $\Delta_{\text{fus}}H^\circ$ ). This is the energy required to change one mole of a substance from solid to liquid at standard conditions. It is an endothermic process.



#### (2) Heat of Vaporisation (boiling)

The standard enthalpy of vaporisation ( $\Delta_{\text{vap}}H^\circ$ ). This is the energy required to change one mole of a substance from liquid to gas at standard conditions. It is an endothermic process.



#### (3) Heat of Sublimation (subliming)

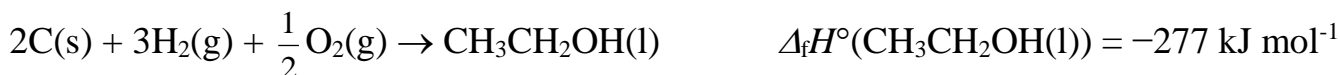
The standard enthalpy of sublimation ( $\Delta_{\text{sub}}H^\circ$ ). This is the energy required to change one mole of a substance from solid directly to gas at standard conditions. It is also an endothermic process.



#### (4) Standard Heat of Formation ( $\Delta_f H^\circ$ )

The standard heat of formation of a substance is the heat evolved or absorbed (either endothermic or exothermic) when one mole of the substance is formed from its elements under standard conditions. This is measured at 298K and 1 atm pressure, the thermochemical symbol is  $\Delta_f H^\circ$ .

e.g. The balanced equation for the thermochemical symbol  $\Delta_f H^\circ(\text{CH}_3\text{CH}_2\text{OH(l)})$  is:



**The enthalpy of formation of all pure elements in their natural states at standard conditions is taken as zero.**

$$\Delta_f H^\circ(\text{Element (in its natural state)}) = 0$$

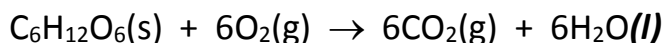
e.g.  $\Delta_f H^\circ(\text{Fe(s)}) = 0 \text{ kJ mol}^{-1}$ ,  $\Delta_f H^\circ(\text{O}_2\text{(g)}) = 0 \text{ kJ mol}^{-1}$ , etc.

These data can be used to calculate the standard enthalpy change of a reaction by using the following formula:

$$\Delta_r H = \text{sum heat of formation of products} - \text{sum heat of formation of reactants}$$

$$\Delta_r H^\circ = \sum \Delta_f H^\circ(\text{P}) - \sum \Delta_f H^\circ(\text{R})$$

e.g. Use the standard enthalpy of formation data to calculate the enthalpy change, in  $\text{kJ mol}^{-1}$ , for the standard enthalpy of combustion of solid glucose. The balanced equation is shown below.



Substance	$\Delta_f H^\circ / \text{kJ mol}^{-1}$
$\text{C}_6\text{H}_{12}\text{O}_6\text{(s)}$	-1260
$\text{H}_2\text{O(l)}$	-286
$\text{CO}_2\text{(g)}$	-394

$$\begin{aligned}\Delta_r H^\circ &= \sum \Delta_f H^\circ(\text{products}) - \sum \Delta_f H^\circ(\text{reactants}) \\ &= [6 \times (-394) + 6 \times (-286)] - [-1260 + 0] \\ &= -2820 \text{ kJ mol}^{-1}\end{aligned}$$



### (5) Standard Heat of Combustion ( $\Delta_c H^\circ$ )

The standard heat of combustion of a substance,  $\Delta_c H^\circ$ , is the energy released (always exothermic) when one mole of the substance is completely burnt in oxygen under standard conditions. Products must be in their stable states under standard conditions.

e.g. The balanced equation for the thermochemical symbol  $\Delta_c H^\circ(\text{CH}_3\text{CH}_2\text{OH(l)})$  is:



The product  $\text{H}_2\text{O(l)}$  is in liquid state not  $\text{H}_2\text{O(g)}$  in gaseous state because at standard conditions  $\text{H}_2\text{O}$  is water not steam.

### (6) Standard Heat of Atomisation ( $\Delta_{\text{atm}} H^\circ$ )

The standard heat of atomisation,  $\Delta_{\text{atm}} H^\circ$ , of an element is defined as the enthalpy change when one mole of gaseous atoms is formed from the element under standard conditions.



This data is equivalent to half of the bond enthalpy of  $\text{H—H}$ .

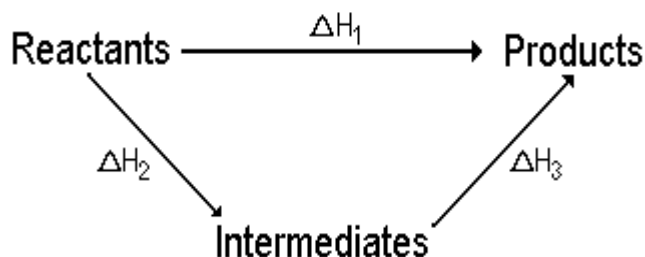


This data is equivalent to half of the bond enthalpy of  $\text{Cl—Cl}$ .

## Hess's Law of Constant Heat Summation

The enthalpy change in forming compound X directly from its constituent elements will be the same as the overall  $\Delta H$  for indirectly forming compound X by combusting those elements and then forming compound X from the combustion products. Thus the enthalpy change for any reaction can be found from enthalpies of formation and combustion using **Hess's law**.

*“The total enthalpy change in a chemical reaction is determined only by the initial and final states of the system, it is independent of the pathway followed”.*

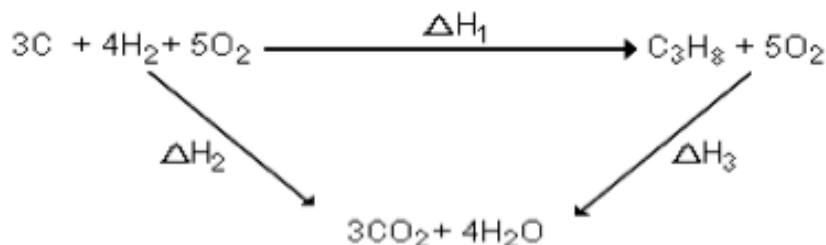


According to Hess's law  $\Delta H_1 = \Delta H_2 + \Delta H_3$

### Example

Calculate the standard enthalpy of formation of  $C_3H_8$  at 298 K given the standard enthalpies of combustion below.

$$\{\Delta_c H^\circ[C(s)] = -393 \text{ kJ mol}^{-1}; \Delta_c H^\circ[H_2(g)] = -286 \text{ kJ mol}^{-1}; \Delta_c H^\circ[C_3H_8(g)] = -2220 \text{ kJ mol}^{-1}\}$$



$$\Delta H_1 = \Delta H_2 - \Delta H_3 \quad (\Delta H_1 = \text{the forward } \Delta H_2 \text{ reaction and the reverse of } \Delta H_3)$$

$$\Delta H_1 = \Delta_f H^\circ[C_3H_8(g)]$$

$$\begin{aligned} \Delta H_2 &= 3 \times \Delta_c H^\circ[C(s)] + 4 \times \Delta_c H^\circ[H_2(g)] \\ &= 3(-393 \text{ kJ mol}^{-1}) + 4(-286 \text{ kJ mol}^{-1}) \\ &= -2323 \text{ kJ mol}^{-1} \end{aligned}$$

$$\Delta H_3 = \Delta_c H^\circ[C_3H_8(g)] = -2220 \text{ kJ mol}^{-1}$$

$$\begin{aligned} \text{Hence: } \Delta_f H^\circ(C_3H_8(g)) &= (-2323 \text{ kJ mol}^{-1}) - (-2220 \text{ kJ mol}^{-1}) \\ &= -103 \text{ kJ mol}^{-1} \end{aligned}$$

### Questions 4

### Questions 5

## Bond Dissociation Energy

The bond energy (bond dissociation energy) is the energy required to break one mole of covalent bonds when the process is carried out in the gas phase.

A covalent bond arises from the attraction of the two nuclei (+ve) for the shared electrons (-ve) between them. To separate these two atoms work must be done against this attractive force. Energy is required for such work to be done and then stored as potential energy in the separated atoms. When the atoms recombine, this energy will be released.

- *Bond breaking is an energy absorbing, or endothermic process.*
- *Bond making is an energy releasing, or exothermic process*

Some bond enthalpy data (average values in  $\text{kJ mol}^{-1}$ ) are shown below.

F – F 159	Cl – Cl 243	Br – Br 193	I – I 151	C – H 413
N – H 391	O – H 463	F – H 561	Cl – H 428	Br – H 362
C – O 358	C – C 348	C = C 614	C $\equiv$ C 839	C – Cl 339
C – Br 285	C = O 740	H – H 436	O = O 498	N $\equiv$ N 945

These bond energies given are **average** values. The actual value of a particular bond may be slightly different depending on the chemical environment of this bond. The same type of bonds in different compounds could have different bond energy values. So the enthalpy change of a chemical reaction calculated by using bond energy data is often an approximate value, and therefore it would not be as accurate as that calculated by using the data of standard enthalpy of formation.

Between the same type of atoms, triple bonds tend to be *stronger* than double bonds *stronger* than single bonds.

e.g. In terms of strength:  $\text{C} \equiv \text{C} > \text{C} = \text{C} > \text{C} - \text{C}$

Between the same type of atoms, single bonds tend to be *longer* than double bonds *longer* than triple bonds.

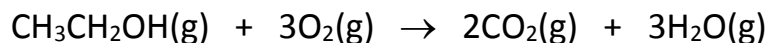
e.g. In terms of length:  $\text{C} - \text{C} > \text{C} = \text{C} > \text{C} \equiv \text{C}$

In general, bond length depends on the atomic radii of the two atoms; bond strength depends on the attraction of the two nuclei on the shared electrons and the repulsion between the two nuclei. If the atoms are too small, the repulsion between the nuclei will become more significant weakening the bond strength, e.g. F – F  $159 \text{ kJ mol}^{-1}$ .

If the relevant bond enthalpy data are obtained, then the enthalpy change of a chemical reaction can be calculated by using the formula below.

$$\Delta_r H = \Sigma \Delta H(\text{bonds broken}) - \Sigma \Delta H(\text{bonds formed})$$

e.g. Calculate the enthalpy change in  $\text{kJ mol}^{-1}$ , for the following reaction using the bond enthalpies provided on the previous page.



Bonds broken

$$5 \times \text{C—H} \quad 5 \times 413$$

$$\text{C—C} \quad 348$$

$$\text{C—O} \quad 358$$

$$\text{O—H} \quad 463$$

$$3 \times \text{O=O} \quad 3 \times 498$$

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$$4728$$

Bonds formed

$$2 \times 2 \text{ C=O} \quad 4 \times 740$$

$$3 \times 2 \text{ O—H} \quad 6 \times 463$$

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$$5738$$

$$\Delta_r H = \Sigma \Delta H(\text{Bonds broken}) - \Sigma \Delta H(\text{Bonds formed})$$

$$= 4728 - 5738$$

$$= -1010 \text{ kJ mol}^{-1}$$

## Questions 6

## Simple Measurements of Enthalpy Changes

This is just a brief introduction to the **two** simple measurements of enthalpy changes of reaction which can be easily carried out in a lab.

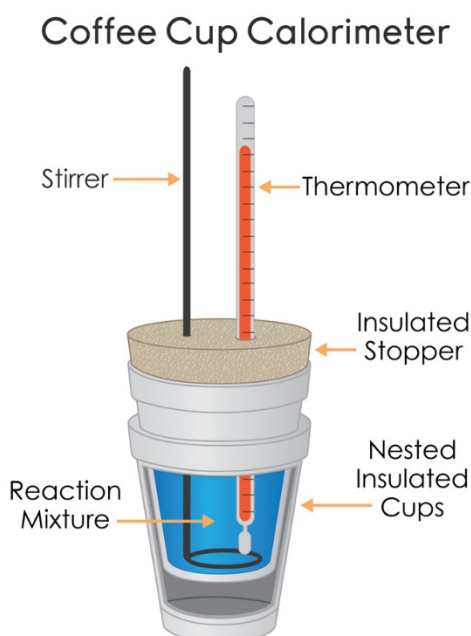
### 1. Enthalpy changes involving solutions or a solution plus a solid

#### *The experiments*

There are a whole range of different enthalpy changes that can be measured by reacting solutions or a solution plus a solid in a simple polystyrene cup. Common examples would be the measurement of the enthalpy change of

- (a) neutralisation of hydrochloric acid and sodium hydroxide solution;
- (b) dissolving anhydrous magnesium sulfate or magnesium sulfate heptahydrate.

The polystyrene cup serves to insulate the reaction mixture, and slows heat losses from the side and bottom. Heat is still lost from the surface of the liquid mixture and that can be reduced by using a polystyrene lid with a hole for a thermometer. You can allow for heat losses during the reaction by plotting a cooling curve.



There are other sources of error in these experiments - in particular in accuracy of the thermometer used and the stirring (when a solid is involved). Since the temperature change probably won't be very great, you would need to use the most accurate thermometer possible in order to keep the percentage error low.

### *The calculations*

Chemical reactions are accompanied by enthalpy changes. The heat evolved or absorbed during a reaction is given by the expression:

$$\text{Heat evolved or absorbed} = \text{mass} \times \text{specific heat} \times \text{temperature change}$$

That can be written in symbols as

$$\text{Heat evolved or absorbed} = c \, m \, \Delta T$$

You will find that the specific heat capacity is sometimes given the symbol "c" and sometimes the symbol "s". The specific heat capacity of a substance is the amount of heat needed to increase the temperature of 1 gram of it by 1 K (which is the same as 1°C). So for water, the specific heat capacity is **4.18 J g<sup>-1</sup> K<sup>-1</sup>**. That means that it takes 4.18 joules to increase the temperature of 1 gram of water by 1 K (or 1°C).

In experiments of this sort, you will quite commonly have measured out a volume of solution rather than a mass, and it is very common to assume that the density of the solution is exactly the same as water (1 g cm<sup>-3</sup>). That means that 100 cm<sup>3</sup> of solution would have a mass of 100 g. That is an approximation, though, and it does introduce some error.

The assumption is also made that the specific heat capacity of a solution is the same as the specific heat capacity of water. That's another approximation, and it also introduces errors into the answer.

What is generally true is that these errors are relatively small compared with errors caused by heat losses.

To do the calculation, you would normally just work out the amount of heat evolved or absorbed in your particular reaction, and then scale it up to give an enthalpy change in **kilo joules per mole**.

### Example

In an experiment, 100 g of pure water was weighed into a styrofoam cup and 7.58 g of highly pure  $\text{MgSO}_4$  was added and stirred until it was completely dissolved. The temperature of the water rose from  $20.0^\circ\text{C}$  to  $31.5^\circ\text{C}$ . Calculate the enthalpy change of solution, in  $\text{kJ mol}^{-1}$ , of  $\text{MgSO}_4(\text{s})$ .



[Assume that the specific heat capacity of the solution is the same as that of water,  $4.18 \text{ J g}^{-1}\text{C}^{-1}$  and ignore the mass of the solute.]

Step 1 Calculate the amount of heat energy, **in joules**, transferred to the water during this dissolving process.

$$\begin{aligned} q &= cm\Delta T \\ &= 4.18 \times 100 \times (31.5 - 20.0) \\ &= 4807 \text{ J} \end{aligned}$$

Step 2 Calculate the amount, **in moles**, of  $\text{MgSO}_4(\text{s})$  dissolved.

$$\begin{aligned} M(\text{MgSO}_4) &= 24.3 + 32.0 + 4 \times 16.0 = 120.3 \text{ g mol}^{-1} \\ n(\text{MgSO}_4) &= \frac{m}{M} = \frac{7.58}{120.3} = 0.0630 \text{ mol} \end{aligned}$$

Step 3 Calculate the enthalpy change of solution, **in  $\text{kJ mol}^{-1}$** , of  $\text{MgSO}_4(\text{s})$ .

$$\Delta H = -\frac{q}{n} = -\frac{4807 \div 1000}{0.0630} = -76.3 \text{ kJ mol}^{-1}$$

**Remember to include a sign with your answer, either + or –, representing endothermic or exothermic.**

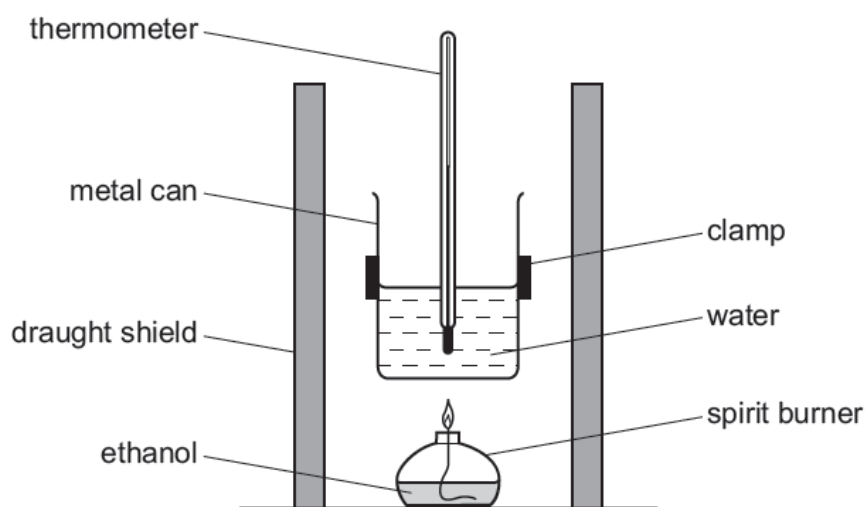
## 2. Enthalpy changes of combustion

As normally measured in a lab at this level, these are less accurate than the simple solution reactions above. Because it is impossible to eliminate heat losses, or to be sure that you have complete combustion. However, we can still use the similar method to calculate enthalpy changes of combustion.

### *Example*

In an experiment to determine the standard enthalpy of combustion of ethanol, the following apparatus was used and some readings were obtained.

Apparatus:



Results:

Mass of water in the metal can	400 g
Mass of ethanol and burner at start	435.2 g
Mass of ethanol and burner at end	432.6 g
Temperature of water at start	22.5°C
Temperature of water at end	57.0°C



Calculate the standard enthalpy of combustion of ethanol,  $\Delta_c H^\circ(\text{C}_2\text{H}_5\text{OH}(l))$ , in  $\text{kJ mol}^{-1}$ . The specific heat capacity of water is  $4.18 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ . Ignore the heat taken in by the metal can.

Step 1 Calculate the amount of heat energy, **in joules**, the water absorbed.

$$\begin{aligned} q &= cm\Delta T = 4.18 \times 400 \times (57.0 - 22.5) \\ &= 4.18 \times 400 \times 34.5 \\ &= 57684 \text{ J} \end{aligned}$$

Step 2 Calculate the amount, **in moles**, of ethanol burnt.

$$\begin{aligned} M(\text{C}_2\text{H}_5\text{OH}) &= 12 \times 2 + 6 + 16 = 46 \text{ g mol}^{-1} \\ n(\text{C}_2\text{H}_5\text{OH}) &= \frac{m}{M} = \frac{435.2 - 432.6}{46} = \frac{2.60}{46} = 0.05652 \text{ mol} \end{aligned}$$

Step 3 Calculate the enthalpy change of combustion, **in  $\text{kJ mol}^{-1}$** , of ethanol.

$$\Delta_c H(\text{C}_2\text{H}_5\text{OH}(l)) = -\frac{q}{n} = -\frac{57684 \div 1000}{0.05652} = -1021 \text{ kJ mol}^{-1}$$

**You must include a negative sign with your answer as enthalpy of combustion is always exothermic.**

## Questions 1

1. What is meant by each of the following terms?
  - (a) Allotropes
  - (b) Absolute zero
  - (c) Sublimation
  - (d) Element
  - (e) Compound
  
2. What kind of particles (atom, ion or molecule) is present in a sample of the following substance?
  - a. copper metal
  - b. water
  - c. gaseous helium
  - d. liquid oxygen
  - e. solid sodium chloride
  - f. carbon dioxide gas

## Questions 2

1.
  - (a) What is measured by temperature?
  
  
  
  
  
  
  
  
  
  
  - (b) Explain the movement of particles at absolute zero.
  
2. Convert the following temperatures to Kelvin.
  - (a) 0 °C
  - (b) 25 °C
  - (c) 100 °C
  - (d) 196 °C

### Questions 3

1. Classify the following processes and reactions as being either endothermic or exothermic:



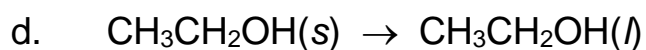
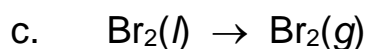
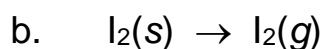
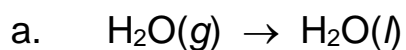
c. Ice turning to water

d. A bush fire

e. Photosynthesis

f. Dissolving concentrated sulfuric acid in water to cause the temperature of the water to rise.

2. Name each of the following physical changes.



(i) Which of these processes are endothermic?

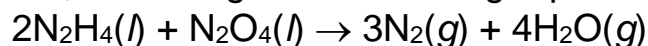
(ii) Which of these processes are exothermic?

#### Questions 4

1. Predict whether the standard enthalpy of formation of each of the following will be zero or positive. Explain your answers.



2. The powerful rocket fuel hydrazine ( $\text{N}_2\text{H}_4$ ), reacts with the oxidising agent, dinitrogen tetroxide, according to the following equation:



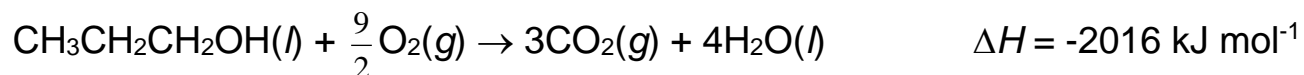
Calculate the enthalpy change for the reaction above, given the following data for standard molar enthalpies of formation:

$$\Delta_f H^\circ(\text{N}_2\text{H}_4(l)) = +50 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ(\text{N}_2\text{O}_4(l)) = -20 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ(\text{H}_2\text{O}(g)) = -242 \text{ kJ mol}^{-1}$$

3. When one mole of propan-1-ol is ignited under standard conditions according to the equation:



- a. Sketch a diagram showing how the total energy content of this reaction system changes during the course of this reaction.

- b. Given the additional data at 25°C

$$\Delta_f H^\circ(\text{H}_2\text{O}(l)) = -285 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ(\text{CO}_2(g)) = -393 \text{ kJ mol}^{-1}$$

Calculate the standard heat of formation for propan-1-ol.

## Questions 5

1. Given the following data

$$\Delta_f H^\circ[\text{CO}_2(g)] = -394 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ[\text{H}_2\text{O}(l)] = -285 \text{ kJ mol}^{-1}$$

$$\Delta_c H^\circ[\text{C}_6\text{H}_6(l)] = -780.1 \text{ kJ mol}^{-1}$$

Calculate the value of  $\Delta_f H^\circ[\text{C}_6\text{H}_6(l)]$

2. The Contact process for the production of sulfuric acid follows three distinct stages:
- (i) The burning of sulfur
  - (ii) The catalytic oxidation of sulfur dioxide
  - (iii) Dissolving  $\text{SO}_3$  in concentrated sulfuric acid and diluting with water.

### DATA

$$\Delta_c H^\circ[\text{S}(s)] = -297 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ[\text{SO}_2(g)] = -297 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ[\text{H}_2\text{O}(l)] = -285 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ[\text{SO}_3(g)] = -396 \text{ kJ mol}^{-1}$$

$$\Delta_f H^\circ[\text{H}_2\text{SO}_4(l)] = -814 \text{ kJ mol}^{-1}$$

- a. Explain in full what is meant by the expression,  $\Delta_c H^\circ[\text{S}(s)] = -297 \text{ kJ mol}^{-1}$ .

- b. Using the definitions for  $\Delta_c H^\circ[\text{S(s)}]$  and  $\Delta_f H^\circ[\text{SO}_2(\text{g})]$  account for the fact that the two values are equal.
- c. Use the  $\Delta_f H^\circ$  values to calculate the enthalpy change for the reaction
- $$\text{H}_2\text{O}(\text{l}) + \text{SO}_3(\text{g}) \rightarrow \text{H}_2\text{SO}_4(\text{l})$$

### Questions 6

Calculate the enthalpy changes for the following reactions using the average bond energies in the table on page 13.

