CHAPTER

2 Energy calculations

'Energy is neither created nor destroyed; it can only be converted from one form to another'; this is one of the famous laws of thermodynamics. Energy powers the way we live, but where does this energy come from? The most concentrated form of energy is held in the bonds of chemicals we live with every day. How can we determine the best choices for fuels to produce this energy? How can we make quantitative predictions that match energy requirements to fuel quantities?

YOU WILL EXAMINE:

- the meaning of enthalpy and the ΔH notation
- energy profile diagrams for exothermic and endothermic reactions
- thermochemical equations and subsequent calculations that can be made using them
- the calculation of the ΔH value for a reaction and from experimental data
- the kinetic molecular theory of gases
- the concept of gas pressure and common units used to measure it
- the absolute (or Kelvin) temperature scale
- Boyle's law and Charles' law
- standard conditions for measuring gas volumes and molar volume
- the universal gas equation
- mass–volume and volume–volume stoichiometry
- the calculation of greenhouse gas amounts associated with the production of energy from carbon-based fuels
- how ΔH values may be estimated from simple laboratory apparatus.

Many modern homes and buildings use 'burning log' fireplaces for their heating. Such fireplaces do not burn logs at all — they merely give the illusion that they are. In many cases, their heat output comes from fuels such as ethanol, natural gas and LPG. In designing such fireplaces, it is important to know the energy output of the chosen fuel so that combustion can produce the required amount of energy. It is also important to control the fuel to oxygen ratio so that unwanted and dangerous combustion products, such as carbon monoxide, are avoided.

Energy changes in chemical reactions

The energy changes that accompany chemical reactions are vital to us. To survive, we depend on the energy content of the food we eat. Our bodies can convert the energy of the chemical bonds in food into other kinds of energy. The quality of lifestyle we lead depends on harnessing energy from different chemical sources, including coal, oil, natural gas and renewable fuels. The study of the energy changes that accompany chemical reactions is called **thermochemistry** or **chemical thermodynamics**. In general, all chemical reactions involve energy changes. The chemical energy stored in a substance has the potential to be converted to heat (which will be discussed in this chapter) or electricity (see chapter 3).

A certain amount of chemical energy is stored within every atom, molecule or ion. This energy is the sum of the potential energy and kinetic energy of the substance and results from:

- the attractions and repulsions present between protons and electrons within the atom
- the attractions and, to some degree, repulsions present between atoms within a molecule
- the motion of the electrons
- the movement of the atoms.

A change in enthalpy occurs when magnesium burns.

The total energy stored in a substance is called the **enthalpy**, or **heat content**, of the substance and is given the symbol H. Unfortunately we cannot directly measure the heat content of a substance. What we can measure is the **change in enthalpy** when the substance undergoes a chemical reaction. Since, in virtually all chemical reactions, the energy of the reactants and products differ, such reactions usually involve some change in enthalpy, which is indicated by a temperature rise or fall. The change in enthalpy during a reaction is known as the **heat of reaction** and is denoted by ΔH.

Let us consider the following reaction:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Magnesium has a certain enthalpy. So too does the oxygen. The total enthalpy of the reactants — magnesium and oxygen — can be represented by the symbol H_r. The product, magnesium oxide, also has some enthalpy — call this H_p.

We can say that:

change in heat content = (heat content of products) – (heat content of reactants)

So, enthalpy change is equal to heat energy produced or absorbed.

$$\Delta H = H_p - H_r$$

Enthalpy is the heat content of a substance. It cannot be measured directly, but the change in enthalpy, known as the heat of reaction, ΔH, can be measured.

The enthalpy change, or heat change, of a reaction depends on the amount of reactants used and the temperature of the initial reactants, compared with that of the final products. To remove these variables in enthalpy studies, the following conditions are assumed:

- the initial temperature of the reactants and final temperature of the products is the same, and is 25 °C
- 1 mole of the substance is involved
- solutions have a concentration of 1 M
- the heat absorbed or released is measured in kilojoules (kJ)
- the pressure is kept at 100.0 kPa.

These conditions are referred to as *standard laboratory conditions* (SLC).

Enthalpy changes are stated at SLC; they refer to the energy needed to return the temperature to 25 °C and the pressure to 100.0 kPa.

Although the usual name for ΔH is heat of reaction (or change in enthalpy), there are some reactions for which specific names have been given.

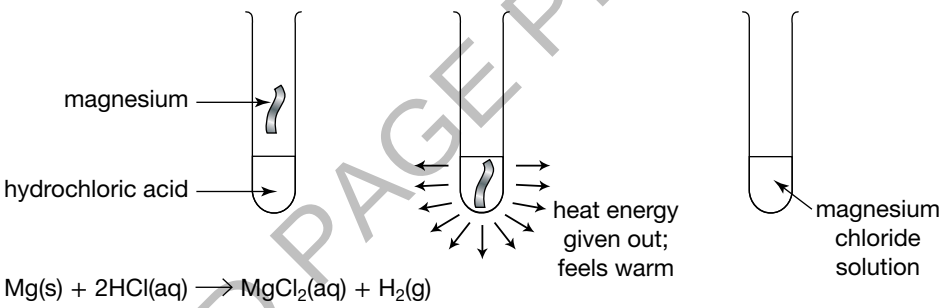
- *Heat of solution* is the change in enthalpy when 1 mole of any substance dissolves in water.
- *Heat of neutralisation* is the change in enthalpy when an acid reacts with a base to form 1 mole of water.
- *Heat of vaporisation* is the change in enthalpy when 1 mole of liquid is converted to a gas.
- *Heat of combustion* is the enthalpy change when a substance burns in air, and is always exothermic.

Exothermic and endothermic reactions

Chemical reactions accompanied by heat energy changes may be divided into two groups: exothermic reactions and endothermic reactions.

Exothermic reactions

Chemical reactions that release heat to the environment are called **exothermic** reactions.



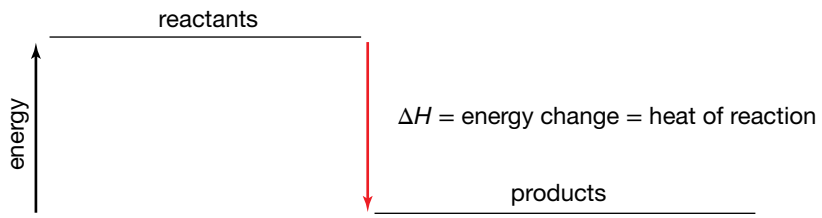
An exothermic reaction — magnesium ribbon dissolving in hydrochloric acid

Exothermic reactions release heat. The enthalpy change, ΔH, is negative.

When a strip of magnesium is placed in hydrochloric acid, heat is given out. This means that the enthalpy of the reactants, H_r, is greater than the enthalpy of the products, H_p. Therefore, the enthalpy change, $\Delta H = H_p - H_r$, is *negative*. The difference in energy between the reactants and the products is released into the environment, usually in the form of heat.

The energy change in a reaction can be drawn as an energy diagram or energy profile. Chemicals with more energy are drawn higher up and those with less energy lower down. The reactants are drawn on the left and the products are drawn on the right. The figure below is an energy diagram for an exothermic reaction.

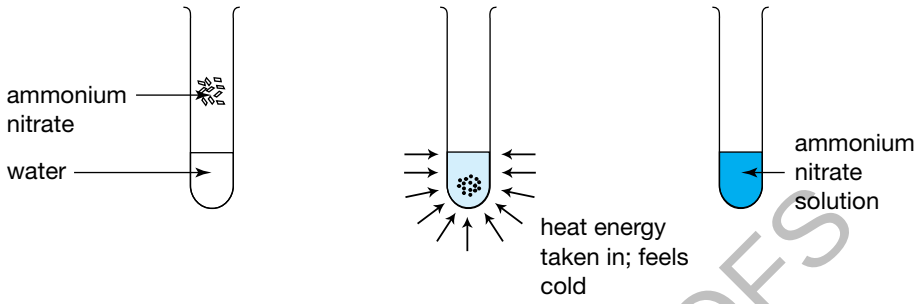
Endothermic reactions absorb heat. The enthalpy change, ΔH, is positive.



In an exothermic reaction, ΔH is always negative because the heat content of the reactants is greater than that of the products. The bonds in the products are more stable than the bonds in the reactants.

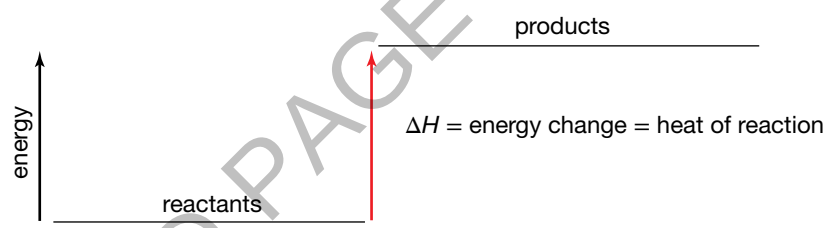
Endothermic reactions

Chemical reactions that absorb heat from the environment are called **endothermic** reactions.



An endothermic reaction — ammonium nitrate dissolving in water

When ammonium nitrate is dissolved in water, heat is absorbed from the environment; the test tube feels cold because the reaction takes in energy from the water and your hand. This means that the enthalpy of the products, H_p, is greater than the enthalpy of the reactants, H_r. Therefore, the enthalpy change, $\Delta H = H_p - H_r$, is *positive*. The difference in energy between the reactants and the products is absorbed from the environment. The figure below is an energy diagram for an endothermic reaction.



In an endothermic reaction, ΔH is always positive because the heat content of the reactants is less than that of the products. The bonds in the products are less stable than the bonds in the reactants.

eBook plus

Digital document
Experiment 2.1 Investigating heat changes in reactions
doc-18833

Revision question

1. **State whether the following are exothermic or endothermic processes.**
   - (a) **Water changing from liquid to gaseous state**
   - (b) **A reaction in which the total enthalpy of the products is greater than that of the reactants**
   - (c) **Burning kerosene in a blow torch**
   - (d) **Burning kerosene in a jet aircraft engine**
   - (e) **A chemical reaction that has a negative ΔH value**
   - (f) **A reaction in which the reactants are at a lower level on an energy profile diagram than the products**

Thermochemical equations

A thermochemical equation is a balanced equation that includes the amount of heat produced or absorbed.

As we saw briefly in chapter 1, an equation that includes the amount of heat produced or absorbed by a reaction is a **thermochemical equation**. As with other chemical equations, charge and mass must balance. However, they must also include the enthalpy change.

In writing a thermochemical equation, the following points should be remembered:

- A positive or negative sign must be included with the ΔH value to indicate whether the reaction is either endothermic or exothermic. If an enthalpy

change is given as $\Delta H = 345\ \text{kJ mol}^{-1}$, the lack of sign does not mean that it is an endothermic reaction.

- Since enthalpy is measured in the units kJ mol^{-1}, the coefficients in the equation represent the amount in moles of each reacting substance. So, the reaction

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) \quad \Delta H = -572\ \text{kJ mol}^{-1}$$

can be read as: when 2 moles of hydrogen react with 1 mole of oxygen, 2 moles of water form and 572 kJ of energy is produced. ΔH refers to the equation as it is written, even though the unit is expressed as kJ mol^{-1}.

- The physical state of matter must be shown, since changes of state require energy changes.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) \quad \Delta H = -572\ \text{kJ mol}^{-1}$$
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \quad \Delta H = -484\ \text{kJ mol}^{-1}$$

In one case, the product is a liquid. In the other case, the product is a gas. So, the condensation of 2 moles of water vapour to 2 moles of liquid water at 25 °C produces 88 kJ of energy, which is the difference between the two enthalpies.

- If the coefficients are doubled, the ΔH value must be doubled.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l) \quad \Delta H = -572\ \text{kJ mol}^{-1}$$
$$4H_2(g) + 2O_2(g) \longrightarrow 4H_2O(l) \quad \Delta H = -1144\ \text{kJ mol}^{-1}$$

The amount of energy produced by a chemical reaction is directly proportional to the amount of substance initially present. If, as in the above example, twice as much reactant is used, then twice as much energy can be produced.

- If a reaction is reversed, ΔH is equal to, but opposite in sign, to that of the forward reaction.

$$2H_2O(l) \longrightarrow 2H_2(g) + O_2(g) \quad \Delta H = +572\ \text{kJ mol}^{-1}$$

You will notice that the key point is that the enthalpy change in a reaction is proportional to the amount of substance that reacts. If these two quantities are measured in an experiment, it is possible to write the accompanying thermochemical equation. This was covered in chapter 1. Sample problem 2.1 provides another example of how this can be done.

studyon

Unit 3 | AOS 1 | Topic 1 | Concept 3

Do more
Combustion equations

Sample problem 2.1

In an experiment, it was found that the combustion of 0.240 g of methanol in excess oxygen yielded 5.42 kJ. Write the thermochemical equation for this reaction.

Solution: The equation for this reaction is:

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l)$$

To match this equation, it is necessary to calculate the heat evolved from 2 moles (64.0 g) of methanol. This can be done by scaling up the results from the experiment as follows.

$$\text{Heat evolved from 2 moles (64.0 g)} = \frac{64.0}{0.240} \times 5.42 = 1.45 \times 10^3\ \text{kJ}$$

The thermochemical equation is therefore:

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l) \quad \Delta H = -1.45 \times 10^3\ \text{kJ mol}^{-1}$$

Sample problem 2.2

The molar heat of combustion of ethanol is tabulated as –1364 kJ mol^{-1}.

(a) Write the thermochemical equation for the combustion of ethanol.

(b) If the density of ethanol is 0.790 g mL^{-1}, calculate the energy evolved in MJ when 1.00 L of ethanol is burned.

Solution: (a) The balanced chemical equation for the combustion of ethanol is:

$$CH_3CH_2OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$$

The given heat value refers to 1 mole, as does the equation above.

The thermochemical equation is therefore:

$$CH_3CH_2OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1364\ \text{kJ mol}^{-1}$$

(b) Since density $= \dfrac{\text{mass}}{\text{volume}}$

mass of ethanol burned $(m) = d(\text{ethanol}) \times V(\text{ethanol})$
$= 0.790 \times 1000$
$= 790\ \text{g}$

$n(\text{ethanol}) = \dfrac{790}{46.0} = 17.17\ \text{mol}$

From the equation, 1 mole of ethanol evolves 1364 kJ.

By direct proportion, 17.17 moles produce x kJ.

Therefore, $x = 17.17 \times \dfrac{1364}{1} = 23\,425\ \text{kJ} = 23.4\ \text{MJ}$.

Sample problem 2.3

The combustion of ethene can be represented by the following thermochemical equation.

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(l) \quad \Delta H = -1409\ \text{kJ mol}^{-1}$$

Calculate the mass of ethene required to produce 500 kJ of heat energy.

Solution: From the equation, 1 mole of C_2H_4 evolves 1409 kJ.

By ratio, x moles are required to evolve 500 kJ.

$$x = 500 \times \frac{1}{1409} = 0.355\ \text{mol}$$

Therefore, $m(C_2H_4) = 0.355 \times 28.0 = 9.94\ \text{g}$.

Sample problem 2.4

The air pollutant sulfur trioxide reacts with water in the atmosphere to produce sulfuric acid according to the equation:

$$SO_3(g) + H_2O(l) \longrightarrow H_2SO_4(aq) \quad \Delta H = -129.6\ \text{kJ mol}^{-1}$$

Calculate the energy released, in kJ, when 0.500 kg $SO_3(g)$ reacts with water.

Solution:

$$n(SO_3) = \frac{m}{M} = \frac{500}{80.1} = 6.24\ \text{mol}$$

From the equation, 1 mole of SO_3 releases 129.6 kJ.

So, 6.24 moles release $6.24 \times 129.6 = 809$ kJ of heat energy.

Revision questions

2. **Calculate the energy released, in kJ, when 3.56 g of carbon undergoes combustion according to the thermochemical equation:**

$$2C(s) + O_2(g) \longrightarrow 2CO(g) \quad \Delta H = -788\ \text{kJ mol}^{-1}$$

3. **The use of hydrogen as a renewable and environmentally friendly fuel is currently the subject of much research. The main product of hydrogen**

combustion is water. The production of liquid water from the reaction between gaseous hydrogen and gaseous oxygen can be represented by the following thermochemical equation.

$$O_2(g) + 2H_2(g) \longrightarrow 2H_2O(l) \quad \Delta H = -572\ \text{kJ mol}^{-1}$$

Calculate how much energy, in kilojoules, would be released or absorbed by the following reactions.

(a) $2O_2(g) + 4H_2(g) \longrightarrow 4H_2O(l)$

(b) $H_2O(l) \longrightarrow \frac{1}{2}O_2(g) + H_2(g)$

4. During the production of sulfuric acid by the Contact process, sulfur dioxide is converted to sulfur trioxide according to the equation:

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g) \quad \Delta H = -197\ \text{kJ mol}^{-1}$$

Calculate the heat energy released in the production of 1.00 tonne (10^6 g) of sulfur trioxide gas.

5. Calculate the energy released when 18.5 g of carbon undergoes combustion in a plentiful supply of air according to the equation:

$$C(s) + O_2(g) \longrightarrow CO_2(g) \quad \Delta H = -394\ \text{kJ mol}^{-1}$$

6. Methanol burns according to the equation:

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l) \quad \Delta H = -1450\ \text{kJ mol}^{-1}$$

Calculate the mass of methanol required to produce 1.000 MJ of energy.

7. Butane and octane are two hydrocarbons commonly used as fuels. The thermochemical equations for these two fuels are shown below.

$$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(l) \quad \Delta H = -5748\ \text{kJ mol}^{-1}$$
$$2C_8H_{18}(g) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(l) \quad \Delta H = -10\,928\ \text{kJ mol}^{-1}$$

(a) Calculate the heat evolved by the combustion of 100 g of butane.

(b) Use your answer to (a) to calculate the mass of octane required to produce the same amount of energy.

8. Explain why it is critical to show symbols of state in thermochemical equations.

Calculating ΔH values from two or more related reactions

The ΔH value of a reaction can be calculated by manipulating and then adding together two or more related reactions and their associated ΔH values.

If the enthalpy change of a desired reaction is not known, it is possible to calculate it from a series of related reactions. The calculated value may then be used to make thermochemical predictions as shown previously. To do this, use some of the dot points mentioned on pages 34–35. The following sample problems show how this is done.

Sample problem 2.5

Calculate the enthalpy change for the incomplete combustion of methane in a limited oxygen supply, given the following two equations.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -889\ \text{kJ mol}^{-1} \quad \text{(equation 1)}$$

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g) \quad \Delta H = -556\ \text{kJ mol}^{-1} \quad \text{(equation 2)}$$

Solution: The required equation is $2CH_4(g) + 3O_2(g) \longrightarrow 2CO(g) + 4H_2O(l)$ (equation 3)

Equations 1 and 2 need to be 'manipulated' so that, when added together, equation 3 is produced. This can be achieved by multiplying equation 1 by 2 and reversing equation 2.

$$2CH_4(g) + 4O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l) \quad \Delta H = -1778\ \text{kJ mol}^{-1}$$

$$2CO_2(g) \longrightarrow 2CO(g) + O_2(g) \quad \Delta H = +556\ \text{kJ mol}^{-1}$$

Adding together results in cancelling to produce the required equation. The ΔH values are also added, thus producing a value of –1778 + 556 = –1222.

Therefore, $2CH_4(g) + 3O_2(g) \longrightarrow 2CO(g) + 4H_2O(l)$ $\Delta H = -1222\ kJ\ mol^{-1}$.

Sample problem 2.6

Fossil fuels such as coal and petroleum contain sulfur as an impurity. This produces sulfur oxides when they are burned. These are atmospheric pollutants. Develop the thermochemical equation for the reaction:

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

given the following equations.

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g) \quad \Delta H = -196\ kJ\ mol^{-1} \qquad \text{(equation 1)}$$

$$2S(s) + 3O_2(g) \longrightarrow 2SO_3(g) \quad \Delta H = -760\ kJ\ mol^{-1} \qquad \text{(equation 2)}$$

Solution: Equation 1 needs to be halved and reversed. Equation 2 needs to be halved. The ΔH values thus become +98 and –380 respectively. Adding the adjusted equations together produces the required thermochemical equation.

$$S(s) + O_2(g) \longrightarrow SO_2(g) \quad \Delta H = -282\ kJ\ mol^{-1}$$

Revision questions

9. Calculate the ΔH value for the incomplete combustion of pentane, according to the equation:

$$2C_5H_{12}(l) + 11O_2(g) \longrightarrow 10CO(g) + 12H_2O(l)$$

given the following equations.

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g) \quad \Delta H = -556\ kJ\ mol^{-1}$$

$$C_5H_{12}(l) + 8O_2(g) \longrightarrow 5CO_2(g) + 6H_2O(l) \quad \Delta H = -3509\ kJ\ mol^{-1}$$

10. Calculate ΔH for the reaction:

$$CH_3CH_2OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(g)$$

given the following equations.

$$CH_3CH_2OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H = -1364\ kJ\ mol^{-1}$$

$$H_2O(l) \longrightarrow H_2O(g) \quad \Delta H = +44.0\ kJ\ mol^{-1}$$

Working with gases

When considering fuels as energy sources, it is often necessary to deal with gases. The products of combustion are nearly always gases, and some of the fuels themselves are gases. While it is possible to measure their masses, it is often more convenient and meaningful to measure their volumes.

Gases have properties and exhibit behaviours that are different from liquids and solids. Scientists use the **kinetic molecular theory of gases** to explain observed gas properties. This consists of five *postulates* (or points) that form a mental picture of how the particles in a gas would look and behave if we could observe them directly. These are applicable to gas samples under 'moderate conditions', which is usually taken to mean pressures that are not much greater than atmospheric pressure and temperatures considerably greater than those at which the gases liquefy.

The kinetic molecular theory states that matter is made up of continuously moving particles.

Hot air balloons use the properties of gases to achieve flight. A large fan is used to blow air into the balloon. Once inflated, the air is heated using a propane burner. The molecules move more rapidly, hitting each other in a random chaotic motion. As they move further apart, the density of the air decreases. The balloon rises because the air inside is warmer, and therefore lighter, than the surrounding air.

The five postulates that make up the kinetic molecular theory of gases can be summarised as follows.

1. Gases are made up of particles moving constantly and at random.
2. Gas particles are very far apart, and the volume of the particles is very small compared with the volume that the gas occupies.
3. The forces of attraction and repulsion between gas particles are practically zero.
4. Gas particles collide with each other and the walls of their container, exerting pressure. The collisions are perfectly elastic. This means that no kinetic energy is lost when they collide.
5. The higher the temperature, the faster the gas particles move, as they have increased kinetic energy.

Specifying the large-scale behaviour of gases

In any consideration of gas behaviour, the *pressure* each gas exerts, the *volume* that it occupies, its *temperature* and the *number of gas molecules* present in the sample must be determined.

Gas pressure

Pressure is the force exerted by gas particles on the walls of a container.

Particles exert a force by colliding with the walls of a container. Each tiny collision adds to all the others to make up the continuous force that we call **pressure**.

The surface of an inflatable airbed, such as the one shown here, exerts a force and tries to collapse. For the airbed to stay inflated, the particles inside the airbed must be able to exert a large enough force to balance the forces exerted by the surface and the external air pressure.

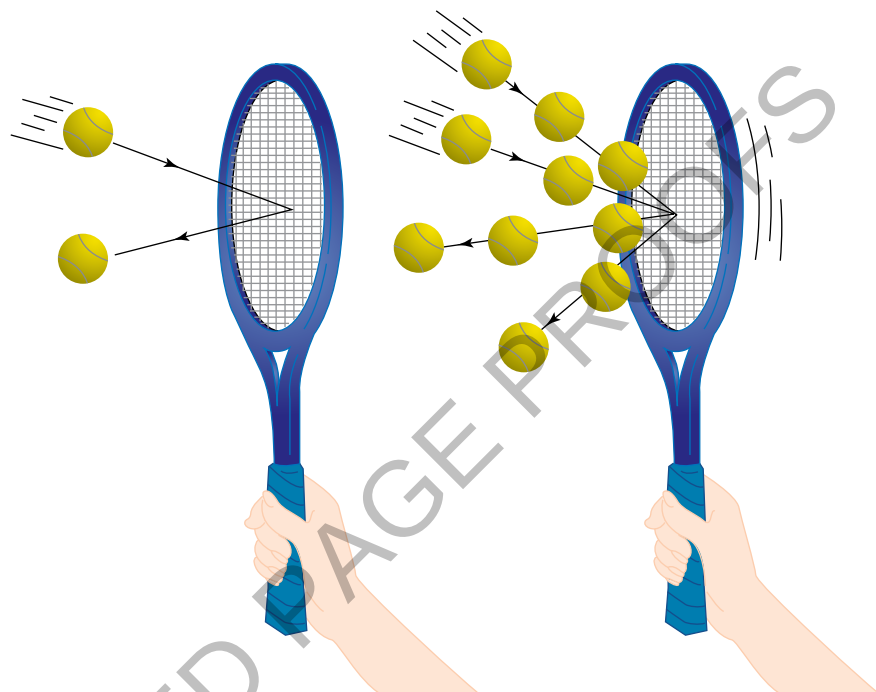
Pressure (P) is defined as *the force exerted per unit area*.

Pressure is force per unit area.

$$P = \frac{\text{force}}{\text{area}}$$

Pressure is measured using various units.

The SI unit of pressure is the **pascal (Pa)**, where 1 pascal is equivalent to a force of 1 newton exerted over an area of 1 square metre (N m^{-2}).



If a tennis racquet is struck by a tennis ball, a small force is felt. If a stream of balls is fired at the racquet, it is felt as a continuous pressure.

Atmospheric pressure

Units of pressure: 760 mmHg = 1.00 atm = 101.3 kPa

The pressure of the atmosphere is measured by a barometer. Atmospheric pressure at sea level is 101 325 Pa. This is usually simplified to read 101.3 kPa (kilopascal). Two older units of pressure are *millimetres of mercury* and *atmospheres*. One millimetre of mercury (1 mmHg) is defined as the pressure needed to support a column of mercury 1 mm high. This unit developed from the early use of mercury barometers. One atmosphere (1 atm) is the pressure required to support 760 mm of mercury (760 mmHg) in a mercury barometer at 25 °C. This is the average atmospheric pressure at sea level. So, 1 atm equals 760 mmHg.

$$760 \text{ mmHg} = 1.00 \text{ atm} = 1.013 \times 10^5 \text{ Pa} = 101.3 \text{ kPa}$$

In weather reports, *bars* and *hectopascals* are used to measure gas pressure.

$$1 \text{ bar} = 100\,000 \text{ Pa} = 100 \text{ kilopascal (kPa)}$$

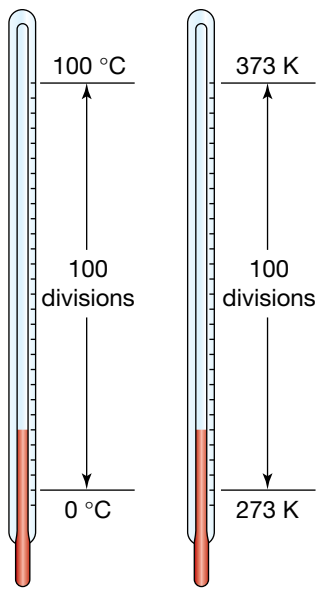
$$1 \text{ millibar (mb)} = \frac{1}{1000} \text{ bar} = 100 \text{ Pa} = 1 \text{ hectopascal (hPa)}$$

Temperature

Units of temperature: 0 °C = 273 K

The Celsius scale takes the freezing point of water as 0 °C and the boiling point of water as 100 °C. The space between these two fixed points is divided into 100 equal intervals, or degrees. Temperatures below the freezing point of water are assigned negative values, such as –10 °C.

Another temperature scale is the **Kelvin**, or absolute, scale. On the Kelvin scale, the freezing point of water is 273 K and its boiling point is 373 K. Notice that a change of 1° on the Celsius scale is the same as that on the Kelvin scale. The zero point on the Kelvin scale, 0 K or **absolute zero**, is –273 °C. The



These thermometers show a comparison of the Celsius and Kelvin temperature scales.

relationship between the temperature on the Celsius scale and that on the Kelvin scale is given by the following equations.

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

For example, to convert 25 °C to the absolute scale:

$$K = 25 + 273$$
$$= 298$$

Note that temperatures given in K do not have a ° sign.

Volume

The volume of a gas is commonly measured in cubic metres (m^3) or litres (L) or millilitres (mL).

$$1\ m^3 = 1000\ L = 1\ 000\ 000\ mL$$

or

$$1\ m^3 = 10^3\ L = 10^6\ mL$$

Units of volume:
$1\ m^3 = 10^3\ L = 10^6\ mL$

Revision question

11. Make the following conversions.

(a) 780 mmHg to atm
(b) 4.0 atm to Pa
(c) 1000 mmHg to Pa
(d) 1250 mmHg to kPa
(e) 200 °C to K
(f) 500 K to °C
(g) 3.0 m^3 to L
(h) 250 L to mL
(i) 1600 mL to L
(j) 3×10^6 mL to L
(k) 5×10^3 mL to m^3
(l) 600 mL to m^3

eBook*plus*

Digital document
Experiment 2.2 The relationship between pressure and volume
doc-18834

Laws to describe the behaviour of gases

The behaviour of gases has been studied for centuries. As a result, a number of laws have evolved to describe their behaviour mathematically. These laws are independent of the type of gas — it does not matter what the gas is (or even if the gas is a mixture such as air), these laws apply in exactly the same way to all gases. Two very useful such laws are Boyle's law and Charles' law.

Boyle's law for pressure–volume changes

Boyle's law is named after the English physicist and chemist who discovered the relationship between pressure and volume for a sample of gas. It states that, for a fixed amount of gas at constant temperature, pressure is inversely proportional to volume. Mathematically, this can be represented as:

$$P \propto \frac{1}{V}$$

from which it can be stated that PV = a constant value.

If P_1V_1 = a constant value and P_2V_2 = a constant value, then:

$$P_1V_1 = P_2V_2$$

Revision questions

12. (a) Set up a spreadsheet and enter the data from the table on the next page into the cells of the first two columns.
(b) In the third column of the spreadsheet, calculate pressure × volume. What pattern do you notice?

(c) In the fourth column of your spreadsheet, convert each pressure measurement (P) into its reciprocal $\left(\frac{1}{P}\right)$. For example, if P has a value of 120, $\left(\frac{1}{P}\right)$ would equal 8.33×10^{-3}.

13. (a) Use the data on your spreadsheet from question 12 to plot a graph of pressure versus volume. Put pressure on the horizontal (x) axis and volume on the vertical (y) axis.

(b) Plot a second graph with the volume on the vertical axis and $\frac{1}{P}$ on the horizontal axis, and extrapolate the graph to the origin.

(c) Compare and account for the two graphs.

Sample results showing the relationship between the pressure and volume of a gas

Pressure (kPa)	Volume (L)
120	0.261
145	0.218
162	0.193
180	0.171
200	0.159
216	0.145
240	0.130
258	0.120

Charles' law for temperature–volume changes

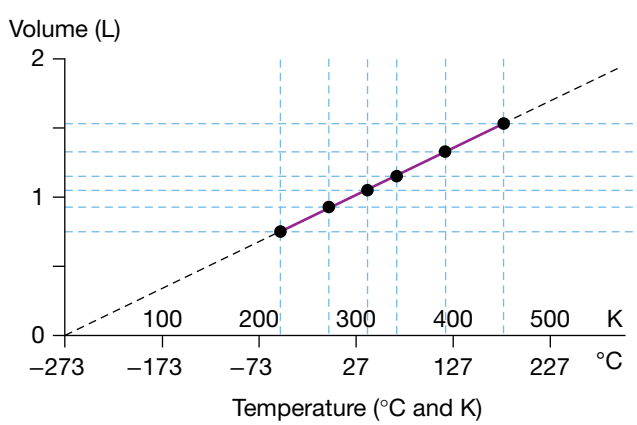
The relationship between temperature and volume was first identified by the French scientist Jacques Charles, after whom the law is named. While the expansion of all materials with increasing temperature is well known, Charles subjected constant amounts of various gases (at constant pressure) to changes in temperature, each time making accurate measurements of the resulting volume.

To understand this law better, consider a typical set of results as shown in table 2.1.

If these results are graphed as shown below, a pattern emerges.

TABLE 2.1 Sample results showing the relationship between the temperature and volume of a gas

Temperature (°C)	Volume (L)
–50	0.75
0	0.92
40	1.05
70	1.15
120	1.32
180	1.52



Representation of Charles' law showing the two temperature scales, Kelvin and degrees Celsius. The graph demonstrates that the volume of a gas varies in direct proportion to absolute temperature.

If the Celsius scale is used to measure temperature, a linear relationship is observed. However, this is not a directly proportional relationship as the graph does not pass through the origin. If, however, the Kelvin scale is used, it does. We can therefore state Charles' law as follows.

For a given amount of gas at constant pressure, volume is directly proportional to the absolute temperature.

Mathematically, this can be represented as, $V \propto T$, (where T is the absolute temperature), from which it can be stated that $\frac{V}{T}$ = a constant value.

If $\frac{V_1}{T_1}$ = a constant value and $\frac{V_2}{T_2}$ = a constant value, then:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

You will notice that the graph on the previous page implies a 'lowest possible temperature' — the temperature at which it intersects the x-axis. This temperature is –273.15 °C or 0 K. It is referred to as 'absolute zero'.

The combined gas equation

Boyle's law and Charles' law may be combined to produce the gas equation:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

This equation applies to a fixed amount of gas. It is particularly useful for making predictions when an amount of gas under a set of initial conditions (P_1, V_1 and T_1) is changed to a second set of conditions (P_2, V_2 and T_2). If any two of the new set of conditions are known, the third can easily be calculated.

Sample problem 2.7

A sample of air at 10 °C and 1.1 atm has a volume of 2.2 L. If the pressure is changed to 1.0 atm and the temperature increases to 15 °C, what volume will it occupy?

Solution: **STEP 1**

List the known information:

$P_1 = 1.1$ atm
$V_1 = 2.2$ L
$T_1 = 10 + 273 = 283$ K

$P_2 = 1.0$ atm
$V_2 = ?$ L
$T_2 = 15 + 273 = 288$ K

STEP 2

Transpose the general gas equation and substitute the values above.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

$$= \frac{1.1 \times 2.2 \times 288}{1.0 \times 283}$$

$$= 2.5 \text{ L}$$

The new volume is, therefore, 2.5 L.

Note that units for pressure and volume must be the same on each side. Temperature must be in kelvin.

Standard conditions for measuring gases

Boyle's law and Charles' law tell us that the volume of a gas sample is sensitive to both temperature and pressure. This makes the comparison of gas volumes very tricky. In order to make these comparisons easier, scientists have established sets of standard conditions. These are accepted worldwide and enable gas volumes to be compared meaningfully without temperature or pressure having unwanted effects.

There are two commonly used such sets of standard conditions.

STP and SLC are standard sets of conditions for comparing gases.

1. **Standard temperature and pressure (STP)**. Standard temperature is 0 °C or 273 K, and standard pressure is 100.0 kPa.
2. **Standard laboratory conditions (SLC)**. Since most experiments are carried out in the laboratory, normal room temperature is substituted for the STP condition. Standard laboratory conditions are a temperature of 25 °C (298 K) and a pressure of 100.0 kPa.

This book will generally refer to SLC only, as STP is not part of the VCE Chemistry study design.

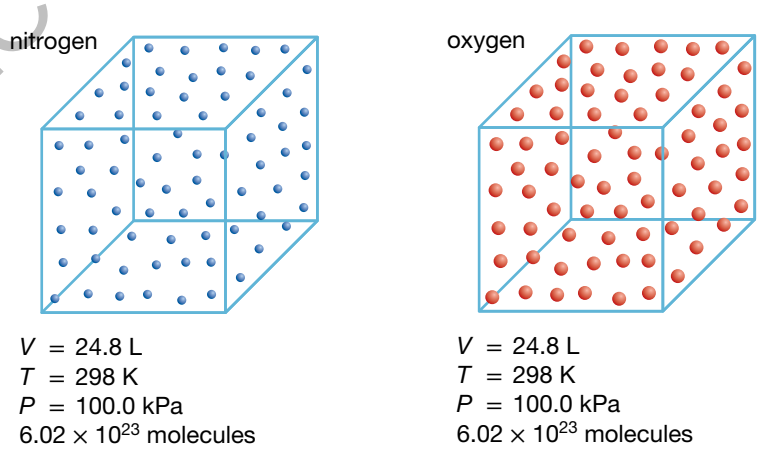
Molar gas volume and Avogadro's hypothesis

The molar gas volume is the volume of 1 mole of any gas at a stated temperature and pressure.
V_M (STP) = 22.7 L
V_M (SLC) = 24.8 L

Avogadro put forward the hypothesis that 'equal volumes of all gases measured at the same temperature and pressure contain the same number of particles.'

This means that, if two gases have the same temperature, pressure and volume, they must contain the same number of moles.

It has been found experimentally that 1 mole of any gas at SLC occupies a volume of 24.8 L. This volume is called the **molar gas volume**.



The standard molar gas volume. Under the same conditions of temperature and pressure, the volume of a gas depends only on the number of molecules it contains, and not on what the particles are. The volume occupied by 28.0 g of nitrogen gas (1 mol or 6.02×10^{23} molecules) at SLC is 24.8 L. The volume occupied by 1 mole of any gas at the same temperature and pressure is 24.8 L. (*Note:* In these diagrams, the number of particles shown is only a fraction of those actually present, since this number is too great to be represented here.)

The molar volume of a gas at SLC is 24.8 L. This means that 1 mole of any gas occupies 24.8 L at 25 °C and 100.0 kPa. The molar volume of a gas varies with temperature and pressure but, at any given temperature and pressure, it is the same for all gases. There is a direct relationship between the number of moles of a gas and its molar volume. This is illustrated by the formula

$$n = \frac{V}{V_m}$$

where n is the number of moles of gas, V is the actual volume and V_m is the molar volume of the gas. Therefore, at SLC:

$$n_{SLC} = \frac{V}{24.8}$$

where V is measured in L.

Sample problem 2.8

How many moles of chlorine gas are present in 48.0 L of the gas at SLC?

Solution:

$$n = \frac{V}{24.8}$$

$$n = \frac{48.0}{24.8} = 1.94 \text{ mol of } Cl_2 \text{ gas}$$

Sample problem 2.9

Find the mass of 1556.5 mL of H_2 gas that was collected at SLC.

Solution:

Why are these bubbles floating upwards? These soap bubbles are filled with hydrogen gas, which is lighter than the same volume of bubbles filled with air. A mole of hydrogen gas would occupy the same volume as a mole of oxygen molecules, but the hydrogen would not weigh as much as the oxygen.

STEP 1
Convert the volume to litres.

1556.5 mL = 1.5565 L

STEP 2
Determine the number of moles of H_2 gas using $n = \frac{V}{24.8}$

$$n = \frac{1.5565}{24.8}$$
$$= 0.0628 \text{ mol}$$

STEP 3
Determine the mass (m) of 0.0628 moles of H_2 gas using $n(H_2) = \frac{m}{M}$

$$0.0628 = \frac{m}{2.0}$$

The mass of H_2 gas collected was 0.13 g.

Sample problem 2.10

(a) 5345 mL of a gas was collected at SLC and weighed. Its mass was 9.50 g. Find the molar mass of the gas.
(b) Given that the gas is one of the main constituents of air, identify the gas.

Solution: (a) **STEP 1**
Convert the volume to litres.
5345 mL = 5.345 L

STEP 2

Determine the number of moles of the gas.

$$n = \frac{V}{24.8}$$
$$= \frac{5.345}{24.8}$$
$$= 0.216 \text{ mol}$$

STEP 3

Determine the molar mass (M) of 0.216 mole of the gas.

$$n = \frac{m}{M}$$
$$0.216 = \frac{9.50}{M}$$
$$M = \frac{9.50}{0.216}$$
$$= 44.0$$

The molar mass of the gas is 44.0 g mol^{-1}.

(b) The main constituents of air are nitrogen (M = 28.0), oxygen (M = 32.0), argon (M = 40.0) and carbon dioxide. The gas must therefore be carbon dioxide as the M of CO_2 = 12 + (2 × 16) = 44.0 g mol^{-1}.

Revision questions

14. Calculate the numbers of moles of the following gases at SLC.
- **(a) 15 L of oxygen, O_2**
- **(b) 25 L of chlorine, Cl_2**

15. Calculate the volumes of the following gases at SLC.
- **(a) 1.3 mol hydrogen, H_2**
- **(b) 3.6 g of methane, CH_4**
- **(c) 0.35 g of argon, Ar**

16. Calculate the masses of the following gas samples. All volumes are measured at SLC.
- **(a) 16.5 L of neon, Ne**
- **(b) 1050 mL of sulfur dioxide, SO_2**

17. What is the mass (in kg) of 850 L of carbon monoxide gas measured at SLC?

18. A 0.953 L quantity of a gas measured at SLC has a mass of 3.20 g. What is the molar mass of the gas? What is the gas?

The universal gas equation

The universal gas equation

$$PV = nRT$$

is used to calculate the amount of a gas at a given volume and temperature.

When Boyle's law and Charles' law are combined with Avogadro's hypothesis, the following equation is obtained.

$$PV = nRT$$

This equation is known as the **general gas equation** or the **universal gas equation**.

When using the universal gas equation, the following units must be used for the **universal gas constant** R = 8.31 J K^{-1} mol^{-1}.

It is important that the correct units are used in the universal gas equation.

- *Pressure* is measured in kilopascals.
- *Volume* is measured in litres.
- *Temperatures* is measured in kelvin.
- The *quantity* of gas is measured in moles

- The *number of moles* can be expressed as:

$$n = \frac{m}{M}$$

where m represents the mass of the substance (in grams) and M represents the molar mass of the substance.

So, the universal gas equation can be rewritten as:

$$PV = \frac{mRT}{M}$$

Sample problem 2.11

Find the volume of 6.30 mol of carbon dioxide gas at 23.0 °C and 550 kPa pressure. (R = 8.31 J K^{-1} mol^{-1})

Solution:

STEP 1

Assign the variables.

$P = 550$ kPa

$V = ?$ L

$n = 6.30$ mol

$T = 23 + 273 = 296$ K

$R = 8.31$ K^{-1} mol^{-1}

STEP 2

Use the universal gas equation to calculate V.

$$PV = nRT$$

$$\therefore V = \frac{nRT}{P}$$

$$= \frac{6.30 \times 8.31 \times 296}{550}$$

$$= 28.2 \text{ L}$$

studyon

Unit 3 | AOS 1 | Topic 1 | Concept 4

Do more
The mole

Revision questions

19. Calculate the volume of gas, in litres, occupied by the following.
- **(a) 3.5 mol of O_2 at 100 kPa and 50 °C**
- **(b) 12.8 mol of CH_4 at 10 atm and 60 °C**
- **(c) 6.5 g of Ar at 50 kPa and 100 °C**
- **(d) 0.56 g of CO_2 at 50 atm and 20 °C**
- **(e) 1.3×10^{-3} g He at 60 kPa and –50 °C**
- **(f) 1.5×10^{21} molecules of Ne at 40 kPa and 200 °C**

20. Calculate the volume occupied by 42.0 g of nitrogen gas at a pressure of 200 kPa and a temperature of 77 °C.

21. A 5.00 L balloon contains 0.200 mol of air at 120 kPa pressure. What is the temperature of air in the balloon?

22. If 55 mol of H_2 gas is placed in a 10 L flask at 7 °C, what would be the pressure in the flask?

Mass–volume stoichiometry

Quantities of gases can be calculated using the four stoichiometric steps.

Many chemical reactions involve gases. For example, barbecues, furnaces and engines, such as the internal combustion engine, burn fuel and produce carbon dioxide and water vapour. In order to ensure that the reactions are

efficient and as complete as possible, calculations of volumes or masses of the gases or fuels required are essential.

Sample problem 2.12

In a gas barbecue, propane is burned in oxygen to form carbon dioxide and water vapour. If 22.0 g of CO_2 is collected and weighed, find the volume of propane at 200 °C and 1.013×10^5 Pa.

Solution:

STEP 1

Write the equation and given information.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

$P = 1.013 \times 10^5$ Pa $\quad m(CO_2) = 22.0$ g

$T = 200 + 273 = 473$ K

$V = ?$ L

STEP 2

Calculate the number of moles of the known quantity of substance.

$$n(CO_2) = \frac{22.0}{44.0} = 0.500 \text{ mol}$$

STEP 3

Use the equation to find the molar ratio of unknown to known quantities, and calculate the number of moles of the required substance. Since 3 moles of CO_2 are produced from 1 mole of C_3H_8

$$n(C_3H_8) = \frac{n(CO_2)}{3} = \frac{0.500}{3} = 0.167 \text{ mol}$$

STEP 4

Find the volume of the propane using $PV = nRT$, where $R = 8.31$ J K^{-1} mol^{-1} and $P = 101\,325$ Pa $= 101.325$ kPa. Then,

$$V(C_3H_8) = \frac{nRT}{P} = \frac{0.167 \times 8.31 \times 473}{101.3}$$

$$V(C_3H_8) = 6.47 \text{ L}$$

Revision questions

23. Ethyne, C_2H_2, is also called acetylene and is used as a welding gas. It is produced from calcium carbide as shown by the ethyne reaction.

$$CaC_2(s) + 2H_2O(l) \longrightarrow Ca(OH)_2(s) + C_2H_2(g)$$

What volume of ethyne measured at 25 °C and 745 mmHg would be produced from 5.00 g of H_2O?

24. The equation for photosynthesis to produce glucose is as follows.

$$6CO_2(g) + 6H_2O(g) \longrightarrow C_6H_{12}O_6(g) + 6O_2(g)$$

What volume of carbon dioxide, measured at 28 °C and 1.3 atm pressure, is needed to produce 1.5 g of glucose?

Sample problem 2.13

Calculate the mass of hexane that is required to produce 100 L of carbon dioxide gas at SLC.

Solution:

STEP 1

Write the equation and identify the given information.

$$2C_6H_{14}(l) + 19O_2(g) \longrightarrow 12CO_2(g) + 14H_2O(l)$$

$m = ?$ g $V = 100$ L Conditions are SLC.

STEP 2

Calculate the number of moles of the known substance — in this case, CO_2 gas. Normally the universal gas equation would be used but, because it is at SLC, we know the molar volume and can therefore use the formula

$$n = \frac{V}{V_m}$$

Therefore, $n(CO_2) = \frac{100}{24.8} = 4.03$ mol.

STEP 3

Using the mole ratios from the equation, the number of moles of hexane can be calculated.

$$n(C_6H_{14}) = \frac{2}{12} \times n(CO_2)$$

Therefore, $n(C_6H_{14}) = \frac{2}{12} \times 4.03 = 0.672$ mol.

STEP 4

The mass of hexane can now be calculated.

$$m(C_6H_{14}) = n(C_6H_{14}) \times M(C_6H_{14}) = 0.672 \times 86.0 = 57.8 \text{ g}$$

Revision questions

25. Oxygen gas can be prepared in the laboratory by the decomposition of potassium nitrate according to the equation:

$$2KNO_3(s) \longrightarrow 2KNO_2(s) + O_2(g)$$

When 14.5 L of O_2 is formed at 1.00 atm and 25.0 °C, what mass of KNO_2 is also formed?

26. Magnesium reacts with hydrochloric acid according to the equation:

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

What mass of magnesium, when reacted with excess hydrochloric acid, would produce 5.0 L of hydrogen gas, measured at 26.0 °C and 1.2 kPa pressure?

Volume–volume stoichiometry

Consider the reaction between nitrogen and hydrogen gas.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

According to this equation, 1 mole of N_2 reacts with 3 moles of H_2 to produce 2 moles of ammonia.

Considering Avogadro's hypothesis, if the gases are at the same pressure and temperature, their *molar* ratios are equal to their *volume* ratios. We therefore use volumes instead of moles and can say that 10 mL of N_2 reacts with 30 mL of H_2 to form 20 mL of ammonia.

If gases are at the same pressure and temperature, their molar ratios are equal to their volume ratios.

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$

1 mol 3 mol $\longrightarrow$ 2 mol

1 vol 3 vol $\longrightarrow$ 2 vol

10 vol 30 vol $\longrightarrow$ 20 vol

Sample problem 2.14

If 100 m^3 of ethene is burned according to

$$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(g)$$

calculate the volume of:
(a) carbon dioxide produced
(b) oxygen consumed.
(Assume all gas volumes are measured at the same temperature and pressure.)

Solution: Because all gas volumes are measured at the same temperature and pressure, the equation may be interpreted in terms of volume ratios.

(a) $V(CO_2) = 2 \times V(C_2H_4)$

Therefore, $V(CO_2) = 2 \times 100 = 200\ m^3$.

(b) $V(O_2) = 3 \times V(C_2H_4)$

Therefore, $V(O_2) = 3 \times 100 = 300\ m^3$.

Revision questions

27. Methane gas burns in air at room temperature and pressure, according to the equation:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

If 25 mL of methane is burned at room temperature and pressure, find the volumes of the following reactants and products:
(a) oxygen (b) carbon dioxide (c) water.

28. At high temperatures, such as those in a car engine during operation, atmospheric nitrogen burns to produce the pollutant nitrogen dioxide, according to the equation:

$$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g)$$

(a) If 20 mL of nitrogen is oxidised, calculate the volume of oxygen needed to produce the pollutant. Assume that temperature and pressure remain constant.
(b) What is the initial volume of reactants in this combustion reaction?
(c) What is the final volume of products in the reaction?
(d) Is there an overall increase or decrease in the volume of gases on completion of the reaction?

Applying volume stoichiometry to thermochemistry

Just as the principles of mass–mass stoichiometry can be extended to include gas volumes, the thermochemical calculations introduced earlier in this chapter can also be extended to include gas volumes. This is very useful because it is often more convenient to measure a gas's volume than its mass. If pressure and temperature are measured as well, the universal gas equation, $pV = nRT$, provides a formula to change between volumes and moles.

Sample problem 2.15

Calculate the heat energy released when 375 mL of methane at 21 °C and 767 mmHg pressure is burned according to:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -889\ kJ\ mol^{-1}$$

Solution:

STEP 1

Calculate the number of moles of methane used. To do this, the information in the question needs to be changed to the appropriate units to enable the universal gas equation to be used.

The universal gas equation, $PV = nRT$, can be used in thermochemical calculations to link energy amounts to volumes of gases involved.

Temperature: 21 °C = 21 + 273 = 294 K

Volume: 375 mL = 0.375 L

$$\text{Pressure: } 767 \text{ mmHg} = \frac{767}{760} \times 101.3 = 102.2 \text{ kPa}$$

$$\text{Therefore, } n(CH_4) = \frac{PV}{RT} = \frac{102.2 \times 0.375}{8.31 \times 294} = 0.0157 \text{ mol.}$$

STEP 2

$$\text{Heat evolved} = 889 \times \frac{0.0157}{1} = 14.0 \text{ kJ}$$

Revision questions

29. Calculate the energy required to convert 2.50 L of carbon dioxide, at SLC, to glucose according to the equation:

$$6CO_2(g) + 6H_2O(l) \longrightarrow C_6H_{12}O_6(aq) + 6O_2(g) \quad \Delta H = +2816 \text{ kJ mol}^{-1}$$

30. Calculate the energy released when 200 mL of diborane, B_2H_6, is burned at 150 °C and 1.50 atmospheres according to the following equation.

$$B_2H_6(g) + 3O_2(g) \longrightarrow B_2O_3(s) + 3H_2O(g) \quad \Delta H = -2035 \text{ kJ mol}^{-1}$$

Fuels and greenhouse gases

The principles of stoichiometry and thermochemistry can be combined to quantify the effect that energy production, especially through the combustion of fossil fuels, has on the environment. As we saw in chapter 1, gases such as carbon dioxide, methane, nitrous oxide and ozone contribute to the enhanced greenhouse effect. Of these, carbon dioxide is produced in the greatest amounts due to the use of fossil fuels as energy sources. The increasing level of these gases has led to international treaties such as the Kyoto Protocol. These obligate signatory countries to reduce their greenhouse gas emissions in an attempt to limit future global warming.

In order to meet targets for reduction, countries need to estimate their energy requirements, both current and future, and compare the energy sources available for the required energy production. A useful unit in such comparisons is the amount of greenhouse gas produced (usually carbon dioxide) per megajoule of energy released. Typical units are g MJ^{-1} and L MJ^{-1}.

Sample problem 2.16

If it is assumed that petrol is entirely octane, and that it burns according to

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(l) \quad \Delta H = -10\,928 \text{ kJ mol}^{-1}$$

calculate the:

(a) mass of carbon dioxide produced per MJ of energy evolved

(b) volume of carbon dioxide produced at 100 kPa and 20 °C per MJ of energy evolved.

Solution: (a) **STEP 1**

Calculate the mass of octane required to evolve 1 MJ of heat energy. From the equation, 2 moles of octane evolves 10.928 MJ.

$M(C_8H_{18}) = 114.0\ \text{g mol}^{-1}$

Therefore, 2×114.0 g evolves 10.928 MJ.

Therefore, 1 MJ requires $\frac{2 \times 114.0}{10.928} = 20.9$ g.

STEP 2

Calculate the mass of CO_2 produced from this amount of octane.

$n(C_8H_{18}) = \frac{m}{M} = \frac{20.9}{114.0} = 0.183$ mol

From the equation, $n(CO_2) = \frac{16}{2} \times 0.183 = 1.464$ mol.

Therefore, $m(CO_2) = 1.464 \times 44.0 = 64.4$ g.

It can therefore be stated that, if petrol is assumed to be pure octane, it produces 64.4 g of carbon dioxide per megajoule of energy released ($64.4\ \text{g MJ}^{-1}$).

(b) $PV = nRT$

$n = 1.464$ mol

$T = 273 + 20 = 293$ K

$P = 100$ kPa

Therefore, $V = \frac{nRT}{P} = \frac{1.464 \times 8.31 \times 293}{100} = 35.6$ L.

Therefore, 35.6 L of CO_2 is produced per megajoule of energy released ($35.6\ \text{L MJ}^{-1}$) at the stated conditions of temperature and pressure.

When comparing the effect of fuels on the environment, it is often useful to consider the carbon dioxide emissions produced per unit of energy. Two common units for making such comparisons are g MJ^{-1} and L MJ^{-1}.

Sample problem 2.17

Methane is produced by a number of natural processes and, in many of these situations, it can easily enter the atmosphere. Methane is a more potent greenhouse gas than carbon dioxide as it is better at absorbing heat. However, it can also be used as a fuel. The equation for burning methane is:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

If methane is captured and then used as a fuel, calculate the net change in greenhouse gas volume that would occur. Assume that all volumes are measured at the same temperature and pressure.

Solution: As temperature and pressure is constant, we can write:

$V(CH_4)$ consumed $= V(CO_2)$ produced

Therefore, there is no net change.

You might think that a process such as capturing and then burning methane would produce no net environmental benefit. However, when it is remembered that methane is a more potent greenhouse gas than carbon dioxide, it can be seen that there is an environmental benefit gained from burning the methane, rather than just letting it escape into the atmosphere. Of course, if carbon dioxide could also be captured and prevented from entering the atmosphere, the environmental benefit would be even greater.

Carbon dioxide is the most significant greenhouse gas due to the large volumes in which it is produced. For this reason, other greenhouse gas emissions are often

measured in 'carbon dioxide equivalent'. That is, all these other gases are converted to the volume of carbon dioxide that would produce the same effect.

Revision question

31. (a) Calculate the net change in mass of greenhouse gas produced by the combustion of 128 g of methane.

(b) Express your answer as a percentage increase or decrease.

(c) Comment on how the units used (mass or volume) may influence the conclusions drawn from calculations such as in this question and in sample problem 2.17.

When electricity generation is considered, the calculations must take into account the efficiency of the generation process. For example, to produce 1 MJ of electrical energy, enough fuel would need to be burned to produce 3.33 MJ of heat energy, assuming that the process is 30% efficient. The mass of carbon dioxide produced in supplying 3.33 MJ would then need to be calculated.

Both transport and the generation of electricity consume large amounts of fossil fuels and, in the process, add large amounts of carbon dioxide to the atmosphere. As nations work to reduce their greenhouse gas emissions, alternative fuels need to be considered. When considering fuels, the amount of carbon dioxide emitted per unit of energy produced is a useful comparison.

Sample problem 2.18

Calculate the carbon dioxide emissions, in kg MJ^{-1}, from a power station that has an overall efficiency of 33.0% and burns black coal with a carbon content of 81.5%. (Heat evolved from black coal = 34.0 kJ g^{-1})

Solution: At 33.0% efficiency, the energy input required to produce 1 MJ of electrical energy output is:

$$\text{Efficiency} = \frac{\text{output energy}}{\text{input energy}} \times 100$$

Therefore, input energy $= \frac{1 \times 100}{33} = 3.03$ MJ.

Mass of black coal required to produce 3.03 MJ $= \frac{3030}{34.0} = 89.1$ g

Mass of carbon in this amount of black coal $= \frac{81.5}{100} \times 89.1 = 72.6$ g

Assuming the carbon produces carbon dioxide according to:

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

$$n(C) = \frac{72.6}{12.0} = 6.05 \text{ mol and } n(C) = n(CO_2)$$

Therefore, $m(CO_2) = 6.05 \times 44.0 = 267 \text{ g} = 0.267 \text{ kg}$.

Therefore, CO_2 production is 0.267 kg MJ^{-1} of electrical energy produced.

Revision questions

32. The molar heat of combustion for ethanol is –1364 kJ mol^{-1}. Calculate the mass of carbon dioxide emitted when ethanol is used to produce 1.00 MJ of heat.

33. (a) Methanol is also a fuel. Its molar heat of combustion is –725 kJ mol^{-1}. What mass of carbon dioxide would be produced using methanol to generate 1.00 MJ of heat?

(b) Comment on your answers to questions 32 and 33(a) in relation to the masses of CO_2 produced.

34. A coal-fired power station using brown coal as its fuel operates at 37.0% overall energy efficiency. The brown coal has a heat value of 16.0 kJ g^{-1} and a carbon content of 29.0%. Assuming that all the carbon present forms carbon dioxide, calculate the carbon dioxide produced per MJ of electrical energy produced in units of:

(a) g MJ^{-1}

(b) L MJ^{-1} (at SLC).

How do we obtain the energy output of a fuel?

The energy available from a fuel may be estimated by using a known amount of fuel to heat a known mass of water. If the temperature change is noted, the formula $Q = mc\Delta T$ may be used to calculate the energy added to the water. Knowing the amount of fuel used, this can then be scaled up to an appropriate reference amount. Common units are kJ g^{-1}, kJ mol^{-1} and MJ $tonne^{-1}$.

So far in this chapter, we have made predictions relating to fuel quantities and heat outputs using thermochemical equations. However, there are two issues that you may have been wondering about.

- How are ΔH values obtained in the first place?
- Thermochemical equations are fine for fuels that are pure substances and can therefore have an equation written for their combustion. But what about fuels that are mixtures? Many common fuels, such as petrol, diesel, wood and coal, fall into this category. How can we obtain heat outputs and make predictions for these fuels?

The answer to both these questions is to burn a small amount of fuel and 'capture' the heat evolved in some way so that it can be measured. The results obtained are then scaled up to a reference amount so that comparisons can be made. The values obtained are quoted in units that take into account the units used to measure the energy output and the particular reference amount being used. Common units are kJ mol^{-1}, kJ g^{-1} and MJ $tonne^{-1}$.

If this is to be done accurately, a device called a calorimeter must be used. There are two main types of calorimeters — solution calorimeters and bomb calorimeters. The features of these and their use are discussed in chapter 12.

An alternative method is to use the fuel to heat a known mass of water and measure the resultant temperature increase. This method produces only approximate results, the degree of accuracy depending on the steps taken to minimise heat loss to the surroundings.

You may remember from unit 1 that the formula:

$$Q = mc\Delta T$$

can be used to calculate the heat required to raise a given mass (m) of a substance of known specific heat (c) by a certain temperature (ΔT). The specific

studyon

Unit 3 | AOS 1 | Topic 1 | Concept 7

Do more
Heats of combustion

heat capacity, c, of a substance is the amount of energy needed to raise the temperature of 1 g of the substance by 1 °C or 1 K. This formula is used to calculate the heat added to the water. Sample problem 2.19 illustrates this method.

Sample problem 2.19

The heat content of kerosene was determined by using a kerosene burner to heat 250 mL of water. It was found that burning 0.323 g of kerosene raised the temperature of the water by 11.2 °C.

Given that the specific heat capacity of water is 4.18 J g^{-1} K^{-1}, calculate the heat energy released from kerosene in kJ g^{-1} and MJ $tonne^{-1}$.

Solution:

STEP 1

Calculate the heat energy (Q) transferred to the water.

250 mL of water = 250 g

$Q = mc\Delta T = 250\text{ g} \times 4.18\text{ J g K}^{-1} \times 11.2\text{ °C}$

$= 11\,704\text{ J} = 11.7\text{ kJ}$

STEP 2

Scale this value up to the chosen reference amount (in this case, 1 g).

$$\text{Heat evolved} = 11.7\text{ kJ} \times \frac{1}{0.323\text{ g}} = 36.2\text{ kJ g}^{-1}$$

To convert this to MJ $tonne^{-1}$, 1 tonne = 10^6 g = 1 000 000 g; 1 MJ = 1000 kJ

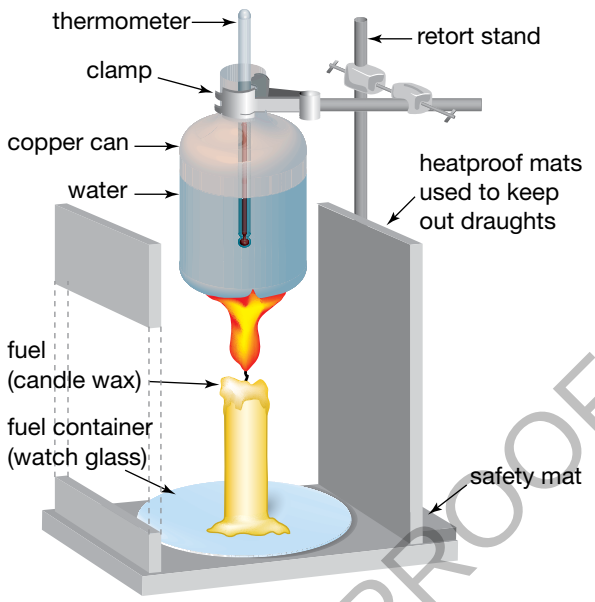
Therefore, $36.2\text{ kJ g}^{-1} = \frac{36.2}{1000} \times 1\,000\,000 = 36\,200\text{ MJ tonne}^{-1}$.

Note: This represents a minimum value. Not all of the heat from the kerosene would have gone into the water. Some of it would have been wasted in heating the surrounding air and the equipment used to hold the water.

Revision questions

An experiment to compare the energy output of candle wax, ethanol and butane was performed by setting up the apparatus shown in the figure at the top of the next page. The ethanol was poured into a crucible, a small wax candle stuck onto a watch glass and a gas cigarette lighter were each used as a 'burner' after being lit. Each burner was weighed before and after it was used to heat 200 g of water. The results are shown in table 2.2 on the next page. (Assume the formula for candle wax is $C_{20}H_{42}$.)

35. **Copy and complete table 2.2.**
36. **Taking the specific heat capacity of water as 4.18 J g^{-1} °C^{-1}, calculate the energy produced from 1.00 gram of each substance. How do the results compare?**
37. **Calculate the heat of combustion (enthalpy per mole of substance used) for each substance. How do the results compare?**
38. **Describe a use for each of the fuels tested. Explain why each fuel is better suited for the purpose it is commonly used for than either of the other two fuels.**
39. **The molar heat of combustion of ethanol has been found experimentally to be –1364 kJ mol^{-1}.**
    - (a) **Is combustion exothermic or endothermic?**
    - (b) **What was the percentage accuracy of the experiment above?**
    - (c) **List the sources of error in the experiment and describe how some of these errors can be minimised.**
40. **Write a report for the experiment.**



Apparatus for testing the energy output of candle wax

Most of our energy is supplied by fossil fuels.

TABLE 2.2 Results of combustion of different fuels

Property	Ethanol	Candle wax	Butane
mass of 'burner' before heating (g)	23.77	32.72	43.94
mass of 'burner' after heating (g)	22.54	32.50	43.71
mass of fuel used (g)			
mass of water (g)	200	200	200
initial temperature of water (°C)	20.0	20.0	20.0
highest temperature of water (°C)	35.0	30.0	29.0
temperature rise (°C)			
molar mass (g mol^{-1})			

Chapter review

Summary

- The unit of energy is the joule (J).
- The study of energy changes that accompany chemical reactions is called thermochemistry or chemical thermodynamics.
- The total energy stored in a substance is called the enthalpy, or heat content, of the substance and is represented by the symbol H.
- The change in enthalpy as a reaction proceeds is known as the heat of reaction and can be determined according to the relationship:

$$\Delta H = H_{\text{products}} - H_{\text{reactants}}$$

- Exothermic reactions release energy to their surroundings and have a negative ΔH value.
- Endothermic reactions absorb energy from their surroundings and have a positive ΔH value.
- Energy diagrams or profiles may be used to visually represent changes in enthalpy.
- Thermochemical equations are chemical equations that, in addition to balancing charge and mass, include the enthalpy change and may be used in stoichiometric calculations to determine the energy changes associated with chemical reactions.
- The kinetic molecular theory of gases helps explain gas properties and begins with five assumptions about gas particles. These particles:
  - are moving constantly and at random
  - experience an increase in kinetic energy and move more quickly when temperature is increased
  - have insignificant attractive or repulsive forces between them
  - are very far apart and their volume is small compared to the volume they occupy
  - collide with one another and the walls of their container, exerting pressure.
- When considering gas behaviour, we are concerned with the pressure a gas exerts, the volume it occupies, its temperature and the number of gas molecules. We express these in SI units.
  - Pressure is measured in pascal (Pa).

$$760 \text{ mmHg} = 1 \text{ atm} = 1.013 \times 10^5 \text{ Pa}$$
$$= 101.3 \text{ kPa}$$

  - Temperature is converted from degrees Celsius (°C) to the absolute or Kelvin scale, where absolute zero is –273 °C.

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

  - Volume is measured in cubic metres (m^3), litres (L) or millilitres (mL).

$$1 \text{ m}^3 = 10^3 \text{ L} = 10^6 \text{ mL}$$

  - Quantity is measured in moles (mol).
- Boyle's law states that the volume of a fixed mass of gas at constant temperature is inversely proportional to the pressure exerted on it.

$$P_1V_1 = P_2V_2$$

- Charles' law states that the volume of a fixed mass of gas at constant pressure is directly proportional to its absolute (Kelvin) temperature.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- Gas volumes are measured using a standard set of fixed conditions, standard laboratory conditions (SLC), where temperature is 25 °C (298 K) and pressure is 100.0 kPa.
- The molar gas volume is the volume that 1 mole of gas occupies. At SLC, this equals 24.8 L. To calculate amounts of gases at SLC, we use the equation:

$$n_{\text{SLC}} = \frac{V}{24.8} \text{ where } V \text{ is measured in litres.}$$

- The universal gas equation combines several of the gas laws and contains the universal gas constant (R).

$PV = nRT$ where the constant R is equal to 8.31 J K^{-1} mol^{-1} and V is measured in litres, T in kelvin and P in kPa.

- The universal gas equation enables stoichiometric calculations that link mases and volumes together (mass–volume calculations).
- Volume–volume stoichiometry can be performed using volumes rather than moles, since, for gases at the same pressure and temperature, mole ratios are equal to volume ratios.

$$\frac{V(\text{unknown})}{V(\text{known})} = \frac{\text{coefficient of unknown}}{\text{coefficient of known}}$$

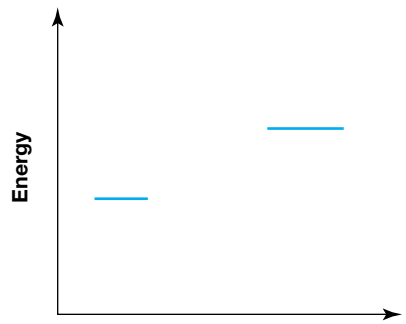
- ΔH values in thermochemical equations can be evaluated from experimental data and from the ΔH values of related equations.
- The environmental effect of a fuel can be measured in terms of its greenhouse gas emissions per megajoule of energy produced.
- The specific heat capacity (symbol c) of a substance is defined as the amount of energy needed to raise the temperature of 1 g of the substance by 1 °C (or 1 K).

- The heat energy in a fuel can be estimated by using a known amount of the fuel to heat a known amount of water. The formula $Q = mc\Delta T$ is used to calculate the heat energy added to the water.

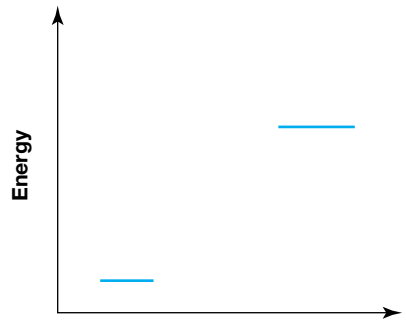
Multiple choice questions

1. Which of the energy profiles below represents the *most exothermic* reaction?

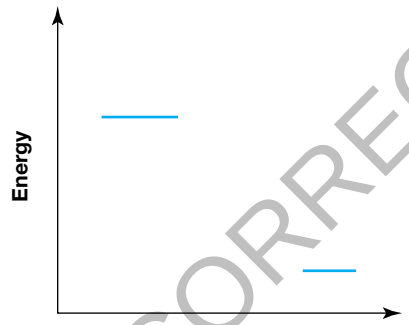
A



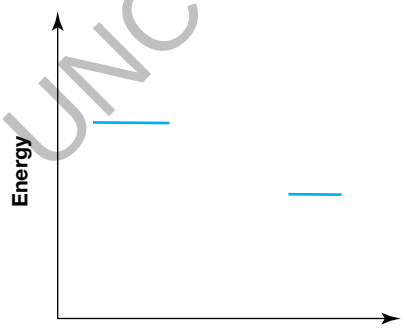
B



C



D



2. Consider the following equation for the formation of hydrogen chloride gas.

$$H_2(g) + Cl_2(g) \longrightarrow 2HCl(g) \quad \Delta H = -184\ \text{kJ mol}^{-1}$$

When 2 moles of hydrogen gas react completely with 2 moles of chlorine gas:

A 184 kJ of energy is released
B 184 kJ of energy is absorbed
C 368 kJ of energy is released
D 368 kJ of energy is absorbed.

3. Which of the following processes is endothermic?

A Burning carbon
B Evaporating kerosene
C Condensing steam
D A reaction with a negative heat of reaction

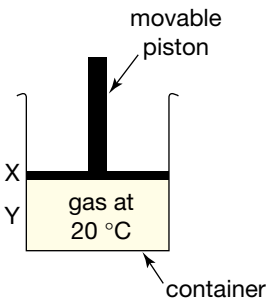
4. 'Dry ice' is solid carbon dioxide. It is stable at very low temperatures but sublimes at room temperature according to the reaction:

$$CO_2(s) + \text{heat} \longrightarrow CO_2(g)$$

Handling dry ice with bare hands can cause severe skin damage because solid carbon dioxide:

A is a strong oxidising agent
B releases considerable heat to the skin while subliming
C absorbs considerable heat from the skin while subliming
D forms a strong acid when dissolved in the moisture of the skin.

5. The diagram below shows a container filled with gas and sealed by a movable piston. There is sufficient gas to support the piston at point X when the temperature is 20 °C. Assume that the piston is locked at X. If the gas is heated, the pressure which it exerts will:



A increase
B decrease
C remain the same
D change in an unpredictable manner.

6. The apparatus in question 5 remains at a constant temperature of 20 °C, but the piston is pushed down to point Y and locked, so that the volume of the container is halved. The average number of molecules striking a unit area of the wall of the container per unit time will:

A double
B halve
C remain the same
D change in an unpredictable manner.

7. The pressure of the gas inside a scuba diver's lungs changes from 100 kPa to 150 kPa. If the diver's lungs initially held 6 L of gas, their volume at this depth would be:

A 3 L
B 4 L
C 6 L
D 9 L.

8. A rigid container holds a fixed volume of gas at a certain temperature and pressure. In order to double the pressure of the gas inside the container, one could:
   - **A** halve the amount of gas in the container
   - **B** halve the amount of gas, but double the absolute temperature
   - **C** double the amount of gas and double the absolute temperature
   - **D** double the absolute temperature.
9. 6.0 L of air at 400 K and 100 kPa is cooled to 300 K at constant pressure. The new volume will be:
   - **A** 2.5 L
   - **B** 4.5 L
   - **C** 6.0 L
   - **D** 8.0 L.
10. Two litres of helium gas was measured at a temperature of 200 K and a pressure of 1 atmosphere. If the kinetic energy (temperature) of the molecules is doubled while the pressure remains constant, what volume will the gas occupy?
   - **A** 0.5 L
   - **B** 1 L
   - **C** 2 L
   - **D** 4 L
11. 10 L of hydrogen gas is collected at 110 kPa and 20 °C. It is then compressed into a 2 L container at a pressure of 1100 kPa. The new temperature will be:
   - **A** 40 °C
   - **B** 200 °C
   - **C** 273 °C
   - **D** 313 °C.
12. 200 mL of gas at 27 °C and 700 mmHg pressure became 210 mL at 1.05 atm pressure. The new temperature is:
   - **A** 360 °C
   - **B** 87 °C
   - **C** 187 °C
   - **D** 260 °C.
13. Tired of working in laboratories that are either freezing or too hot, Jenny has proposed a third set of standard conditions. She has defined these as a temperature of 15 °C and a pressure of 100 kPa. Under these conditions, the molar volume of a gas would be:
   - **A** less than 22.7 L
   - **B** 22.7 L
   - **C** between 22.7 L and 24.8 L
   - **D** greater than 24.8 L.
14. Methane, $CH_4(g)$, burns in air to form carbon dioxide and water according to the following equation.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -889\ \text{kJ mol}^{-1}$$

   Which of the following statements about this reaction is false?
   - **A** The reaction is exothermic.
   - **B** The enthalpy (heat content) of the products is greater than the reactants, as shown by the sign of ΔH.
   - **C** Heat is released to the environment by the reaction.
   - **D** The volume occupied by the gaseous products is equal to that of the reactants, if all are measured at the same temperature and pressure.
15. Consider the following reaction.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \quad \Delta H = -484\ \text{kJ mol}^{-1}$$

   Energy is released when hydrogen burns in oxygen because:
   - **A** the net strength of the chemical bonds within the reactant molecules is greater than the net strength of the chemical bonds within the product molecules
   - **B** the net strength of the chemical bonds within the reactant molecules is less than the net strength of the chemical bonds within the product molecules
   - **C** there are fewer product molecules than reactant molecules
   - **D** the reactants are elements while the product is a compound.
16. Consider the following reaction.

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \quad \Delta H = -484\ \text{kJ mol}^{-1}$$

   Compared with the reaction given above, the heat evolved in the reaction

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

   will necessarily be:
   - **A** greater, because extra heat is evolved in forming liquid water from gaseous water
   - **B** less, because extra heat is absorbed in forming liquid water from gaseous water
   - **C** the same, since the heat evolved is the same provided the chemical formulas of the reactants and products remain unchanged
   - **D** the same, provided the same amounts of reactants are used.

Questions 17 and 18 refer to the following thermochemical equation for the combustion of ethyne, commonly known as acetylene.

$$2C_2H_2(g) + 5O_2(g) \longrightarrow 4CO_2(g) + 2H_2O(l) \quad \Delta H = -2600\ \text{kJ mol}^{-1}$$

17. The combustion of 0.26 g of ethyne will:
   - **A** absorb 13 J of energy
   - **B** evolve 13 J of energy
   - **C** absorb 13 kJ of energy
   - **D** evolve 13 kJ of energy.
18. The carbon dioxide contribution from the use of ethyne is:
   - **A** 17 kg MJ^{-1}
   - **B** 34 kg MJ^{-1}
   - **C** 68 kg MJ^{-1}
   - **D** 134 kg MJ^{-1}.

19. The amount of energy required to raise 150 g of water by 40 °C is:
   A 25 J
   B 150 J
   C 25 kJ
   D 150 kJ.
20. If burning 4.5 g of a fuel raises the temperature of 75 g of water by 30 °C, the heat of combustion of this fuel is:
   A 0.0075 kJ g^{-1}
   B 0.057 kJ g^{-1}
   C 2.1 kJ g^{-1}
   D 2.4 kJ g^{-1}.

Review questions

Energy changes in chemical reactions

1. What is meant by the 'enthalpy change' of a chemical reaction?
2. State whether each of the following reactions is exothermic or endothermic.
   (a) $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$ $\Delta H = -185\ kJ\ mol^{-1}$
   (b) $H_2(g) + CO_2(g) \longrightarrow CO(g) + H_2O(g)$ $\Delta H = +41\ kJ\ mol^{-1}$
   (c) $SnO_2(s) + 2C(s) \longrightarrow Sn(s) + 2CO(g)$ $\Delta H = +360\ kJ\ mol^{-1}$
   (d) $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$ $\Delta H = -92.2\ kJ\ mol^{-1}$
3. Draw an energy profile diagram for each of the reactions in question 2. In each diagram, label the reactants and the products as given by the equation, and show the ΔH value and its magnitude.
4. Compare endothermic and exothermic reactions with respect to:
   (a) the sign of ΔH
   (b) heat exchange with the surroundings
   (c) energy levels of reactants and products.
5. Energy is lost during an exothermic reaction. Explain why this observation is consistent with the Law of Conservation of Energy.

Writing thermochemical equations

6. Ethanol burns in oxygen to produce water and either carbon dioxide or carbon monoxide. The particular oxide produced depends on whether the oxygen supply is plentiful or limited. The molar heats of combustion for these two reactions are –1364 kJ mol^{-1} and –1192 kJ mol^{-1} respectively.
   (a) Write the thermochemical equation for the combustion of ethanol to produce carbon dioxide.
   (b) Write the thermochemical equation for the combustion of ethanol to produce carbon monoxide.
   (c) Use the equations from (a) and (b) to explain how the amount of oxygen consumed influences the oxide produced.
   (d) Calculate ΔH for the following reaction.
   $$2CO(g) + O_2(g) \longrightarrow 2CO_2(g).$$
7. Data tables give the heat output from the complete combustion of ethane and ethene as 51.9 kJ g^{-1} and 50.3 kJ g^{-1} respectively. Write thermochemical equations for the complete combustion of these fuels, showing ΔH values in units of kJ mol^{-1}.
8. Dodecane has the formula $C_{12}H_{26}$. In an experiment, 3.10 g of dodecane was burned in a plentiful supply of oxygen. 137.0 kJ of heat was released. Write the thermochemical equation for the complete combustion of dodecane.
9. While researching on the internet, a student found that candle wax is a mixture of hydrocarbons, but, as an approximation, it may be assumed to be pentacosane.
   (a) To what homologous series does pentacosane belong?
   (b) Write the formula for pentacosane given that its molecules contain 25 carbon atoms.
   (c) Write the equation for the combustion of pentacosane in a plentiful supply of oxygen.
   In an experiment, the student measured the heat given off by a burning candle. She discovered that 153 kJ of energy was evolved and that the mass of the candle decreased from 75.384 to 71.735 g.
   (d) On the assumption that candle wax may be approximated as pentacosane, determine the heat of combustion in kJ mol^{-1} and then write the thermochemical equation for its combustion.
   On further investigation, she found a source claiming tetracosane, $C_{24}H_{50}$, was a better approximation for candle wax.
   (e) How would this change affect the value of ΔH?

Measuring gases

10. Use the kinetic molecular theory to explain:
   (a) why gases exert pressure
   (b) why, for a fixed amount of gas at constant volume, pressure increases with temperature.
11. (a) Convert 1.34 atm to mmHg and kPa.
   (b) Convert 365 mmHg to Pa and atm.
   (c) Convert 102 576 Pa to atm and mmHg.
12. Convert the following Celsius temperatures to Kelvin temperatures.
   (a) 100 °C
   (b) –20 °C
   (c) –200 °C
   (d) 345 °C
13. Convert the following Kelvin temperatures to Celsius.
   (a) 300 K
   (b) 427 K
   (c) 173 K
   (d) 392 K
   (e) 73 K

14. Convert the following volumes to the second unit given.
(a) 125 L to m^3
(b) 125 L to mL
(c) 300 mL to L
(d) 2.6 mL to L
(e) 2 m^3 to L
(f) 3×10^3 mL to L

Standard conditions

15. Calculate the number of moles of the following gases at SLC.
(a) 1.5 L oxygen, O_2
(b) 2.56 L of hydrogen, H_2
(c) 250 mL of nitrogen, N_2

16. Calculate the volume of the following gases at SLC.
(a) 1.53 mol hydrogen, H_2
(b) 13.6 g of methane, CH_4
(c) 2.5×10^{30} molecules of nitrogen, N_2

17. Calculate the mass of the following gas samples. All volumes are measured at SLC.
(a) 150 mL of oxygen, O_2
(b) 4.5 L of carbon dioxide, CO_2

18. (a) How many moles of nitrogen are contained in 26.0 L of nitrogen gas at SLC?
(b) How many individual nitrogen molecules are present in 26.0 L of the gas?

Gas laws

19. What is absolute zero? Can one actually reach this temperature? Explain.

20. The pressure on 13.0 L of neon gas is increased from 90.0 kPa to 360 kPa. If the temperature remains constant, find the new volume of the gas.

21. 1.00 L of hydrogen gas in a cylinder under 5.0×10^4 kPa fills a balloon at 150 kPa at the same temperature. Determine the volume of the balloon.

22. The pressure on a gas remains constant. Its volume is 700 mL. The temperature is 27 °C. Calculate the temperature needed to change the volume to:
(a) 14.0 mL
(b) 420.0 mL.

The universal gas equation

23. Calculate the number of moles of gas present in each of the following gas samples.
(a) 32.3 L of argon at 102.0 kPa and 15 °C
(b) 24.3 L of nitrogen at 13.2 atm and 35 °C
(c) 12.3 L of helium at 755 mmHg and –10 °C
(d) 24.8 L of argon at 100.0 kPa and 25 °C

24. Calculate the volumes occupied by the following gas samples.
(a) 0.560 mol of carbon dioxide gas at 101.5 kPa and 17 °C
(b) 15.5 mol of ethyne gas at 10.35 atm and 20 °C
(c) 0.800 g of oxygen gas at 98.9 Pa and 11 °C
(d) 5.279 g of nitrogen gas at 100 kPa and –50 °C

25. An aerosol can of deodorant has a volume of 120 mL. The contents exert a pressure of 9.0×10^5 Pa at 27 °C.
(a) Calculate the number of moles of gas present in the can.
(b) How many particles are present in the can of deodorant?
(c) If the contents of the can are transferred to a 200 mL container, what would be the temperature if the pressure drops to 6.0×10^5 Pa?

26. Calculate the temperature, in °C, of 0.25 mol of a gas that occupies a volume of 4.155×10^{-3} m^3 at a pressure of 150 kPa.

27. A certain mass of unknown gas (molar mass = 30.0 g mol^{-1}) occupies 16.64 L at a pressure of 125 mmHg and a temperature of –73 °C. Calculate the mass of gas present in the container.

28. Calculate the pressure needed to contain 3.0 mol of ammonia in a container of volume 20 L at a temperature of 27 °C.

29. An empty 200 mL flask has a mass of 84.845 g. It is filled with a gas at 17.0 °C and 770 mmHg pressure and then weighs 85.084 g. Calculate the molar mass of the gas. What is the gas?

30. A gas is at SLC and has a volume of 11.2 L. If its temperature is changed to 127 °C, and its pressure changes to 253.3 kPa, calculate its new volume.

31. Each time Nicole breathes, she inhales about 400 mL of air. Oxygen makes up about 20% by volume of air. How many oxygen molecules does she inhale in one breath at 25 °C and 1.0×10^4 Pa pressure?

Mass–volume calculations

32. Magnesium burns according to the equation:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

What mass of magnesium combines with 5.80 L of oxygen measured at SLC?

33. Butan-1-ol (density = 0.81 g mL^{-1}) burns according to the following equation.

$$CH_3CH_2CH_2CH_2OH(l) + 6O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$$

When 10.0 mL of butan-1-ol is burned, calculate:
(a) the mass of water produced
(b) the volume of carbon dioxide produced at SLC
(c) the volume of carbon dioxide produced at 200 °C and 1.2 atm pressure.

The butan-1-ol is used to produce 100 mL of carbon dioxide at SLC.
(d) Calculate the volume of butan-1-ol needed.

34. A number of technologies either exist or are being developed for removing carbon dioxide from flue gases. Some of these are also being investigated for the removal of carbon dioxide directly from the air. One such process uses calcium oxide and steam (at 400 °C) to remove carbon dioxide. The overall equation for this reaction is:

$$CaO(s) + CO_2(g) \xrightarrow[\text{steam}]{400\ °C} CaCO_3(s)$$

(a) Calculate the maximum volume of carbon dioxide (at SLC) that can be removed per 1.00 tonne of calcium oxide.
(b) Calculate the maximum mass of calcium carbonate that would be formed per tonne of calcium oxide.

Volume–volume stoichiometry

35. Methane gas, CH_4, is a major constituent of the natural gas in Bass Strait. It readily burns in air to produce carbon dioxide and water vapour. If all gas volumes are measured at SLC calculate:
(a) the volume of oxygen gas that is required for the complete combustion of 12.5 L of methane gas
(b) the volume of air required for reaction, given that air contains 20% oxygen.

36. When propane, C_3H_8, is burned in a limited supply of oxygen, carbon monoxide is among its products. If 4.25 L of propane gas, measured at room temperature and pressure, is burned in a limited air supply, what volume of carbon monoxide gas would be produced at the same conditions?

37. Hydrogen reacts with oxygen in the reaction:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

What volume of water vapour is produced when 4.0 L of hydrogen reacts with 10 L of oxygen? What volume of oxygen remains unreacted? (Assume that all volumes are measured at the same temperature and pressure.)

38. In the Haber process, nitrogen and hydrogen are combined under specific conditions of temperature and pressure to industrially manufacture ammonia.
(a) Write the balanced equation.
(b) What volume of nitrogen combines with 45 m^3 of hydrogen?
(c) What volume of ammonia would be produced if the reaction went to completion?

Calculations using thermochemical equations

39. Iron can react with carbon dioxide according to the following equation.

$$2Fe(s) + 3CO_2(g) \longrightarrow Fe_2O_3(s) + 3CO(g) \qquad \Delta H = +26.3\ kJ\ mol^{-1}$$

(a) Explain why this equation may be described as a thermochemical equation.
(b) Calculate the heat change, in kilojoules, when 2 g of iron reacts completely with excess carbon dioxide.
(c) Calculate the heat change, in kilojoules, when 200 mL of carbon dioxide is used at SLC.
(d) What mass of iron would be needed to absorb 1000 kJ of heat energy?
(e) What volume of carbon dioxide would be needed at SLC to absorb 1000 kJ of energy?

40. For the reaction:

$$S(s) + O_2(g) \longrightarrow SO_2(g) \quad \Delta H = -296\ kJ\ mol^{-1}$$

(a) how much heat is evolved when 2.00 moles of sulfur are burned in excess oxygen
(b) how much heat is evolved when 200 g of sulfur is burned in excess oxygen
(c) how much heat is evolved when 450 mL of oxygen at SLC reacts with excess sulfur
(d) how much heat is evolved when 200 g of sulfur reacts with 2 litres of oxygen at 80 °C and 101.3 kPa pressure?

41. Given the following two equations:

$$\tfrac{1}{2}N_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow NO(g) \quad \Delta H = 90\ kJ\ mol^{-1}$$

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g) \quad \Delta H = -114\ kJ\ mol^{-1}$$

calculate the enthalpy change for the following reaction:

$$2NO_2(g) \longrightarrow N_2(g) + 2O_2(g)$$

42. Butane, C_4H_{10}, is a gaseous fuel. For the combustion of butane, $\Delta H = -2874\ kJ\ mol^{-1}$:
(a) write a balanced thermochemical equation
(b) calculate the quantity of heat that would be evolved from the combustion of 12 L of butane measured at SLC.

43. Hydrazine, N_2H_4, is a liquid fuel that has been used for many years in the engines of space probes. It was famously used in the terminal-descent engines that successfully landed the *Curiosity* rover on the surface of Mars in 2012. When passed over a suitable catalyst, it decomposes quickly in a multi-step exothermic chemical reaction. The overall equation for this process is:

$$N_2H_4(l) \longrightarrow N_2(g) + 2H_2(g) \quad \Delta H = -50.3\ kJ\ mol^{-1}$$

(a) Calculate the energy released per kilogram of hydrazine in the equation above.
(b) Given that the average temperature and pressure of the Martian atmosphere are –60 °C and 600 Pa respectively, calculate the total volume of gas added to the Martian atmosphere if 50 kg of hydrazine is used.

44. Octane burns according to the following equation.

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(l) \quad \Delta H = -10\,928 \text{ kJ mol}^{-1}$$

(a) Given that the density of octane is 0.70 g mL^{-1}, calculate the volume of carbon dioxide produced (at SLC) per litre of octane.
(b) Calculate the mass of carbon dioxide evolved in the production of 1.00 MJ of energy through the reaction above.
(c) Express your answer to (b) as litres of carbon dioxide (at SLC) per MJ.

45. Petrol and LPG are two fuels commonly used in Australia. It is claimed that LPG is 'better for the environment' as it releases less carbon dioxide. It is also attractive to motorists because, even though more litres are used, it is cheaper than petrol. As an approximation, petrol may be assumed to be octane, whereas LPG is a mixture of propane and butane. Some relevant data is shown in the table below.

Fuel	Molar enthalpy of combustion (kJ mol^{-1})	Density (g mL^{-1})
propane	–2217	0.51 (as LPG)
butane	–2874	0.51 (as LPG)
octane	–5464	0.70 (as petrol)

(a) Calculate the mass of LPG (assuming it to be propane) required to produce 1.00 MJ of heat energy.
(b) Calculate the mass of carbon dioxide produced from (a) and express your answer as g MJ^{-1}.
(c) Calculate the mass of petrol (assuming it to be octane) required to produce 1.00 MJ of heat energy.
(d) Calculate the mass of carbon dioxide produced from (c) and express your answer as g MJ^{-1}.
(e) Hence state the net reduction (in g MJ^{-1}) of carbon dioxide emission when LPG is used in preference to petrol.
(f) Repeat (a), (b) and (e) for LPG if it is assumed to be butane.
(g) Calculate the volume of LPG (assuming it to be propane) required to produce the same energy as 1.00 litre of petrol (assuming it to be octane).
(h) Is it true that LPG is better than petrol? Use your answers to (a)–(g) to explain your response.

Estimating heat energy in fuels

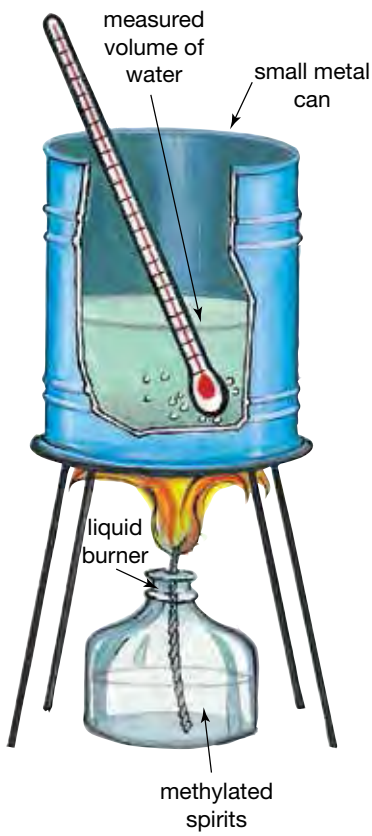
46. The specific heat capacity of aluminium is 0.900 J °C^{-1} g^{-1}. Calculate the energy needed to raise the temperature of a 7.20×10^2 g block of aluminium by 10.0 °C.

47. It takes 78.2 J of energy to raise the temperature of 45.6 g of lead from 19.2 °C to 32.5 °C. Calculate the specific heat capacity of lead.

48. A student investigated the heat output from some solid fuel blocks designed for use with model steam engines. In one experiment, it was found that the temperature of 80.0 mL of water increased from 23.2 °C to 68.5 °C, and 0.300 g of fuel was burned in the process. Calculate the heat content of this fuel in kJ g^{-1}.

49. It is useful to know how much energy can be obtained from different fuels in order to determine which would be the best fuel for a particular purpose. The apparatus in the figure below can be constructed in the laboratory to measure the heat given out when a fuel such as ethanol is burned.

The heat produced when the fuel burns is absorbed by the water in the metal can. The temperature can be measured so, given that the specific heat of water is 4.18 J and the density of water is 1.00 g mL^{-1}, the heat of combustion may be determined.



Apparatus for measuring the heat of combustion of a fuel

The results of one experiment are shown in the following table.

Volume of water in metal can =	200 mL
Thus, mass of water in can =	200 g
Rise in temperature of water =	11.0 °C
Mass of ethanol burned =	0.500 g

(a) From these results, calculate the heat produced when 1 gram of ethanol is burned.
(b) Calculate the heat of combustion for ethanol using the above results.
(c) An accurate value for the heat produced when 1 mole of ethanol burns is 1364 kJ mol^{-1}. Calculate the percentage accuracy of this experiment.
(d) Outline the sources of error in the experiment and then suggest how the design of the experiment could be improved so that more accurate heats of combustion for different fuels may be determined.

50. A student wished to estimate the ΔH value for the reaction:

$$C_5H_{12}(l) + 8O_2(g) \longrightarrow 5CO_2(g) + 6H_2O(l)$$

using an identical experimental set-up to that in question 49. Using pentane in the liquid burner, it was found that the combustion of 0.735 g caused a temperature increase of 35.0 °C.
(a) Estimate the ΔH value for the equation above using the data obtained.
(b) Given that the molar enthalpy of combustion of pentane is –3509 kJ mol^{-1}, comment on the value obtained in (a).

Exam practice questions

studyon

Sit VCAA exam

In a chemistry examination, you will be required to answer a number of multiple choice and extended response questions.

Extended response questions

1. Nitroglycerine, $C_3N_3H_5O_9$, is a dangerous explosive that releases all gaseous products according to the equation:

$$4C_3N_3H_5O_9(l) \longrightarrow 6N_2(g) + 12CO_2(g) + 10H_2O(g) + O_2(g)$$

Calculate the total volume of gaseous products that results when 1.00 kg of liquid nitroglycerine explodes at a temperature of 250 °C and a pressure of 300 kPa. **3 marks**

2. Kerosene is a hydrocarbon fuel that may be used in lamps, jet engines and camp stoves. It has a heat of combustion of 44 100 kJ kg^{-1}.

(a) Explain why the heat evolved from the combustion of kerosene is measured in kJ kg^{-1} rather than kJ mol^{-1}. **1 mark**

(b) A cup of billy tea contains 250 g of water. How many cups of tea can be made if 12.5 mL of kerosene is used to heat the water? (Assume that the temperature of the water increases from 20.0 °C to 100.0 °C, the specific heat capacity of water is 4.18 J g^{-1} $°C^{-1}$ and the heat of combustion of kerosene is 37 000 kJ L^{-1}.) **2 marks**

3. Consider the following two equations.

$$2CH_3OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(l) \quad \Delta H = -1450 \text{ kJ mol}^{-1}$$

$$2CH_3CH_2CH_2OH(l) + 9O_2(g) \longrightarrow 6CO_2(g) + 8H_2O(l) \quad \Delta H = -4032 \text{ kJ mol}^{-1}$$

(a) Calculate the mass of methanol required to produce 1.00 MJ of heat energy. **2 marks**

(b) Calculate the mass of 1-propanol required to produce 1.00 MJ of heat energy. **2 marks**

(c) Assuming that methanol and 1-propanol are made from non-renewable resources, calculate the net mass reduction (in g MJ^{-1}) of carbon dioxide when methanol is used as a fuel in preference to 1-propanol. **3 marks**

(d) Calculate the net volume reduction (in L MJ^{-1}) from (c), assuming the carbon dioxide is at 101.3 kPa and 15 °C. **3 marks**