

# Mastering Chemistry

- Book 4A
- Topic 10 Rate of Reaction



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# Content

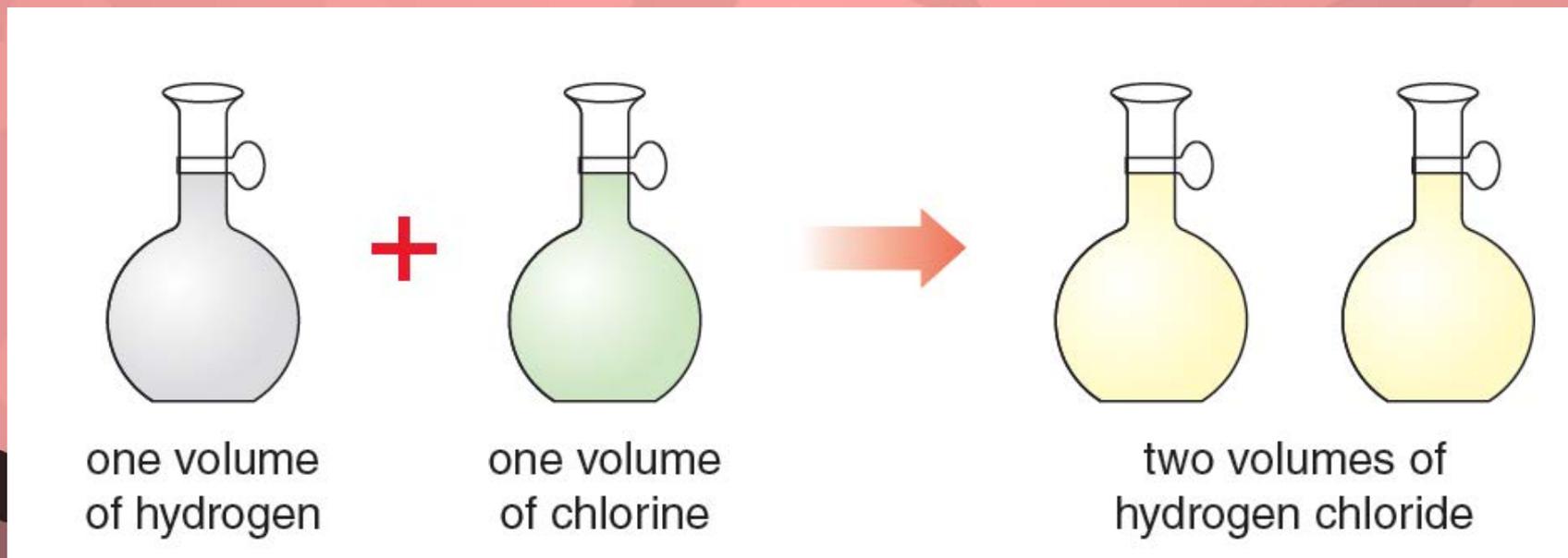
- ➔ **Key terms**
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## 38.1 Gas volume calculations (p.71)

- ◆ A French chemist called Gay-Lussac studied chemical reactions between gases.
- ◆ He found that one volume of hydrogen always reacted with exactly the same volume of chlorine to form two volumes of hydrogen chloride.





## 38.1 Gas volume calculations (p.71)

- ◆ This means that, for example, if the volumes of both gases are measured at the same temperature and pressure, 50 cm<sup>3</sup> of oxygen will contain exactly the same number of molecules as 50 cm<sup>3</sup> of carbon dioxide, despite the fact that the carbon dioxide molecule is larger and weighs more.
- ◆ The number of molecules in one mole of any gas is always  $6.02 \times 10^{23}$  and equal numbers of moles of gases contain equal numbers of molecules.

**Equal volumes of gases, measured at the same temperature and pressure, contain the same number of moles of gases.**



## 38.1 Gas volume calculations (p.71)

1 volume of  
hydrogen

reacts  
with

1 volume of  
chlorine

to  
form

2 volumes  
of hydrogen  
chloride

means

1 mole of  
hydrogen

reacts  
with

1 mole of  
chlorine

to  
form

2 moles of  
hydrogen  
chloride

and

1 molecule of  
hydrogen

reacts  
with

1 molecule of  
chlorine

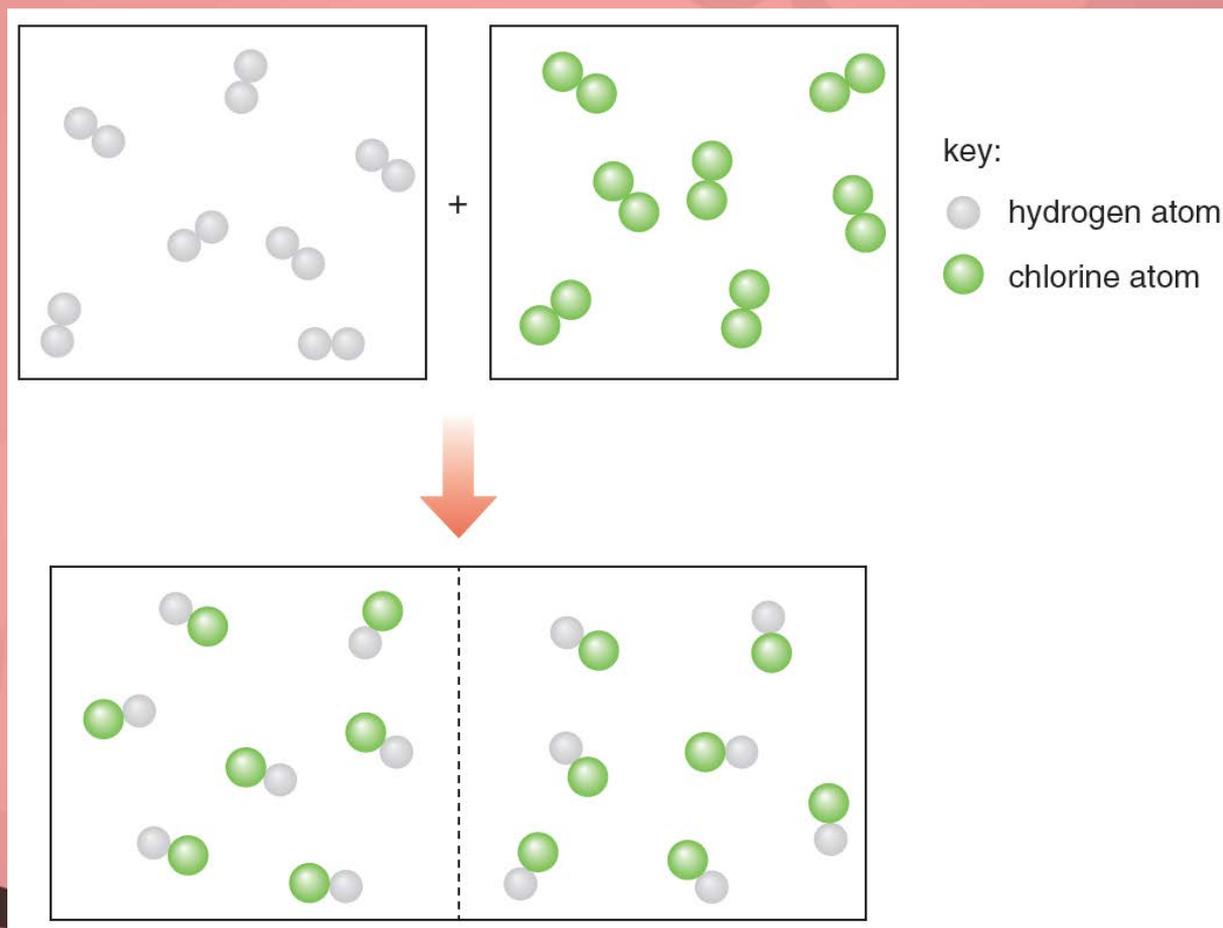
to  
form

2 molecules  
of hydrogen  
chloride



## 38.1 Gas volume calculations (p.71)

- The formation of two volumes of hydrogen chloride from one volume of hydrogen and one volume of chlorine:





## 38.2 Molar volume of a gas (p.73)

- ◆ It follows from Avogadro's law that the volume occupied by one mole of any gas must be the same under the same conditions.

**The volume occupied by one mole of a gas is called the molar volume (摩爾體積).**



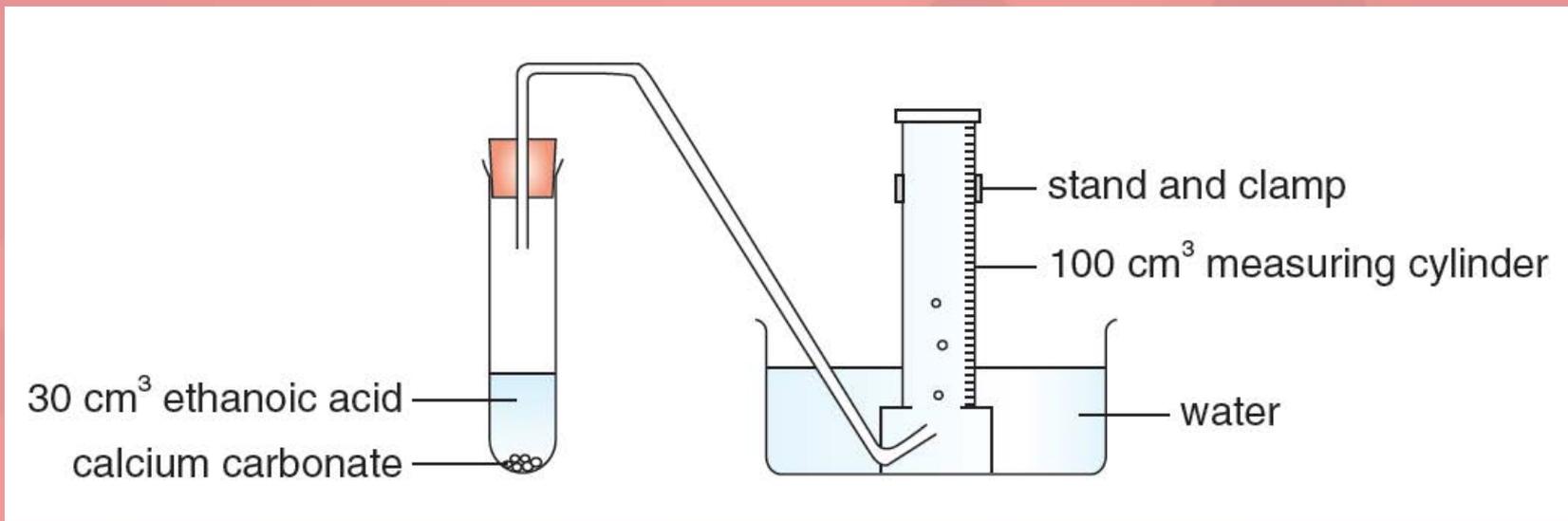
## 38.2 Molar volume of a gas (p.73)

### Determining the molar volume of carbon dioxide

- ◆ Follow the procedure below to determine the molar volume of carbon dioxide:
  - 1 Place 30 cm<sup>3</sup> of 1 mol dm<sup>-3</sup> ethanoic acid in a boiling tube.
  - 2 Set up the apparatus as shown in Fig. 38.3.
  - 3 Place approximately 0.05 g of calcium carbonate in a test tube.
  - 4 Weigh the test tube and its contents accurately.
  - 5 Remove the stopper from the boiling tube in the set-up shown in Fig. 38.3. Tip the calcium carbonate into the tube. Quickly stopper the boiling tube.
  - 6 Once the reaction is over, measure the volume of gas collected in the measuring cylinder.
  - 7 Reweigh the test tube that had contained the calcium carbonate.
  - 8 Repeat the experiment six more times, increasing the mass of calcium carbonate by about 0.05 g each time. Do not exceed 0.40 g of calcium carbonate.



# 38.2 Molar volume of a gas (p.73)



**Determining the molar volume of carbon dioxide**



## 38.2 Molar volume of a gas (p.73)

- ◆ The table below shows the results.

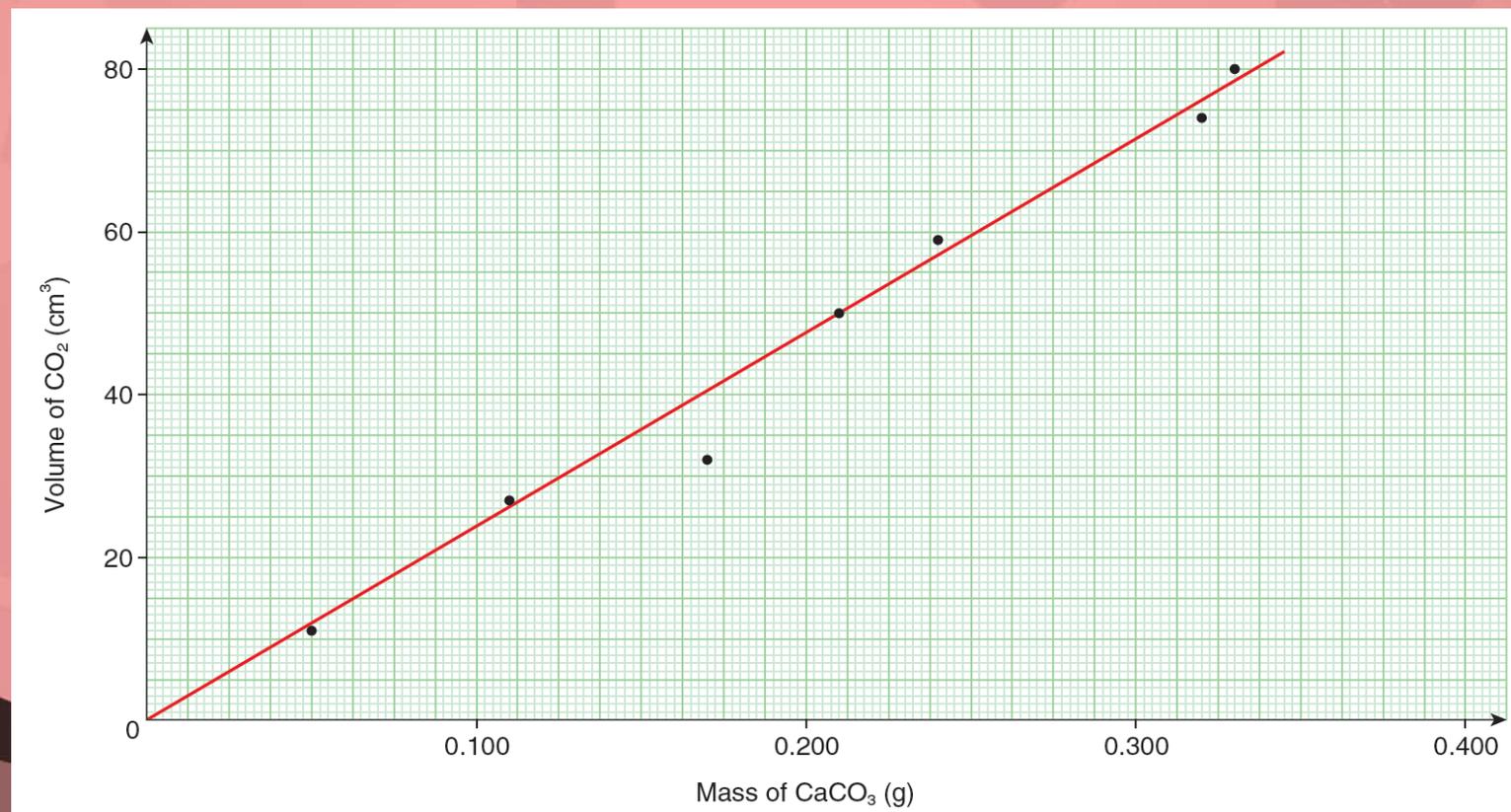
▶ **Table 38.1 Results obtained from the reaction between calcium carbonate and 1 mol dm<sup>-3</sup> ethanoic acid**

Mass of calcium carbonate (g)	Volume of carbon dioxide collected (cm <sup>3</sup> )
0.050	11
0.110	27
0.170	32
0.210	50
0.240	59
0.320	74
0.330	80



## 38.2 Molar volume of a gas (p.73)

- ◆ The volume of carbon dioxide collected is plotted against the mass of calcium carbonate. A straight line passing through the origin is obtained. From the graph, you can see that  $60 \text{ cm}^3$  of carbon dioxide can be made from  $0.250 \text{ g}$  of calcium carbonate.





## 38.2 Molar volume of a gas (p.73)

- ◆ Calcium carbonate reacts with ethanoic acid according to the equation below:



- ◆ Number of moles of 0.250 g  $\text{CaCO}_3 = \frac{0.250 \text{ g}}{100.1 \text{ g mol}^{-1}} = 0.00250 \text{ mol}$

According to the equation, 1 mole of  $\text{CaCO}_3$  makes 1 mole of  $\text{CO}_2$ .  
i.e. number of moles of  $\text{CO}_2$  made in this reaction = 0.00250 mol

$$\text{Molar volume of CO}_2 = \frac{60 \text{ cm}^3}{0.00250 \text{ mol}} = 24\,000 \text{ cm}^3 \text{ mol}^{-1}$$

**At room temperature and pressure, the molar volume of any gas is  $24.0 \text{ dm}^3 \text{ mol}^{-1}$  (or  $24\,000 \text{ cm}^3 \text{ mol}^{-1}$ )**



## 38.2 Molar volume of a gas (p.73)

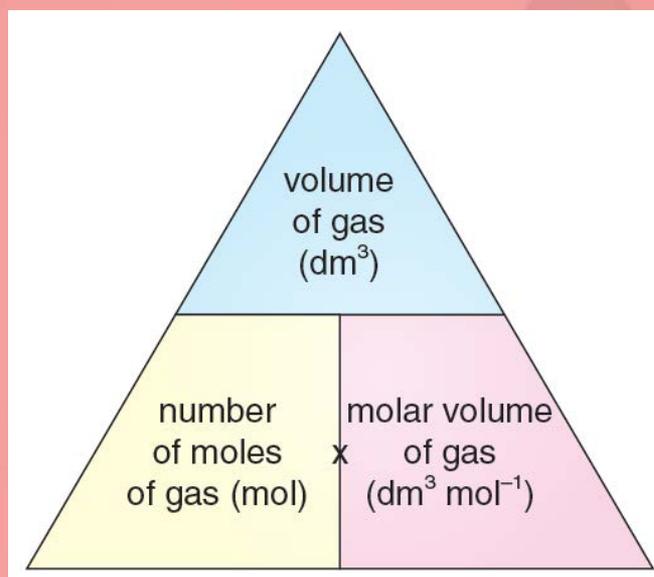
- ◆ To give an idea of scale,  $24.0 \text{ dm}^3$  is equivalent to the amount of air contained in four footballs.
- ◆ So 1.0 mole of carbon dioxide occupies  $24.0 \text{ dm}^3$  at room temperature and pressure. Hence 2.0 moles of carbon dioxide occupy  $48.0 \text{ dm}^3$  and 0.50 mole of carbon dioxide occupies  $12.0 \text{ dm}^3$ . You can work out the volume of a gas by multiplying the number of moles by the molar volume.

**Volume of gas (at room temperature and pressure)  
= number of moles of gas x molar volume of gas (at  
room temperature and pressure)**



## 38.2 Molar volume of a gas (p.73)

- The relationship between the number of moles of a gas and its volume is summarised in the figure below. You can work out the amount, in moles, of a gas by dividing the volume by the molar volume.



$$\text{Number of moles of gas} = \frac{\text{volume of gas (at r.t. and pressure)}}{\text{molar volume of gas (at r.t. and pressure)}}$$



## 38.2 Molar volume of a gas (p.73)

### Q (Example 38.1)

What is the volume of 0.150 mole of carbon dioxide, measured at room temperature and pressure?

(Molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

### A

Volume of  $\text{CO}_2$  (at room temperature and pressure)

= number of moles of  $\text{CO}_2$  x molar volume of gas (at room temperature and pressure)

=  $0.150 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$

=  $3.60 \text{ dm}^3$

$\therefore$  the volume of 0.150 mole of carbon dioxide is  $3.60 \text{ dm}^3$ .



## 38.2 Molar volume of a gas (p.73)

### Q (Example 38.2)

What is the number of molecules in 792 cm<sup>3</sup> of oxygen, measured at room temperature and pressure?

(Molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>;  
Avogadro constant = 6.02 x 10<sup>23</sup> mol<sup>-1</sup>)

### A

$$\begin{aligned}\text{Number of moles of oxygen} &= \frac{\text{volume of oxygen (at r.t.and pressure)}}{\text{molar volume of gas (at r.t.and pressure)}} \\ &= \frac{792 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} = 0.0330 \text{ mol}\end{aligned}$$

Number of oxygen molecules = 0.0330 mol x 6.02 x 10<sup>23</sup> mol<sup>-1</sup> = 1.99 x 10<sup>22</sup>  
∴ 792 cm<sup>3</sup> of oxygen contain 1.99 x 10<sup>22</sup> molecules.



## 38.2 Molar volume of a gas (p.73)

### Practice 38.1

- 1 A flask contains 420 cm<sup>3</sup> of butane at room temperature and pressure.  
What is the number of moles of gas in the flask?  
(Molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

$$\text{Number of moles of butane} = \frac{\text{volume of butane}}{\text{molar volume of gas}} = \frac{420 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}}$$

$$= 0.0175 \text{ mol}$$

∴ 0.0175 mole of gas is in the flask.



## 38.2 Molar volume of a gas (p.73)

2 What volume do 7.0 g of nitrogen gas occupy at room temperature and pressure?

(Relative atomic mass: N = 14.0; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

$$\text{Number of moles of N}_2 = \frac{7.0 \text{ g}}{2 \times 14.0 \text{ g mol}^{-1}} = 0.250 \text{ mol}$$

$$\begin{aligned} \text{Volume of N}_2 &= \text{number of moles of N}_2 \times \text{molar volume of gas} \\ &= 0.250 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1} = 6.00 \text{ dm}^3 \\ \therefore \text{the nitrogen gas occupies } 6.00 \text{ dm}^3. \end{aligned}$$



## 38.2 Molar volume of a gas (p.73)

3 At room temperature and pressure, it is found that 1.08 dm<sup>3</sup> of a gas has a mass of 3.96 g. What is the molar mass of the gas?

(Molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

$$\text{Number of moles of gas} = \frac{1.08 \text{ dm}^3}{24.0 \text{ dm}^3 \text{ mol}^{-1}}$$

$$= 0.0450 \text{ mol}$$

$$\text{Molar mass of gas} = \frac{3.96 \text{ g}}{0.0450 \text{ mol}} = 88.0 \text{ g mol}^{-1}$$

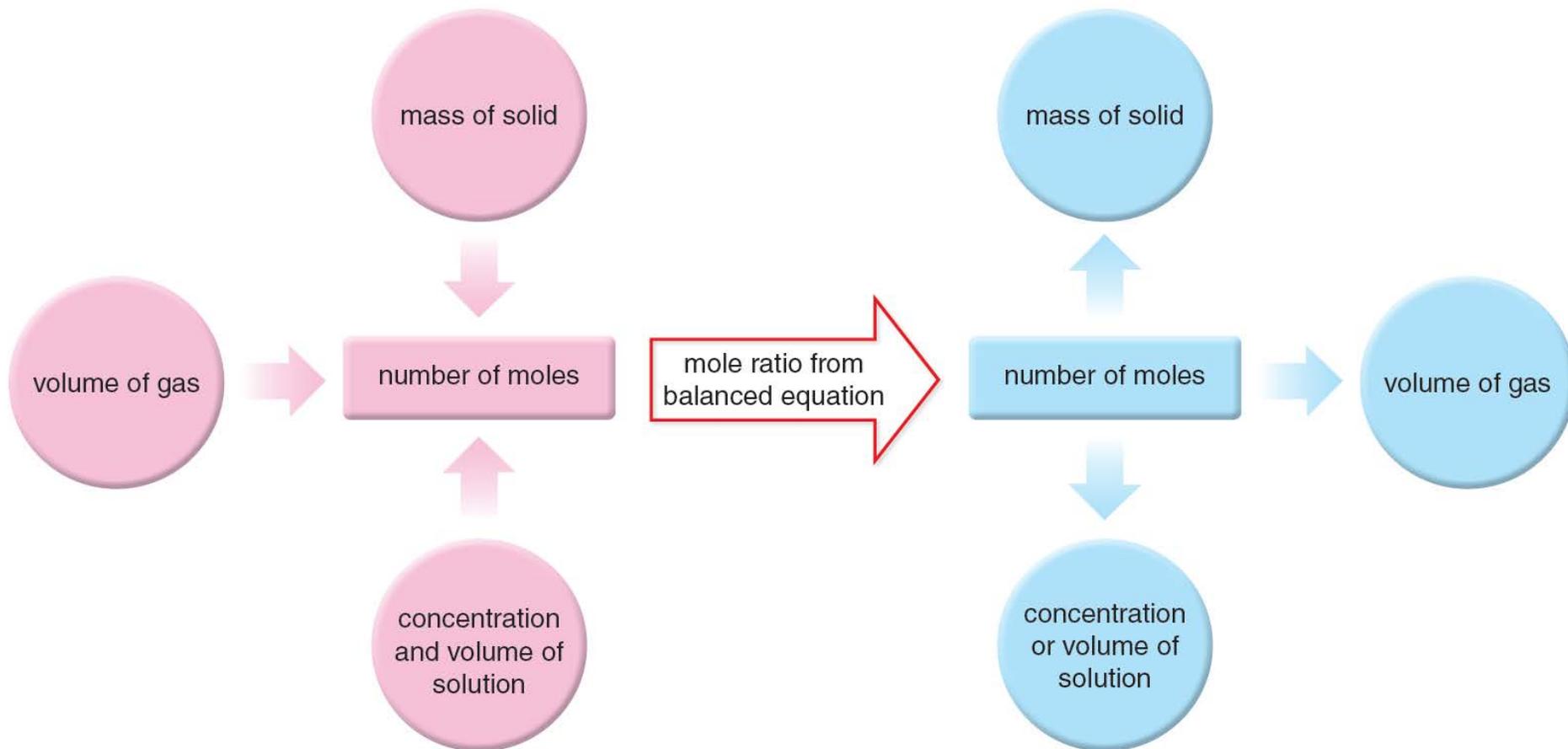
∴ the molar mass of the gas is 88.0 g mol<sup>-1</sup>.



# 38.3 Calculations involving gases and solids or solutions (p.78)

information given in the question

information the question asks





## 38.3 Calculations involving gases and solids or solutions (p.78)

### Q (Example 38.3)

In an experiment carried out under room conditions, 10.0 cm<sup>3</sup> of H<sub>2</sub>O<sub>2</sub>(aq) were allowed to decompose into O<sub>2</sub>(g) and H<sub>2</sub>O(l) in the presence of a catalyst. A total of 96.0 cm<sup>3</sup> of gas was collected.

Calculate the initial concentration of the H<sub>2</sub>O<sub>2</sub>(aq), in mol dm<sup>-3</sup> (Molar volume of gas at room conditions = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

### A

H<sub>2</sub>O<sub>2</sub>(aq) decomposes according to the equation:



$$\text{Number of moles of O}_2 = \frac{\text{volume of O}_2 \text{ (at r.t. and pressure)}}{\text{molar volume of gas (at r.t. and pressure)}} = \frac{96.0 \text{ cm}^3}{24\,000 \text{ cm}^3}$$

$$= 4.00 \times 10^{-3} \text{ mol}$$



## 38.3 Calculations involving gases and solids or solutions (p.78)

According to the equation, 2 moles of  $\text{H}_2\text{O}_2$  decompose to give 1 mole of  $\text{O}_2$ .

i.e. number of moles of  $\text{H}_2\text{O}_2$  decomposed =  $2 \times 4.00 \times 10^{-3} \text{ mol}$

=  $8.00 \times 10^{-3} \text{ mol} = \text{concentration of } \text{H}_2\text{O}_2 \times \frac{10.0}{1000} \text{ dm}^3$

Concentration of  $\text{H}_2\text{O}_2 = 8.00 \times 10^{-3} \text{ mol} \times \frac{1000}{10.0} \text{ dm}^{-3} = 0.800 \text{ mol dm}^{-3}$

$\therefore$  the initial concentration of the  $\text{H}_2\text{O}_2(\text{aq})$  is  $0.800 \text{ mol dm}^{-3}$ .



## 38.3 Calculations involving gases and solids or solutions (p.78)

### Q (Example 38.4)

A hydrocarbon burns completely in oxygen to give 1.44 g of water and 1 920 cm<sup>3</sup> of carbon dioxide, at room temperature and pressure.

What is the empirical formula of this hydrocarbon?

(Relative atomic masses: H = 1.0, C = 12.0; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

### A

$$\text{Number of moles of CO}_2 = \frac{\text{volume of CO}_2 \text{ (at r.t. and pressure)}}{\text{molar volume of gas (at r.t. and pressure)}}$$

$$= \frac{1\,920 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} = 0.0800 \text{ mol}$$

i.e. number of moles of C in the hydrocarbon = 0.0800 mol



## 38.3 Calculations involving gases and solids or solutions (p.78)

$$\text{Number of moles of H in 1.44 g of water} = 2 \times \frac{1.44 \text{ g}}{18.0 \text{ g mol}^{-1}} = 0.160 \text{ mol}$$

Mole ratio of C:H = 0.0800 : 0.160 = 1 : 2

∴ the empirical formula of this hydrocarbon is CH<sub>2</sub>.



## 38.3 Calculations involving gases and solids or solutions (p.78)

### Q (Example 38.5)

What is the theoretical volume of carbon dioxide that can be obtained, at room temperature and pressure, when 1.40 g of  $\text{Na}_2\text{CO}_3(\text{s})$  react with 50.0  $\text{cm}^3$  of 1.00  $\text{mol dm}^{-3}$   $\text{HCl}(\text{aq})$ ?

(Relative atomic masses: C = 12.0, O = 16.0, Na = 23.0; molar volume of gas at room temperature and pressure = 24.0  $\text{dm}^3 \text{mol}^{-3}$ )

### A

$\text{Na}_2\text{CO}_3(\text{s})$  reacts with  $\text{HCl}(\text{aq})$  according to the equation below:



$$\text{Number of moles of Na}_2\text{CO}_3 = \frac{1.40 \text{ g}}{106.0 \text{ g mol}^{-1}} = 0.0132 \text{ mol}$$

$$\text{Number of moles of HCl} = 1.00 \text{ mol dm}^{-3} \times \frac{50.0}{1000} \text{ dm}^3 = 0.0500 \text{ mol}$$



## 38.3 Calculations involving gases and solids or solutions (p.78)

According to the equation, 1 mole of  $\text{Na}_2\text{CO}_3$  reacts with 2 moles of  $\text{HCl}$  to form 1 mole of  $\text{CO}_2$ .

In this reaction, 0.0132 mole of  $\text{Na}_2\text{CO}_3$  reacts with 0.0264 mole of  $\text{HCl}$ . Thus,  $\text{HCl}$  is in excess and  $\text{Na}_2\text{CO}_3$  is the limiting reactant.

Number of moles of  $\text{CO}_2$  obtained = 0.0132 mol

Volume of  $\text{CO}_2$  obtained (at r.t. and pressure)

= number of moles of  $\text{CO}_2$  x molar volume of gas (at r.t. and pressure)

= 0.0132 mol x 24 000  $\text{cm}^3 \text{ mol}^{-1}$

= 317  $\text{cm}^3$

$\therefore$  317  $\text{cm}^3$  of carbon dioxide can be obtained at room temperature and pressure.



## 38.3 Calculations involving gases and solids or solutions (p.78)

### Practice 38.2

1 Potassium chlorate decomposes when heated.



What is the mass of potassium chlorate that decomposes to produce 180.0 cm<sup>3</sup> of oxygen, measured at room temperature and pressure?

(Relative atomic masses: O = 16.0, Cl = 35.5, K = 39.1; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)



## 38.3 Calculations involving gases and solids or solutions (p.78)

$$\begin{aligned}\text{Number of moles of O}_2 &= \frac{\text{volume of O}_2}{\text{molar volume of gas}} \\ &= \frac{180.0 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} \\ &= 7.50 \times 10^{-3} \text{ mol}\end{aligned}$$

2 moles of  $\text{KClO}_3$  decompose to produce 3 moles of  $\text{O}_2$ .

i.e. number of moles of  $\text{KClO}_3$  that decompose

$$= \frac{2}{3} \times 7.50 \times 10^{-3} \text{ mol}$$

$$= 5.00 \times 10^{-3} \text{ mol}$$

$$\begin{aligned}\text{Mass of KClO}_3 \text{ that decomposes} &= 5.00 \times 10^{-3} \text{ mol} \times 122.6 \text{ g mol}^{-1} \\ &= 0.613 \text{ g}\end{aligned}$$

$\therefore$  0.613 g of  $\text{KClO}_3$  decomposes.



## 38.3 Calculations involving gases and solids or solutions (p.78)

2 65.0 cm<sup>3</sup> of a sample of sulphuric acid reacts with excess zinc to form 234 cm<sup>3</sup> of hydrogen, measured at room temperature and pressure.

What is the concentration of the sulphuric acid?

(Molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)



## 38.3 Calculations involving gases and solids or solutions (p.78)



$$\begin{aligned}\text{Number of moles of H}_2 &= \frac{\text{volume of H}_2}{\text{molar volume of gas}} \\ &= \frac{234 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} \\ &= 9.75 \times 10^{-3} \text{ mol}\end{aligned}$$

1 mole of  $\text{H}_2\text{SO}_4$  reacts with zinc to form 1 mole of  $\text{H}_2$ .

i.e. number of moles of  $\text{H}_2\text{SO}_4 = 9.75 \times 10^{-3} \text{ mol}$

$$\begin{aligned}\text{Concentration of sulphuric acid} &= \frac{9.75 \times 10^{-3} \text{ mol}}{\frac{65.0}{1\,000} \text{ dm}^3} \\ &= 0.150 \text{ mol dm}^{-3}\end{aligned}$$

$\therefore$  concentration of the sulphuric acid is  $0.150 \text{ mol dm}^{-3}$ .



## 38.3 Calculations involving gases and solids or solutions (p.78)

3 Tin reacts with nitric acid according to the equation below:



0.500 g of an impure sample of tin reacts with excess nitric acid to form 269 cm<sup>3</sup> of nitrogen dioxide, measured at room temperature and pressure.

What is the percentage by mass of tin in the sample?

(Relative atomic mass: Sn = 118.7; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)



## 38.3 Calculations involving gases and solids or solutions (p.78)

$$\begin{aligned}\text{Number of moles of NO}_2 &= \frac{\text{volume of NO}_2}{\text{molar volume of gas}} \\ &= \frac{269 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} \\ &= 0.0112 \text{ mol}\end{aligned}$$

1 mole of Sn reacts with nitric acid to form 4 moles of NO<sub>2</sub>.

$$\begin{aligned}\text{i.e. number of moles of Sn that react} &= \frac{0.0112}{4} \text{ mol} \\ &= 2.80 \times 10^{-3} \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Mass of Sn in sample} &= 2.80 \times 10^{-3} \text{ mol} \times 118.7 \text{ g mol}^{-1} \\ &= 0.332 \text{ g}\end{aligned}$$

$$\begin{aligned}\text{Percentage by mass of Sn in sample} &= \frac{0.332 \text{ g}}{0.500 \text{ g}} \times 100\% \\ &= 66.4\%\end{aligned}$$

∴ the percentage by mass of tin in the sample is 66.4%.



## 38.4 Calculations using reacting volumes of gases (p.84)

- ◆ Gas volume calculations are straightforward when all the relevant substances are gases. In these cases, the ratio of the gas volumes in the reaction is the same as the ratio of the numbers of moles in the equation.
- ◆ This is the case because the volume of a gas, under given conditions of temperature and pressure, depends only on the number of moles of the gas and not on the type of gas.



## 38.4 Calculations using reacting volumes of gases (p.84)

### Q (Example 38.6)

Propane and oxygen react to form carbon dioxide and water according to the equation below:



What is the volume of oxygen required for the complete combustion of 408 cm<sup>3</sup> of propane (measured at r.t. and pressure)?



## 38.4 Calculations using reacting volumes of gases (p.84)

A

Method 1

$$\begin{aligned}\text{Number of moles of } C_3H_8 &= \frac{\text{volume of } C_3H_8 \text{ (at r.t. and pressure)}}{\text{molar volume of gas (at r.t. and pressure)}} \\ &= \frac{408 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} = 0.0170 \text{ mol}\end{aligned}$$

According to the equation, 1 mole of  $C_3H_8$  requires 5 moles of  $O_2$  for complete combustion.

i.e. number of moles of  $O_2$  required =  $5 \times 0.0170 \text{ mol} = 0.0850 \text{ mol}$

Volume of  $O_2$  required (at r.t. and pressure)  
= number of moles  $O_2$  x molar volume of gas (at r.t. and pressure)  
=  $0.0850 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1} = 2.04 \text{ dm}^3$



## 38.4 Calculations using reacting volumes of gases (p.84)

### Method 2

According to the equation, 1 mole of  $C_3H_8$  requires 5 moles of  $O_2$  for complete combustion.

i.e. 24 000  $cm^3$  of  $C_3H_8$  require  $5 \times 24\ 000\ cm^3$  of  $O_2$  for complete combustion.

$$\text{Volume of } O_2 \text{ required (at r.t. and pressure)} = 408\ cm^3 \times \frac{5 \times 24\ 000\ cm^3}{24\ 000\ cm^3}$$
$$= 2\ 040\ cm^3 = 2.04\ dm^3$$

$\therefore$  2.04  $dm^3$  of oxygen are required for the complete combustion of 408  $cm^3$  of propane.



## 38.4 Calculations using reacting volumes of gases (p.84)

### Q (Example 38.7)

Hydrogen sulphide and oxygen react to form sulphur dioxide and water according to the equation below:



250 cm<sup>3</sup> of hydrogen sulphide and 600 cm<sup>3</sup> of oxygen are mixed, and allowed to react.

Deduce the volume of the remaining gas, measured at room temperature and pressure, at the end of the reaction.



## 38.4 Calculations using reacting volumes of gases (p.84)

### A

According to the equation, 2 moles of  $\text{H}_2\text{S}$  react with 3 moles of  $\text{O}_2$  to form 2 moles of  $\text{SO}_2$ .

i.e.  $250 \text{ cm}^3$  of  $\text{H}_2\text{S}$  react with  $\frac{3}{2} \times 250 \text{ cm}^3$  (i.e.  $375 \text{ cm}^3$ ) of  $\text{O}_2$  to form  $250 \text{ cm}^3$  of  $\text{SO}_2$ .

Volume of  $\text{O}_2$  left after reaction =  $(600 - 375) \text{ cm}^3 = 225 \text{ cm}^3$

Volume of remaining gas = volume of  $\text{O}_2$  left + volume of  $\text{SO}_2$  formed  
=  $(225 + 250) \text{ cm}^3 = 475 \text{ cm}^3$

$\therefore$  the volume of the remaining gas is  $475 \text{ cm}^3$ .



## 38.4 Calculations using reacting volumes of gases (p.84)

### Q (Example 38.8)

40 cm<sup>3</sup> of a gaseous hydrocarbon, C<sub>x</sub>H<sub>y</sub>, were mixed with an excess of oxygen (160 cm<sup>3</sup>) and exploded.

The final volume of the resulting mixture was 120 cm<sup>3</sup> at the end of the reaction. This mixture was shaken with a little sodium hydroxide solution. The volume reduced to 40 cm<sup>3</sup>.

(All volumes were measured at room temperature and pressure.)  
Deduce the chemical formula of the hydrocarbon.



## 38.4 Calculations using reacting volumes of gases (p.84)

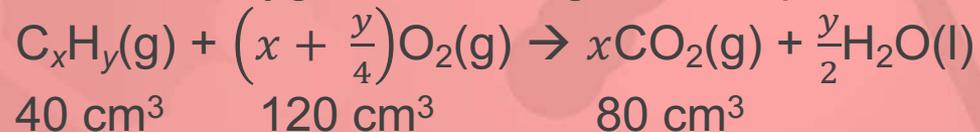
A

Volume of CO<sub>2</sub> in resulting mixture = volume of gas absorbed by NaOH(aq)  
 = (120 - 40) cm<sup>3</sup> = 80 cm<sup>3</sup>

Volume of O<sub>2</sub> in resulting mixture = 40 cm<sup>3</sup>

Volume of O<sub>2</sub> consumed = (160 - 40) cm<sup>3</sup> = 120 cm<sup>3</sup>

C<sub>x</sub>H<sub>y</sub> reacted with oxygen according to the equation below:



i.e. 1 mole of C<sub>x</sub>H<sub>y</sub> reacts with 3 moles of O<sub>2</sub> to form 2 moles of CO<sub>2</sub>.

The value of x is 2.

$$x + \frac{y}{4} = 3, \quad y = 4$$

∴ the chemical formula of the hydrocarbon is C<sub>2</sub>H<sub>4</sub>.



## 38.4 Calculations using reacting volumes of gases (p.84)

### Practice 38.3

- 1 Assuming that air contains 20% of oxygen by volume, how much air is required to burn completely 100 cm<sup>3</sup> of but-1-ene (C<sub>4</sub>H<sub>8</sub>)?  
(All volumes are measured at the same temperature and pressure.)



According to the equation, 1 mole of C<sub>4</sub>H<sub>8</sub> requires 6 moles of O<sub>2</sub> for complete combustion.  
i.e. 24 000 cm<sup>3</sup> of C<sub>4</sub>H<sub>8</sub> require 6 x 24 000 cm<sup>3</sup> of O<sub>2</sub> for complete combustion.



$$\begin{aligned} \text{Volume of O}_2 \text{ required} &= 100\text{ cm}^3 \times \frac{6 \times 24\,000\text{ cm}^3}{24\,000\text{ cm}^3} \\ &= 600\text{ cm}^3 \end{aligned}$$

$$\begin{aligned} \text{Volume of air required} &= \frac{600\text{ cm}^3}{20.0\%} \\ &= 3\,000\text{ cm}^3 \end{aligned}$$

∴ 3 000 cm<sup>3</sup> of air are required.



## 38.4 Calculations using reacting volumes of gases (p.84)

2 Sulphur dioxide and oxygen react to form sulphur trioxide according to the equation below:

$$2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$$

200 cm<sup>3</sup> of SO<sub>2</sub>(g) and 200 cm<sup>3</sup> of O<sub>2</sub>(g) are mixed and allowed to react. Deduce the volume of the remaining gas.  
(All volumes are measured at the same temperature and pressure.)

According to the equation, 2 moles of SO<sub>2</sub> react with 1 mole of O<sub>2</sub> to form 2 moles of SO<sub>3</sub>.

i.e. 200 cm<sup>3</sup> of SO<sub>2</sub> react with  $\frac{200}{2}$  cm<sup>3</sup> (i.e. 100 cm<sup>3</sup>) of O<sub>2</sub> to form 200 cm<sup>3</sup> of SO<sub>3</sub>.

Volume of O<sub>2</sub> left after reaction = (200 – 100) cm<sup>3</sup>  
= 100 cm<sup>3</sup>

Volume of remaining gas = volume of O<sub>2</sub> left + volume of SO<sub>3</sub> formed  
= (100 + 200) cm<sup>3</sup>  
= 300 cm<sup>3</sup>

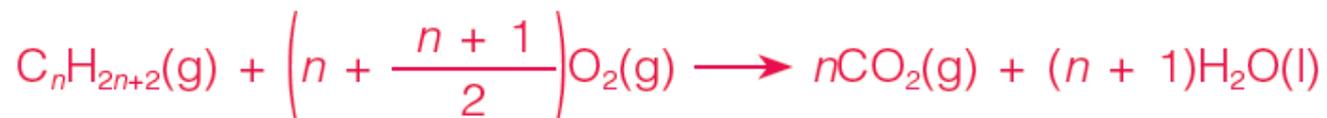
∴ the volume of the remaining gas is 300 cm<sup>3</sup>.



## 38.4 Calculations using reacting volumes of gases (p.84)

3 Alkane X burns completely in oxygen. The volume of oxygen required and the volume of carbon dioxide formed is in the ratio of 8 : 5. Deduce the molecular formula of X. (All volumes are measured at the same temperature and pressure.)

Suppose the molecular formula of X is  $C_nH_{2n+2}$ .



$$\frac{\text{Volume of oxygen}}{\text{Volume of } CO_2} = \frac{8}{5} = \frac{n + \frac{n+1}{2}}{n}$$

$$8n = 5n + \frac{5n+5}{2}$$

$$n = 5$$

$\therefore$  the molecular formula of alkane X is  $C_5H_{12}$ .



## Key terms (p.88)

Avogadro's law	亞佛加德羅定律	molar volume	摩爾體積



## Summary (p.89)

- 1 Avogadro's law states that:
  - a) Equal volumes of gases, measured at the same temperature and pressure, contain the same number of molecules.
  - b) Equal volumes of gases at the same temperature and pressure, contain the same number of moles of gases.
- 2 The volume occupied by one mole of a gas is called the molar volume.
- 3 At room temperature and pressure, the molar volume of any gas is  $24.0 \text{ dm}^3 \text{ mol}^{-1}$  (or  $24\,000 \text{ cm}^3 \text{ mol}^{-1}$ ).



## Summary (p.89)

4

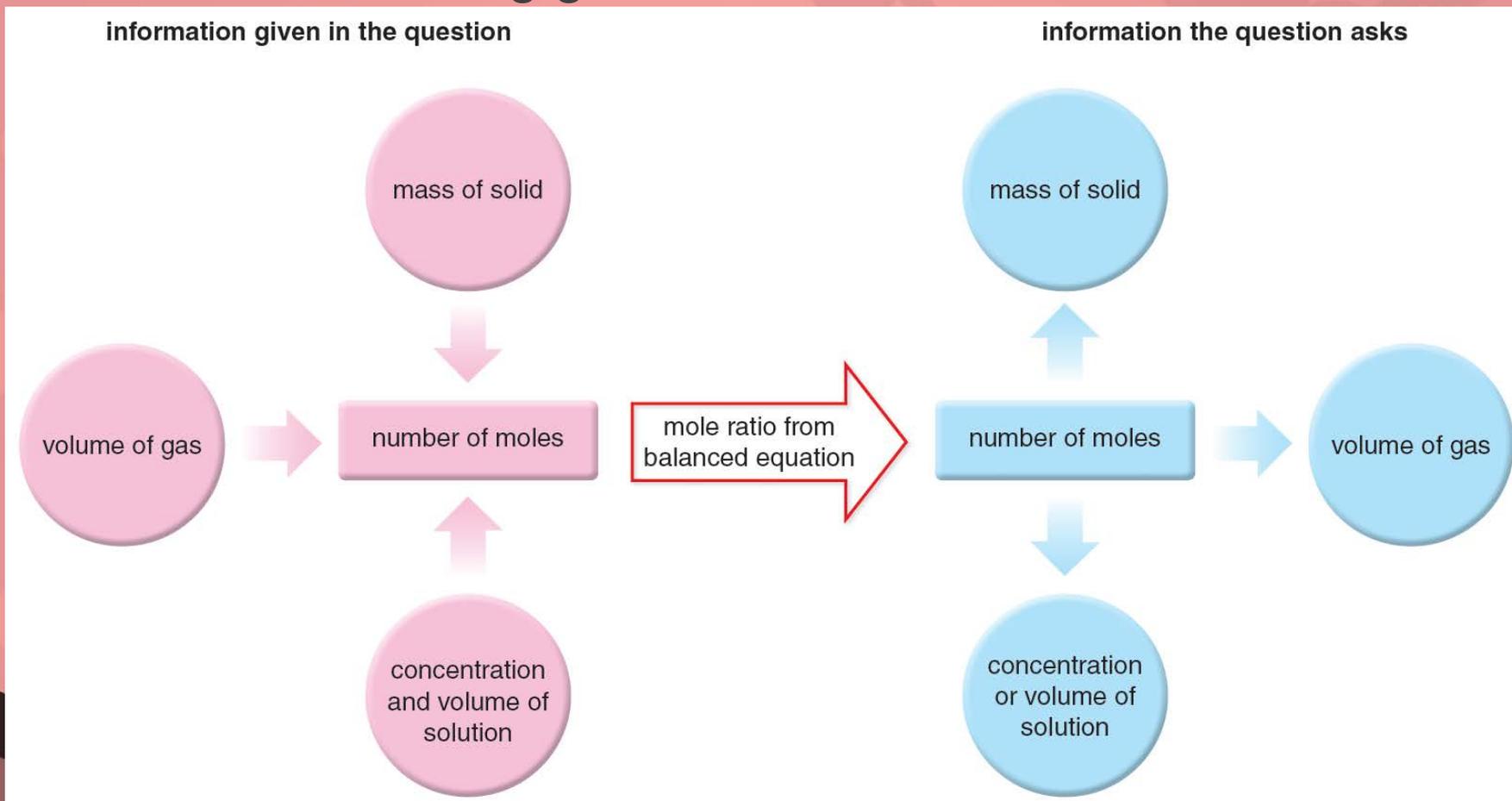
a) Volume of gas (at room temperature and pressure)  
= number of moles of gas x  
molar volume of gas (at room temperature and pressure)

b) Number of moles of gas =  $\frac{\text{volume of gas(at r.t.and pressure)}}{\text{molar volume of gas (at r.t.and pressure)}}$



## Summary (p.89)

5 The following flow diagram summarises the steps of calculations involving gases and solids or solutions:





## Unit Exercise (p.90)

**Note: Questions are rated according to ascending level of difficulty (from 1 to 5):**



**question targeted at level 3 and above;**



**question targeted at level 4 and above;**



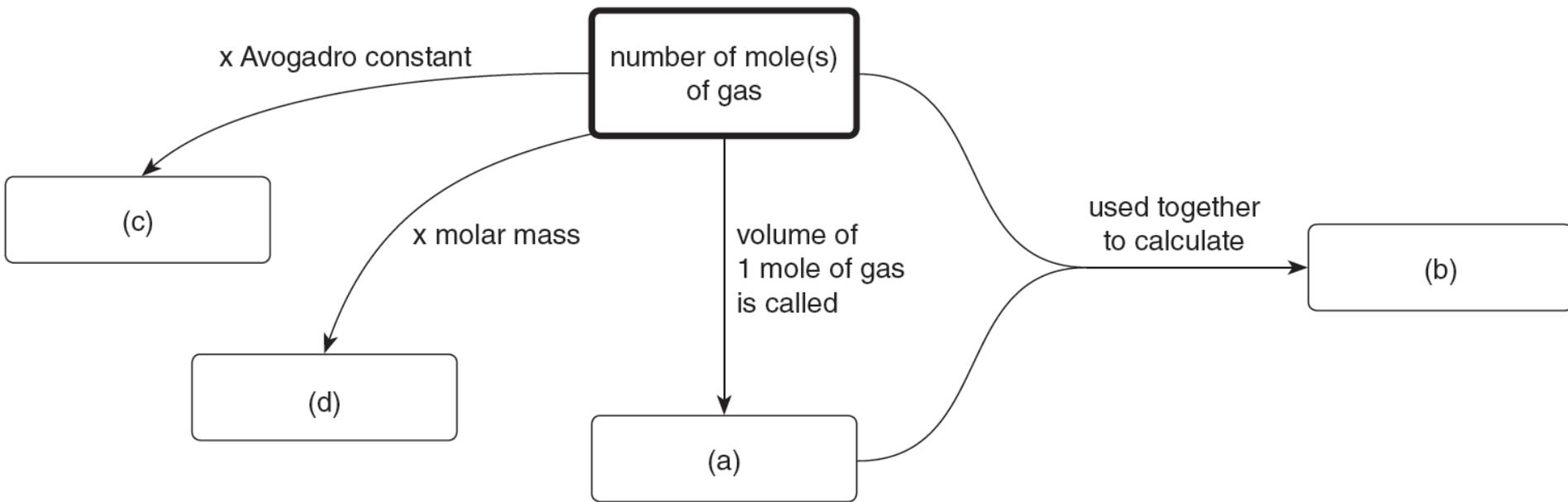
**question targeted at level 5.**

**' \* ' indicates 1 mark is given for effective communication.**

# Unit Exercise (p.90)

## PART I KNOWLEDGE AND UNDERSTANDING

1 Complete the following concept map.



- a) molar volume
- b) volume of gas
- c) number of particles
- d) mass of gas

 Unit Exercise (p.90)**PART II MULTIPLE CHOICE QUESTIONS**

2 Which of the following gas samples occupies the SMALLEST volume at room temperature and pressure?

(Relative atomic masses: C = 12.0, N = 14.0, O = 16.0)

- A 1 gram of carbon dioxide
- B 1 gram of nitrogen
- C 1 gram of carbon monoxide
- D 1 gram of oxygen

**Answer: A**

	Molar mass (g mol <sup>-1</sup> )	Number of moles of 1 g of sample (mol)
CO <sub>2</sub>	44.0	$\frac{1}{44.0}$
N <sub>2</sub>	28.0	$\frac{1}{28.0}$
CO	28.0	$\frac{1}{28.0}$
O <sub>2</sub>	32.0	$\frac{1}{32.0}$

∴ 1 g of CO<sub>2</sub> occupies the SMALLEST volume.



## Unit Exercise (p.90)

- 3 How many moles of ATOMS are present in 576 cm<sup>3</sup> of carbon dioxide at room temperature and pressure?  
(Molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)
- A 0.0720 mol
  - B 0.0417 mol
  - C 0.0240 mol
  - D 0.0138 mol

**Answer: A**

$$\begin{aligned}\text{Number of moles of CO}_2 &= \frac{\text{volume of CO}_2}{\text{molar volume of gas}} \\ &= \frac{576 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} \\ &= 0.0240 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Number of moles of atoms} &= 3 \times 0.0240 \text{ mol} \\ &= 0.0720 \text{ mol}\end{aligned}$$

 Unit Exercise (p.90)

4 Sodium nitrate decomposes on heating.



What is the maximum volume of oxygen, measured at room temperature and pressure, that can be obtained by heating 0.800 mole of sodium nitrate?

(Molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

- A  $30.0 \text{ dm}^3$
- B  $19.2 \text{ dm}^3$
- C  $9.60 \text{ dm}^3$
- D  $12.0 \text{ dm}^3$

**Answer: C**

Heating 0.800 mole of  $\text{NaNO}_3$  gives 0.400 mole of  $\text{O}_2$ .

Volume of  $\text{O}_2$  = number of moles of  $\text{O}_2$  x molar volume of gas  
=  $0.400 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$   
=  $9.60 \text{ dm}^3$



## Unit Exercise (p.90)

5 A chemical explosion is the result of a very rapid reaction that generates a large quantity of heat, and usually a large quantity of gas.



The explosive RDX ( $C_3H_6N_6O_6$ ) decomposes according to the equation below during an explosion:



**Answer: A**

What is the volume of gas, measured at room temperature and pressure, when 1.02 g of RDX decompose?

(Relative atomic masses: H = 1.0, C = 12.0, N = 14.0, O = 16.0; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

- A  $662 \text{ cm}^3$
- B  $828 \text{ cm}^3$
- C  $994 \text{ cm}^3$
- D  $1\,460 \text{ cm}^3$

$$\text{Number of moles of RDX} = \frac{1.02 \text{ g}}{222.0 \text{ g mol}^{-1}}$$

1 mole of RDX gives 6 moles of gas during an explosion.

$$\begin{aligned} \text{i.e. number of moles of gas} &= 6 \times \frac{1.02}{222.0} \text{ mol} \\ &= 0.0276 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Volume of gas} &= \text{number of moles of gas} \times \text{molar volume of gas} \\ &= 0.0276 \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1} \\ &= 662 \text{ cm}^3 \end{aligned}$$



## Unit Exercise (p.90)

Answer: B

6 The first reaction that occurs when an airbag is set off is:



6.5 g of  $\text{NaN}_3$  completely decompose.

(Relative atomic masses: N = 14.0, Na = 23.0; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

Which of the following statements is correct?

A 2.3 g of products are formed.

B  $3.6 \text{ dm}^3$  of gas are formed (at room temperature and pressure).

C 4.6 g of sodium is formed.

D The volume of nitrogen formed is 1.5 times the volume of sodium formed.

*(OCR Advanced Subsidiary Level, Chem. B (Salters), H033, Sample Question Paper, 2016, 11(a))*



## Unit Exercise (p.90)

$$\begin{aligned}\text{Number of moles of NaN}_3 \text{ that decompose} &= \frac{6.5 \text{ g}}{65.0 \text{ g mol}^{-1}} \\ &= 0.10 \text{ mol}\end{aligned}$$

2 moles of  $\text{NaN}_3$  decompose to give 3 moles of gas.

$$\begin{aligned}\text{i.e. number of moles of gas} &= \frac{3}{2} \times 0.10 \text{ mol} \\ &= 0.15 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Volume of gas} &= \text{number of moles of gas} \times \text{molar volume of gas} \\ &= 0.15 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1} \\ &= 3.6 \text{ dm}^3\end{aligned}$$



## Unit Exercise (p.90)

7 An oxide of metal M reacts completely with carbon to give 8.93 g of M and 2.88 dm<sup>3</sup> of carbon dioxide, measured at room temperature and pressure. (Relative atomic mass: M = 55.8; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

**Answer: C**

- A M<sub>2</sub>O
- B MO<sub>2</sub>
- C M<sub>2</sub>O<sub>3</sub>
- D M<sub>3</sub>O<sub>4</sub>

$$\begin{aligned} \text{Number of moles of CO}_2 &= \frac{\text{volume of CO}_2}{\text{molar volume of gas}} \\ &= \frac{2.88 \text{ dm}^3}{24.0 \text{ dm}^3 \text{ mol}^{-1}} \\ &= 0.120 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{i.e. number of moles of O in oxide} &= 2 \times 0.120 \text{ mol} \\ &= 0.240 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Number of moles of M in oxide} &= \frac{8.93 \text{ g}}{55.8 \text{ g mol}^{-1}} \\ &= 0.160 \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{Mole ratio of M : O} &= 0.160 : 0.240 \\ &= 2 : 3 \end{aligned}$$

$\therefore$  empirical formula of the compound is M<sub>2</sub>O<sub>3</sub>.



## Unit Exercise (p.90)

8 Oxygen (O<sub>2</sub>) can be converted into ozone (O<sub>3</sub>) by passing it through an electric discharge.   $3\text{O}_2(\text{g}) \rightarrow 2\text{O}_3(\text{g})$

In an experiment, a sample of 200 cm<sup>3</sup> of oxygen was used, but only 15% of the oxygen was converted into ozone.

(All volumes were measured at the same temperature and pressure.)

What was the total volume of gas present at the end of the experiment?

A 200 cm<sup>3</sup>

B 190 cm<sup>3</sup>

C 170 cm<sup>3</sup>

D 130 cm<sup>3</sup>

**Answer: B**

**Explanation:**

Volume of O<sub>2</sub> converted into O<sub>3</sub> = 200 cm<sup>3</sup> x 15% = 30 cm<sup>3</sup>

Volume of O<sub>3</sub> =  $\frac{2}{3} \times 30 \text{ cm}^3 = 20 \text{ cm}^3$

Total volume of gas present = volume of O<sub>2</sub> remaining +  
volume of O<sub>3</sub>

= (200 – 30) cm<sup>3</sup> + 20 cm<sup>3</sup>

= 190 cm<sup>3</sup>



## Unit Exercise (p.90)

9 In an experiment, 1.96 g of potassium chlorate were heated. 480 cm<sup>3</sup> of oxygen were collected at room temperature and pressure.



(Relative atomic masses: O = 16.0, Cl = 35.5, K = 39.1; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

What was the percentage yield of oxygen?

**Answer: C**

A 60.8%

**Explanation:**

B 75.6%

$$\text{Number of moles of KClO}_3 = \frac{1.96 \text{ g}}{122.6 \text{ g mol}^{-1}} = 0.0160 \text{ mol}$$

C 83.3%

2 moles of KClO<sub>3</sub> give 3 moles of O<sub>2</sub> when heated.

D 90.0%

$$\text{i.e. number of moles of O}_2 = \frac{3}{2} \times 0.0160 \text{ mol} = 0.0240 \text{ mol}$$

Theoretical yield of O<sub>2</sub> =

number of moles of O<sub>2</sub> x molar volume of O<sub>2</sub>

$$= 0.0240 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1} = 0.576 \text{ dm}^3$$

Percentage yield of O<sub>2</sub> =  $\frac{480 \text{ cm}^3}{576 \text{ cm}^3} \times 100\% = 83.3\%$

$$= \frac{480 \text{ cm}^3}{576 \text{ cm}^3} \times 100\% = 83.3\%$$



## Unit Exercise (p.90)

10 Complete combustion of 100 cm<sup>3</sup> of carbon compound X requires 450 cm<sup>3</sup> of oxygen. (All volumes are measured at the same temperature and pressure.)



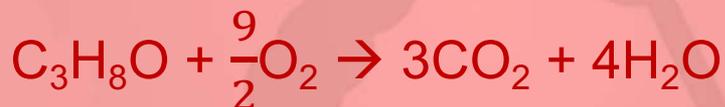
What is X?

Answer: C

- A C<sub>3</sub>H<sub>8</sub>
- B C<sub>4</sub>H<sub>8</sub>
- C C<sub>3</sub>H<sub>8</sub>O
- D C<sub>4</sub>H<sub>8</sub>O

Explanation:

C<sub>3</sub>H<sub>8</sub>O burns completely according to the equation below.



∴ 1 mole of C<sub>3</sub>H<sub>8</sub>O requires  $\frac{9}{2}$  moles of O<sub>2</sub> for complete combustion.

i.e. 100 cm<sup>3</sup> of C<sub>3</sub>H<sub>8</sub>O require 450 cm<sup>3</sup> of O<sub>2</sub> for complete combustion.



## Unit Exercise (p.90)

11 Hydrogen reacts with oxygen to form steam.



200 cm<sup>3</sup> of hydrogen react with 75 cm<sup>3</sup> of oxygen.

What is the volume of steam formed?

(All volumes are measured at the same temperature and pressure.)

**Answer: C**

A 75 cm<sup>3</sup>

B 100 cm<sup>3</sup>

C 150 cm<sup>3</sup>

D 200 cm<sup>3</sup>

**Explanation:**

150 cm<sup>3</sup> of hydrogen react with 75 cm<sup>3</sup> of oxygen.

150 cm<sup>3</sup> of steam form.



## Unit Exercise (p.90)

12 What is the total volume of gas remaining after 50 cm<sup>3</sup> of ethane are burnt completely in 200 cm<sup>3</sup> of oxygen?



(All volumes are measured at the same temperature (above 100 °C) and pressure.)



**Answer: D**

**Explanation:**

A 100 cm<sup>3</sup>

B 125 cm<sup>3</sup>

C 250 cm<sup>3</sup>

D 275 cm<sup>3</sup>

According to the equation, 2 moles of C<sub>2</sub>H<sub>6</sub> react with 7 moles of O<sub>2</sub> to form 4 moles of CO<sub>2</sub> and 6 moles of H<sub>2</sub>O.

i.e. 50 cm<sup>3</sup> of C<sub>2</sub>H<sub>6</sub> react with  $\frac{7}{2} \times 50 \text{ cm}^3$  (i.e. 175 cm<sup>3</sup>) of O<sub>2</sub> to form 100 cm<sup>3</sup> of CO<sub>2</sub> and 150 cm<sup>3</sup> of H<sub>2</sub>O.

Volume of O<sub>2</sub> left after reaction = (200 – 175) cm<sup>3</sup> = 25 cm<sup>3</sup>

Volume of gas remaining = volume of O<sub>2</sub> left + volume of CO<sub>2</sub> formed + volume of H<sub>2</sub>O formed

= (25 + 100 + 150) cm<sup>3</sup>

= 275 cm<sup>3</sup>

# Unit Exercise (p.90)

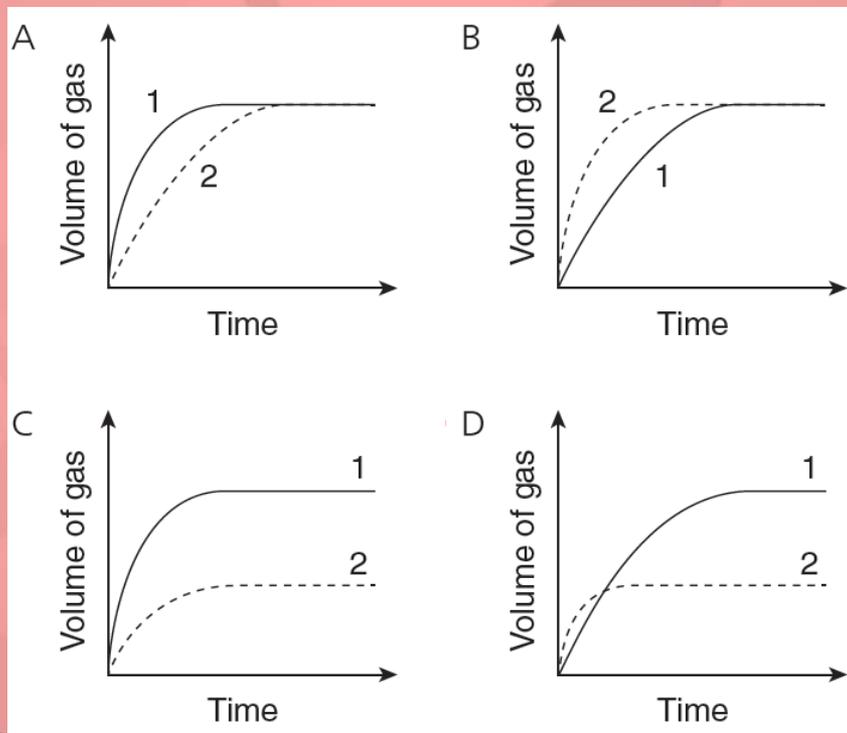
13 Two experiments were carried out to show the reaction between calcium carbonate and excess dilute hydrochloric acid. The volume of carbon dioxide formed was measured in the course of the reaction.



*Experiment 1* used 100 g of calcium carbonate in large lumps, whereas *Experiment 2* used 50 g of calcium carbonate in fine powder.

Which of the following graphs is correct?

Answer: D



 Unit Exercise (p.90)

Explanation:

*Experiment 1* used calcium carbonate in large lumps while *Experiment 2* used calcium carbonate in fine powder. Hence the initial rate of reaction in *Experiment 1* was lower than that in *Experiment 2*.

*Experiment 1* used 100 g of calcium carbonate while *Experiment 2* used 50 g of calcium carbonate. Hence the volume of gas formed in *Experiment 1* was greater than that in *Experiment 2*.

 Unit Exercise (p.90)**PART III STRUCTURED QUESTIONS**

- 14 Zinc reacts with dilute hydrochloric acid to form hydrogen. 9.81 g of zinc react completely with excess hydrochloric acid. What is the volume of hydrogen formed, measured at room temperature and pressure?  
(Relative atomic mass: Zn = 65.4; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)



$$\text{Number of moles of Zn} = \frac{9.81 \text{ g}}{65.4 \text{ g mol}^{-1}} = 0.150 \text{ mol}$$

According to the equation, 1 mole of Zn reacts with hydrochloric acid to form 1 mole of H<sub>2</sub>.

i.e. number of moles of H<sub>2</sub> = 0.150 mol (1)

Volume of H<sub>2</sub> = number of moles of H<sub>2</sub> x molar volume of gas

$$= 0.150 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$$

$$= 3.60 \text{ dm}^3 \text{ (1)}$$

∴ 3.60 dm<sup>3</sup> of hydrogen formed.

 Unit Exercise (p.90)

15 Magnesium carbonate is the ingredient of some indigestion tablets.

Magnesium carbonate reacts with stomach acid according to the equation below:



A student used an indigestion tablet containing 0.220 g of magnesium carbonate to react with excess hydrochloric acid. Calculate the volume of carbon dioxide gas produced at room temperature and pressure.

(Relative atomic masses: C = 12.0, O = 16.0, Mg = 24.3; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

$$\text{Number of moles of MgCO}_3 = \frac{0.220 \text{ g}}{84.3 \text{ g mol}^{-1}} = 2.61 \times 10^{-3} \text{ mol}$$

1 mole of  $\text{MgCO}_3$  reacts with stomach acid to produce 1 mole of  $\text{CO}_2$ .

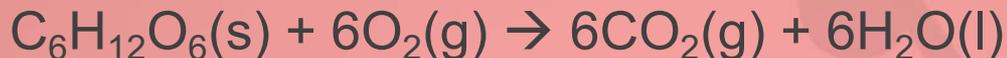
i.e. number of moles of  $\text{CO}_2 = 2.61 \times 10^{-3} \text{ mol}$  (1)

$$\begin{aligned} \text{Volume of CO}_2 &= \text{number of moles of CO}_2 \times \text{molar volume of gas} \\ &= 2.61 \times 10^{-3} \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1} \\ &= 62.6 \text{ cm}^3 \text{ (1)} \end{aligned}$$

$\therefore 62.6 \text{ cm}^3$  of carbon dioxide gas were produced.

 Unit Exercise (p.90)

16 When glucose is metabolised in the body, it essentially reacts with oxygen to produce carbon dioxide and water.



There are 29.0 g of carbohydrate in a chocolate bar, which will be used up as glucose in the body.

(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

a) What volume of oxygen, measured at room temperature and pressure, is needed to react with 29.0 g of glucose?

$$\text{Number of moles of } \text{C}_6\text{H}_{12}\text{O}_6 = \frac{29.0 \text{ g}}{180.0 \text{ g mol}^{-1}} = 0.161 \text{ mol}$$

1 mole of  $\text{C}_6\text{H}_{12}\text{O}_6$  reacts with 6 moles of  $\text{O}_2$ .

i.e. number of moles of  $\text{O}_2 = 6 \times 0.161 \text{ mol} = 0.966 \text{ mol}$  (1)

Volume of  $\text{O}_2 = \text{number of moles of } \text{O}_2 \times \text{molar volume of gas}$

$$= 0.966 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$$

$$= 23.2 \text{ dm}^3 \text{ (1)}$$

$\therefore 23.2 \text{ dm}^3$  of oxygen are needed.

 Unit Exercise (p.90)

b) What volume of air is needed, if air contains 21% of oxygen?

$$\text{Volume of air} = \frac{23.2 \text{ dm}^3}{21\%}$$

$$= 110 \text{ dm}^3 (1)$$

$\therefore$  110 dm<sup>3</sup> of air are needed.

 Unit Exercise (p.90)

17 When hydrochloric acid is added to aluminium sulphide, hydrogen sulphide is evolved.  $\text{Al}_2\text{S}_3(\text{s}) + 6\text{HCl}(\text{aq}) \rightarrow 2\text{AlCl}_3(\text{aq}) + 3\text{H}_2\text{S}(\text{g})$

A sample of  $\text{Al}_2\text{S}_3(\text{s})$  reacts with excess  $\text{HCl}(\text{aq})$  to form  $14.4 \text{ dm}^3$  of  $\text{H}_2\text{S}(\text{g})$ , measured at room temperature and pressure.

Calculate the mass of  $\text{Al}_2\text{S}_3(\text{s})$  reacted.

(Relative atomic masses: Al = 27.0, S = 32.1; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

$$\text{Number of moles of H}_2\text{S} = \frac{\text{volume of H}_2\text{S}}{\text{molar volume of gas}} = \frac{14.4 \text{ dm}^3}{24.0 \text{ dm}^3 \text{ mol}^{-1}}$$

$$= 0.600 \text{ mol (1)}$$

1 mole of  $\text{Al}_2\text{S}_3$  reacts with hydrochloric acid to form 3 moles of  $\text{H}_2\text{S}$ .

$$\text{i.e. number of moles of Al}_2\text{S}_3 = \frac{0.600}{3} \text{ mol} = 0.200 \text{ mol (1)}$$

$$\text{Mass of Al}_2\text{S}_3 = 0.200 \text{ mol} \times 150.3 \text{ g mol}^{-1}$$

$$= 30.1 \text{ g (1)}$$

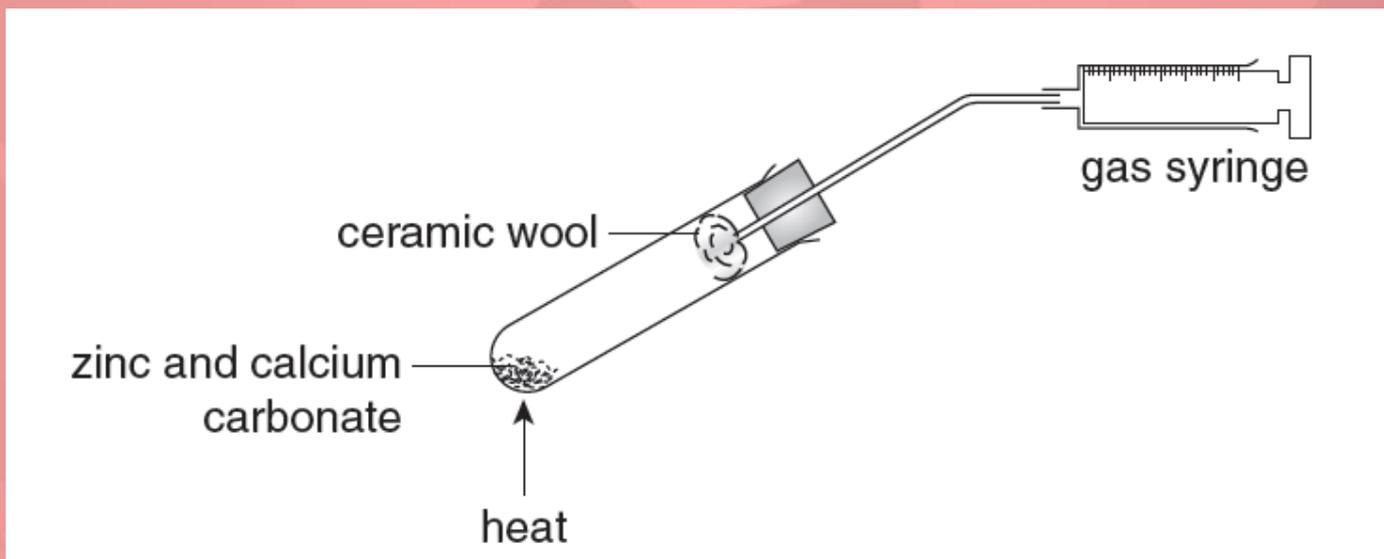
$\therefore 30.1 \text{ g}$  of  $\text{Al}_2\text{S}_3$  reacted.

 Unit Exercise (p.90)

18 Carbon monoxide can be made in the laboratory by heating a mixture of  zinc metal and calcium carbonate. An equation for this reaction is shown below.

$$\text{Zn(s)} + \text{CaCO}_3\text{(s)} \rightarrow \text{ZnO(s)} + \text{CaO(s)} + \text{CO(g)}$$

A student carried out the reaction of zinc (Zn) and calcium carbonate ( $\text{CaCO}_3$ ) in a fume cupboard. The student measured the volume of gas produced.





## Unit Exercise (p.90)

A mixture containing 0.27 g of powdered zinc and 0.38 g of powdered  $\text{CaCO}_3$  was heated strongly for two minutes.

The volume of gas collected in the  $100 \text{ cm}^3$  syringe was then measured.

The experiment was then repeated.

a) Calculate the maximum volume of carbon monoxide, measured at room temperature and pressure, that could be produced by heating this mixture of Zn and  $\text{CaCO}_3$ .

(Relative atomic masses: C = 12.0, O = 16.0, Ca = 40.1, Zn = 65.4; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

Show ALL your working.

 Unit Exercise (p.90)

$$\text{Number of moles of Zn} = \frac{0.27 \text{ g}}{65.4 \text{ g mol}^{-1}} = 4.1 \times 10^{-3} \text{ mol}$$

$$\text{Number of moles of CaCO}_3 = \frac{0.38 \text{ g}}{100.1 \text{ g mol}^{-1}} = 3.8 \times 10^{-3} \text{ mol}$$

1 mole of Zn reacts with 1 mole of CaCO<sub>3</sub>.

In this reaction,  $3.8 \times 10^{-3}$  mol of Zn reacted with  $3.8 \times 10^{-3}$  mol of CaCO<sub>3</sub>.

Thus, CaCO<sub>3</sub> is the limiting reactant. (1)

$$\text{Number of moles of CO} = 3.8 \times 10^{-3} \text{ mol (1)}$$

$$\text{Volume of CO} = \text{number of moles CO} \times \text{molar volume of gas}$$

$$= 3.8 \times 10^{-3} \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1}$$

$$= 91 \text{ cm}^3 \text{ (1)}$$



## Unit Exercise (p.90)

b) The student did NOT obtain the volume of gas predicted in (a) using this procedure.

Apart from further repeats, suggest TWO improvements to the practical procedure that would allow the student to obtain a more accurate result.

*(OCR Advanced Subsidiary Level, Chem. A, H032, Sample Question Paper, 2014, 21(c))*

Heat until the syringe plunger stops moving / no more gas is produced. (1)  
Wait until the gas has cooled (to room temperature) before measuring the volume. (1)



## Unit Exercise (p.90)

19 Traditional gunpowder is a mixture of potassium nitrate, sulphur powder and charcoal powder. When this gunpowder is ignited, the reaction occurred can be represented by the equation below:



A sample of gunpowder was prepared using the following quantities:

8.00 g of potassium nitrate

1.00 g of carbon

1.00 g of sulphur

Calculate the total volume of gas, at room temperature and pressure, produced from the explosion of this mixture.

(Relative atomic masses: C = 12.0, N = 14.0, O = 16.0, S = 32.1, K = 39.1; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )



## Unit Exercise (p.90)

$$\text{Number of moles of KNO}_3 = \frac{8.00 \text{ g}}{101.1 \text{ g mol}^{-1}} = 0.0791 \text{ mol}$$

$$\text{Number of moles of S} = \frac{1.00 \text{ g}}{32.1 \text{ g mol}^{-1}} = 0.0311 \text{ mol}$$

$$\text{Number of moles of C} = \frac{1.00 \text{ g}}{12.0 \text{ g mol}^{-1}} = 0.0833 \text{ mol}$$

2 moles of  $\text{KNO}_3$  react with 1 mole of S and 3 moles of C.

In this reaction,  $2 \times 0.0278$  mole of  $\text{KNO}_3$  reacts with 0.0278 mole of S and 0.0833 mole of C. (1)

$$\text{Total number of moles of gas} = 4 \times 0.0278 \text{ mol} = 0.111 \text{ mol (1)}$$

Total volume of gas

= number of moles of gas  $\times$  molar volume of gas

$$= 0.111 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$$

$$= 2.66 \text{ dm}^3 \text{ (1)}$$

$\therefore$  2.66  $\text{dm}^3$  of gas were produced.



## Unit Exercise (p.90)

20 Nitrogen ( $\text{N}_2$ ) and hydrogen ( $\text{H}_2$ ) react to form ammonia ( $\text{NH}_3$ ) gas.

a) Write the chemical equation for the reaction involved.



b) Calculate the volume of nitrogen, measured at room temperature and pressure, needed to react completely with 6.75 moles of hydrogen.

1 mole of  $\text{N}_2$  reacts with 3 moles of  $\text{H}_2$ .

$$\text{i.e. number of moles of } \text{N}_2 = \frac{6.75}{3} \text{ mol}$$

$$= 2.25 \text{ mol} \quad (1)$$

Volume of  $\text{N}_2$  = number of moles of  $\text{N}_2$  x molar volume of gas

$$= 2.25 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$$

$$= 54.0 \text{ dm}^3 \quad (1)$$

$\therefore$  54.0  $\text{dm}^3$  of nitrogen are needed.

 Unit Exercise (p.90)

- c) Calculate the volume of hydrogen, measured at room temperature and pressure, needed to produce 4.50 moles of ammonia.  
(Molar volume of gas at room conditions =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

3 moles of  $\text{H}_2$  react with nitrogen to produce 2 moles of  $\text{NH}_3$ .

$$\text{Number of moles of } \text{H}_2 = \frac{3}{2} \times 4.50 \text{ mol}$$

$$= 6.75 \text{ mol (1)}$$

Volume of  $\text{H}_2$  = number of moles of  $\text{H}_2$  x molar volume of gas

$$= 6.75 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1}$$

$$= 162 \text{ dm}^3 \text{ (1)}$$

$\therefore$  162  $\text{dm}^3$  of hydrogen are needed.



## Unit Exercise (p.90)



21 The volume of a gas mixture of propane ( $C_3H_8$ ) and butane ( $C_4H_{10}$ ) was  $3.00 \text{ dm}^3$ . Complete combustion of the mixture produced  $10.0 \text{ dm}^3$  of carbon dioxide. What was the composition of each gas in the gaseous mixture by volume?

(All volumes are measured at the same temperature and pressure.)

Let  $x \text{ dm}^3$  be the volume of  $C_3H_8$  in the mixture.

Then  $(3.00 - x) \text{ dm}^3$  is the volume of  $C_4H_{10}$  in the mixture.



$x \text{ dm}^3$

$3x \text{ dm}^3$



$(3.00 - x)$

$4(3.00 - x) \text{ dm}^3$

Total volume of  $CO_2 = 10.0 \text{ dm}^3$

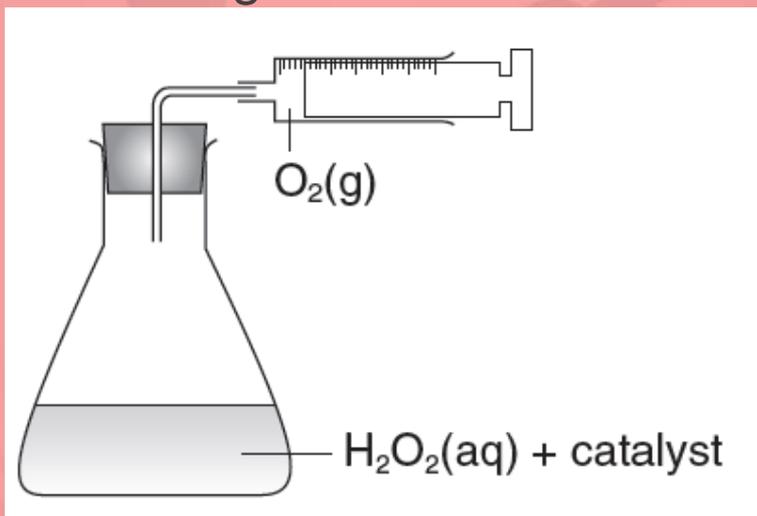
$$= [3x + 4(3.00 - x)] \text{ dm}^3 \quad (1)$$

$$x = 2.00 \quad (1)$$

$\therefore$  the gaseous mixture contained  $2.00 \text{ dm}^3$  of propane and  $1.00 \text{ dm}^3$  of butane.

 Unit Exercise (p.90)

- 22 In an experiment performed under room conditions as shown below,  $5.00 \text{ cm}^3$  of  $\text{H}_2\text{O}_2(\text{aq})$  decomposed into  $\text{O}_2(\text{g})$  and  $\text{H}_2\text{O}(\text{l})$  in the presence of a catalyst.  $\text{O}_2(\text{g})$  was continuously released from the start of the experiment until the third minute when a total of  $60 \text{ cm}^3$  of gas was collected. After that, no more gas was collected.

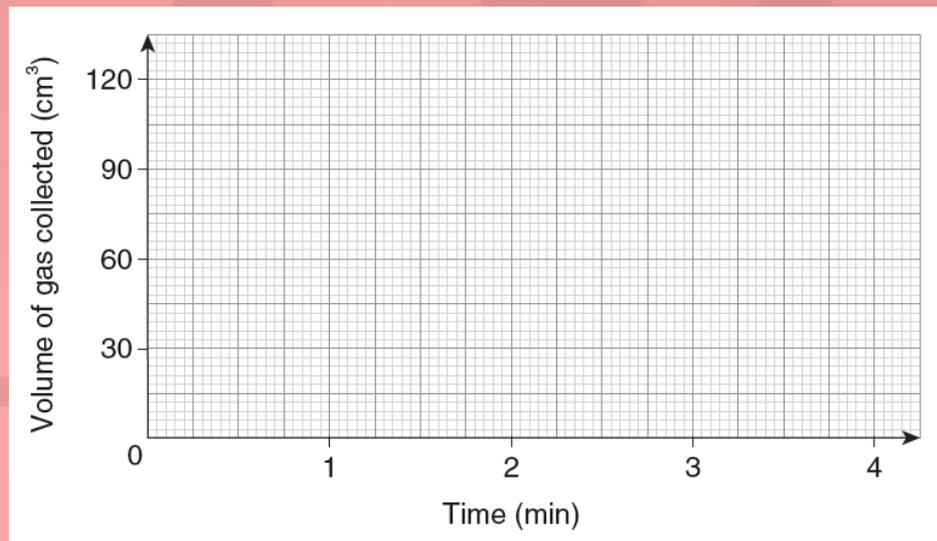


- a) Calculate the initial concentration of the  $\text{H}_2\text{O}_2(\text{aq})$ , in  $\text{mol dm}^{-3}$ .  
(Molar volume of gas at room conditions =  $24 \text{ dm}^3 \text{ mol}^{-1}$ )

 Unit Exercise (p.90)

- b) In the graph below, sketch the variation of the volume of gas collected with time in the first 4 minutes.

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).



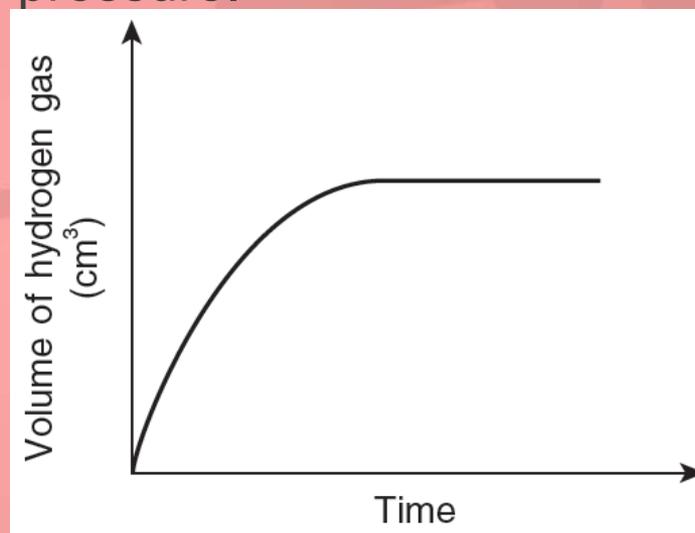
- c) The experiment is repeated using a  $\text{H}_2\text{O}_2(\text{aq})$  at a higher temperature but other conditions remain unchanged. Explain whether the total volume of gas obtained would still be  $60 \text{ cm}^3$ . (The volume of gas is measured at room conditions.)
- d) Suggest another method that can be used to follow the progress of this reaction.

(HKDSE, Paper 1B, 2017, 10)



## Unit Exercise (p.90)

23 A student carried out an experiment to determine the rate of the reaction  between 1.00 g of zinc powder and 100.0 cm<sup>3</sup> of 1.00 mol dm<sup>-3</sup> HCl(aq) at 298 K. The graph below shows the volume of hydrogen gas released, measured at room temperature and pressure.



a) Calculate the total volume of hydrogen gas released, measured at room temperature and pressure.

(Relative atomic mass: Zn = 65.4; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

 Unit Exercise (p.90)

$$\text{Number of moles of Zn} = \frac{1.00 \text{ g}}{65.4 \text{ g mol}^{-1}} = 0.0153 \text{ mol}$$

$$\text{Number of moles of HCl} = 1.00 \text{ mol dm}^{-3} \times \frac{100.0}{1\,000} \text{ dm}^3 = 0.100 \text{ mol}$$

According to the equation, 1 mole of Zn reacts with 2 moles of HCl.

In this reaction, 0.0153 mole of Zn reacted with 0.0306 mole of HCl.

Thus, Zn was the limiting reactant. (1)

$$\text{Number of moles of H}_2 = 0.0153 \text{ mol (1)}$$

$$\text{Volume of H}_2 = \text{number of moles of H}_2 \times \text{molar volume of gas}$$

$$= 0.0153 \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1}$$

$$= 367 \text{ cm}^3 \text{ (1)}$$

∴ 367 cm<sup>3</sup> of hydrogen gas were released.

 Unit Exercise (p.90)

b) The experiment was repeated using  $100.0 \text{ cm}^3$  of  $1.00 \text{ mol dm}^{-3}$   $\text{H}_2\text{SO}_4(\text{aq})$  while the other conditions remained constant.

i) Explain the effect of this change on the initial rate of reaction.

ii) Explain whether the total volume of gas released would be the same as the first experiment.

(The volume was measured at room temperature and pressure.)

b) i) HCl(aq) is a monobasic acid while  $\text{H}_2\text{SO}_4(\text{aq})$  is a dibasic acid. (1)

The concentration of  $\text{H}^+(\text{aq})$  ions in  $1.00 \text{ mol dm}^{-3}$   $\text{H}_2\text{SO}_4(\text{aq})$  is higher than that in  $1.00 \text{ mol dm}^{-3}$  HCl(aq). (1)

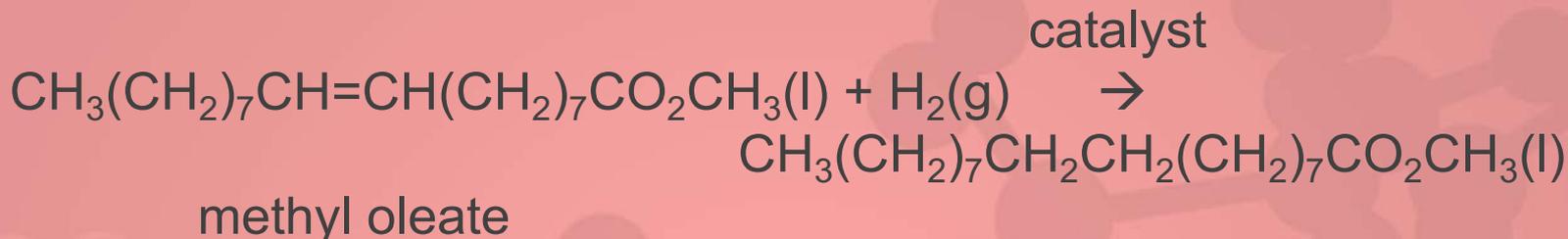
The initial rate of reaction increases.

ii) The same

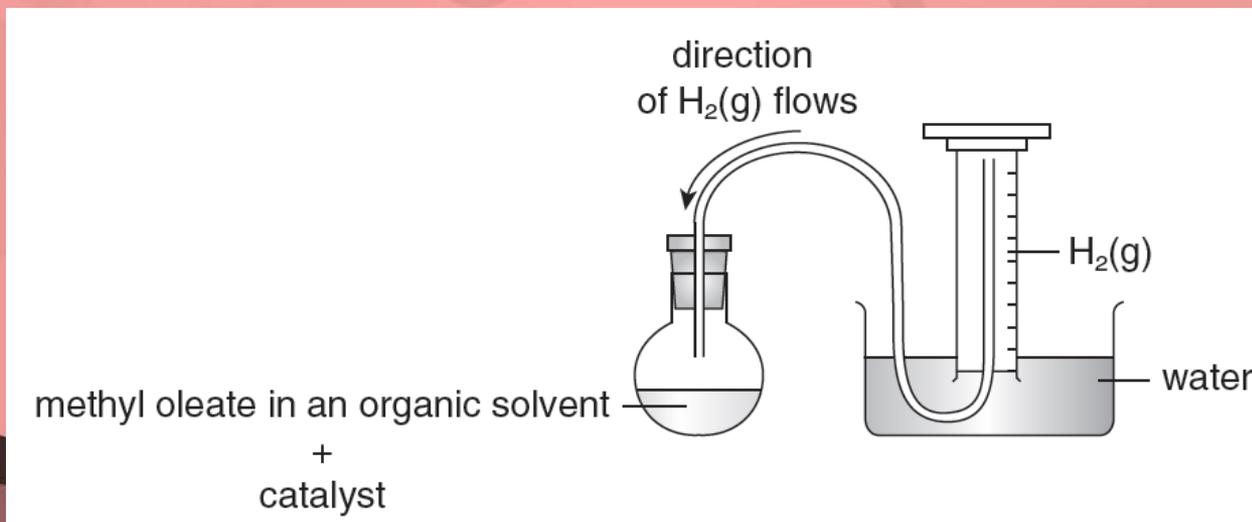
Zinc was the limiting reactant. (1)

## Unit Exercise (p.90)

24 Consider the reaction below:



At room temperature and pressure, a micro-scale experiment was performed using the set-up shown below in which 0.080 g of methyl oleate in an organic solvent was allowed to react with excess  $\text{H}_2(\text{g})$ . The  $\text{H}_2(\text{g})$  flowed from the inverted measuring cylinder to the reacting flask through the tubing.

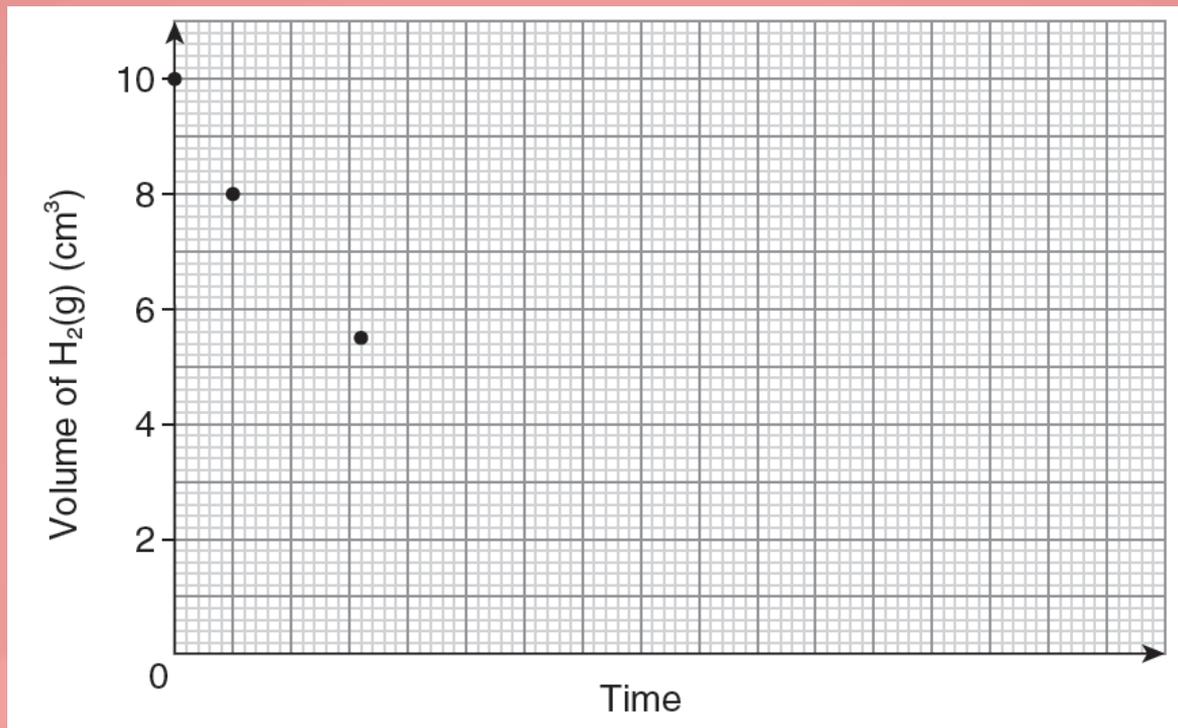


 Unit Exercise (p.90)

- a) State ONE advantage of conducting this reaction in a micro-scale experiment.
- b) Explain why the right end of the tubing was placed at the uppermost position of the inverted measuring cylinder.
- c) State an expected observation in the inverted measuring cylinder during the reaction.
- d) Calculate the theoretical volume of  $\text{H}_2(\text{g})$  needed for the reaction to complete at room temperature and pressure.  
(Molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ ; relative molecular mass: methyl oleate = 296.0)
- e) i) Sketch, in the graph below, the variation of the volume of  $\text{H}_2(\text{g})$  in the measuring cylinder with time from start until the completion of the reaction. You should label this sketch as 'A'. (The measuring cylinder initially contained  $10.0 \text{ cm}^3$  of  $\text{H}_2(\text{g})$ . The first few points have been given in the graph to facilitate the sketch.)



## Unit Exercise (p.90)



- ii) In the same graph above, give another sketch as required in (i) but only using 0.040 g of methyl oleate for the reaction while the other conditions remain unchanged. You should label this sketch as 'B'.

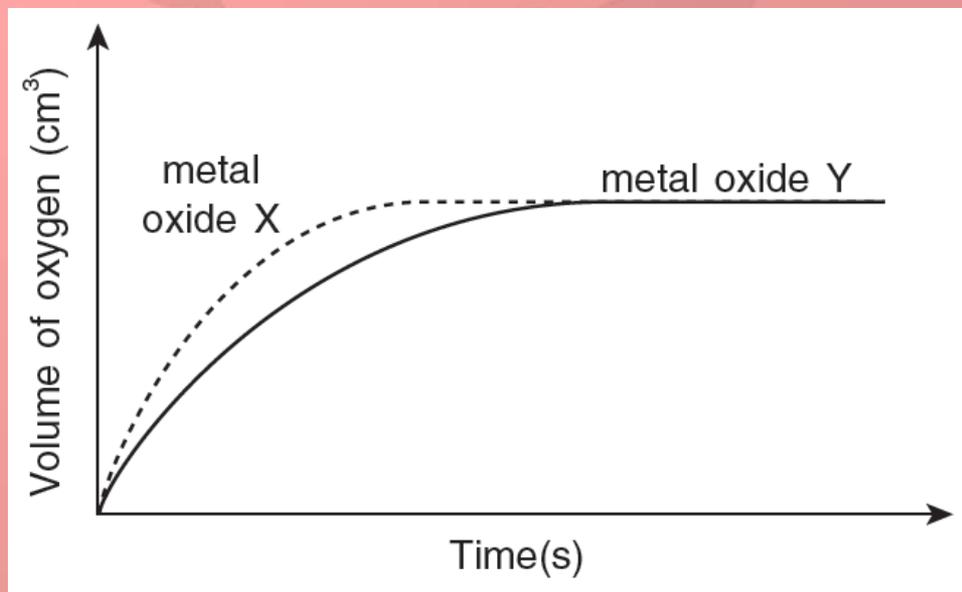
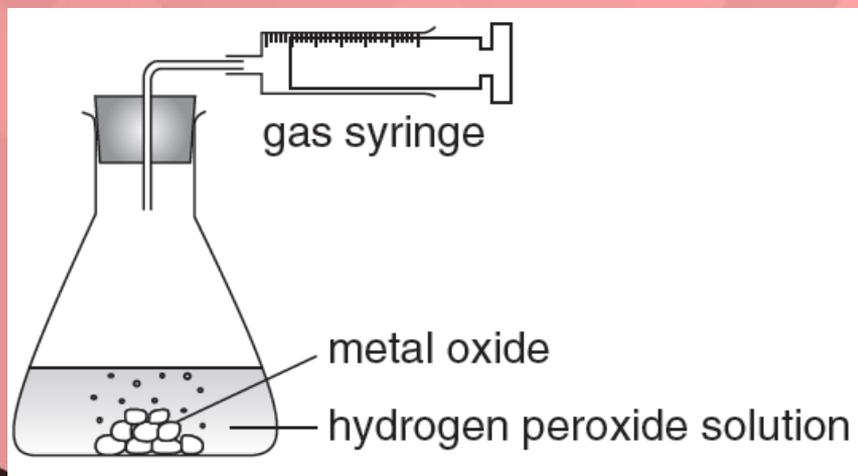
*(HKDSE, Paper 1B, 2015, 9)*

Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

## Unit Exercise (p.90)

25 Two trials of an experiment were carried out to study the effect of using two metal oxides, X and Y, to catalyse the decomposition of  $50.0 \text{ cm}^3$  of  $0.130 \text{ mol dm}^{-3}$  hydrogen peroxide solution at  $25^\circ \text{C}$ .

The volume of oxygen released was measured using the experimental set-up shown below. The curve below shows the volume, measured at room temperature and pressure, of the oxygen released in the experiment.



 Unit Exercise (p.90)

- a) Write the chemical equation for the decomposition of hydrogen peroxide solution.  $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$
- b) Calculate the total volume of oxygen produced, measured at room temperature and pressure, in each case.  
(Molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

$$\text{Number of moles of } \text{H}_2\text{O}_2 = 0.130 \text{ mol dm}^{-3} \times \frac{50.0}{1000} \text{ dm}^3 = 6.50 \times 10^{-3} \text{ mol}$$

According to the equation, 2 moles of  $\text{H}_2\text{O}_2$  decompose to produce 1 mole of  $\text{O}_2$ .

$$\text{i.e. number of moles of } \text{O}_2 = \frac{6.50 \times 10^{-3}}{2} \text{ mol} = 3.25 \times 10^{-3} \text{ mol (1)}$$

$$\begin{aligned} \text{Volume of } \text{O}_2 &= \text{number of moles of } \text{O}_2 \times \text{molar volume of gas} \\ &= 3.25 \times 10^{-3} \text{ mol} \times 24\,000 \text{ cm}^3 \text{ mol}^{-1} \\ &= 78.0 \text{ cm}^3 \text{ (1)} \end{aligned}$$

$\therefore$  78.0 cm<sup>3</sup> of oxygen were produced.

 Unit Exercise (p.90)

c) State ONE property of the two catalysts which the student should keep the same for a fair comparison.

Any one of the following:

- Mass (1)
- State of subdivision (1)
- Particle size (1)

d) State, give a reason, which of the above oxide is the more efficient catalyst.

(Assume that a fair comparison is carried out.)

Metal oxide X is more efficient because the decomposition is faster. (1)



## Topic Exercise (p.99)

**Note: Questions are rated according to ascending level of difficulty (from 1 to 5):**



**question targeted at level 3 and above;**



**question targeted at level 4 and above;**



**question targeted at level 5.**

**' \* ' indicates 1 mark is given for effective communication.**

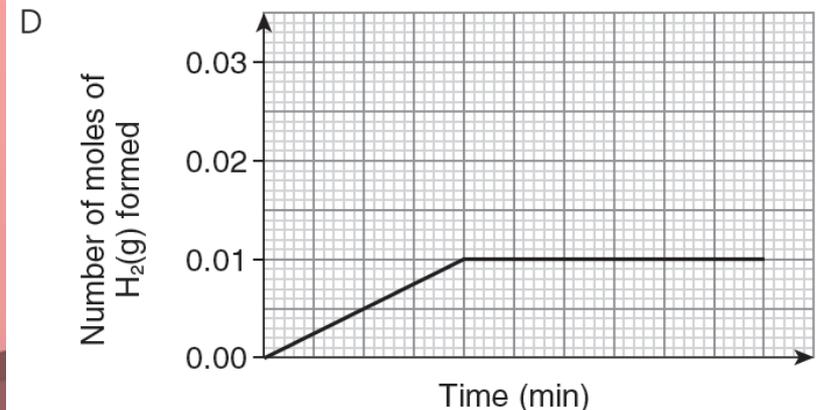
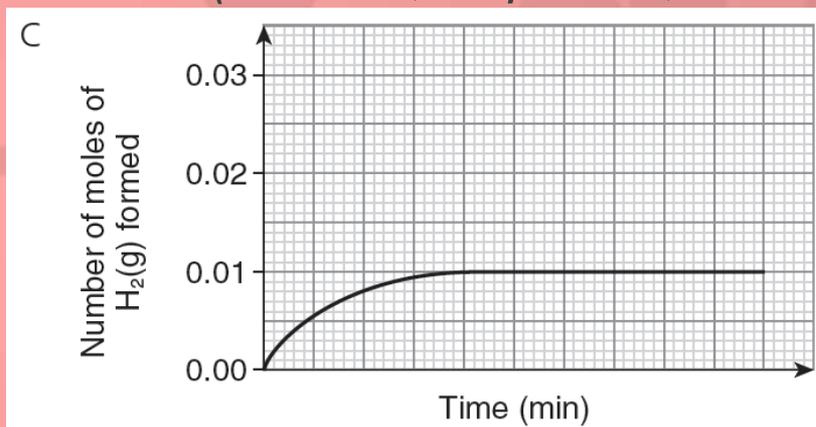
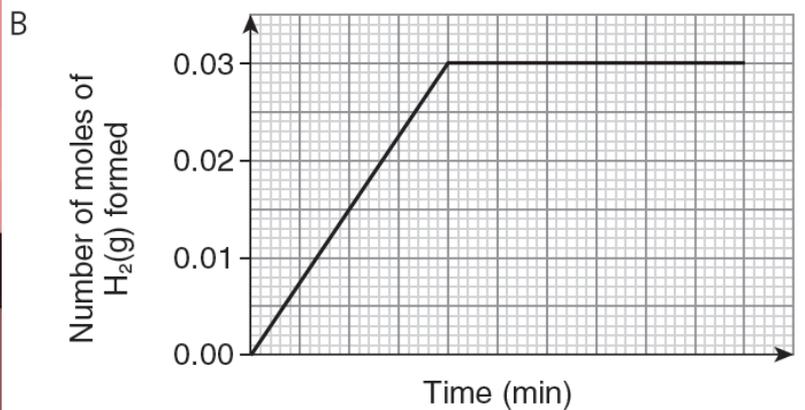
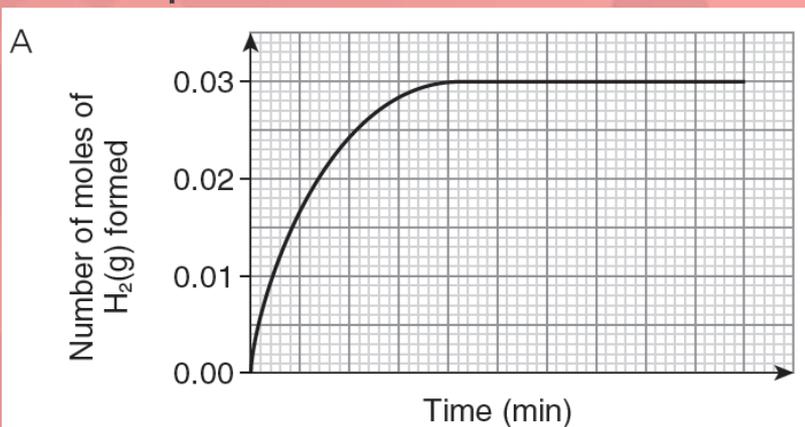


## Topic Exercise (p.99)

Answer: C

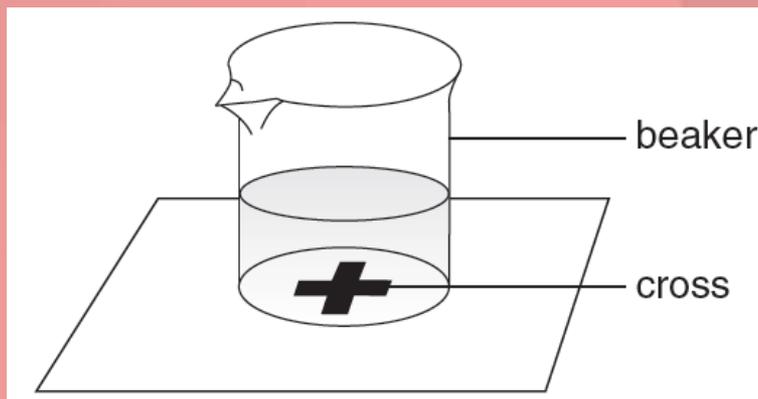
## PART I MULTIPLE CHOICE QUESTIONS

- 1 In an experiment, 0.03 mol of Mg(s) is allowed to react with 20.0 cm<sup>3</sup> of 1.0 M HCl(aq). Which of the following graphs best represents the results of the experiment? (HKDSE, Paper 1A, 2016, 25)



 Topic Exercise (p.99)

Direction: Questions 2 and 3 refer to the following set-up.



2 A(aq) and B(aq) react to form a turbid mixture. Three trials of an  experiment were performed to study the rate of the reaction. In each trial, A(aq) was mixed with  $\text{H}_2\text{O}(\text{l})$  in the beaker. After that, B(aq) was added to the mixture, and immediately started to measure the time needed for the cross to become invisible when viewed from above. The table below shows the relevant data.

Answer: C



## Topic Exercise (p.99)

Trial	Volume used (cm <sup>3</sup> )			Time (s)
	A(aq)	H <sub>2</sub> O(l)	B(aq)	
1	10.0	20.0	10.0	82
2	10.0	10.0	20.0	41
3	20.0	10.0	10.0	82

Which of the following statements concerning the rate of the reaction is correct?

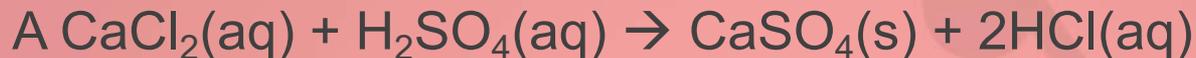
- A It depends on [A(aq)], and also depends on [B(aq)].
- B It increases with [A(aq)], but does not increase with [B(aq)].
- C It increases with [B(aq)], but does not increase with [A(aq)].
- D It does not depend on [A(aq)], and also does not depend on [B(aq)].

(HKDSE, Paper 1A, 2017, 27)

**Answer: C**

 Topic Exercise (p.99)

3 Of which of the following reactions can the rate be studied by the above set-up?



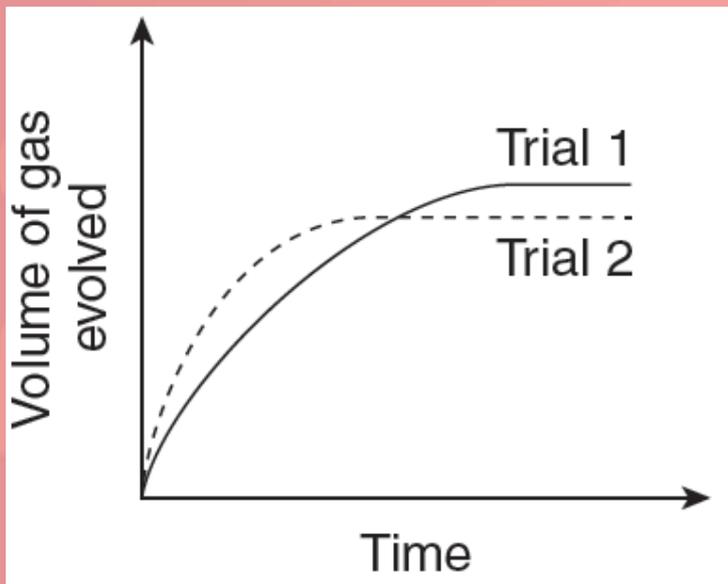
(HKDSE, Paper 1A, 2017, 28)

Answer: D



## Topic Exercise (p.99)

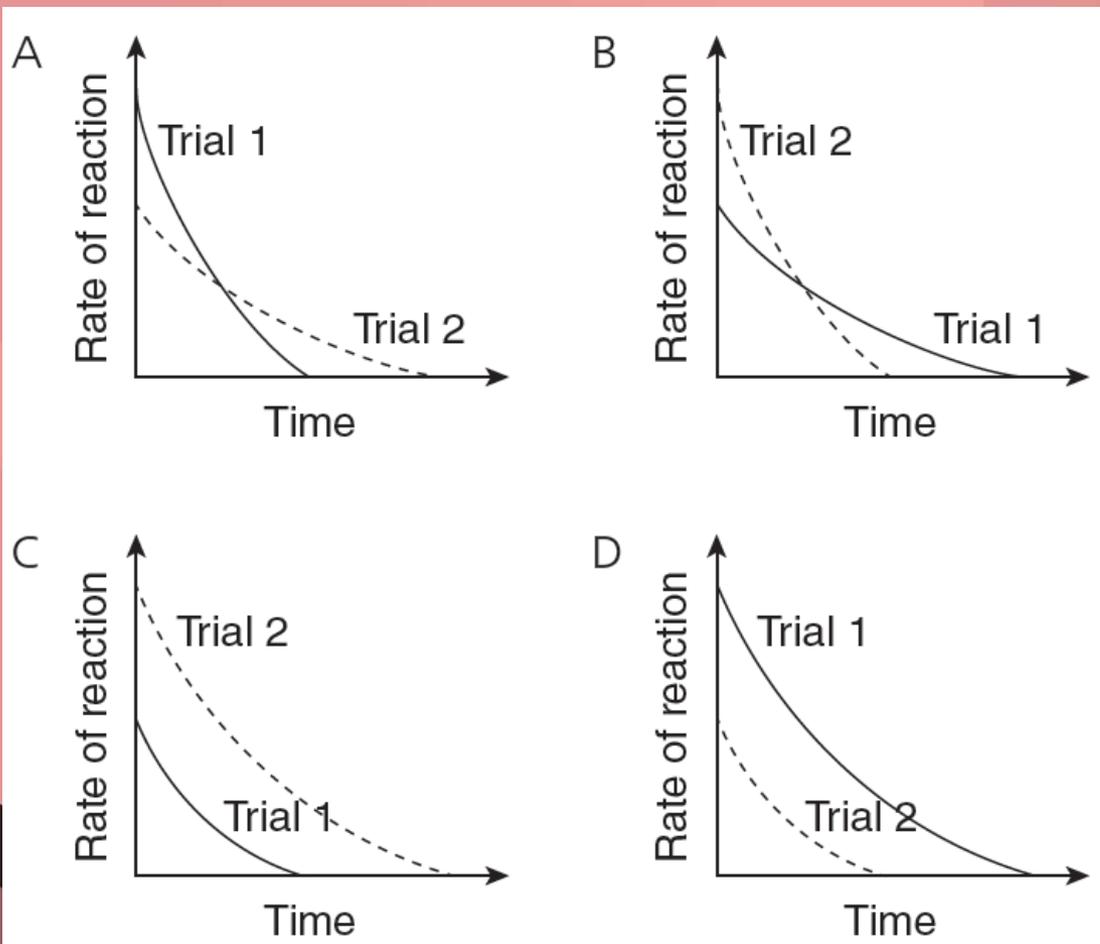
4 In two trials of an experiment, a substance is decomposed and the gas evolved is measured. The graph below shows the results obtained.



Answer: B

 Topic Exercise (p.99)

Which of the following graphs shows how the rate of reaction varies with time in each trial?



**Answer: B**

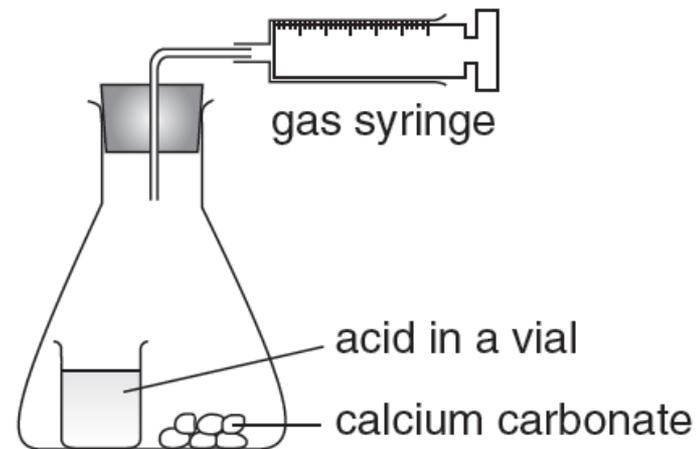
**Explanation:**

The initial rate of reaction in *Trial 2* is higher than that in *Trial 1*.

The reaction in *Trial 2* completes earlier than the reaction in *Trial 1*.

 Topic Exercise (p.99)

 5 The set-up shown below is used to study the rate of the reaction between calcium carbonate and different acids.



The conical flask is shaken to overturn the vial in order to start the reaction. The initial rate of the reaction with respect to the gas liberated is determined. Which of the following pairs of chemicals, upon mixing at the same temperature, has the HIGHEST initial rate of the reaction?

- A 1 g of  $\text{CaCO}_3$  lumps and  $10 \text{ cm}^3$  of  $1 \text{ mol dm}^{-3}$   $\text{HCl}(\text{aq})$
- B 1 g of  $\text{CaCO}_3$  powder and  $10 \text{ cm}^3$  of  $1 \text{ mol dm}^{-3}$   $\text{HCl}(\text{aq})$
- C 1 g of  $\text{CaCO}_3$  lumps and  $10 \text{ cm}^3$  of  $1 \text{ mol dm}^{-3}$   $\text{CH}_3\text{COOH}(\text{aq})$
- D 1 g of  $\text{CaCO}_3$  powder and  $10 \text{ cm}^3$  of  $1 \text{ mol dm}^{-3}$   $\text{CH}_3\text{COOH}(\text{aq})$

**Answer: B**



## Topic Exercise (p.99)

6  $\text{H}_2\text{O}_2(\text{aq})$  decomposes into  $\text{H}_2\text{O}(\text{l})$  and  $\text{O}_2(\text{g})$  in the presence of  $\text{MnO}_2(\text{s})$ . Two experiments are performed to study this decomposition under the same conditions, except that  $50 \text{ cm}^3$  of  $2 \text{ M H}_2\text{O}_2(\text{aq})$  is used in *Experiment (1)*, while  $100 \text{ cm}^3$  of  $1 \text{ M H}_2\text{O}_2(\text{aq})$  is used in *Experiment (2)*. Which of the following combinations is correct?



	<u>Rate of formation of <math>\text{O}_2(\text{g})</math> at the start</u>	<u>Total volume of <math>\text{O}_2(\text{g})</math> formed</u>
A	<i>Experiment (1) &gt; Experiment (2)</i>	<i>Experiment (1) = Experiment (2)</i>
B	<i>Experiment (1) &gt; Experiment (2)</i>	<i>Experiment (1) &gt; Experiment (2)</i>
C	<i>Experiment (1) = Experiment (2)</i>	<i>Experiment (1) = Experiment (2)</i>
D	<i>Experiment (1) = Experiment (2)</i>	<i>Experiment (1) &gt; Experiment (2)</i>

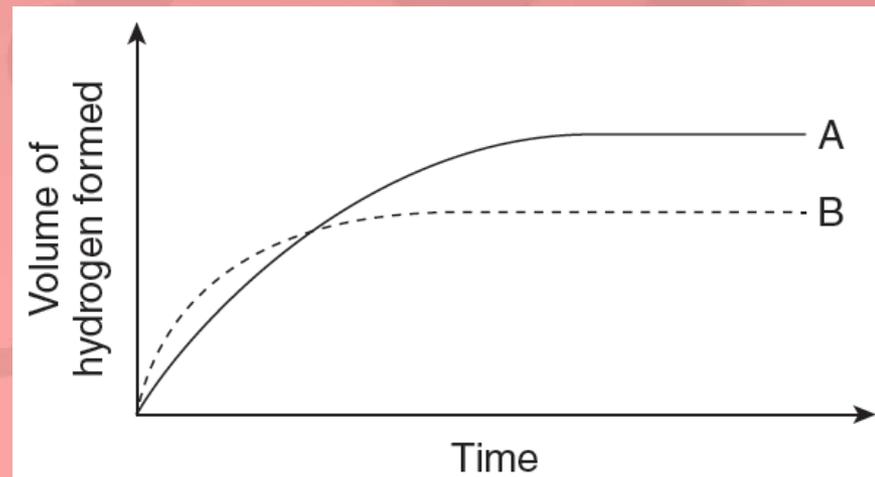
**Answer: D**

(HKDSE, Paper 1A, 2014, 25)



## Topic Exercise (p.99)

- 7 100 cm<sup>3</sup> of 1.0 M HCl(aq) reacts with excess zinc granules giving curve A in the graph below.



Answer: D

Which of the following changes may give curve B?

- A Increase the temperature by 5 °C.
- B Use the same mass of zinc powder instead of zinc granules.
- C Use 200 cm<sup>3</sup> of 0.80 M HCl(aq) instead of 100 cm<sup>3</sup> of 1.0 M HCl(aq).
- D Use 50 cm<sup>3</sup> of 1.50 M HCl(aq) instead of 100 cm<sup>3</sup> of 1.0 M HCl(aq).

(HKDSE, Paper 1A, 2018, 25)

## Topic Exercise (p.99)

- 8 The emergency oxygen system in a passenger aircraft uses the decomposition of sodium chlorate to produce oxygen.



At the temperature and pressure inside the aircraft, each adult passenger needs  $1.60 \text{ dm}^3$  of oxygen per minute.

(Relative atomic masses: O = 16.0, Na = 23.0, Cl = 35.5; molar volume of gas at temperature and pressure inside the aircraft =  $32.2 \text{ dm}^3 \text{ mol}^{-1}$ )

What mass of sodium chlorate is required to provide the required volume of oxygen for each adult passenger per minute?

**Answer: A**



A 3.53 g Number of moles of  $\text{O}_2$  each adult passenger needed per minute

$$\text{B } 4.41 \text{ g} = \frac{\text{volume of } \text{O}_2}{\text{molar volume of gas}} = \frac{1.60 \text{ dm}^3}{32.2 \text{ dm}^3 \text{ mol}^{-1}} = 0.0497 \text{ mol}$$

C 5.29 g According to the equation, 2 moles of  $\text{NaClO}_3$  produces 3 moles of  $\text{O}_2$ .

D 17.0 g i.e. number of moles of  $\text{NaClO}_3$  required per minute

$$= \frac{2}{3} \times 0.0497 \text{ mol} = 0.0331 \text{ mol}$$

$$\text{Mass of } \text{NaClO}_3 \text{ required per minute} = 0.0331 \text{ mol} \times 106.5 \text{ g mol}^{-1} = 3.53 \text{ g}$$



## Topic Exercise (p.99)



- 9 X is a compound of aluminium and carbon. 0.180 g of X reacts with excess water to produce a gas, and an aluminium compound without carbon. The gas burns completely in oxygen to form water and 90.0 cm<sup>3</sup> of carbon dioxide, measured at room temperature and pressure. (Relative atomic masses: C = 12.0, Al = 27.0; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

What could be the chemical formula of X?

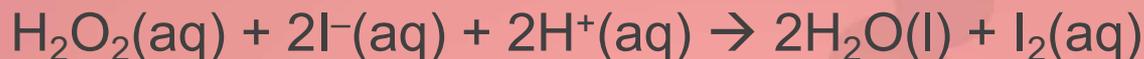
- A Al<sub>2</sub>C<sub>3</sub>
- B Al<sub>3</sub>C<sub>4</sub>
- C Al<sub>4</sub>C<sub>3</sub>
- D Al<sub>5</sub>C<sub>3</sub>

Answer: C



## Topic Exercise (p.99)

10 Consider the reaction between hydrogen peroxide and iodide ions:



Which of the following can be used to follow the progress of the reaction?

(1) Colorimetry

**Answer: A**

(2) Collecting and measuring the volume of water produced

(3) Quenching, followed by titration with a standard acid

A (1) only

B (2) only

C (1) and (3) only

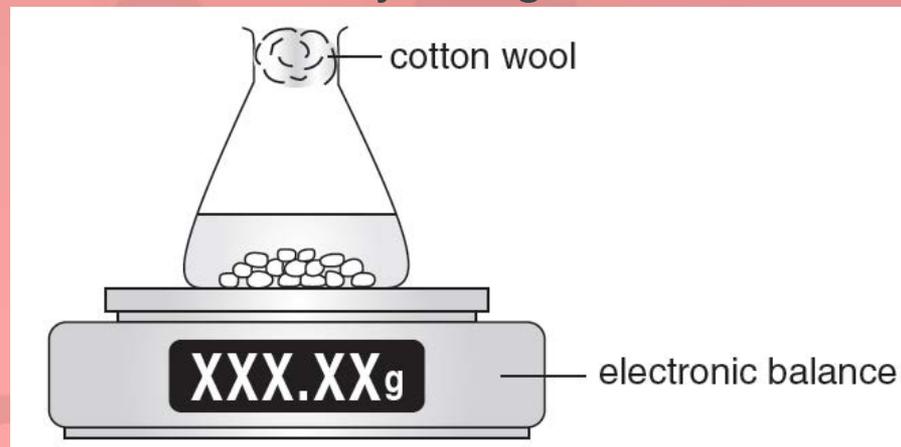
D (2) and (3) only

**Explanation:**

(1) The intensity of the brown colour of iodine in the reaction mixture increases as the reaction proceeds. Thus, the progress of the reaction can be followed by colorimetry.

 Topic Exercise (p.99)

- 11 The progress of some reactions can be followed by using the experimental set-up shown below.



For which of the following reactions is / are this set-up suitable?

- (1)  $\text{FeO(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2\text{(aq)} + \text{H}_2\text{O(l)}$   
(2)  $\text{FeCO}_3\text{(s)} + 2\text{HCl(aq)} \rightarrow \text{FeCl}_2\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$   
(3)  $\text{FeCl}_2\text{(aq)} + 2\text{NaOH(aq)} \rightarrow \text{Fe(OH)}_2\text{(s)} + 2\text{NaCl(aq)}$

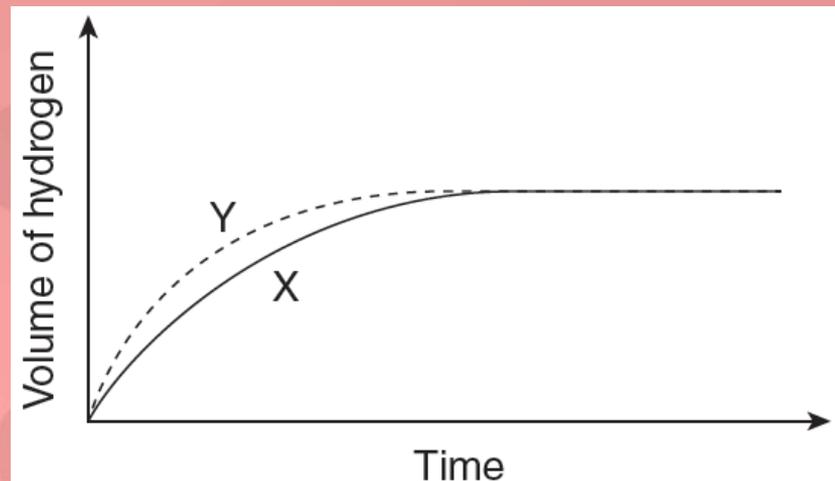
- A (1) only  
B (2) only  
C (1) and (3) only  
D (2) and (3) only

**Answer: B**



## Topic Exercise (p.99)

- 12 In an experiment using a dilute acid and a metal to produce hydrogen, the volume of hydrogen produced is measured at regular time intervals (curve X on the graph). The experiment is repeated but with some conditions changed (curve Y on the graph).



Which changes in conditions could result in curve Y?

- (1) Increasing the concentration of the acid
- (2) Increasing the particle size of the metal
- (3) Increasing the temperature

- A (1) and (2) only  
B (1) and (3) only  
C (2) and (3) only  
D (1), (2) and (3)

**Answer: B**

**Explanation:**

(2) Increasing the particle size of the metal could decrease the rate of the reaction.



## Topic Exercise (p.99)

**Directions :** Each question (Questions 13–15) consists of two separate statements. Decide whether each of the two statements is true or false; if both are true, then decide whether or not the second statement is a correct explanation of the first statement. Then select one option from A to D according to the following table :

- A Both statements are true and the 2nd statement is a correct explanation of the 1st statement.
- B Both statements are true but the 2nd statement is NOT a correct explanation of the 1st statement.
- C The 1st statement is false but the 2nd statement is true.
- D Both statements are false.



## Topic Exercise (p.99)

### 1<sup>st</sup> statement

13 When equal mass of Mg and Ca is added separately to excess dilute HCl(aq), Ca will produce a greater volume of gas than Mg.

**Explanation:** *Refer to next slide.*

14 1 mol dm<sup>-3</sup> H<sub>2</sub>SO<sub>4</sub>(aq) reacts faster with 1 g of zinc granules than 1 g of zinc powder.

**Explanation:** The surface area of 1 g of zinc granules is smaller than that of 1 g of zinc powder.

15 At room conditions, the volume of 1 mol of SO<sub>2</sub>(g) is larger than that of 1 mol of N<sub>2</sub>(g).

### 2<sup>nd</sup> statement

Ca is more reactive than Mg.

**Answer: C**

The surface area of 1 g of zinc granules is larger than that of 1 g of zinc powder.

**Answer: D**

The number of atoms constituting 1 mol of SO<sub>2</sub>(g) is greater than that constituting 1 mol of N<sub>2</sub>(g).

**Answer: C**

(HKDSE, Paper 1A, 2015, 36)



## Topic Exercise (p.99)

### 13 Explanation:

When equal mass of Mg and Ca is added separately to excess dilute HCl(aq), Mg will produce a greater volume of gas than Ca.

Suppose 100 g of Mg and 100 g of Ca are added separately to excess dilute HCl(aq).



$$\text{Number of moles of Mg in 100 g} = \frac{100 \text{ g}}{24.3 \text{ g mol}^{-1}} = 4.12 \text{ mol}$$

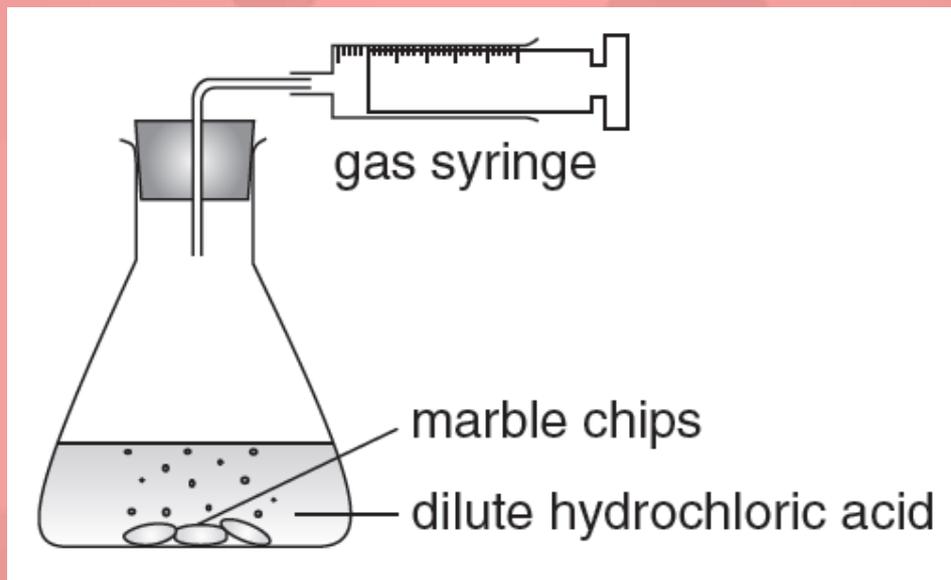
$$\text{Number of moles of Ca in 100 g} = \frac{100 \text{ g}}{40.1 \text{ g mol}^{-1}} = 2.49 \text{ mol}$$

100 g of Mg will produce 4.12 moles of hydrogen while 100 g of Ca will produce 2.49 moles of hydrogen.

Hence the difference in the amount of hydrogen produced by Mg and Ca is due to the difference between the molar masses of the two metals.

 Topic Exercise (p.99)**PART II STRUCTURED QUESTIONS**

16 Marble chips react with dilute hydrochloric acid. The progress of the reaction is followed by measuring the volume of gas given off at regular time intervals. The experimental set-up used is shown below.





## Topic Exercise (p.99)

a) The effect of change in the surface area of marble chips on the rate of this reaction is investigated.

i) Name TWO variables that should be kept constant in the investigation.

Any two of the following:

- Concentration / volume of acid (1)
- Mass of marble chips (1)
- Temperature (1)

ii) Explain how using smaller marble chips affects the rate of this reaction, when all the other conditions remain the same.

The smaller marble chips have a greater surface area, giving a greater area over which collisions can occur. (1)

So there are more effective collisions in a unit volume per unit time.

Thus, the rate of reaction increases.

b) Explain how increasing the concentration of the hydrochloric acid affects the rate of this reaction, when all the other conditions remain the same.



## Topic Exercise (p.99)

- b) Explain how increasing the concentration of the hydrochloric acid affects the rate of this reaction, when all the other conditions remain the same.

Increasing the concentration of the acid means increasing the number of hydrogen ions per unit volume.

The hydrogen ions are more crowded and collide more often with particles on the surfaces of marble chips. (1)

The chance of collision increases, so there will be more effective collisions in a unit volume per unit time.

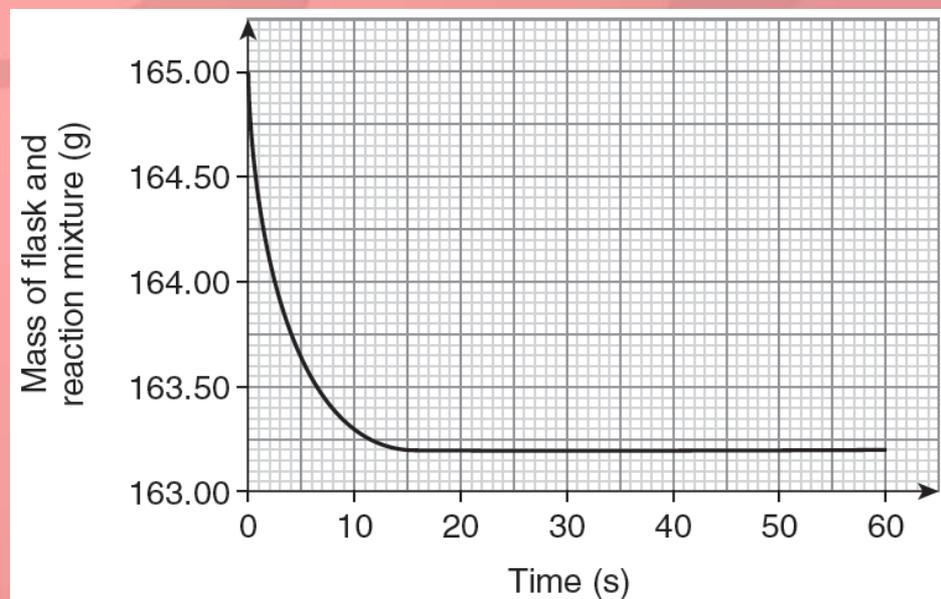
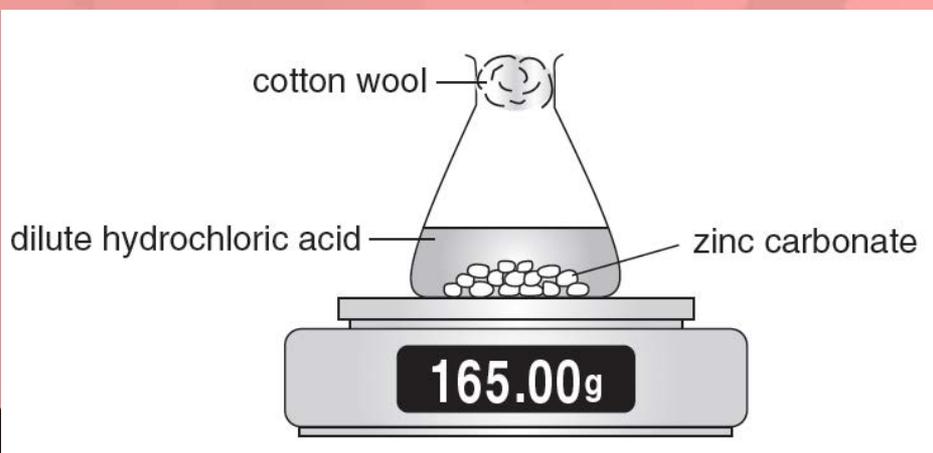
Thus, the rate of reaction increases.

 Topic Exercise (p.99)

17 Zinc carbonate reacts with dilute hydrochloric acid according to the following equation:  $\text{ZnCO}_3(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$



A student used the set-up shown below to follow the progress of the reaction. Zinc carbonate disappeared at the end of the reaction. The graph below shows the results obtained using a sample of zinc carbonate and  $50 \text{ cm}^3$  of  $1 \text{ mol dm}^{-3}$  hydrochloric acid.





## Topic Exercise (p.99)

a) Why did the mass decrease during the reaction?

Carbon dioxide gas escaped. (1)

b) Calculate the average rate of consumption of zinc carbonate in the first 10 seconds.

Mass of  $\text{CO}_2$  formed in the first 10 seconds =  $(165.00 - 163.30) \text{ g} = 1.70 \text{ g}$  (1)

Number of moles of  $\text{CO}_2$  formed in the first 10 seconds =  $\frac{1.70 \text{ g}}{44.0 \text{ g mol}^{-1}}$

= 0.0386 mol

Average rate of formation of  $\text{CO}_2$  in the first 10 seconds =  $\frac{0.0386 \text{ mol}}{10 \text{ s}}$

= 0.00386 mol  $\text{s}^{-1}$  (1)

= average rate of consumption of  $\text{ZnCO}_3$  in the first 10 seconds



## Topic Exercise (p.99)

c) Suggest how the instantaneous rate of consumption of zinc carbonate at the 10th second can be determined from the graph.

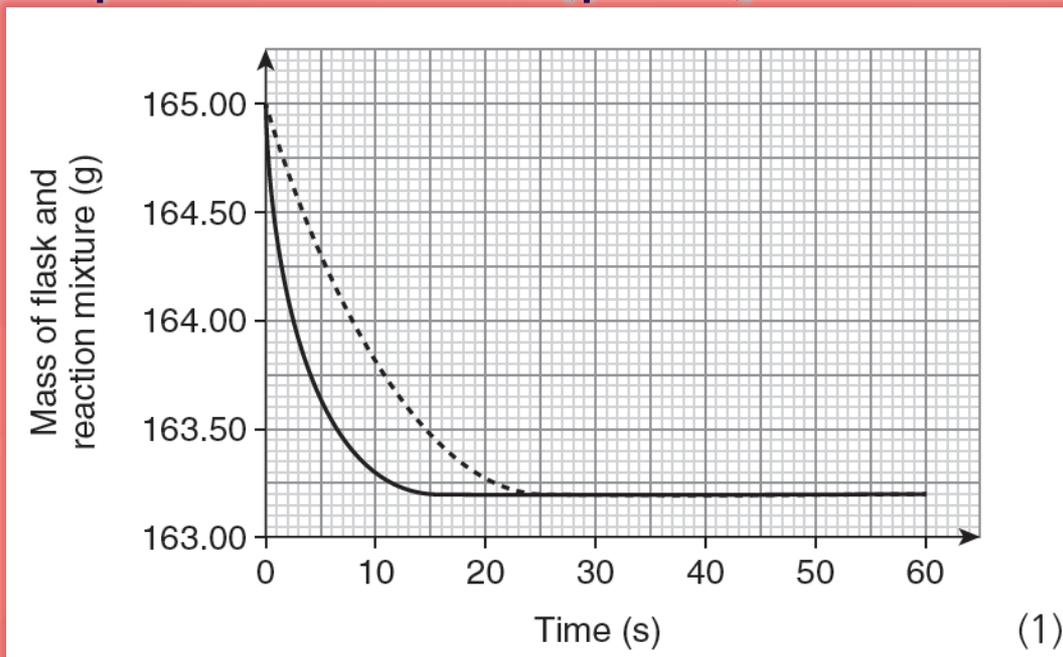
**Determine the slope of tangent to the curve at  $t = 10$  s. (1)**

d) The experiment was repeated using the same volume of  $1 \text{ mol dm}^{-3}$  ethanoic acid. Zinc carbonate also disappeared at the end of the reaction. Sketch a curve on the same graph to show the variation of the mass of the flask and the reaction mixture with time.

Explain your answer.



## Topic Exercise (p.99)



Zinc carbonate is the limiting reactant in both experiments. Thus, the same amount of carbon dioxide gas is formed in both experiments. The loss in mass of the reaction mixture is the same in both experiments. (1)

Hydrochloric acid is a strong acid while ethanoic acid is a weak acid.  
1.00 mol dm<sup>-3</sup> ethanoic acid has a lower concentration of hydrogen ions than 1.00 mol dm<sup>-3</sup> hydrochloric acid. (1)



## Topic Exercise (p.99)

18 Consider the reaction represented by the equation below:



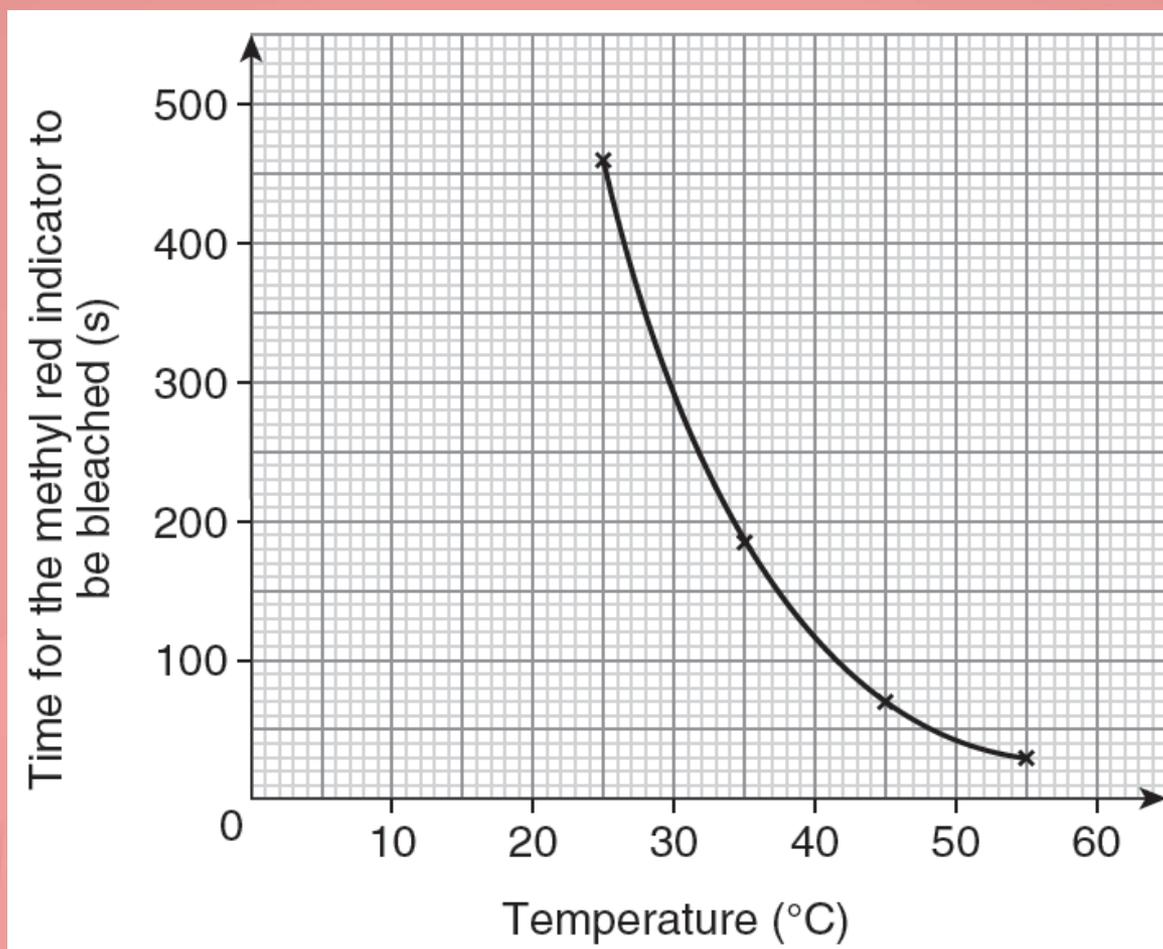
The progress of the reaction may be followed by adding a fixed amount of phenol together with some methyl red indicator. The bromine produced during the reaction reacts very rapidly with phenol.

Once all the phenol is consumed, any further bromine bleaches the indicator immediately. Therefore, the time for the reaction to proceed to a given point may be determined.

A student studied the effect of temperature on the rate of this reaction. The results obtained are shown in the graph below.



## Topic Exercise (p.99)



 Topic Exercise (p.99)

a) What conclusion can you draw from the graph? Explain how temperature affects the rate of reaction in terms of the behaviour of particles.

As the temperature increases, the time for the methyl red indicator to be bleached decreases. /

As the temperature increases, the rate of reaction increases. (1)

At a higher temperature, the reactant particles have more kinetic energy and move faster:

- the reactant particles involve in more collisions, resulting in more effective collisions; (1)
- a larger fraction of the colliding particles will have kinetic energy equal to or greater than the activation energy and hence a higher percentage of collisions can result in a reaction. (1)

b) A second student carried out the same experiment using a lower concentration of  $\text{BrO}_3^-$ (aq). Draw the curve you would expect him to obtain on the same grid.

The curve is above the original curve and less steep. (1)



## Topic Exercise (p.99)



19 A student investigated the effect of temperature on the decomposition of hydrogen peroxide solution in the presence of manganese(IV) oxide. The student measured the time taken to collect  $5 \text{ cm}^3$  of oxygen gas. The student did the investigation at two different temperatures. All the other variables were kept constant. The results obtained are shown in the table below.

Temperature of hydrogen peroxide solution ( $^{\circ}\text{C}$ )	Time taken to collect $5 \text{ cm}^3$ of oxygen gas (s)
20	50
30	34

 Topic Exercise (p.99)

a) State the function of manganese(IV) oxide.

As a catalyst (1)

b) The teacher said that the student should repeat the investigation to get more results. Suggest why. Any one of the following:

- Wider range of temperatures (1)
- Repeat at the same temperature to improve accuracy / reliability (1)
- Reveal anomalous result (1)
- So an average can be obtained (1)

c) The student concluded that:

‘the rate of the reaction increases when the temperature is increased’.

Explain this in terms of the collisions of particles.

At a higher temperature, the reactant particles have more kinetic energy and move faster:

- the reactant particles involve in more collisions, resulting in more effective collisions; (1)
- a larger fraction of the colliding particles will have kinetic energy equal to or greater than the activation energy and hence a higher percentage of collisions can result in a reaction. (1)



## Topic Exercise (p.99)

- 20) Under certain conditions, a pink compound X reacts with NaOH(aq) to give a colourless product. Three trials of an experiment were conducted to study the kinetics of the reaction. Firstly, three NaOH(aq) solutions were prepared by mixing different volumes of 2.0 M NaOH(aq) and H<sub>2</sub>O(l) at 25 °C. After that, one drop of X was added to each of them and the time needed for the pink colour to disappear was recorded. The relevant data is shown below.



Trial	Volume of 2.0 M NaOH(aq) used (cm <sup>3</sup> )	Volume of H <sub>2</sub> O(l) used (cm <sup>3</sup> )	Time needed for the pink colour to disappear (s)
1	5.0	0	61
2	4.0	1.0	76
3	3.0	2.0	101

 Topic Exercise (p.99)

- Why is it necessary to make the total volume of the reaction mixtures the same for the trials?
- Given that at 25 °C,  $[\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})] = 1.0 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ , calculate the pH of the NaOH(aq) solution prepared in *Trial 2*.
- Based on the information provided, deduce ONE factor which affects the rate of this reaction.
- Detection of colour change using naked eye is not accurate enough. Suggest an instrumental method that can be used to more accurately detect the colour change.

(HKDSE, Paper 1B, 2016, 11)

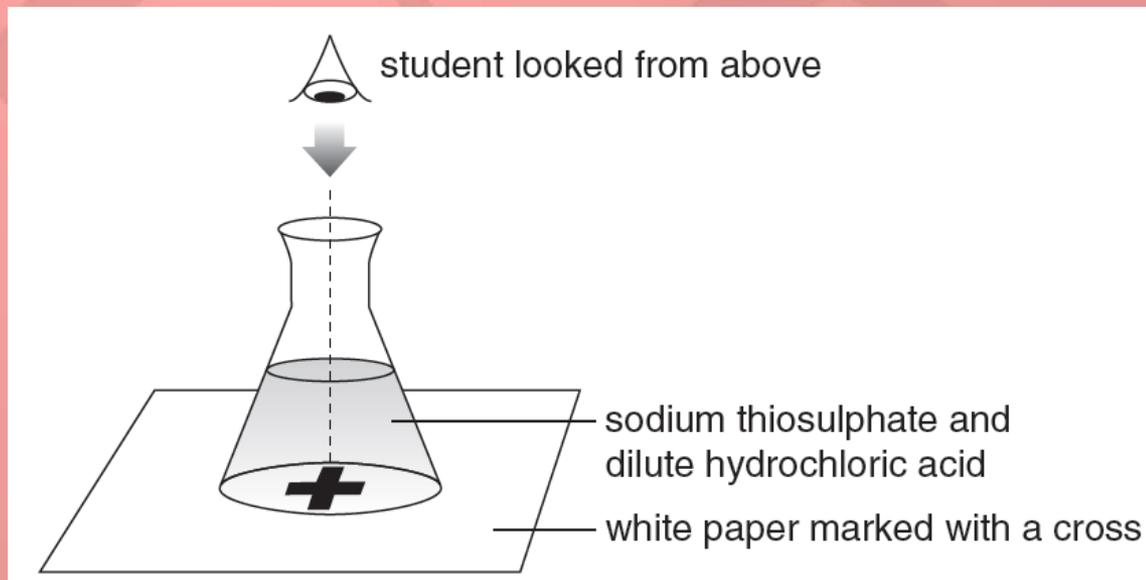
Answers for the questions of the public examinations in Hong Kong are not provided (if applicable).

 Topic Exercise (p.99)

 21 A student investigated the rate of the reaction between sodium thiosulphate solution and dilute hydrochloric acid as shown below. The reaction mixture turned turbid when the reactants were mixed. The student carried out the experiment at different temperatures by warming the reactants before mixing. She timed how long it took for the cross to become invisible each time.

a) Name the product that made the reaction mixture turbid.

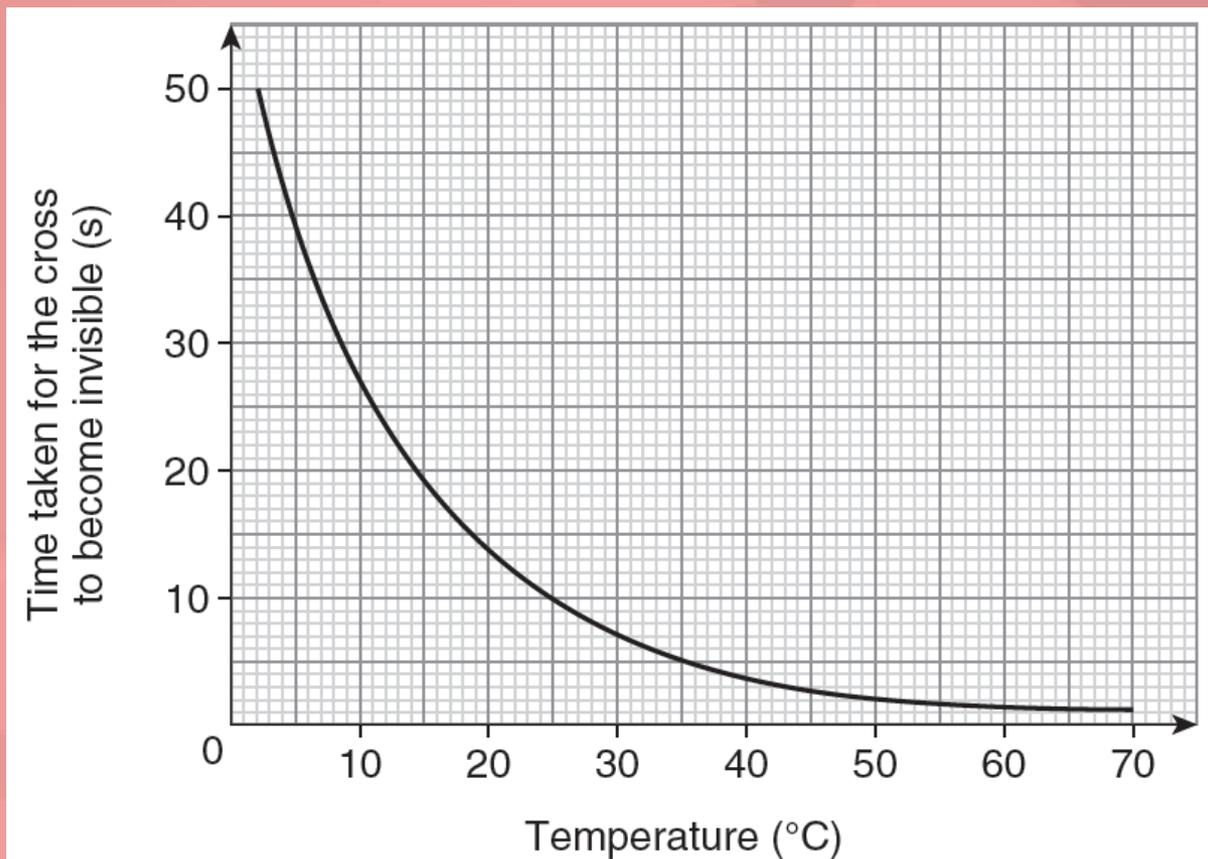
Sulphur (1)





## Topic Exercise (p.99)

b) The student's results are shown below.



 Topic Exercise (p.99)

i) What conclusion can you draw from the graph?

As the temperature increases, the time taken for the cross to become invisible decreases. /

As the temperature increases, the rate of reaction increases. (1)

ii) For many reactions, the rate of reaction doubles with every 10 °C increase in temperature. Is this statement correct for this reaction? Justify your answer.

Yes

At 10 °C, rate of reaction  $\propto \frac{1}{27 \text{ s}}$

At 20 °C, rate of reaction  $\propto \frac{1}{14 \text{ s}}$

Thus, the rate doubles with a 10 °C increase in temperature.

iii) The teacher suggested that the student's results were less accurate at 70 °C than at 40 °C. Explain why.

At 70 °C, the time taken for the cross to become invisible was very short (less than 2 seconds). (1)



## Topic Exercise (p.99)



22 You are provided with common laboratory apparatus, hydrogen peroxide solution, manganese(IV) oxide and dish detergent. Outline how you would perform a fair comparison in studying the effect of different concentrations of hydrogen peroxide solution on the rate of catalysed decomposition of hydrogen peroxide solution.



## Topic Exercise (p.99)

### Procedure

- 1 Add a little dish detergent to 50 cm<sup>3</sup> of hydrogen peroxide solution in a 500 cm<sup>3</sup> measuring cylinder. Add 0.5 g of manganese(IV) oxide to the mixture. Start the stopwatch. (1)
- 2 A foam rises up the cylinder. Stop the stopwatch when the foam rises to the top (or other marked point) of the cylinder. (1)
- 3 Dilute the hydrogen peroxide solution to different concentrations by adding water. (1)
- 4 Repeat the experiment using hydrogen peroxide solutions of different concentrations. Compare the times taken for the foam to rise to the top (or other marked point) of the cylinder. (1)

### Conditions for performing a fair comparison

Any one of the following:

- Use the same amount of dish detergent / hydrogen peroxide solution / catalyst. (1)
- Carry out the experiment under the same experimental conditions, such as the same temperature or pressure. (1)

Communication mark (1)



## Topic Exercise (p.99)

23 Catalysts are important in many industrial processes.

a) i) Write an equation for a reaction in an industrial process that uses a catalyst.

ii) Name the catalyst used in the reaction.

Any suitable example is acceptable.



iron



vanadium(V) oxide

b) Give ONE reason why using a catalyst reduces costs.

Any one of the following:

- Increases the rate of reaction. (1)

- Reduces the energy required. (1)

- Lower temperature can be used. (1)

- Catalyst is not consumed at the end. (1)

 Topic Exercise (p.99)

24 The explosive nitroglycerine ( $\text{C}_3\text{H}_5\text{O}_9\text{N}_3$ ) decomposes according to the equation below:



Assuming that all the products are measured at room temperature and pressure, calculate the total volume of gas produced from 1.00 kg of nitroglycerine.

(Relative atomic masses: H = 1.0, C = 12.0, N = 14.0, O = 16.0; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )

$$\text{Number of moles of } \text{C}_3\text{H}_5\text{O}_9\text{N}_3 = \frac{1\,000 \text{ g}}{227.0 \text{ g mol}^{-1}} = 4.41 \text{ mol}$$

4 moles of  $\text{C}_3\text{H}_5\text{O}_9\text{N}_3$  produce 19 moles of gas.

$$\text{Number of moles of gas produced} = \frac{19}{4} \times 4.41 \text{ mol} = 20.9 \text{ mol (1)}$$

$$\begin{aligned} \text{Volume of gas produced} &= \text{number of moles of gas} \times \text{molar volume of gas} \\ &= 20.9 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1} \end{aligned}$$

$$= 502 \text{ dm}^3 \text{ (1)}$$

$\therefore$  502  $\text{dm}^3$  of gas are produced.



## Topic Exercise (p.99)



25 The nitrate of strontium is often added to firework mixtures to produce red flames. The equation for the decomposition of one such mixture is as follows:



A 10.0 g mixture of  $\text{Sr}(\text{NO}_3)_2$  and carbon, in a mole ratio of 1 : 3, is allowed to decompose. What is the volume of gas given off, measured at room temperature and pressure?

(Relative atomic masses: C = 12.0, N = 14.0, O = 16.0, Sr = 87.6; molar volume of gas at room temperature and pressure =  $24.0 \text{ dm}^3 \text{ mol}^{-1}$ )



## Topic Exercise (p.99)

$$\begin{aligned} \text{Mass of 1 mole of Sr(NO}_3)_2 + 3 \text{ moles of C} \\ &= (87.6 + 2 \times 14.0 + 6 \times 16.0 + 3 \times 12.0) \text{ g} \\ &= 247.6 \text{ g} \end{aligned}$$

1 mole of  $\text{Sr(NO}_3)_2$  reacts with 3 moles of C to give 4 moles of gas.

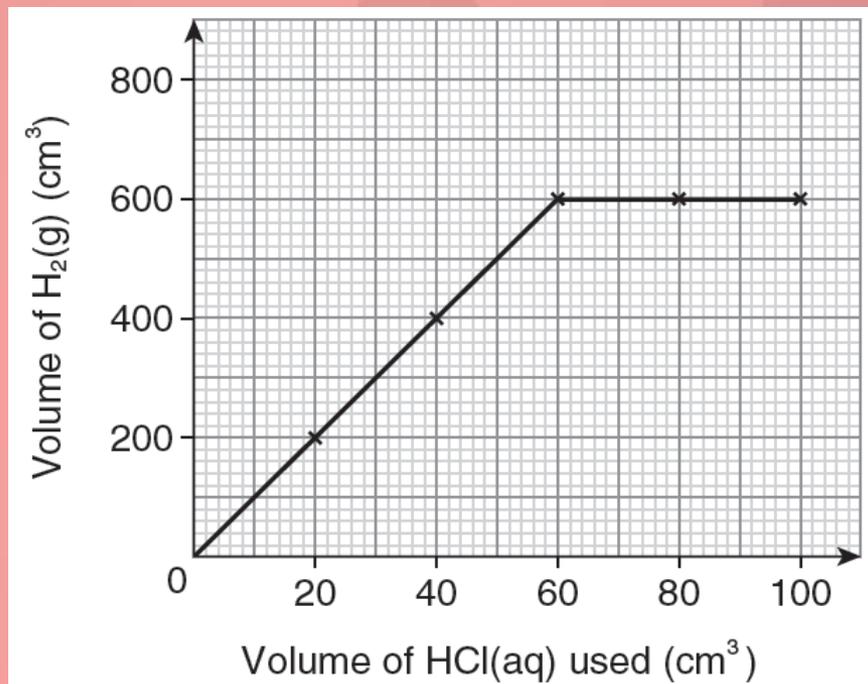
$$\begin{aligned} \text{Number of moles of gas given off} &= 4 \times \frac{10.0}{247.6} \text{ mol} \\ &= 0.162 \text{ mol (1)} \end{aligned}$$

$$\begin{aligned} \text{Volume of gas given off} &= \text{number of moles of gas} \times \text{molar volume of gas} \\ &= 0.162 \text{ mol} \times 24.0 \text{ dm}^3 \text{ mol}^{-1} \\ &= 3.89 \text{ dm}^3 \text{ (1)} \\ \therefore 3.89 \text{ dm}^3 \text{ of gas are given off.} \end{aligned}$$



## Topic Exercise (p.99)

- 26 The volume of hydrogen formed, at room temperature and pressure, as hydrochloric acid was added to a pure sample of aluminium is in the graph shown below.



(Relative atomic mass: Al = 27.0; molar volume of gas at room temperature and pressure = 24.0 dm<sup>3</sup> mol<sup>-1</sup>)

Calculate the mass of aluminium in this pure sample.



## Topic Exercise (p.185)



$$\begin{aligned} \text{Number of moles of H}_2 &= \frac{\text{volume of H}_2}{\text{molar volume of gas}} = \frac{600 \text{ cm}^3}{24\,000 \text{ cm}^3 \text{ mol}^{-1}} \\ &= 0.0250 \text{ mol (1)} \end{aligned}$$

According to the equation, 1 mole of Al reacts with HCl to give  $\frac{3}{2}$  moles of H<sub>2</sub>.

$$\text{i.e. number of moles of Al} = \frac{2}{3} \times 0.0250 \text{ mol (1)}$$

$$= 0.0167 \text{ mol}$$

$$\text{Mass of Al in sample} = 0.0167 \text{ mol} \times 27.0 \text{ g mol}^{-1} = 0.451 \text{ g (1)}$$

∴ the mass of aluminium in the pure sample is 0.451 g.