# Topic Exploration Pack

# Gas volumes

[Instructions and answers for teachers 1](#_Toc454782098)

[Learning outcomes 2](#_Toc454782099)

[Introduction 2](#_Toc454782100)

[Teacher preparation 3](#_Toc454782101)

[Activity 1 – Moles: True or False? 3](#_Toc454782102)

[Activity 2 – Gas calculation extension questions 5](#_Toc454782103)

[Activity 3 – PhET – Ideal gas behaviour 8](#_Toc454782104)

[Gas volumes – Test yourself questions – Mark scheme 10](#_Toc454782105)

[Additional guidance: 12](#_Toc454782106)

[Learner activity 15](#_Toc454782109)

[Activity 1 15](#_Toc454782110)

[Activity 2 16](#_Toc454782111)

[Activity 3 18](#_Toc454782112)

[Gas volume – Test yourself questions 20](#_Toc454782113)

## Instructions and answers for teachers

These instructions cover the student activity section which can be found on [page 14](#_Student_Activity). This Topic Exploration Pack supports OCR A Level Chemistry A.

**When distributing the activity section to the students either as a printed copy or as a Word file you will need to remove the teacher instructions section.**

### Learning outcomes

2.1.3(a) explanation and use of the terms (ii) *mole* (symbol ‘mol’), as the unit for amount of substance; (iii) *the Avogadro constant*, *N*A, (the number of particles per mole, 6.02 × 1023 mol–1); (v) *molar gas volume* (gas volume per mole, units dm3 mol–1)

2.1.3(e) calculations, using amount of substance in mol, involving (ii) gas volume

2.1.3(f) the ideal gas equation: *pV* = *nRT*

2.1.3(g) use of stoichiometric relationships in calculations

### Introduction

**Prior knowledge and student misconceptions**

Gas volume calculations are included in Module 2 of the specification, alongside other amount of substance calculations. It is expected that calculations involving gas volumes will be introduced once learners have been given a good introduction to basic stoichiometric calculations. At this point, they will have covered many mole calculations, but may still be harbouring misconceptions. Learners often find the mole concept abstract, and will tend to rely on rote learning steps to complete calculations. Many learners also struggle with the concept of ‘empty space’ in a gas and as such cannot understand why volume is independent of particle size. The ‘Moles True or False’ activity mentioned below is designed to elicit the common confusions and misconceptions learners may have at this stage.

**Gas volume calculations**

Learners from different backgrounds may or may not have covered gas volume calculations prior to A Level. (Molar gas volume calculations are covered in the reformed (9–1) GCSE Chemistry specification at Higher Level, but not in GCSE Combined Science. The relationship between temperature and pressure of gases is studied in GCSE Physics and Combined Science.) They will also have very different numerical skills. All learners will need practice with unit conversions and it is important to try to get them to visualise the quantities under consideration rather than just learning a set of conversion factors. Learners also struggle with the relationship between cm3 and dm3 as the prefixes are unhelpful in determining the conversion factor. Encouraging learners to visualise a dm3 as a cube of 10 cm × 10 cm × 10 cm often helps, and this sets them up for converting to m3 at a later date. A block of wood with these dimensions and other every day objects (such as a milk bottle) can help with this visualisation.

Learners who are mathematically very able will complete more straightforward calculations quickly and there is little to be gained from giving them ‘more of the same’ – for this reason a set of more challenging and open-ended questions is included with this pack.

**The ideal gas law**

Although the ideal gas equation contains a relatively large number of variables for the first year of A Level, actual manipulation of the equation is fairly straightforward. What learners are more likely to struggle with is the substitution of correct units. It is worth spending lots of time going over the units used and practising conversions before learners become too embroiled in the calculations themselves.

Practical activities which involve measurements of molar gas volumes are a useful introduction to the ideal gas equation. An experiment in which a known mass of magnesium is reacted with acid and the volume of evolved hydrogen measured can prove a useful starting point – learners can use the molar volume at RTP (24.0 dm3) to calculate the expected volume of hydrogen, and this can then lead to a discussion of why the experimental results may vary from the calculation.

**Ideal gas equation and molar volume**

Learners are often guilty of very ‘black or white’ thinking when it comes to chemical calculations and do not always fully appreciate where assumptions or estimations have been applied. Once learners are confident with the use of 24.0 dm3 as molar volume, they could do a practical activity to calculate molar volume and see how close to this value they are able to get. This allows other factors such as temperature to be taken into account and the ideal gas equation introduced.

From the ideal gas equation, learners can try calculating the molar volume of a gas at 0 °C compared to 25 °C. They will find that the 25 °C value is 24.8 dm3 (assuming standard pressure of 100 000 Pa). This is a good point for discussion about the phrase ‘room temperature and pressure’. Up until this point learners are probably unaware that the molar gas volume is only an approximate value for ‘average conditions’. ‘Room temperature and pressure’ as used within the specification does not imply any specific temperature or pressure values (it is distinct from ‘standard conditions’). To evaluate the use of 24.0 dm3 mol–1, learners might finally calculate the temperature at which 1 mol of gas would take up this volume at the average sea-level pressure of 101 325 Pa (or alternatively look up the atmospheric pressure in the weather report on the day). It works out to 19.5 °C, which is not an unreasonable ‘room temperature’.

### Teacher preparation

Separate activity sheets are provided for learners. Teacher instructions and answers are given below.

### Activity 1 – Moles: True or False?

This activity can be used at the start of the topic to elicit ideas about learners’ current level of understanding and likely misconceptions. You may then wish to revisit it at the end of the topic to establish whether or not understanding has improved or whether any areas of conceptual weakness remain. Answers and possible examples / explanations are given below:

1. **The total number and type of atoms present are the same at the start and end of a reaction.**

*True*. Learners should understand that this is the basis behind the need to balance chemical equations, which they will have encountered in previous topics and hopefully at GCSE level.

1. **The number of moles is the same at the start and the end of a reaction.**

*False*. This can be a real sticking point for learners and if they do have any misconceptions in this area it will make it very difficult for them at a later stage when they encounter equilibrium calculations that require deduction of how the moles of reactants and products change relative to each other. There are many simple example reactions that can be used to dispel this misconception, in particular thermal decomposition reactions and combustion of large hydrocarbons. A simple demonstration using lego blocks as atoms can show how rearrangement can result in a difference in the total number of ‘pieces’ in a chemical reaction.

1. **The total mass of reactants is equal to the total mass of products for any reaction.**

*True*. This relates directly to the first statement and is the basis of reacting mass calculations, which learners should have encountered in previous topics. However, examples where particular reagents are in excess may have confused learners so it is worth pointing out that ‘left over’ reagents are not part of the balanced chemical equation.

Another area for confusion here are reactions where gases make up part of the reactants or products, so it is worth checking that students understand that the mass simply ends up in a different place rather than being ‘lost’ or ‘gained’ during the reaction.

1. **The total volume of gas is the same at the start and the end of a reaction.**

*False.* If learners are confused about this point then combustion of solid or liquid hydrocarbons make good examples; thermal decomposition is also a useful example. Learners may have also become confused at some point about equilibrium and gaseous reactions, believing that shifts in equilibrium position somehow result in a ‘balancing out’ of the total moles of gas.

1. **The number of moles is directly proportional to the number of particles for any substance.**

*True.* This statement is designed to test learners’ understanding of Avogadro’s number and that the amount of substance relates to a number of particles, regardless as to whether those particles are molecules, individual atoms, individual ions or the repeating unit of a lattice structure.

1. **Increasing concentration means more particles in the same volume.**

*True*. This is an important point to make as this topic starts to include the idea of concentration as measured in percentage volume or ppm for gases; learners should not think of concentration as only applying to solutions.

1. **One mole of methane molecules (CH4) contains 1/5 mole of carbon atoms and 4/5 mole of hydrogen atoms.**

*False.* Learners who struggled to distinguish between statements 1 and 2 will also find this difficult, and it can lead them to great difficulties later on when carrying out reacting mass or equilibrium calculations. Balanced equations for the decomposition of substances may help; if learners are really struggling with the concept of moles not being conserved it may help to introduce ‘real life’ ideas that are completely removed from the subject. For example, dismantling a bicycle will result in two wheels, one seat and one handlebar. In this way, taking apart one item has resulted in four items with no change in total mass. Since moles are a measure of the ‘number’ of items, they do not have to be conserved. Additionally, learners could be asked to compare number of items (pennies, dice etc.) with the total mass of items to see that the two concepts are not the same.

1. **One mole of methane molecules (CH4) contains 1 mole of carbon atoms and 4 moles of hydrogen atoms.**

*True*. A learner having difficulty with this statement will most likely have also struggled with statement 7 – see comments above for how to approach this.

1. **One mole of methane (CH4) contains 1 mole of carbon and 2 moles of hydrogen.**

*False*. This statement is false as it doesn’t contain enough information. It would be true if the questions had asked about a relation between carbon “atoms” and hydrogen “molecules”, though even then learners should realise a mole of methane does not contain any hydrogen molecules.

1. **One mole of methane gas at room temperature and pressure occupies the same volume as one mole of hydrogen gas under the same conditions.**

*True*. If used at the start of the topic, learners may be uncertain of this. They may have unclear ideas about the meaning of pressure as applied to gases, so this will be a useful statement to lead onto gas calculations. If they have encountered gas calculations before it is still important to discuss this as learners may have become easily confused with statement 11 below.

1. **One mole of methane gas at room temperature and pressure has the same mass as one mole of hydrogen gas under the same conditions.**

*False*. Learners should know that this is not true due to their previous experience with molar mass calculations; however they may have become confused between molar mass and molar volume at some stage.

### Activity 2 – Gas calculation extension questions

It is suggested that these problems are only attempted once learners are competent at tackling straightforward examples of calculations involving molar gas volume, molar mass, and the ideal gas equation. They can be tackled individually but are better approached in pairs or small groups as learners will benefit enormously from working as a team and talking through their approach to the problems.

Some of the problems are very open ended and can be approached from several different directions – this may lead learners to panic if they are only comfortable with very clear steps in calculations. Before leading them straight to the answers, try to prompt them to write down all the information they are given alongside the familiar equations (molar gas volume, molar mass, and *pV* = *nRT* if it is relevant). They should also be encouraged to use units to help them to derive relationships they are less familiar with (such as density).

#### Sample answers / notes on individual problems

N.B. The open nature of these calculations means that the calculation methods provided here may not always be the same as those demonstrated by learners, or your preferred method!

#### Percentage empty space in air

The information given actually makes this problem fairly straightforward as it can be answered with a few mole calculations. There are a number of ways to approach it but the simplest is to compare the volume of one mole of liquid with one mole of gas, with the assumption being that the volume of a liquid is entirely due to its component molecules (i.e. that there are no spaces between particles).

* The volume of one mole of nitrogen gas at RTP is 24 000 cm3.
* Since the density of liquid nitrogen is 0.807 g cm–3 and its molar mass is 28 g mol–1, the volume of one mole of liquid nitrogen is equal to 28 / 0.807 = 34.7 cm3.
* The percentage volume of nitrogen gas that is occupied by molecules is therefore just 0.15%, and the percentage empty space is 99.85%.

This answer will often surprise learners and they may even assume their calculations must be wrong – it is worth pointing out that this is one of the reasons that the molar gas volume is a constant for all gases no matter the molecular volume.

#### Car engine problem

This problem is one of the more straightforward ones as a lot of starting information is provided. A sample calculation is given below:

* One mile will require 100 cm3 octane, which is equivalent to 70.3 g.
* The molar mass of octane is 114 g mol–1, so the amount of octane required is 70.3/114 = 0.62 mol
* In the balanced equation one mole of octane reacts with 12.5 moles of oxygen gas, so the amount of oxygen required is 7.7 mol
* The volume of **pure oxygen** required is therefore 7.7 × 24 = 185 dm3
* However, air is only about 20% oxygen so the volume of air would be five times this, 925 dm3. (Learners may have been much more precise about the percentage composition of air – it is worth pointing out that so many assumptions are already made that this would have little effect on the accuracy of the final value).

If learners find this easy it is worth asking them to list all the approximations that are made, and to estimate the relative impacts of each assumption on the accuracy of their answer.

#### Argon problem

This problem of course relies heavily on the size of the room the learners are in! You could get them to pace the room and then measure the size of their paces using metre rules. Remind them that we are only working on an estimate and that there is no need to account for the complicated shapes that may arise from a science laboratory. Alternatively, you could get them to estimate for a space such as a corridor which would be easier to pace and does not contain so many obstacles!

The proportion of argon in air is not provided in the question. You could encourage learners to look up the information if time is available.

* If the room is about 200 m3, this is 2 × 108 cm3. Argon makes up about 1% of the air in the room, meaning that there are 2 million cubic centimetres of argon in the room.
* The learners can now either adopt an easy approximation (2 000 000 / 24 000 = 83.3 moles argon) or they can use the ideal gas equation. Using the simple approximation the mass of argon in this size of room is 3.3 kg.
* It might be useful to get learners to try the calculation with and without the ideal gas equation, and then ask them if they think the error introduced by *not* taking temperature into account is significant compared to the other errors (for example, in their estimation of room volume).

#### Bicycle tyre problem

This calculation does require knowledge of the ideal gas equation, since the pressure difference inside a tyre is given. However, there is no need for learners to plug all the variables into the equation – since the calculation is for room temperature it can simply be assumed that there will be 8 times the number of moles that would be present under standard pressure. In the initial calculation learners need to take care over units and in the second part they can use either the nitrogen approximation used previously, or calculate an average molar mass for air using percentage composition.

1. Using the data in the question, *r* = 1.25 cm, *D* = 70 cm, *V* = 1080 cm3. At room temperature and pressure this would equate to 1080/24000 = 0.045 moles gas per tyre; at eight times this pressure there would be 0.045 × 8 = 0.36 moles of gas per tyre and 0.72 moles in two tyres.

N.B. If the ideal gas equation is used fully at 298 K, the answer would equal 0.70 moles.

1. If the air is assumed to be 78% nitrogen, 21% oxygen and 1% argon the average molar mass would be approximately 29.2 g mol–1. The mass of air in the two types would be approximately 0.72 × 29.2 = 21.0 g.
2. If the tyres were filled with helium, the mass of gas would be 0.72 × 4 = 2.88 g. The reduction in mass would then be 21.0 – 2.88 = 18.1 g.

### Activity 3 – PhET – Ideal gas behaviour

This activity will require use of a PC to run the simulation. It can be carried out individually during lesson time if ICT facilities are available; alternatively it could be set as a homework exercise or worked through with teacher guidance using a projector. The best scenario is for learners to work through it in an ICT suite with teacher guidance; however you should check that your facilities allow the simulation to run. The address for the simulation is <https://phet.colorado.edu/en/simulation/legacy/gas-properties>.

If learners are working through the worksheet themselves then they can use the following instructions. Additional notes for teachers are provided in italics.

**Using the simulation:**

There are a large number of parameters on the right hand side so try to only change one thing at a time while observing effects. There are controls that can be operated on the diagram as follows:

* Temperature can be changed by clicking the ‘add’ and ‘remove’ signs on the gas stove.
* Gas particles can be added using the pump, and removed by opening the lid on the top of the container.
* The volume of the container can be changed by sliding the man on the left of the container backwards and forwards.
* Keep gravity set to zero – we are considering the behaviour of gases in an ‘ideal’ situation!

There is a reset button on the bottom right hand side for when you need to start with a new set of conditions.

**1. Investigating the effect of the number of molecules on pressure**

Use either the pump or the ‘gas in chamber’ controls to add gas molecules to the container. Record the number of molecules against the pressure (read on the pressure gauge).

1. What is the **numerical** relationship between pressure and number of gas molecules?
2. By inference, what do you think the relationship is between pressure and moles of gas?
3. Does it make any difference whether the gas molecules are ‘light’ or ‘heavy’? Can you explain your answer?

***Answers:*** *The relationship between number of molecules and pressure is directly proportional; since number of moles is proportional to the number of molecules then number of moles is also directly proportional to pressure. The nature of the molecules does not make a difference as it is the force and frequency of molecules hitting the side of the container which creates pressure (this is investigated in more detail later on).*

**2. Investigating the effect of gas volume on pressure**

Set up the container so that it contains 100 gas molecules, with a temperature of 300 K. Use the controls on the top right of the screen to set the temperature as constant. Make a note of the pressure.

1. Use the slider to reduce the volume of the container (you will have to do this by eye). What happens to the pressure when you decrease the volume by half? What is the relationship between pressure and volume?
2. Try moving the slider back and forth and watching the thermometer. What happens to the temperature initially when you compress the gas? Why do you think this is?

***Answers:*** *The pressure doubles when the volume decreases by half – the two variables are inversely proportional. The next point is harder for learners to understand – in order to compress a gas, work has to be done (energy supplied). This means that the particles will have more energy when compressed into a smaller volume. In this simulation temperature has been kept constant so that only the effect of volume changes can be observed, so the mixture is cooled back to the original temperature immediately after compression.*

**3. Investigating the effect of temperature on pressure**

Reset the simulation and again add 100 molecules to the container. Set the volume to constant.

1. Record the pressure for a range of temperatures between 100 and 1000 K, and plot your results on a graph.
2. What is the numerical relationship between pressure and temperature in kelvin?
3. What do you think would happen to the pressure if the temperature was zero kelvin? Use the simulation to test your prediction, and try to explain what happened to the particles to cause this effect.
4. Why do you think it might be important to measure the temperature in kelvin, rather than degrees Celsius, when carrying out these investigations?

***Answers:*** *The graph should produce a roughly linear relationship between temperature and pressure, indicating that the two variables are directly proportional. This means that the pressure is zero when the temperature is zero – this is because the particles have no kinetic energy at absolute zero and no collisions with the container walls take place. If the temperature was measured in degrees Celsius, the proportionality would disappear as the temperature would pass through zero and into negative numbers.*

**4. Further questions / extension**

1. Try to describe what the pressure gauge is actually detecting (you may want to look up what gas pressure is a measure of). Explain why changing temperature, volume and number of molecules have the effect that they do.
2. Compare a container filled with 100 light particles, and a container filled with 100 heavy particles at the same temperature. Do the containers have the same pressure? Are the particles moving at the same speed? Do the particles have the same energy? Can you explain the similarities and differences between the two containers, using your knowledge of energy and forces?

***Answers:*** *Pressure is a measure of the force and frequency with which molecules hit the walls of a container. Increasing temperature increases both force (particles have more energy) and frequency of collisions. Decreasing the volume (with temperature constant) and increasing the number of moles of gas both increase the frequency of collisions.*

*The second question is a bit more difficult! A visual comparison of the light and heavy species in the simulation shows us that the particles of the heavier gas are moving much slower – therefore they will collide less frequently with the walls of the container. However, the reason that the heavier particles move slower is because they have* ***the same average kinetic energy*** *as the lighter particles (this is also how the time of flight mass spectrometer works). Although they hit the walls of the container less frequently, their greater mass means that they hit the walls with greater force.*

### Gas volumes – Test yourself questions – Mark scheme

#### Section A: Multiple choice

1. Correct answer: C

*The molar mass of CaCO3 is 100 g mol–1 so 10 grams is equal to 0.1 mol. This equates to 2400 cm3 of CO2.*

1. Correct answer: D

*The best way to think of this is to imagine a cubic metre as a cube with sides of length 100 cm (as there are 100 cm in a metre). 100 × 100 × 100 = 1 000 000.*

1. Correct answer: B

*The ideal gas assumes that there are no intermolecular attractions between gas molecules – the stronger these forces of attraction are, the less likely the gas will display ‘ideal’ behaviour.*

1. Correct answer: C

*The units of pressure are pascals because this is derived from SI units (1 Pa = 1 N m–2)*

1. Correct answer: C

*This can be answered either with some lengthy calculations or with a bit of common sense. A describes 0.1 moles of a gas at 298 K and 100 kPa; B has half the number of moles at half the pressure; D has twice the number of moles at twice the pressure. C is the only answer at standard temperature rather than room temperature, hence the difference in volume.*

#### Section B: Longer answers

*All answers given to 3 s.f.*

1. a) *n*(glucose) = 28 / 180 = 0.155 mol

*n*(O2) = 0.155 × 6 = 0.933 mol

mass O2 = 0.933 × 32 = **29.9 g**

b) 0.933 mol (or answer for moles oxygen from part a) × 24.0 = **22.4 dm3**

c) 22.4 dm3 = 21% *V*

therefore 22.4 = (21/100) *V*; *V* = 22.4 /(21/100) = **107 dm3**

1. a) *n*(C3H5O9N3) = 1000 / 227 = 4.4 mol

2 moles of C3H5O9N3 produce 9.5 moles of gas (carbon dioxide, oxygen and nitrogen – water is assumed a liquid at RTP) therefore 4.4 mol produce 9.5 × 2.2 = 20.9 moles of gas

Total volume = 20.9 × 24 = **502 dm3**

b) The volume of gas would be much larger than calculated. Water is a gas at higher temperature, so there would be 5 more moles of gas produced per 2 mol nitroglycerine. Also, according to the ideal gas equation (*pV = nRT*), the volume taken up by the gas is greater at higher temperature than at room temperature.

1. a)  *n*(CO2) produced = 0.001

Conversion of temperature and pressure:

20 °C = 293 K 101 kPa = 101 000 Pa

Substitution into ideal gas equation to get volume in m3:

*V* = *nRT*/*P* = 0.001 × 8.314 × 293 / 101 000 = 2.41 × 10–5 m3

Conversion into cm3: 2.41 × 10–5 × 1 000 000 = **24.1 cm3**

b) Using molar gas volume: 0.001 × 24 000 = 24.0 cm3

Difference between results = 24.1 – 24.0 = 0.1 cm3

Percentage difference = (0.1 / 24.0) × 100 = **0.417%**

1. a) *Conversion of temperature, pressure and volume units*:

40.5 cm3 = 4.05 × 10–5 m3 50 °C = 323 K 106 kPa = 106 000 Pa

*Substitution into ideal gas equation to get number of moles:*

*n* = *pV*/*RT* = (106 000 × 4.05 × 10–5) / (8.314 × 323) = 0.00160 mol

b) mass of gas = 182.7393 – 182.5601 = 0.1792 g

molar mass of gas = 0.1792 / 0.00160 = 112 g mol–1

c) Mass of CH2 = 14; 112 / 14 = 8 therefore the formula is C8H16

### Additional guidance:

Teachers should expect, and indeed encourage, a wide range of questions when handling difficult concepts such as amount of substance or ideal gas behaviour. For dealing with the more complex topic areas, the Chemguide website is an invaluable tool for providing clear and concise information:

<http://www.chemguide.co.uk/physical/kt/idealgases.html#top>

The link given above is for the Ideal Gases section, but there are also useful sections on other calculations and gas collection methods.

Learners will have very variable ranges of ability and those that lack confidence will really struggle with some of the more demanding problems. Encourage learners to gain confidence in small calculation steps first before attempting multi-step calculations; the use of rough paper or individual whiteboards will give learners more confidence as otherwise they will be reluctant to commit their first ideas to paper. Be very explicit whether you want learners to work individually (in which case silence should be encouraged so that learners are not tempted to ask for help from their peers) to gain competence in familiar tasks, or whether you want learners to discuss ideas when tackling less familiar problems that require some creativity.

The ‘Knockhardy’ website provides useful PowerPoints and notes on a range of topics and the PowerPoints are designed so that learners can step through them at their own pace – this is particularly useful for consolidation of mole calculations.

<http://www.knockhardy.org.uk/ppoints.htm>

<http://www.knockhardy.org.uk/sci.htm>

**Engaging worksheets and activities:**

The RSC has a huge range of resources for lesson starters, plenaries and extension activities. Additionally they provide a useful practical worksheet for the measurement of molar gas volume. Links to the most relevant sections of the website are included below:

Quantitative Chemistry ‘Starters for 10’: <http://www.rsc.org/learn-chemistry/resource/res00000954/starters-for-ten?cmpid=CMP00001406>

Experiment to measure one mole of hydrogen gas: <http://www.rsc.org/learn-chemistry/resource/res00000452/the-volume-of-1-mole-of-hydrogen-gas>

Experiment to determine the molecular mass of butane using the ideal gas equation: <http://www.rsc.org/learn-chemistry/resource/res00001720/determining-the-relative-molecular-mass-of-butane>

Gridlocks games (includes mole calculations and units of volume): <http://www.rsc.org/learn-chemistry/resource/res00000878/gridlocks-can-you-unlock-the-grid#!cmpid=CMP00001043>

Assessment for learning – short list of calculations to assess student understanding in basic mole calculations: <http://www.rsc.org/learn-chemistry/resource/res00000096/afl-calculations-in-chemistry>

**Videos:**

‘Fuse School’ is an excellent Chemistry You Tube channel. Links to the most relevant videos are provided below – they assume little prior knowledge and are ideal as recaps or introductions to new topics.

Molar volume of gases: <https://www.youtube.com/watch?v=UCmYSIjOnUA&list=PLW0gavSzhMlTfVVUm_0GQeV3Gc9yenOXd&index=16>

Calculating molar volume using experimental data: <https://www.youtube.com/watch?v=lRiOW9GiFcE&list=PLW0gavSzhMlTfVVUm_0GQeV3Gc9yenOXd&index=18>

Calculating gas volume using equations: <https://www.youtube.com/watch?v=dcqa4md4pAg&index=17&list=PLW0gavSzhMlTfVVUm_0GQeV3Gc9yenOXd>

Collecting and identifying gases: <https://www.youtube.com/watch?v=EOBRZxn4JIA>

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# Topic Exploration Pack

# Gas volumes

## Learner activity

### Activity 1

**Moles: True or False?**

First decide whether the following statements are true or false. They need some thought! Then try to come up with a written explanation for each answer, using an example to prove your point!

1. The total number and type of atoms present are the same at the start and end of a reaction.
2. The number of moles is the same at the start and the end of a reaction.
3. The total mass of reactants is equal to the total mass of products for any reaction.
4. The total volume of gas is the same at the start and the end of a reaction.
5. The number of moles is directly proportional to the number of particles for any substance.
6. Increasing concentration means more particles in the same volume.
7. One mole of methane molecules (CH4) contains 1/5 mole of carbon atoms and 4/5 mole of hydrogen atoms.
8. One mole of methane molecules (CH4) contains 1 mole of carbon atoms and 4 moles of hydrogen atoms.
9. One mole of methane (CH4) contains 1 mole of carbon and 2 moles of hydrogen.
10. One mole of methane gas at room temperature and pressure occupies the same volume as one mole of hydrogen gas under the same conditions.
11. One mole of methane gas at room temperature and pressure has the same mass as one mole of hydrogen gas under the same conditions.

### Activity 2

**Gas calculations: Extension questions**

These questions are designed to challenge you and to make you think a little harder than the more straightforward calculations you have been used to. Remember to refer to all the equations and constants you have covered so far, and make notes of all the calculations as you make them.

**Percentage empty space in air**

What percentage of air is empty space at room temperature and pressure?

*You can answer this question by approximating air to be completely composed of nitrogen. Liquid nitrogen has a density of 0.807 g cm–3. You will of course need other familiar constants and equations to work this out. N.B. we are ignoring the empty space* within *atoms themselves – only the space between molecules is being calculated!*

**Car engine problem**

Approximately what volume of air is required to enter a car engine as it travels one mile?

*Some of the information needed to complete this problem is listed below; there are other equations and assumptions that you will have covered already in your course.*

1. *Petrol is a mixture of hydrocarbons but can be assumed to be mostly octane.*
2. *Octane (C8H18) has a density of 0.703 g cm–3.*
3. *A normal car will consume petrol at a rate of about ten miles per litre.*
4. *Although air is heated and pressurised as it enters the car engine, we can assume that the normal molar gas volume applies to the air before it enters.*
5. *Air is about 20% oxygen.*

**Argon problem**

What is the approximate mass of argon in the room you are in at the moment?

*You will need to estimate the volume of the room – it may be easier to measure lengths in cm so that your final answer is in cm3. You could estimate the length of one of your paces and then count paces – the height estimation will be less accurate!*

*You may also need to account for the temperature of the room if it is not ‘room temperature’ – how much difference do you think this will make to your final answer?*

**Bicycle tyre problem**

The shape of a bicycle inner tube is called a torus. The tyre radius is significantly smaller than the wheel diameter, and its volume can be approximated according to the following equation:

*V* = π2*r*2*D*

*r* = the radius of the tyre

*D* = the diameter of the wheel

When inflated, a typical bicycle tyre has a wheel diameter of 70 cm and a tyre radius of 12.5 mm.

The pressure inside an inflated tyre is about eight times higher than atmospheric pressure.

1. Estimate how many moles of gas are in **two** inflated tyres at room temperature.
2. Estimate the mass of gas in two tyres inflated with dry air.
3. For elite cyclists, even small savings in mass can have significant impacts in races – this is the idea of marginal gains. Estimate the reduction in mass of the bicycle if the tyres were inflated with helium instead of air.

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Volume of a torus derivation: <http://whistleralley.com/torus/torus.htm>

### Activity 3

**PhET – Ideal gas behaviour**

This activity requires a PC to run (the simulation may not work on tablets if they do not have Java enabled). The address for the simulation is <https://phet.colorado.edu/en/simulation/legacy/gas-properties>.

**Using the simulation:**

There are a large number of parameters on the right hand side so try to only change one thing at a time while observing effects. There are controls that can be operated on the diagram as follows:

* Temperature can be changed by clicking the ‘add’ and ‘remove’ signs on the gas stove.
* Gas particles can be added using the pump, and removed by opening the lid on the top of the container.
* The volume of the container can be changed by sliding the man on the left of the container backwards and forwards.
* Keep gravity set to zero – we are considering the behaviour of gases in an ‘ideal’ situation!

There is a reset button on the bottom right hand side for when you need to start with a new set of conditions.

1. **Investigating the effect of the number of molecules on pressure**

Use either the pump or the ‘gas in chamber’ controls to add gas molecules to the container. Record the number of molecules against the pressure (read on the pressure gauge).

1. What is the **numerical** relationship between pressure and number of gas molecules?
2. By inference, what do you think the relationship is between pressure and moles of gas?
3. Does it make any difference whether the gas molecules are ‘light’ or ‘heavy’? Can you explain your answer?
4. **Investigating the effect of gas volume on pressure**

Set up the container so that it contains 100 gas molecules, with a temperature of 300 K. Use the controls on the top right of the screen to set the temperature as constant. Make a note of the pressure.

1. Use the slider to reduce the volume of the container (you will have to do this by eye). What happens to the pressure when you decrease the volume by half?
2. Try moving the slider back and forth and watching the thermometer. What happens to the temperature initially when you compress the gas? Why do you think this is?
3. **Investigating the effect of temperature on pressure**

Reset the simulation and again add 100 molecules to the container. Set the volume to constant.

1. Record the pressure for a range of temperatures between 100 and 1000 K, and plot your results on a graph.
2. What is the numerical relationship between pressure and temperature in kelvin?
3. What do you think would happen to the pressure if the temperature was zero kelvin? Use the simulation to test your prediction, and try to explain what happened to the particles to cause this effect.
4. Why do you think it might be important to measure the temperature in degrees kelvin, rather than degrees Celsius, when carrying out these investigations?

**4. Further questions / extension**

1. Try to describe what the pressure gauge is actually detecting (you may want to look up what gas pressure is a measure of). Explain why changing temperature, volume and number of molecules have the effect that they do.
2. Compare a container filled with 100 light particles, and a container filled with 100 heavy particles at the same temperature. Do the containers have the same pressure? Are the particles moving at the same speed? Do the particles have the same energy? Can you explain the similarities and differences between the two containers, using your knowledge of energy and forces?

### Gas volume – Test yourself questions

*You will need to refer to a data sheet that includes the gas constant R, the Avogadro constant and the molar volume at RTP.*

#### Section A: Multiple choice

1. Calcium carbonate decomposes according to the equation shown below.

CaCO3(s) → CaO(s) + CO2(g)

What volume of carbon dioxide would be produced from 10 grams of calcium carbonate fully decomposing?

|  |  |
| --- | --- |
| **A** | 44 dm3 |
| **B** | 4.4 dm3 |
| **C** | 2400 cm3 |
| **D** | 240 dm3 |

Your answer

1. How many cm3 are in one cubic metre?

|  |  |  |
| --- | --- | --- |
| **A** | 100 |  |
| **B** | 1000 |  |
| **C** | 100 000 |  |
| **D** | 1 000 000 |  |

Your answer

1. Which of these is **not** an assumption made by the ideal gas equation?

|  |  |  |
| --- | --- | --- |
| **A** | Gas molecules behave as perfect spheres. |  |
| **B** | There are strong intermolecular attractions between gas molecules. |  |
| **C** | The volume occupied by the molecules is negligible relative to the volume of the container. |  |
| **D** | The temperature of the gas is proportional to the average kinetic energy of the molecules. |  |

Your answer

1. The units of pressure in the ideal gas equation are:

|  |  |  |
| --- | --- | --- |
| **A** | atmospheres |  |
| **B** | kilopascals |  |
| **C** | pascals |  |
| **D** | millibars |  |

Your answer

1. Which of the following does **not** occupy the same volume as the other three?

|  |  |  |
| --- | --- | --- |
| **A** | 1.6 grams of methane, CH4, at 298 K and 100 kPa pressure |  |
| **B** | 0.05 moles of oxygen, O2, at 298 K and 50 kPa pressure |  |
| **C** | 4.4 grams carbon dioxide, CO2, at 273 K and 100 kPa pressure |  |
| **D** | 8.8 grams propane, C3H8, at 298 K and 200 kPa pressure |  |

Your answer

#### Section B: Longer answers

1. When glucose is metabolised in the body it is essentially reacting with oxygen to produce carbon dioxide and water:

C6H12O6 + 6O2 → 6CO2 + 6H2O

There are about 28 g of carbohydrate in a chocolate bar, most of which will end up as glucose in the body.

1. How many grams of oxygen gas would be needed to react with 28 g of glucose?
2. What volume would this amount of oxygen occupy at room temperature and pressure?
3. What volume of air is this equivalent to, if air is 21% oxygen?
4. The explosive nitroglycerine has the chemical formula C3H5O9N3. Assume that the equation for the decomposition of nitroglycerine is as follows:

2C3H5O9N3 → 6CO2 + ½O2 + 3N2 + 5H2O

1. Assuming all the products are produced at room temperature and pressure, calculate the total volume of gas produced from 1 kg of nitroglycerine.
2. In actual fact the reaction is highly exothermic and the temperature of the products is well above room temperature. How would this affect your answer to part a? Explain your answer (there is no need to include a calculation).
3. Calcium carbonate reacts with hydrochloric acid according to the following equation:

CaCO3(s) + 2HC*l*(aq) → CaC*l*2(aq) + H2O(l) + CO2(g)

1. Use the ideal gas equation to calculate the volume of carbon dioxide released, in cm3, from the complete reaction of 0.002 mol hydrochloric acid with excess calcium carbonate powder when the temperature is 20 °C and the pressure is 101 kPa.
2. A student calculated the volume released by assuming a molar volume of 24.0 dm3. Calculate the percentage difference between the two results.
3. The molar mass of a volatile compound can be found by allowing a small sample of the liquid to evaporate and fill an empty weighed syringe at a known temperature and pressure. The ideal gas equation can be used to calculate the amount of gas in mol. Results from such an experiment are shown below:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Mass of empty syringe / g** | **Mass of syringe with compound / g** | **Volume of syringe after leaving compound / cm3** | **Temperature recorded / °C** | **Pressure recorded / kPa** |
| 182.5601 | 182.7393 | 40.5 | 50.0 | 106.0 |

1. Calculate the amount of gas, in mol, in the syringe, using the ideal gas equation.
2. Calculate the mass of gas in the syringe and hence calculate its molecular mass.
3. The empirical formula of the compound was found to be CH2. Deduce its molecular formula.