Mrs Lawrence's
YEAR 11 CHEMISTRY

Module 2
Introduction to Quantitative Chemistry

Student Learning Guide

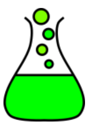
Notes to the student:

- This is a guide only and must be supported by you accessing a wide variety of information sources and by **hammering the teacher with lots of questions**.
- Don't let a concept go by… keep pestering.
- In your Lab, you have access to:
  - This workbook (your main resource)
  - Chemistry in Focus
  - SI Chemical Data
  - Surfing NSW Chemistry
  - Nelson Chemistry Units
  - Internet

An online copy of Chemistry in Focus is available at **NelsonNet.com.au**
Username: Jacob.fryatt1@det.nsw.edu.au
Password: TareeHS

- Make sure you are staying connected via Facebook "THS Chemistry 2018"
- Get work checked regularly – but at least at the end of every section.
- Mark off your own version of the syllabus as we work through the booklet
- You will be getting homework… get used to doing it!

Have fun ☺



Got a Question? Ask me!
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Taree High School

ACKNOWLEDGEMENT & REFERENCES:

In compiling this resource, extracts from a range of prepared materials were utilised. Original sources include:

Harrison, C. (2018). *Module 2: Introduction to Quantitative Chemistry Workbook.* NSW: Pitt Water House.

NESA (2017). *Chemistry Stage 6 Syllabus.* Retrieved 27th April 2018 from https://syllabus.nesa.nsw.edu.au/chemistry-stage6/content/2237/

Schell, M. & Hogan, M. (2018). *Surfing NSW Chemistry Modules 1 & 2.* Marrickville: Science Press.

Smith, D. (2003). Conquering Chemistry Preliminary Course: Blackline Masters. North Ryde: McGraw Hill Australia.

Smith, R. (2004). *Conquering Chemistry Preliminary Course.* Melbourne: Nelson Cengage Learning Australia.

Smith, R. & Davis, A. (2018). *Chemistry In Focus. Melbourne: Nelson Cengage Learning Australia.*

Diagrams used are sourced through google images.

"Most great learning happens in groups. Collaboration is the stuff of growth."

-Sir Ken Robinson Ph.D.

Module 2: Introduction to Quantitative Chemistry

Outcomes

A student:

- designs and evaluates investigations in order to obtain primary and secondary data and information CH11/12-2
- selects and processes appropriate qualitative and quantitative data and information using a range of appropriate media CH11/12-4
- solves scientific problems using primary and secondary data, critical thinking skills and scientific processes CH11/12-6
- describes, applies and quantitatively analyses the mole concept and stoichiometric relationships CH11-9

Content Focus

Students are introduced to the quantitative nature of chemistry. Chemists must be able to quantify reactions in order to make predictions about yields and communicate with specific audiences for specific purposes using nomenclature, genres and modes unique to the discipline. Using the mole concept, students will have the opportunity to select and use appropriate mathematical representations to solve problems, make predictions and calculate the mass of reactants and products, whether solid, liquid or gas.

Students further develop their understanding of the universal language of chemistry. They are introduced to the idea that science is a global enterprise that relies on clear communication, international conventions, peer review and reproducibility.

Working Scientifically

In this module, students focus on designing and evaluating investigations that enable them to obtain quantitative data to help them solve problems related to quantitative chemistry. Students should be provided with opportunities to engage with all the Working Scientifically skills throughout the course.

Content

Chemical Reactions and Stoichiometry

Inquiry question: What happens in chemical reactions?

Students:

- conduct practical investigations to observe and measure the quantitative relationships of chemical reactions, including but not limited to:
  - masses of solids and/or liquids in chemical reactions
  - volumes of gases in chemical reactions (ACSCH046)
- relate stoichiometry to the law of conservation of mass in chemical reactions by investigating:
  - balancing chemical equations (ACSCH039)
  - solving problems regarding mass changes in chemical reactions (ACSCH046)

Mole Concept

Inquiry question: How are measurements made in chemistry?

Students:

- conduct a practical investigation to demonstrate and calculate the molar mass (mass of one mole) of:
  - an element
  - a compound (ACSCH046)
- conduct an investigation to determine that chemicals react in simple whole number ratios by moles
- explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)
  - moles of elements and compounds $n = \frac{m}{M}$ (*n* = chemical amount in moles, *m* = mass in grams, *MM* = molar mass in $gmol^{-1}$)
  - percentage composition calculations and empirical formulae
  - limiting reagent reactions

Concentration and Molarity

Inquiry question: How are chemicals in solutions measured?

Students:

- conduct practical investigations to determine the concentrations of solutions and investigate the different ways in which concentrations are measured (ACSCH046, ACSCH063)
- manipulate variables and solve problems to calculate concentration, mass or volume using:
  - $c = \frac{n}{v}$ (molarity formula) (ACSCH063)
  - dilutions (*number of moles before dilution* = *number of moles of sample after dilution)*
- conduct an investigation to make a standard solution and perform a dilution

Gas Laws

Inquiry question: How does the Ideal Gas Law relate to all other Gas Laws?

Students:

- conduct investigations and solve problems to determine the relationship between the Ideal Gas Law and:
  - Gay-Lussac's Law (temperature)
  - Boyle's Law
  - Charles' Law
  - Avogadro's Law (ACSCH060)

- **conduct practical investigations to observe and measure the quantitative relationships of chemical reactions, including but not limited to:**
  - **masses of solids and/or liquids in chemical reactions**
  - **volumes of gases in chemical reactions (ACSCH046)**

The practical component of this outcome is covered throughout Module 2

The three states of matter are as follows:

Solid

Has fixed shape and volume

Liquid

Takes shape of container
Forms horizontal surface
Has fixed volume

Gas

Expands to fill container

Due to the various properties of the states, it is appropriate to use different methods and technologies to measure each.

MEASURING SOLIDS:

Within the laboratory, we can use electronic balances (aka: electronic scales) are used to weigh solid substances. Typically, when weighing out a sample, we can TARE the scales to exclude the weight of the glassware, before we add the substance.

What is meant by the term ‘tare’:

____

____

Describe how to use electronic balances to calculate the weight of the solid within the flask above:

____

____

____

____

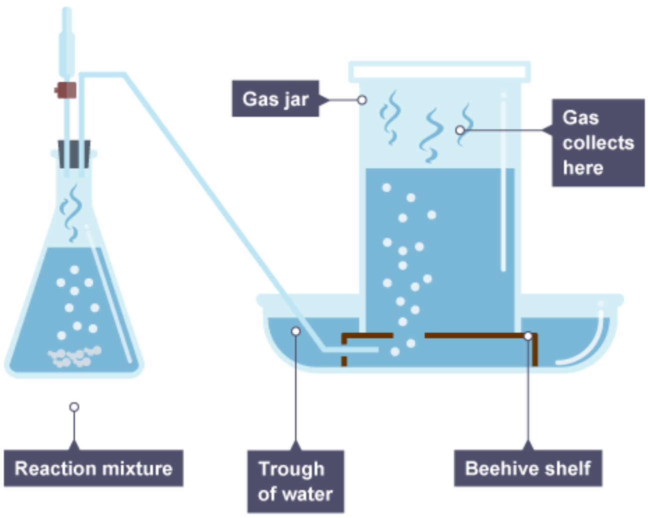
MEASURING LIQUIDS:

Read through Conquering Chemistry Prelim 'Apparatus for Measuring Solutions' pg 214-215

Complete the following table to summarise how to measure liquids:

Apparatus Name	Description	Sketch	Advantages	Disadvantages
Measuring Cylinder				
Pipettes				
Burettes				
Volumetric Flasks				

MEASURING GASES:



Explain how the following apparatus can be used to collect and measure gas liberated during a chemical reaction:

____

____

____

____

____

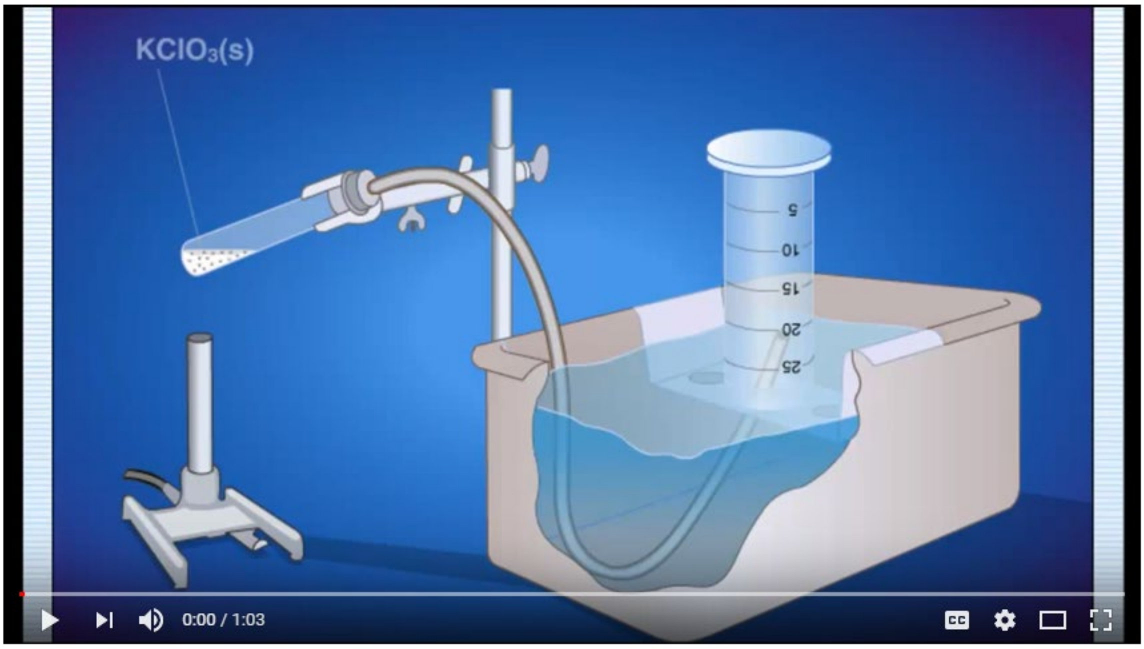
____

What is the purpose of the water in this apparatus?

____

____

Watch YouTube: 'Collecting a gas over water animation' at https://youtu.be/jg4c0rVSckg



Collecting A Gas Over Water Animation

What is the chemical reaction utilised in this demonstration?

____

MnO_2 is added to the original solid. What is its purpose?

____

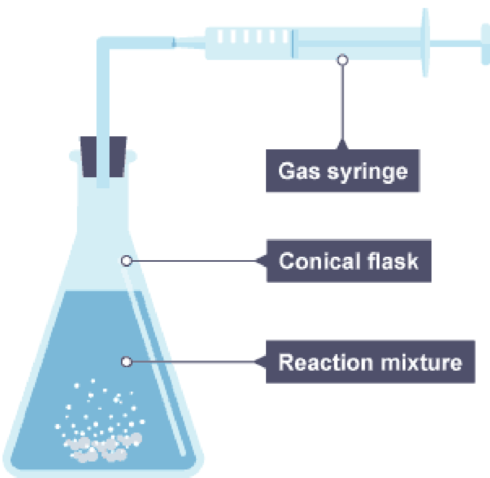
____

What is a limitation of using water displacement to measure the gas liberated?

____

____

Another method of measuring the gas volume liberated in an experiment is shown below:



Describe how to use this method to measure the gas liberated:

____

____

What is one advantage of using this method?

____

What is one limitation of using this method?

____

Experiment: Measuring Gases

Design and conduct an experiment to utilise a gas jar to collect the gas liberated from a chemical reaction

Aim:

____

____

Materials:

____

____

Method:

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Risk Assessment:

Risk	Precaution

Chemical MSDS:

____

____

Chemical Reactions:

Word equation:

Chemical Equation:

Signs of Chemical Reaction:

____

____

Results:

Conclusion:

____

____

- **relate stoichiometry to the law of conservation of mass in chemical reactions by investigating:**
  - **balancing chemical equations (ACSCH039)**
  - **solving problems regarding mass changes in chemical reactions (ACSCH046)**

Read through Conquering Chemistry Prelim 'Physical and Chemical Changes' pg 67-69

Define the following terms:

Reactants:

____

____

Products:

____

____

Therefore the following applies…

Reactants → Products

Where…

1. Mass is conserved:

   ____

   ____

2. The number of atoms of each type is conserved:

   ____

   ____

These facts are often summarised by the LAW OF CONSERVATION, which states:

____

____

Explain how new products are formed from reactants if no new matter is made?

____

____

____

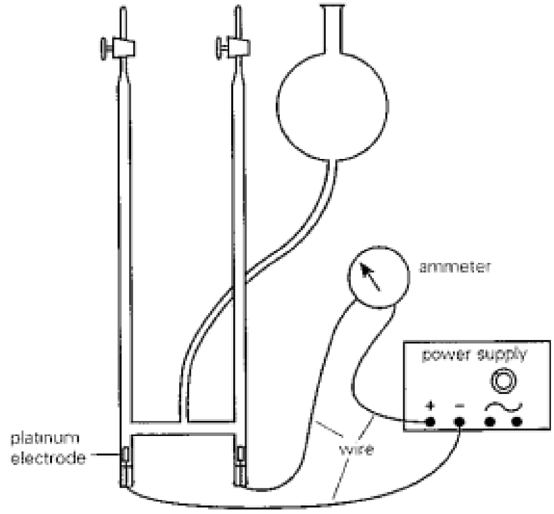


Diagram A shows the decomposition of water

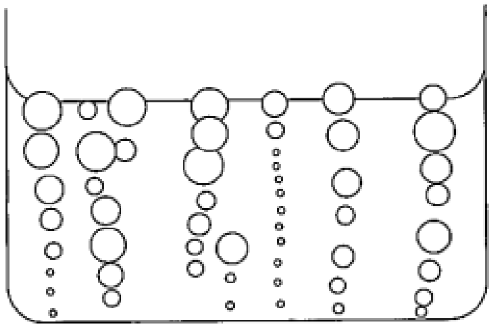
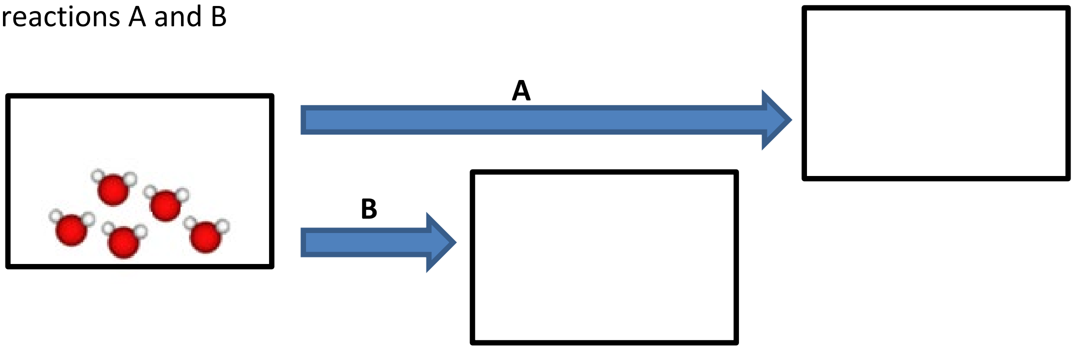


Diagram B shows water boiling

1. Diagram A shows a _________________ change while Diagram B shows a _________________ change.
2. Which of the two changes involves the greater energy change? _________________
3. Complete the following diagram showing the changes in the particle arrangement for reactions A and B



4. Process A needs 286kJ energy, while Process B needs only 44 kJ of energy for the same amount of water. Explain the difference in these values in terms of the changes involved.

____

____

____

5. Explain the relationship between the amount of energy needed to separate the atoms in a compound and the strength of the bond between atoms in that compound:

____

____

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Experiment: Investigation 7.1 – Mass in Chemical Reactions

This resource is available on NelsonNet 'Chemistry in Focus' pg 141-142

Aim: To determine whether mass is conserved during a chemical reaction

Materials:

- Calcium chloride (dried) _________________
- Sodium Hydrogen Carbonate _________________
- 10ml measuring cylinder
- Water _________________
- Universal indicator in dropper bottle and colour chart
- 2 teaspoons
- Glass vial with flat bottom
- Ziploc bag
- Electronic balance

Risk assessment:

Risk	Precaution

Method:

1. Add 1 teaspoon of calcium hydroxide to the bag
2. Add 1 teaspoon of sodium hydrogen carbonate to the bag
3. Place 10ml water and four drops of universal indicator in the glass vial
4. Carefully stand the glass vial in the bag so that it doesn't tip over
5. Seal the bag
6. Weigh the sealed bag and contents
7. Invert the bag so that the vial empties, allowing the universal indicator and water to mix with the solids
8. Record observations
9. Reweigh the sealed bag

Results:

Measurement	Observation Notes
Initial Mass (g)	
Observations of Reaction	
Final Mass (g)	
Change in Mass (g)	

Equations:

Word Equation:

Chemical Equation:

Discussion:

How could you tell that a reaction was occurring?

____

What was the purpose of doing the investigation in a sealed bag?

____

____

Justify your results:

____

____

Would you expect the results to remain the same if it were conducted in the open atmosphere? Justify your response

____

____

What sources of error are there in this experiment?

____

____

____

____

How could you minimise these?

Describe two methods that you could use to ensure your data is reliable:

Conclusion:

SUMMARY OF CONCEPTS: THE LAW OF CONSERVATION

In the reaction:

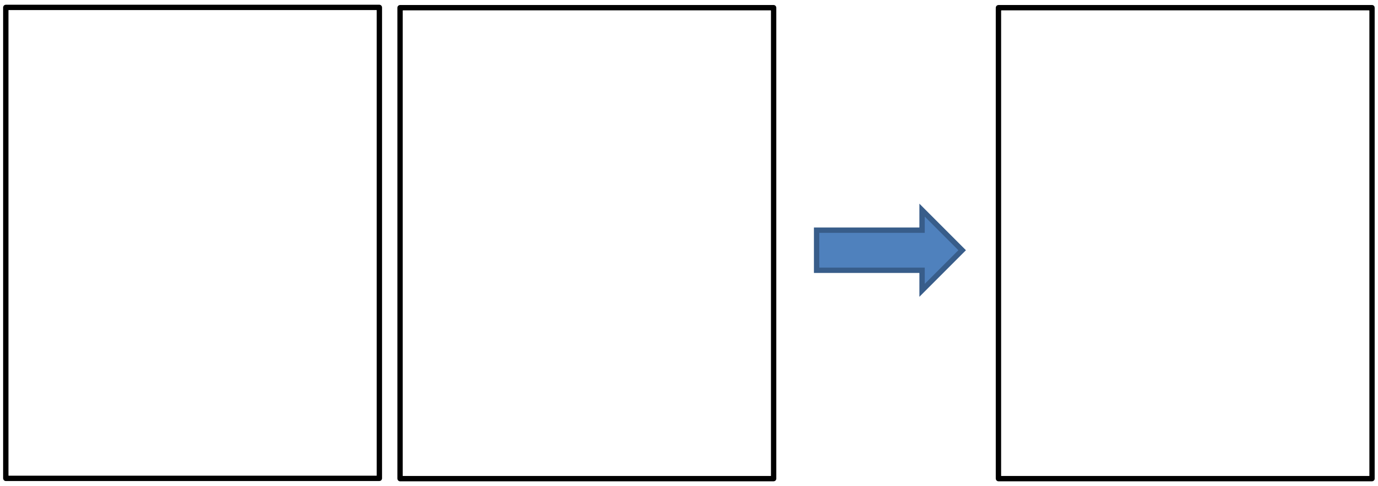
$$4Li_{(s)} + O_{2(g)} \rightarrow 2Li_2O_{(s)}$$

a. what is the product?

b. what are the reactants?

c. what does the "(s)" after the formula of lithium oxide signify?

d. represent each side of the arrow with diagrams to show the rearrangement of the atoms as the reaction takes place



CHEMICAL EQUATIONS:

Read through Conquering Chemistry Prelim '3.6 Equations for Chemical Reactions' pg 74-75

Fill in the missing words:

Chemists use ________________ to describe what happens in chemical ________________ .

To start with we write ________________ equations:

- Word equation for heating copper nitrate:

Reactants: ____

Products: ____

- Word equation for Phosphorus reacting with chlorine:

Reactants: ____

Products: ____

After we have written all the reactants and products as a word equation, we can use the periodic table to write in symbols, making a ________________ equation.

- Chemical equation for Phosphorus reacting with chlorine:

To represent the Law of ________________ of matter, we need to ensure that a chemical equation has the same ________________ of ________________ of each element on each ________________ of the ________________.

When this is true, we can say that the chemical equation is ________________. To do this, we put numbers in front of the ________________ until the number of ________________ of each ________________ on both sides of the ________________ are ________________.

- BALANCED Chemical equation for Phosphorus reacting with chlorine

The final thing we need to include in our equation is the ________________ state for each species present.

We use the following symbols:

- ________________ is for solid
- ________________ is for liquid (pure form of substance, not for mixtures)
- ________________ is for gas
- ________________ is for aqueous (meaning dissolved into a solution or mixtures)

Go through and add the appropriate symbols to the equation above.

Exercises:

1. Write word equations and chemical equations that convey the same information as the following statements:

a) Two molecules of ethane, each which contains two carbon and 6 hydrogen atoms, react with seven molecules of oxygen, each containing two atoms of oxygen, to form four molecules of carbon dioxide and six molecules of water

Word Equation

Chemical Equation:

b) One molecule of dinitrogen pentoxide, which contains two nitrogen atoms and five oxygen atoms, combines with one molecule of water to form two molecules of nitric acid, which contains one hydrogen, one nitrogen and three oxygen atoms each.

Word Equation

Chemical Equation:

2. Write in words all the information that is conveyed by each of the following equations:

a) $C_4H_{8\,(g)} + 6O_{2\,(g)} \rightarrow 4CO_{2\,(g)} + 4H_2O_{(l)}$

____

____

____

____

Draw a sketch to represent this reaction:

b) $Fe_2O_{3(s)} + 3CO_{(g)} \rightarrow 2Fe_{(s)} + 3CO_{2\,(g)}$

____

____

____

____

Draw a sketch to represent this reaction:

3. Balance the following chemical reactions. Show all your working. Remember you can only change coefficients (subscripts must remain the same).

$\square$ Ca + $\square$ O_2 → $\square$ CaO

Ca = Ca =

O = O =

$\square$ N_2 + $\square$ H_2 → $\square$ NH_3

N = N =

H = H =

$\square$ Cu_2O + $\square$ C → $\square$ Cu + $\square$ CO_2

Cu = Cu =

O = O =

C = C =

$\square$ H_2O_2 → $\square$ H_2O + $\square$ O_2

H = H =

O = O =

4. Balance the following equations – show your working (h. is a challenge!).

a.	Na	+	MgF_2	→	NaF	+	Mg
b.	Mg	+	HCl	→	$MgCl_2$	+	H_2
c.	Cl_2	+	KI	→	KCl	+	I_2
d.	NaCl			→	Na	+	Cl_2
e.	Na	+	O_2	→	Na_2O		
f.	Na	+	HCl	→	H_2	+	NaCl
g.	K	+	Cl_2	→	KCl		

EXTENSION QUESTION:

h.	C_2H_6	+	O_2	→	CO_2	+	H_2O

WORKSHEET 14

CHEMICAL EQUATIONS

Syllabus reference 8.2.3

1 Write chemical equations to convey the same information as the following statements.

a A molecule of phosphorus which contains four atoms reacts with three molecules of oxygen (each containing two atoms) to form one molecule of a compound containing four phosphorus atoms and six oxygen atoms.

b A molecule of acetylene (ethyne) contains two carbon atoms and two hydrogen atoms. Oxygen gas exists as diatomic molecules. Two molecules of acetylene react with five molecules of oxygen gas to form four molecules of carbon dioxide and two molecules of water.

2 Write in words all the information contained in the following chemical equations.

a $CH_4 + 2Cl_2 \rightarrow CH_2Cl_2 + 2HCl$

(CH_4 is called methane, CH_2Cl_2 is dichloromethane and HCl is hydrogen chloride.)

b $2H_2SO_4 + S \rightarrow 3SO_2 + 2H_2O$

(H_2SO_4 is sulfuric acid and SO_2 is sulfur dioxide.)

3 Balance the following equations.

a $Mg + O_2 \rightarrow MgO$

b $Zn + HCl \rightarrow ZnCl_2 + H_2$

c $Zn + AgNO_3 \rightarrow Zn(NO_3)_2 + Ag$

d $AlCl_3 + NaOH \rightarrow Al(OH)_3 + NaCl$

e $Al(OH)_3 + HCl \rightarrow AlCl_3 + H_2O$

f $Pb(NO_3)_2 + KI \rightarrow PbI_2 + KNO_3$

g $C + H_2SO_4 \rightarrow CO_2 + SO_2 + H_2O$

4 Write balanced chemical equations for the following reactions.

a burning carbon in limited oxygen gas to form carbon monoxide

b heating silver carbonate to form silver oxide and carbon dioxide

c reacting magnesium with hydrochloric acid to form magnesium chloride and hydrogen gas

d reacting sodium hydroxide with sulfuric acid to form sodium sulfate and water

e reacting silver nitrate with potassium bromide to form potassium nitrate and silver bromide

64 Revision – Valency Or Combining Power

The **valency** of an element refers to its **combining power** and it can be used to predict the formulas of chemical compounds. Valency/combining power of an element depends on the number of electrons in the outer shell of the atoms of that element. To write formulas you need to understand about the valency/combining power of elements.

The valency/combining power of elements is **related to the electron configuration of their atoms** which you have already studied, and the position of elements in the periodic table can help you to determine the valency of many elements.

Notice that valency is just a number – it is **not positive or negative.** An ion can be positive or negative, but not valency. For example, the sodium ion Na^+ has a positive charge, but the valency of sodium is said to be 1 (not +1).

Valency and groups of the periodic table

The valency is the **same for all the elements in any group**. For example, all the group 1 metals (alkali metals) have a valency of 1. This is because they all have one electron in the outer shell of their atoms. They tend to lose this one electron, to form charged particles with stable outer valence electron shells. A charged particle formed when an atom gains or loses one or more electrons is called an ion.

Valency and periods of the periodic table

Across a period, the valency of elements **rises to a maximum and then decreases**, as shown in Table 64.1.

For groups 1 to 4 (14), the valency is the same as the group number.

For groups 5 (15) and 6 (16), the valency is either the group number or 8 minus the group number.

For group 7 (17) the valency is 1.

For group 8 (18) the valency is 0.

The pattern of valencies across a period can be explained by the atomic structure of the elements. The important thing is how many electrons each atom can give up, accept or share in order to end up with a stable outer valence shell of 4 pairs of electrons.

Table 64.1 Valency of elements in period 3.

Group	1	2	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	8 (18)
Symbol	Na	Mg	Al	Si	P	S	Cl	Ar
Common valency	1	2	3	4	3 or 5	2 or 6	1	0
Example of compound	NaCl	$MgCl_2$	$AlCl_3$	SiH_4	PCl_3 and PCl_5	H_2S and SO_3	HCl	None

Using the periodic table, it is easy to work out the valency for elements in groups 1 to 8 (18). For the other elements, you will just have to learn their valencies. Table 64.2 summarises the valencies and symbols of the elements you will use most often. Notice that the most common valency is 2 and some transition elements have more than one possible valency. Be sure to revise these valencies thoroughly. You need to know them in order to write formulas.

Table 64.2 Valencies (combining powers) of some elements.

Valency 1	Valency 2	Valency 3	Valency 4
Metals	**Metals**	**Metals**	**Metals**
Sodium Na	Beryllium Be	Boron B	Tin Sn
Lithium Li	Magnesium Mg	Aluminium Al	Lead Pb
Potassium K	Calcium Ca	Iron Fe	
Copper Cu	Strontium Sr	Chromium Cr	
Silver Ag	Barium Ba		
Mercury Hg	Manganese Mn		
	Cobalt Co		
	Nickel Ni		
	Copper Cu		
	Zinc Zn		
	Tin Sn		
	Mercury Hg		
	Lead Pb		
	Iron Fe		
Non-metals	**Non-metals**	**Non-metals**	**Non-metals**
Fluorine F	Oxygen O	Nitrogen N	Carbon C
Chlorine Cl	Sulfur S	Phosphorus P	**Semi-metals**
Bromine Br			Silicon Si
Iodine I			
Hydrogen H			

Elements in groups 1, 2 and 3 are metals which form compounds by losing the electrons in the outer electron shell of their atoms. For example, magnesium in group 2 has 2 electrons in the outer electron shell of its atoms. Its atoms can lose these 2 electrons, so magnesium has a valency of 2.

Elements in group 4 (14) have a valency of 4. In order to achieve a stable, outer, valence shell structure, carbon and silicon share up to four electrons, whereas tin and lead can lose up to four electrons.

Elements in groups 5 (15) to 7 (17) form compounds by either sharing or receiving electrons to stabilise outer electron shells of their atoms. Group 5 (15) elements have a valency of 3 or 5. Group 6 (16) elements have a valency of 6 or 2. Group 7 (17) elements have a valency of 1.

Elements in group 8 (18) have stable outer shells with four pairs of electrons (an octet), so they do not lose, accept or share electrons and have a valency of zero.

The **transition metals**, can show more than one valency, although the most common valency is 2. For these elements, the valency is included, using Roman numerals, in the names of their compounds. For example, there are two oxides of copper which are named to show the valency of copper in each – copper(I) oxide (Cu_2O) and copper(II) oxide (CuO).

Using valency

The formula of a compound can be predicted if the valencies of the elements or ions involved are known. For example, Table 64.3 shows the valencies of elements in period 3 and the resulting formulas of their oxides.

Table 64.3 Formulas of oxides of period 3.

Group	Valency of element combining with oxygen (valency 2)	Predicted formula of oxide
1	1	Na_2O
2	2	MgO
3 (13)	3	Al_2O_3
4 (14)	4	SiO_2
5 (15)	3 or 5	P_2O_3 or P_2O_5
7 (17)	1	Cl_2O
8 (18)	0	nil

QUESTIONS

1. (a) What is meant by the 'valency' of an element?
   (b) Copy and complete the following table to summarise the valency of elements in different groups of the periodic table.

Group	Number of electrons lost, gained or shared	Combining power or valency
1	1 lost	1
2		
3 (13)		
4 (14)	4 shared or lost	
5 (15)	3 gained or shared	
6 (16)		
7 (17)		
8 (18)		

2. Account for the valency of group 8 (18) elements being zero.
3. Describe the trend in the valency of the elements across period 3 of the periodic table.
4. (a) What statement can be made about the valency of the metals in group 2 of the periodic table?
   (b) Relate the valency of group 2 metals to their electron configurations.
5. (a) Identify two groups in the periodic table which have a common valency of 1.
   (b) In terms of electrons, explain why both sodium and chlorine have a valency of 1, as shown in the compound sodium chloride, NaCl.
6. Consider the metal barium, element 56 in the periodic table.
   (a) Identify the group to which barium belongs.
   (b) How many electrons does barium have in the outer electron shell of its atoms?
   (c) What is the valency of barium?
   (d) Deduce the charge on an ion of barium.
7. (a) In a compound of formula XH_3, determine the valency of element X.
   (b) In a compound of formula YO, what is the valency of element Y?
   (c) Given that element Z forms a chloride with formula ZCl_3 and element Q forms an oxide with formula Q_2O_3, what is the formula of the compound formed by the elements Z and Q?
8. (a) Explain why group 6 (16) elements can have a valency of either 2 or 6.
   (b) State valencies for the following elements.
      (i) Sodium.
      (ii) Aluminium
      (iii) Calcium.
      (iv) Magnesium.
9. Name the compounds which have the following formulas.
   (a) Cs_2S (b) $ZnCl_2$
   (c) MnO (d) $FeBr_2$
   (e) KCl (f) AlI_3
   (g) Cu_2O (h) CuO
10. Check your knowledge with this quick quiz.
   (a) The combining power of an element is called its ....
   (b) The valency is the same for all the elements in a (group/period) of the periodic table.
   (c) Across a period of the periodic table, the valency of the elements at first ...., reaches a maximum and then ....
   (d) Name the group of elements which have a valency of 0.
   (e) Identify two groups of the periodic table containing elements with a valency of 2.

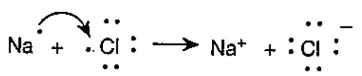
65 Revision – Writing and Naming Formulas

Stoichiometry involves calculations, and many chemistry calculations are based on formulas, so it is important that you are able to write formulas correctly. This chapter provides an opportunity for you to practise this skill.

Atoms of two or more different elements combine together to form compounds. The formula of a compound tells us what it is made of – the elements present and the ratio in which they are combined together.

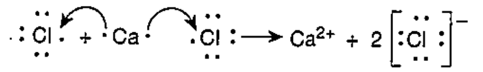
Writing formulas is easy if you know the symbols and valencies of elements thoroughly. Learning them may be boring, but it will pay off. Remember, the number of electrons in the outer shell of atoms determines the valency of elements and therefore the formulas of compounds. Here are some examples.

Sodium chloride has the formula NaCl. Sodium and chloride ions combine in the ratio of 1 : 1 to make the compound sodium chloride. This happens because Na and Cl each have a valency of 1.

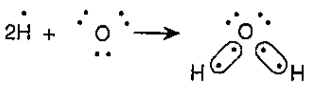


The sodium and chloride ions formed are attracted because of their opposite charges, forming the ionic compound sodium chloride.

Calcium chloride. When calcium joins with chlorine, the calcium atom has two electrons to give up (valency 2). A chloride atom only needs one electron to fill its outer shell (valency 1), making it stable. So calcium gives its second electron to another chlorine atom. This means that every calcium ion can join with two chloride ions. Ca^{2+} and Cl^- ions combine in the ratio 1 : 2 and the formula of calcium chloride is $CaCl_2$.



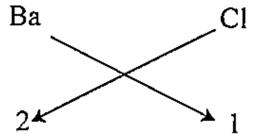
Formulas work the same way whether the electrons are transferred or shared. Oxygen needs two electrons (valency 2) so when it forms water, it combines with two hydrogens (valency 1), sharing one electron from each so that each atom gets a stable valence shell.



Crossing over. To check that you have worked out a formula correctly you can use the **'cross over' method**. For example, to write the formula for barium chloride:

- Write the symbols of the reacting elements.
- Two lines below, write the valency of each element and then draw diagonal arrows as shown.

  Draw arrows diagonally:



This gives you the formula $BaCl_2$.

Polyatomic ions. If the formula contains a polyatomic ion, and you need more than one of these ions in the formula, you must put the ion in brackets. For example, in magnesium hydroxide, there are magnesium ions (valency 2) and hydroxide ions (valency 1). Magnesium ions join with hydroxide ions in the ratio 1 : 2. We write this as $Mg(OH)_2$. It would be wrong to write $MgOH_2$ as this would indicate two hydrogen atoms rather than two OH groups. Similarly, barium nitrate would be written as $Ba(NO_3)_2$.

Naming compounds

There are rules to help you name compounds.

Rule 1 – If the **compound is made from a metal and a non-metal**, then:

- The first name of the compound is the name of the metal, e.g. NaCl is sodium chloride (not chlorine sodium).
- The ending of the non-metal is changed from … ine to … ide, e.g. NaCl is sodium chloride (not sodium chlorine). Other examples are:

  Oxygen becomes oxide.

  Sulfur becomes sulfide.

  Fluorine becomes fluoride.

  Hydrogen becomes hydride.

  Nitrogen becomes nitride.

Rule 2 – If the **compound is made from two non-metals**, then:

- The one furthest to the left in the periodic table is written first, e.g. SiO_2 is silicon dioxide (silicon is from group 4 and oxygen is from group 6). If they are in the same group, put the lower one first, e.g. SO_2 is sulfur dioxide.
- The second non-metal has its ending changed to –ide, e.g. HCl is hydrogen chloride.
- A prefix is used to indicate the number of atoms of each element in the molecule, e.g. H_2S is called dihydrogen sulfide. Mono can be used to indicate one atom present, but is often omitted. A prefix must be used if there could be ambiguity, e.g. carbon oxide could be carbon monoxide CO or carbon dioxide CO_2.

Rule 3 – Transition metals. Some transition metals have more than one possible valency. So the name of the compound includes the valency in brackets, e.g. iron can have a valency of 2 or 3 – so FeO is iron(II) oxide; Fe_2O_3 is iron(III) oxide. Other examples are lead and tin which can have valencies of 2 or 4 and copper with valencies of 1 or 2.

Rule 4 – You have to learn the formulas and the valencies of polyatomic ions. The most commonly used ones are in Table 65.1 below. Learn them well. You must be able to think – sulfate ... SO_4^{2-}... valency 2.

Table 65.1 Formulas and valencies of some polyatomic ions.

Valency 1	Valency 2	Valency 3
Hydroxide ion OH^- Nitrate ion NO_3^- Acetate ion $CHCOO^-$ Ammonium ion NH_4^+ Hydrogen carbonate HCO_3^-	Carbonate ion CO_3^{2-} Sulfate ion SO_4^{2-}	Phosphate ion PO_4^{3-}

To use these, just treat them as if each polyatomic ion is a single particle. Here are three examples.

Calcium carbonate. Calcium with a valency of 2, joins with a carbonate ion (valency of 2).

They have the same valency, the formula will be $CaCO_3$.

Sodium sulfate. Two atoms of sodium (with valency 1) join with one sulfate ion (valency 2). The formula will be Na_2SO_4.

Calcium nitrate. Calcium has a valency of 2, but nitrate (NO_3^-) has a valency of only 1. You need two nitrates. The formula will be $Ca(NO_3)_2$. The nitrate has to be in a bracket because you need two of the whole nitrate unit. Without the bracket it would be $CaNO_{32}$ – obviously not correct!

Rule 5 – Some compounds go by **common names**. For example, H_2O is technically dihydrogen oxide, but we just call it water; NH_3 is nitrogen trihydride, but we call it ammonia.

QUESTIONS

1. Write names and formulas for compounds made from the following combinations of elements.
   (a) Silver and iodine.
   (b) Barium and oxygen.
   (c) Nickel and chlorine.
   (d) Zinc and fluorine.
   (e) Strontium and bromine.
   (f) Cobalt and nitrogen.
2. Write chemical names for the following compounds.
   (a) CS_2 (b) AgCl
   (c) PbS (d) $ZnCl_2$
   (e) HgO (f) CoF_2
   (g) N_2O_3 (h) H_2O
   (i) BeF_2
3. Write formulas for:
   (a) Copper(II) chloride
   (b) Diphosphorus trioxide.
   (c) Iron(II) sulfide.
   (d) Ammonia.
   (e) Ammonium chloride.
   (f) Aluminium nitride.
4. State names and valencies for each of the following polyatomic ions.
   (a) SO_4^{2-} (b) PO_4^{3-}
   (c) NO_3^- (d) NH_4^+
   (e) OH^- (f) HCO_3^-
5. The following formulas all contain polyatomic ions. Name the compounds.
   (a) NaOH
   (b) $BaSO_4$
   (c) K_2CO_3
   (d) $Mg(OH)_2$
   (e) $Ca(NO_3)_2$
   (f) $Al_2(SO_4)_3$
   (g) Li_2SO_4
   (h) $AlPO_4$
   (i) $Zn(OH)_2$
   (j) CH_3COONa
6. Determine the value of x in each of the following formulas of compounds.
   (a) H_xCO_3 (b) H_xPO_4
   (c) H_xSO_4 (d) H_xNO_3
   (e) H_xF
7. Copy and complete the following table to show the formulas of compounds that would be formed by all possible combinations of ions. Some are done to help get you started.

Ion	Cl^-	OH^-	NO_3^-	SO_4^{2-}	CO_3^{2-}	O^{2-}
Na^+						
K^+				K_2SO_4		
Mg^{2+}				$MgSO_4$		
Ca^{2+}						
Cu^+			$CuNO_3$			
Cu^{2+}			$Cu(NO_3)_2$			
Al^{3+}						

66 Writing and Balancing Equations

To write equations for a chemical reaction you need to know, or be able to work out:

- The **names of reactants and products** of the reaction.
- The **formulas of reactants and products**.
- Understand that the **law of conservation of mass** applies to all chemical reactions. You end up with the same number and types of atoms after a reaction as you started with and thus the mass stays the same.

Follow through these two examples of writing and balancing equations and then use the questions to practise doing it by yourself. If you have trouble, consult your teacher.

Example 1: Decomposing copper(II) carbonate to copper oxide and carbon dioxide

Step 1 – Word equation

Write the equation in words showing the reactants and products.

Copper(II) carbonate → copper(II) oxide + carbon dioxide

Step 2 – Formulas

Under each substance in the equation, write its formula if it is a compound or its symbol if it is an element.

Copper(II) carbonate	→	copper(II) oxide	+	carbon dioxide
$CuCO_3$	→	CuO	+	CO_2

Once you are sure you have all the formulas correct, do NOT change them again. You balance an equation by changing the numbers in front of the formulas – called the **coefficients** – not by changing the formulas. The coefficients represent the ratio of reacting moles – you will learn about moles soon.

Step 3 – Balance the equation

Count up how many of each type of atom you have on each side of the equation. Compare the number of each type of atom on the left of the arrow with the number on the right.

Copper carbonate	→	copper oxide	+	carbon dioxide
$CuCO_3$	→	CuO	+	CO_2
Cu – 1		Cu – 1		
C – 1				C – 1
O – 3		O – 1		O – 2

Check that you have the same numbers of each type of atom on each side of the equation. If so, your equation is complete – it is **balanced**.

Balancing an equation is working out the ratios in which substances combine during a reaction so that we end up with the same number of each sort of atom on either side of the arrow.

Step 4 – Add states

If you know the states of reactants and products, you can indicate this as follows.

(s) indicates a solid.

(g) indicates a gas.

(aq) indicates aqueous (a substance in water solution).

(l) indicates a pure liquid such as water or a concentrated acid.

Your balanced equation is:

$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

Example 2: Burning hydrogen

Step 1 – Word equation

You know that hydrogen burns in oxygen to form water.

Hydrogen + oxygen → water

Step 2 – Formulas

You might be tempted to write the following incorrect symbols in this reaction – think about why they are wrong.

$H + O \rightarrow H_2O$ or $2H + O \rightarrow H_2O$

These are not correct because hydrogen and oxygen can never exist as single atoms (H and O). They will always join up to form diatomic molecules H_2 and O_2. So the equation must be written as:

$H_2 + O_2 \rightarrow H_2O$

Step 3 – Balance the equation

You can see that the equation is not balanced; you started with two atoms of oxygen, but there is only one at the end (in the water). Writing H_2O_2 for water would balance the equation, but H_2O_2 is not the correct formula for water. H_2O_2 is hydrogen peroxide.

We have to balance the equation by changing the numbers in front of the reactants and/or products, not by changing formulas.

H_2	+	O_2	→	$2H_2O$
H – 2		O – 2		H – 4, O – 2

Now the oxygens are balanced, but not the hydrogens. We must have used twice as much hydrogen on the left, so we add a 2 in front of H_2.

$2H_2$	+	O_2	→	$2H_2O$
H – 4		O – 2		H – 4, O – 2

The equation is now balanced.

Step 4 – Add states of the reactants and products.

(s) – solid; (l) – liquid; (g) – gas; (aq) – dissolved in water.

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

QUESTIONS

1. (a) Recall the law of conservation of matter.
   (b) Are new atoms made in a chemical reaction? Explain.
2. Write word equations for the following reactions.
   (a) $2PbO(s) \rightarrow 2Pb(s) + O_2(g)$
   (b) $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
   (c) $H_2(g) + I_2(g) \rightarrow 2HI(g)$
   (d) $2Ag(s) + Cl_2(g) \rightarrow 2AgCl(s)$
   (e) $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$
   (f) $Cu_2S(s) + O_2(g) \rightarrow 2Cu(s) + SO_2(g)$
   (g) $4P(s) + 5O_2(g) \rightarrow 2P_2O_5(g)$
   (h) $Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
3. Balance the following equations. All the formulas are correct, you just have to add coefficients (numbers) in front of the formulas where they are needed.
   (a) $HgO(s) \rightarrow Hg(s) + O_2(g)$
   (b) $Fe(s) + O_2(g) \rightarrow FeO(s)$
   (c) $Na(s) + Cl_2(g) \rightarrow NaCl(s)$
   (d) $N_2(g) + O_2(g) \rightarrow NO(g)$
   (e) $H_2(g) + Cl_2(g) \rightarrow HCl(g)$
   (f) $NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$
4. Write equations for the following described reactions, showing the states of the reactants and products.
   (a) Calcium hydroxide solution (limewater) reacts with carbon dioxide gas to form solid calcium carbonate and liquid water.
   (b) Solid calcium carbonate decomposes when heated, forming a white solid called calcium oxide and carbon dioxide gas.
   (c) Solid zinc sulfate decomposes when heated, forming solid zinc oxide and sulfur trioxide gas.
   (d) Solid iron sulfide is formed when hot iron reacts with solid sulfur.
   (e) Hydrogen bromide is a gas which decomposes on heating, forming the gases hydrogen and bromine.
   (f) Sodium hydroxide reacts with dilute nitric acid to produce sodium nitrate solution and liquid water.
   (g) Hydrogen burns in air, combining with oxygen to form liquid water.
   (h) Aluminium reacts with hydrochloric acid to give aluminium chloride solution and hydrogen gas.
5. For each of the following reactions, write the equation in words and then write a balanced chemical equation.
   (a) Solid silver oxide decomposes when heated to form the metal silver and diatomic molecules of oxygen gas.
   (b) During the combustion of magnesium metal, the metal reacts with oxygen in the air and forms a white powder called magnesium oxide.
   (c) During the extraction of copper from its ores, solid copper(II) sulfide undergoes combustion and forms solid copper and the gas sulfur dioxide.
   (d) The metal aluminium slowly reacts with oxygen in air to form a thin layer of solid aluminium oxide.
   (e) Barium reacts with dilute hydrochloric acid forming hydrogen gas and a solution of barium chloride.
6. Explain why it is not correct to balance an equation by changing the formulas of reactants or products.
7. Write equations for the following reactions.
   (a) Calcium + oxygen → calcium oxide
   (b) Chlorine + oxygen → dichlorine monoxide
   (c) Magnesium + water → hydrogen + magnesium hydroxide
   (d) Sodium + water → hydrogen + sodium hydroxide
   (e) Magnesium + chlorine → magnesium chloride
8. Check your knowledge with this quick quiz.
   (a) What symbol would you use to indicate that a substance in an equation is dissolved in water?
   (b) State the ratio of atoms in sulfur dioxide.
   (c) State the first step in writing a symbolic equation.
   (d) In an equation, how would you write two molecules of gaseous carbon dioxide?
   (e) Is it acceptable to balance an equation by changing formulas of reactants or products?

End of Section: Chemical Reactions and Stoichiometry

Signed: ______________________ Resubmit? ☐ No ☐ Yes by due date:__________

Mole Concept

Inquiry question: How are measurements made in chemistry?

- **conduct a practical investigation to demonstrate and calculate the molar mass (mass of one mole) of:**
  - **an element**
  - **a compound (ACSCH046)**
- **explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)**
  - **moles of elements and compounds $n = \frac{m}{M}$ (*n* = chemical amount in moles, *m* = mass in grams, *MM* = molar mass in $gmol^{-1}$)**

Read through Conquering Chemistry Prelim '5.4 The Mole' pg 134

What standard measure is a mole of any substance measured against?

____

Define a 'mole': ____

____

From this concept, we get the Avogadro's Constant (N_A):

N_A = ____ atoms

Avogadro's constant states that there is 6.02×10^{23} particles per mole of any substance.

Explain why one mole of carbon dioxide will weigh more than one mole of pure carbon.

____

____

____

Sketch each molecule and provide a molar mass for each:

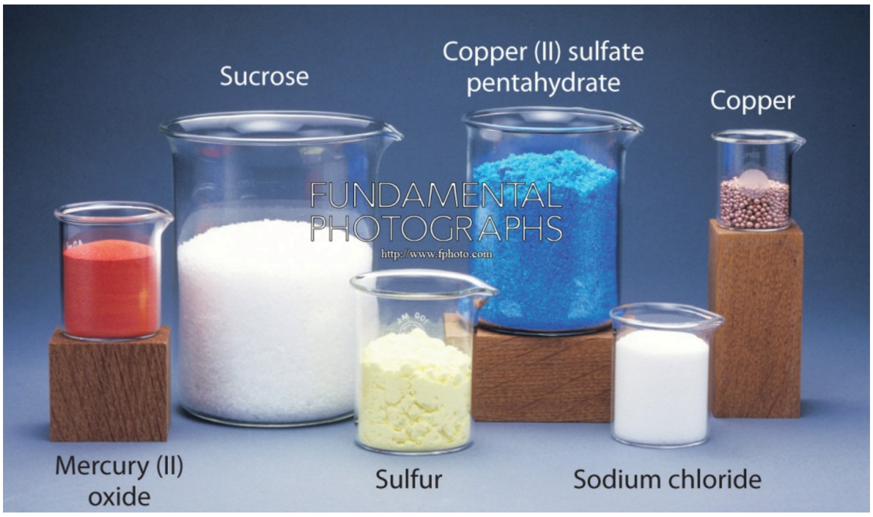
Carbon: ____________/mole

Carbon Dioxide: ____________/mole

Note: The mass of ONE MOLE of:

- element = atomic weight
- compound = molecular weight

To demonstrate, each sample in the image is one mole: That is each sample contains 6.02 x 10^{23} particles.



Depending on the particle size, atomic weight and packing density, each sample will take up a different volume. But <u>EVERY SAMPLE contains 6.02 x 10^{23} particles.</u>

Moles of Gases

Read through Conquering Chemistry Prelim '5.5 Moles of Gaseous Elements' pg 136

Fill in the missing words:

A mole of oxygen atoms is quite clearly _______g

A mole of oxygen molecules is _______ g, because the formula for oxygen gas is ______, as it contains ____ atoms of oxygen per molecule.

Complete the following examples:

For chlorine 1mol = _______g, whereas chlorine gas (_____) 1mol = _______g

For oxygen 1mol = _______g, whereas ozone (O_3) 1mol = _______g

For sulphur 1mol = _______g, whereas sulphur gas (S_8) 1mol = _______g

KEY CONCEPTS

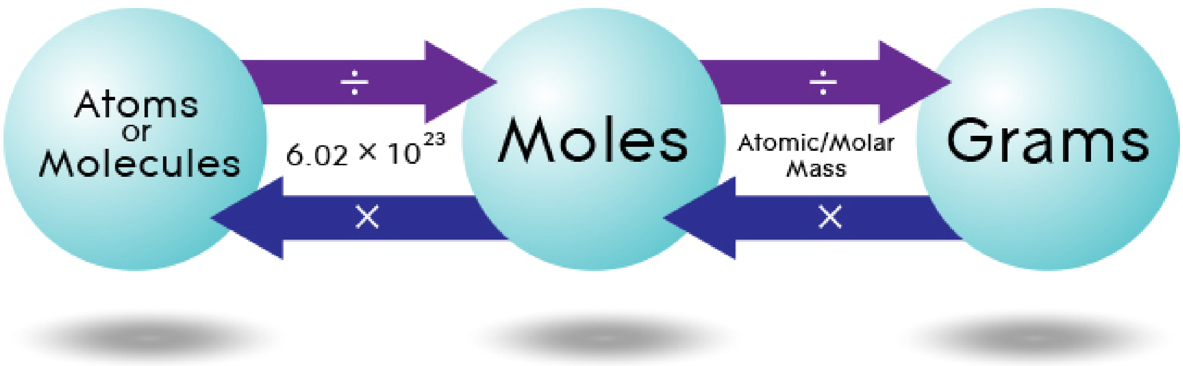
- A **mole** of a substance is the quantity that contains as many elementary units (e.g. atoms, ions or molecules) as there are atoms in exactly 12 g of the carbon-12 isotope.
- The **Avogadro constant** (symbol N_A) is the number of atoms in exactly 12 g of the carbon-12 isotope.
  $N_A = 6.022 \times 10^{23}$ particles per mole
- A mole of atoms of an element is the mass that in grams is numerically equal to the relative atomic mass but a mole of nitrogen gas is twice the relative atomic mass (because nitrogen gas is N_2).
- A mole of a compound is the mass that in grams is numerically equal to the relative molecular mass or relative formula mass.
- The molar mass is the mass of a mole of the substance. It can be used for both elements and compounds.

Converting Between Mass, Moles and Numbers of Particles

There are two ways of looking at the 'mole' as a concept.

You can think of the mole as:

- a set number , (6.02×10^{23}) particles of any substance
- a standard weight (either atomic or molecular mass) for a substance



Both of these are correct, and will be useful for different applications throughout the course.

Read through Conquering Chemistry Prelim '5.6 Mass to Moles' pg 136 and 'Moles to number of atoms or molecules' pg 136-137

The key equation for converting mass into the number of moles is

Number of moles = ____________________

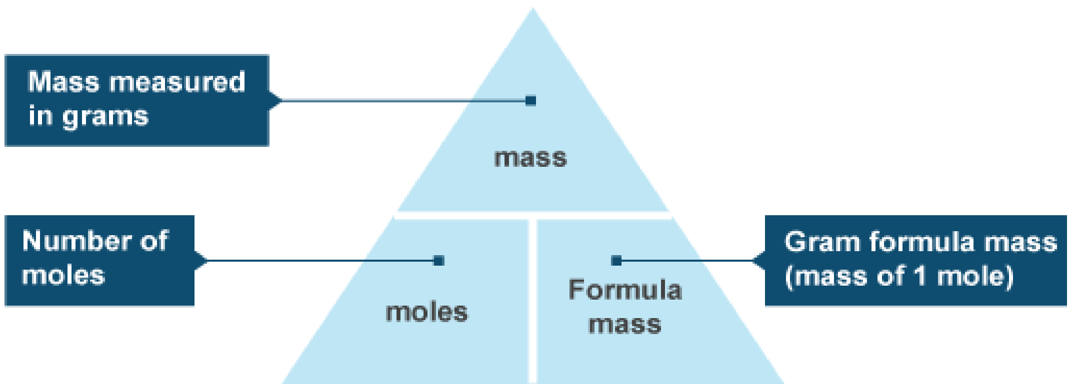
= ____________________

= ____________________

KEY EQUATION:

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

We use this as:



Exercises:

Complete the questions within Conquering Chemistry Prelim 'Exercises' pg 138

CHEMICAL CALCULATIONS

Syllabus reference 8.3.5

1 Calculate the relative molecular or formula mass of each of the following.

a sulfur dioxide (SO_2)

b potassium nitrate (KNO_3)

c zinc hydroxide ($Zn(OH)_2$)

d sulfuric acid (H_2SO_4)

e copper sulfate pentahydrate ($CuSO_4.5H_2O$)

2 Calculate the number of moles in each of the following.

a 3.01×10^{24} carbon atoms

b 1.05×10^{22} carbon monoxide molecules

c 1.204×10^{28} ethanoic acid molecules (CH_3COOH)

3 Work out the number of:

a atoms in 0.30 mole of nickel metal

b molecules in 4.0 moles of ethane (C_2H_6)

c formula units in 7.5 moles of potassium bromide (KBr)

4 How many moles are contained in each substance given below?

a 1.01 g of neon gas (Ne)

b 2.538 g of iodine molecules (I_2)

c 5.10 g aluminium oxide (Al_2O_3)

5 Calculate the mass of:

a 2.05 moles of rubidium atoms

b 1.5 moles of sulfuric acid molecules

c 1.0 mole of glucose ($C_6H_{12}O_6$)

Experiment: Investigation 7.2 – Mole Problems

This resource is available on NelsonNet 'Chemistry in Focus' pg 154-155

Aim: To solve problems involving quantities of element and compounds to develop an understanding of moles

Materials:

Chemicals:

- Sulphur
- Sodium chloride
- Water
- Sugar cube or sachet of sugar
- Aluminium foil
- Copper wire
- Magnesium ribbon

Method:

For each of the questions, devise a procedure for answering it. Fill in the template pages that follow for each inquiry.

1. How many atoms of sulphur are in a level teaspoon of sulphur?
2. How many sodium ions are in a level teaspoon of sodium chloride?
3. How many teaspoons of water are in one mole of water?
4. How many atoms of aluminium are in a 3cm2 of aluminium foil?
5. What length of copper wire corresponds to 1mol of copper
6. What length of magnesium tape contains 1mol of magnesium
7. How many atoms of carbon are there the sugar sample?

Before you start:

Fill in the table below as a starting point to summarise your information:

Name of sample	Element or Compound?	Atomic/Molecular Formula	Atomic/Molecular Mass	Particles in one mole
Sulfur				
Sodium Chloride				
Water				
Sugar Cube		$C_{12}H_{22}O_{11}$		6.02×10^{23} particles
Aluminium Foil				
Copper Wire				
Magnesium Ribbon				

Aim: ____

Materials:

________________________________ ________________________________

________________________________ ________________________________

________________________________ ________________________________

Method:

____

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____

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Results:

Discussion of Accuracy/Calculations:

____

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Conclusion:____

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Aim: ____

Materials:

____________________________________ ____

____________________________________ ____

____________________________________ ____

Method:

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Results:

Discussion of Accuracy/Calculations:

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Conclusion:____

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Aim: ____

Materials:

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Results:

Discussion of Accuracy/Calculations:

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Conclusion:____

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Aim: ____

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Discussion of Accuracy/Calculations:

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Conclusion:____

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Results:

Discussion of Accuracy/Calculations:

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Conclusion:____

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- **explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)**
  - **percentage composition calculations and empirical formulae**

Read through Conquering Chemistry Prelim '5.7 Per Cent Composition' pg 138-139

Fill in the missing words:

The chemical formula of a ________________tells us directly the ________________ in which the ________________ are present. In aluminium oxide, ________________, the Al and O atoms are in the ration ________________. Sometimes we need to know the ratio by the mass in which the elements are present, this called the ________________ composition by ________________.

Knowing the per cent composition, an ________________ can determine how much ________________ can be extracted from a tonne of ________________ oxide, or a farmer can decide which of ________________ fertilisers contains the most ________________ per ________________ of a compound.

Write out example 5 to show how to calculate the per cent composition:

____

____

____

____

____

____

$$\%\boldsymbol{A\ in\ a\ compound} = \frac{\text{mass of A in 1 mole of the compound}}{\text{mass of one mole of a compound}} \text{ x } 100$$

EXERCISES:

Complete the questions within Conquering Chemistry Prelim 'Exercises' pg 139-140

72 Empirical and Molecular Formulas

The formula of a substance tells us about its composition. However, there are two types of formulas – empirical and molecular formulas – and they tell us different information about the substance they are representing.

Empirical formula

An empirical formula represents the **simplest ratio of atoms or ions in a compound**. For example, the formula of the ionic compound sodium chloride is NaCl and this tells us that sodium chloride is made up of an ionic lattice composed of sodium ions and chloride ions **in the ratio 1 : 1**. An empirical formula does **not** tell you how many ions are present – there is no sodium chloride unit made of just one sodium ion and one chloride ion.

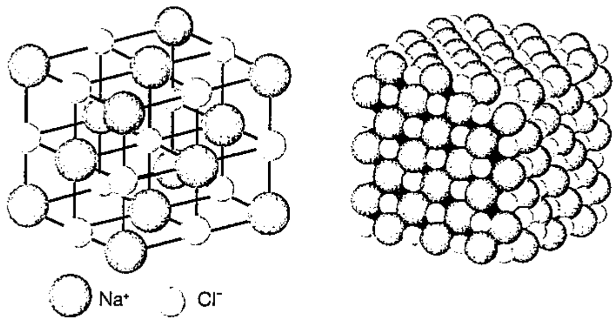

Figure 72.1 A crystal of sodium chloride represented by the formula NaCl.

Ionic compounds can only have an empirical formula. Ionic compounds cannot have a molecular formula as they do not exist as molecules.

Also giant covalent structures, such as silicon dioxide (sand) can only have an empirical formula.

Molecular formula

A molecular formula tells you the **number of atoms in a molecule**. For example, the formula of the covalent compound ethanol is C_2H_5OH. This formula tells us that ethanol exists as molecules which each contain two atoms of carbon combined with six atoms of hydrogen and one atom of oxygen.

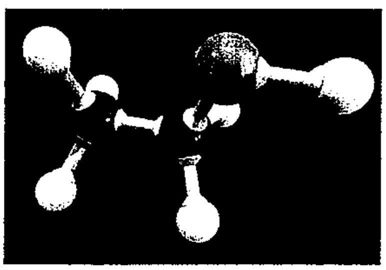

Figure 72.2 A molecule of ethanol represented by the formula C_2H_5OH.

Notice that **covalent compounds can have both a molecular and an empirical formula.** For example, ethane has a molecular formula of C_2H_6 and an empirical formula of CH_3. Its molecules each consist of two carbon atoms and six hydrogen atoms, and the simplest ratio of its atoms is C : H = 1 : 3 as shown in the empirical formula. Notice that the molecular formula is a multiple of the empirical formula (×2).

Sometimes the molecular and empirical formulas are the same, for example in water H_2O, carbon dioxide CO_2 and methane CH_4. In these cases the ratio of atoms is as simple as possible.

Formulas and percentage composition

The empirical formula is important in analysis. It is used to determine the percentage composition of compounds. Conversely if the percentage composition of a compound is known, this can be used to determine its formula.

Calculating percentage composition from empirical formula

Example: Calculate the percentage composition of potassium oxide, which has an empirical formula of K_2O.

Answer:

- Calculate the molar mass.
  MM of $K_2O = 2 \times 39.2 + 16.0 = 94.2$ grams.
- Calculate the percentage of each element in the molar mass.
  $\% K = \frac{78.2}{94.2} \times 100 = 83.0\%$
  $\% O = \frac{16.0}{94.2} \times 100 = 17.0\%$

Calculating empirical formula from percentage composition

This is the reverse of the above process.

Example: Find the empirical formula of a compound which consists of 83.0% potassium and 17.0% oxygen.

Answer:

- Find the ratio of atoms of the elements present by dividing the percentage of each element by its atomic mass.
  K = 83% and $\frac{83}{39.1} = 2.12$
  O = 17% and $\frac{17}{16.0} = 1.06$
  The ratio of K : O in potassium oxide is 2.12 : 1.06
- The ratio of atoms must be in whole numbers.
  So a mathematical technique used is to divide by the smaller number (in this case 1.06).
  2.12 : 1.06 is the same as $\frac{2.16}{1.06} : \frac{1.06}{1.06} = 2 : 1$
  So, the ratio of elements present is K : O = 2 : 1

And the empirical formula is K_2O_1 which is written as K_2O.

Calculating molecular formulas

For covalent compounds, you can use the empirical formula plus the molar mass to determine the molecular formula.

Example: The percentage composition of a hydrocarbon (a compound made of hydrogen and carbon only) is 80% carbon and 20% hydrogen.

(a) Find its empirical formula.

(b) Calculate its molecular formula if the molar mass of this compound is 30 grams.

Answer:

(a) $C : H = \frac{80}{12.01} : \frac{20}{1.008}$
$= 1 : 3$

So the empirical formula is CH_3.

(b) Mass of empirical formula CH_3 = 15.018 g

The molar mass = 30

So molecular formula = 2 × empirical formula.

Molecular formula is C_2H_6.

(The compound is ethane.)

QUESTIONS

1. Distinguish between an empirical and a molecular formula.
2. (a) Explain why magnesium oxide has an empirical formula but not a molecular formula.
   (b) A student found this statement in the opening chapter of a chemistry textbook, 'The molecular formula of sodium carbonate is Na_2CO_3.' Discuss this statement.
3. Copy and complete the table to show molecular and empirical formulas for some covalent compounds.

Name of compound	Molecular formula	Empirical formula
Carbon dioxide		
Butane	C_4H_{10}	
Ethane	C_2H_6	
Water		

4. The empirical formula of a covalent compound called benzene is CH and its molar mass is 78 grams. Determine the molecular formula of benzene.
5. Calculate the percentage by mass of the hydrogen and oxygen that make up water.
6. The percentage composition of the hydrocarbon called ethene is 85.7% carbo and 14.3% hydrogen.
   (a) Determine the empirical formula of ethene.
   (b) It the molar mass of ethene is 28 grams, determine its molecular formula.
7. A sample of barium metal with a mass of 1.00 g is heated in pure oxygen gas. The mass of the oxide produced is 1.117 g. Find the empirical formula of this oxide.
8. In an experiment, zinc is heated in a stream of pure oxygen. 3.27 g zinc combines with 0:80 g oxygen. Find the empirical formula of the oxide produced.
9. A sample of a compound is analysed and found to contain 58.84 g barium, 13.74 g sulfur and 27.43 g oxygen. Determine its empirical formula.
10. Determine the name and empirical formula of compounds with the following percentage compositions.
    (a) 36.5% sodium, 25.4% sulfur and 38.1% oxygen.
    (b) 26.2% nitrogen, 7.5% hydrogen and 66.3% chlorine.
11. A compound is made by heating together cobalt and sulfur. A 100 g sample of the compound contains 55.06 grams of cobalt.
    (a) What is the percentage composition of the compound?
    (b) Determine the empirical formula and name of the compound.
12. Find the percentage composition of the ionic compound potassium nitrate (KNO_3).
13. Describe the experiment you carried out to determine the empirical formula of an ionic compound.
14. In an experiment, the complete combustion of 10.00 grams of magnesium was carried out. The mass of the product formed was found to be 16.58 grams.
    (a) Write a word equation for this chemical reaction.
    (b) Calculate the mass of oxygen used by the combustion of the magnesium.
    (c) Determine the empirical formula of the product.
    (d) Does this product have a molecular formula? Explain why or why not.
15. Check your knowledge with this quick quiz.
    (a) The simplest ratio of atoms in a compound is shown by a(n) (empirical/molecular) formula.
    (b) The number of atoms in a molecule of a compound is shown by a(n) (empirical/ molecular) formula.
    (c) An ionic compound can only have a(n) (empirical/molecular) formula.
    (d) For covalent compounds, the molecular formula is the same or a .... of its empirical formula.

6 Calculate the percentage of:

a calcium in calcium hydroxide ($Ca(OH)_2$)

b oxygen in aluminium carbonate ($Al_2(CO_3)_3$)

7 Calculate the percentage of each element in:

a nickel(II) sulfate ($NiSO_4$)

b phosphoric acid (H_3PO_4)

8 In an experiment to find the formula of a compound the following results were obtained:

Mass of crucible = 35.03 g
Mass of crucible + copper = 38.205 g
Mass of crucible + copper oxide = 39.005 g

Find the empirical formula of the compound.

9 In an experiment 0.98 g of magnesium combined with chlorine gas to form 3.8 g of magnesium chloride. Show that these results support an empirical formula of $MgCl_2$ for this compound.

- **conduct an investigation to determine that chemicals react in simple whole number ratios by moles**

EXPERIMENT: DETERMINING THE EMPIRICAL FORMULA OF MAGNESIUM OXIDE

Syllabus reference 8.3.5

INTRODUCTION

The empirical formula of a substance is the simplest whole number ratio of the number of atoms of each element in the compound. This can be calculated knowing the mass of each element and using this to calculate the number of moles of each element.

In this experiment you will carry out the reaction:

$$\text{magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$$

By determining the mass of each element and using the relationship:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}}$$

The number of moles of each element can be determined and from this the empirical formula obtained.

AIM

....

EQUIPMENT

- 1 strip of magnesium ribbon (about 20 cm)
- crucible and lid
- pipe clay triangle
- tripod
- matches
- steel wool
- crucible tongs
- Bunsen burner
- balance

SAFETY: Safety glasses must be worn at all times. Burning magnesium emits a bright light: do not look directly at it.

PROCEDURE

1 Heat the crucible and lid strongly for 5 min over a hot Bunsen flame. Allow to cool.

2 Weigh the crucible and lid.

3 Thoroughly clean the surface of the magnesium ribbon with steel wool.

4 Record the appearance of the cleaned magnesium.

5 Coil the piece of magnesium ribbon so that it fits inside the crucible. Weigh the crucible, lid and magnesium ribbon.

6 Place a pipe clay triangle over the Bunsen burner and carefully heat the crucible without the lid until the magnesium begins to glow then put the lid on with tongs before it bursts into flames.

WARNING: Do not look directly at the burning magnesium.

7 Once the lid is in place heat strongly for about 10 mins lifting the lid occasionally to admit oxygen. Try to prevent any magnesium oxide smoke escaping.

8 Remove the lid and heat for a further 5 min to ensure complete reaction.

9 Replace the lid and allow to cool. Reweigh the crucible, lid and magnesium oxide.

Results:

	Description	Calculate by...	Measurement
A	Mass of Empty crucible and lid		
B	Mass of crucible, lid and magnesium		
C	Mass of magnesium		
D	Mass of crucible, lid and magnesium oxide		
E	Mass of magnesium oxide		
F	Mass of oxygen		

Discussion Questions:

1. Calculate the number of moles of magnesium and oxygen:

____

____

____

____

2. What is the ratio of moles of magnesium to moles of oxygen ___________________
3. Based on your results, what is the empirical formula of magnesium oxide?

____

____

4. Compare the class results. Are your results reliable?

____

____

5. The actual formula of magnesium oxide is MgO. The ratio is 1:1. How does this compare with your results?

____

____

6. What are the likely sources of error?

____

____

____

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Conclusion:____

____

- **explore the concept of the mole and relate this to Avogadro's constant to describe, calculate and manipulate masses, chemical amounts and numbers of particles in: (ACSCH007, ACSCH039)**
  - **limiting reagent reactions**

75 Limiting Reagent Reactions

If reactants are present in stoichiometric proportions, then all of each reactant will be used up.

However, if there is not enough of one reactant, to use up all of the other reactant, then the one that runs out is called the **limiting reagent.**

The limiting reagent **limits the reaction.**

The **other reactant** will be present in **excess** and some will be left over when the reaction stops.

For example, in the combustion of hydrogen:

$2H_2(g)$	+	$O_2(g)$	$\rightarrow$	$2H_2O(l)$
2 moles		1 mole		2 moles
4 grams		32 grams		36 grams

If only 0.5 mole of oxygen is available, with 2 moles of hydrogen gas, then oxygen will be a limiting reagent. When that oxygen is used up, the reaction will stop. There will still be 1 mole of hydrogen left unreacted, as hydrogen was present in excess.

Here is an example of a typical problem involving a limiting reagent.

Example: Iron(II) sulfide can be synthesised by heating a mixture of iron and sulfur in the absence of air.

(a) Write an equation for this reaction.

(b) 3.0 grams of iron is mixed with 1.4 grams of sulfur. Calculate the number of moles of each reactant present in the mixture.

(c) Are these reacting elements present in stoichiometric proportions? Explain.

(d) Identify which element is present in excess and which will limit the reaction.

(e) How much sulfur would be needed to use up all of the iron?

Answer:

(a) $Fe(s) + S(s) \rightarrow FeS(s)$

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

Mol Fe $= \frac{3.0}{55.8} = 0.05376$ moles.

Mol S $= \frac{1.4}{32.1} = 0.04361$ moles.

(b) Fe : S reacting = 0.05376 : 0.04361 which is a ratio of approximately 5 : 4.

(c) The reacting elements are not present in stoichiometric proportions.

If they were in stoichiometric proportions, Fe : S would be 1 : 1.

(d) Iron is present in excess and sulfur is the limiting reactant. When the sulfur has been used up, there will still be some iron left unreacted.

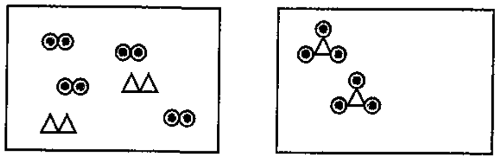
(e) There is 0.05376 mol of Fe, which could react with 0.05376 mol of S.

Mass of 0.05376 mol of S $= n \times MM$
$= 0.05376 \times 32.1$
$= 1.7$ grams of sulfur.

QUESTIONS

1. What is meant by a limiting reagent in a chemical reaction?
2. Does the following diagram accurately represent the reaction shown by the equation? Explain.

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



3. Magnesium reacts with water as shown by the equation:
$Mg(s) + 2H_2O(l) \rightarrow H_2(g) + Mg(OH)_2(aq)$
0.8 mol of hot water is added to 0.3 mol of magnesium.
   (a) Identify the limiting reagent and justify your choice.
   (b) How many moles of hydrogen gas will be produced by this reaction?
4. 3.29 grams of zinc reacts with 2.74 grams of hydrochloric acid. The equation for this reaction is:
$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
   (a) Calculate the moles of reactants available.
   (b) Identify the reactant that is the limiting reagent and justify your answer.
   (c) Calculate the mass of zinc chloride produced.
5. 1.46 grams of hydrochloric acid is added to 1.70 grams of magnesium carbonate and they react as shown by the equation.
$MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$
   (a) Calculate the moles of reactants used.
   (b) Is there a limiting reagent? Justify your answer.
6. Manganese dioxide reacts with concentrated hydrochloric acid as shown in the equation.
$MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O(l)$
If 2.74 g of manganese dioxide is mixed with 2.17 g of hydrochloric acid, determine the following.
   (a) Which reactant is the limiting agent?
   (b) What mass of chlorine will be produced?
   (c) What mass of excess reagent will remain after the reaction ceases?

CALCULATING REACTING QUANTITIES

Syllabus reference 8.3.5

1 The equation for the combustion of butane is:

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

Calculate the number of moles of:

a CO_2 produced in the combustion of 1 mole of C_4H_{10}

b H_2O produced in the combustion of 3 moles of C_4H_{10}

c O_2 consumed in the combustion of 0.6 moles of C_4H_{10}

2 How many moles of oxygen gas would be obtained if:

a 0.80 moles of potassium chlorate ($KClO_3$) is decomposed by heating into potassium chloride and oxygen?

b 0.24 moles of mercury(II) oxide is decomposed by heating into mercury and oxygen gas?

3 0.30 moles of magnesium oxide is completely dissolved in nitric acid (HNO_3) to form a solution of magnesium nitrate. Calculate the number of moles and masses of:

a nitric acid required

b magnesium nitrate formed

4 Excess sulfuric acid was added to 3.00 g of calcium carbonate ($CaCO_3$) to form calcium sulfate, carbon dioxide and water. Calculate the masses of:

a sulfuric acid consumed
b carbon dioxide produced
c calcium sulfate formed

5 Dilute hydrochloric acid was added to zinc until no further reaction occurred.

a Write a balanced equation for the reaction.
b If 0.012 mole of zinc chloride was formed what mass of zinc reacted?
c What mass of zinc chloride would be produced if 0.040 mole of hydrochloric acid reacted with excess zinc?

6 What volume of oxygen is required for the complete combustion of 125 mL of butene gas (C_4H_8) and what volume of carbon dioxide is formed? (All volumes are measured at the same temperature and pressure.)

7 One method of extracting nickel from concentrated ores is to mix finely divided ore with carbon monoxide. Nickel reacts to form gaseous nickel carbonyl. When excess nickel was reacted with 500 mL CO, 125 mL nickel carbonyl was formed at the same temperature and pressure. Assuming all the CO reacted, what is the formula of nickel carbonyl? Explain how you deduced this.

8 Butane is used as a fuel in camping stoves. The equation for the complete combustion of butane is:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

100 mL of butane is burnt in an excess of oxygen. Calculate:

a the volume of oxygen consumed

b the volume of carbon dioxide produced.

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

9 A silver coin was analysed and found to contain only silver and copper. 1.580 g of this coin was dissolved in concentrated nitric acid and the resultant solution diluted. Reaction of the solution with excess hydrochloric acid yielded 1.050 g of silver chloride. Calculate the percentage of silver in the sample.

10 A sample of iron wire reacts with oxygen to form iron(III) oxide. If 3.58 g of the wire yields 5.00 g of the iron(III) oxide, calculate the percentage purity of the iron.

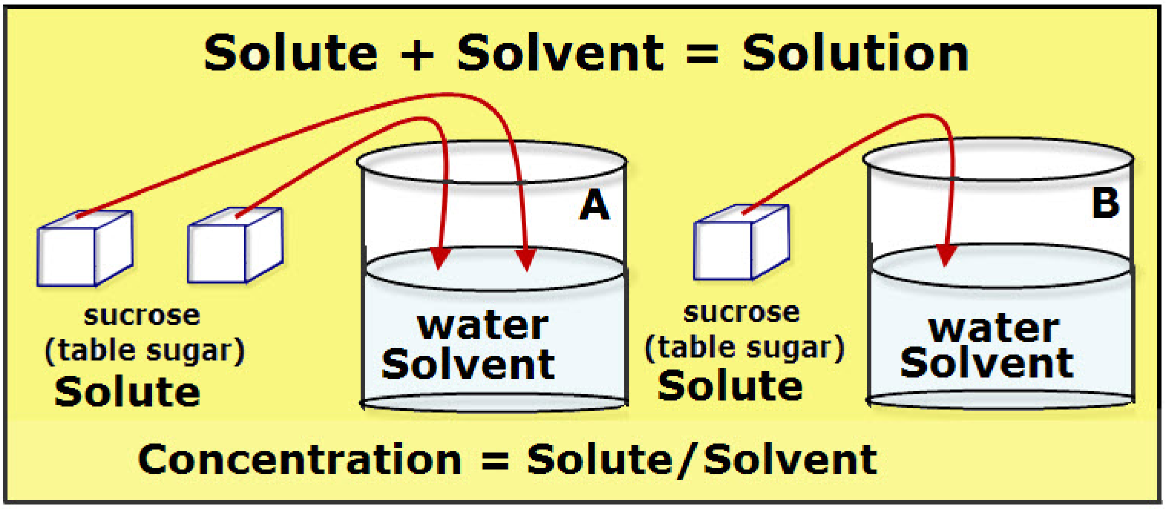
End of Section: Mole Concept

Signed: ____________________ Resubmit? ☐ No ☐ Yes by due date: __________

Concentration and Molarity

Inquiry question: How are chemicals in solutions measured?

Revision of Solutions:



Define the terms:

Solute:____

____

Solvent:____

____

Solution:____

____

We have previously discussed measuring the amount of a substance in chemistry using the standard measurement of a 'mole'. This is the same when we consider concentration. We can think of the concentration as being the number of moles per litre of solvent.

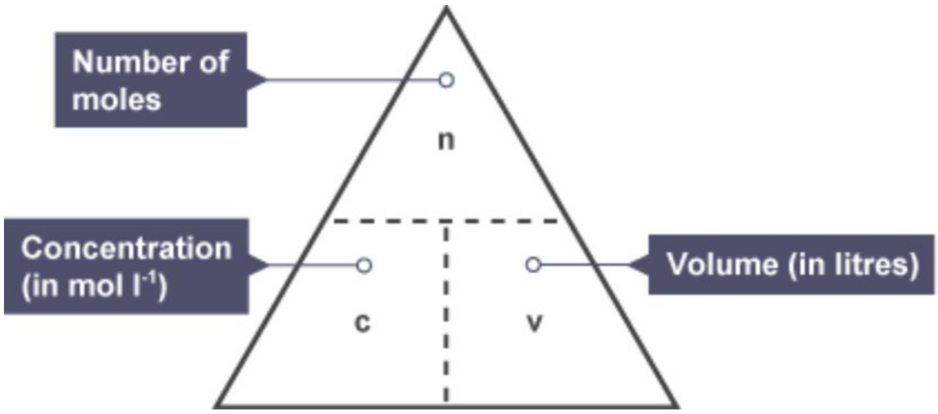
From this:

$$\boxed{\text{Concentration Mol/L}} \rightarrow C = \frac{n}{v}$$

Number of moles (mol) → n

Volume (L) → v

Using the Equation:



1. Write an analogy + draw a picture to represent the concept of a 4g of solid sodium hydroxide are dissolved in 500mL of water.

   Calculate the molarity of the solution.

____

____

2. What is the concentration of gL^{-1}, of a solution containing 5.00g of glucose in 250mL of solution?

____

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____

____

3. What is the concentration in gL^{-1}, of a 60mL solution that contains 5.0g of sugar?

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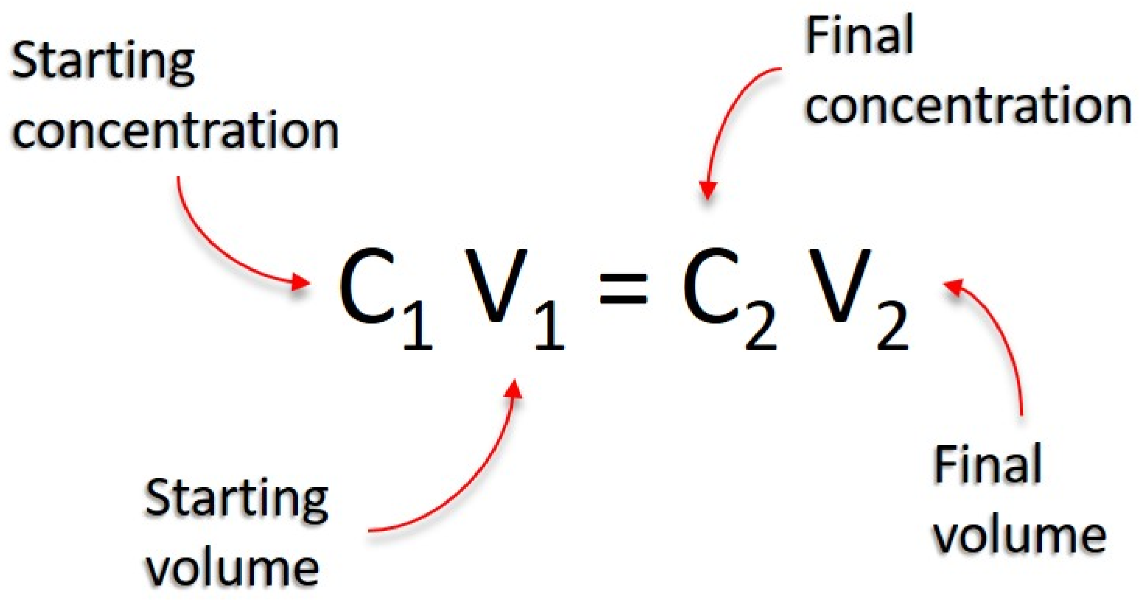
REVISION OF DILUTION:

Watch YouTube 'Molarity and Dilution' at https://youtu.be/QYK3Aj-IUIs

Increasing the solvent volume within a solution will ________________________ the concentration as long as the ________________________ of the solute remain the ________________________.

Knowing that we have not changed the number of moles, we can say that n = n. Therefore, if n = cv, that we can then use the formula:



These terms are inversely related, meaning that if we double the volume…

____

- **conduct an investigation to make a standard solution and perform a dilution**

Background Information:

The concentration of a solution is a measure of the amount of solute present in an amount of solvent. A solution may be prepared by dissolving a known mass or volume of solute in an known amount of solvent.

In this investigation the colour intensity of the potassium permanganate indicates the concentration of the solution.

Aim: To prepare a solution of known concentration and dilute it to a specified concentration.

Equipment:

- Approximately 0.1g potassium permanganate
- 250mL beaker
- 25ml pipette
- 4 x 250mL volumetric flasks
- Stirring rod
- Filter funnel
- Water

Method:

1. Weigh out approximately 0.1g potassium permanganate in one of the beakers
2. Add just enough water to dissolve the crystals and transfer the solution to a 250mL volumetric flask
3. Fill the flask to the 250mL mark with water and mix thoroughly. Record your results
4. Pipette 25mL of the first solution into a clean volumetric flask and fill it to 250mL. Record your results
5. Repeat step 4 until you have made a set of 4 dilutions.

Risk Assessment:

Risk	**Precaution**

Results:

Complete the table:

Solution	Relative Colour Intensity	Initial Volume	Final Volume	Mass $KMnO_4$	Moles $KMnO_4$	Concentration Mol/L
1						
2						
3						
4						

Discussion Questions:

1. What effect does dilution have on:

a) colour intensity:

____

____

b) concentration of the solution:

____

____

c) mass of the solute in the diluted solution, compared to the original solution:

____

____

2. Represent the process above using a flow chart:

3. How do you prepare a solution twice as concentrated as Solution 2?

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Conclusion:

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Summary of Concepts:

Standard solution:

____

____

Dilution:

____

____

Concentration/Molarity:

____

____

As the volume of a solution increases the molarity of the solution ____________________

- **conduct practical investigations to determine the concentrations of solutions and investigate the different ways in which concentrations are measured (ACSCH046, ACSCH063)**

MEASURES OF CONCENTRATION

Syllabus reference 8.4.4

Concentration is defined as the amount of solute in a specified amount of solvent. There are many ways of expressing concentration depending on the purpose and circumstances.

The four most common ways are described here.

A Mass per unit volume (grams/litre)
This method is commonly used when describing solubility or making solutions involving dissolving a solid in a solvent, usually water; for example, dissolving a specified amount of soluble fertiliser in water before applying it to plants.

Example: 2 g of fertiliser is dissolved in enough water to make a 50 mL solution. Calculate the concentration of the fertiliser in g/L.

Solution: $m = 2$ g $\quad\quad V = 0.050$ L

$$\text{Concentration} = \frac{m}{V} = \frac{2}{0.050} = 40 \text{ g/L}$$

B Per cent by mass, %(w/w)
This method uses a comparison of weight of solute compared to weight of solution, converted to per cent for ease of comparison. Labels on household cleaners often express concentrations of active ingredients in this way. It provides a useful way of comparing the strengths of cleaners e.g. bleaches.

Example: Calculate the percentage by mass of 35 g NaCl dissolved in 100 g of water.

Solution: $m_{\text{solute}} = 35$ g $\quad\quad m_{\text{solution}} = 100 + 35 = 135$ g

$$\text{Concentration} = \left(\frac{35}{135}\right) \times 100 = 26\ \%(\text{w/w})$$

C Per cent by volume, %(v/v)
This is commonly used when liquids are dissolved in other liquids. The ratio of the liquids is calculated and converted to a percentage. Alcohol content in drinks is expressed this way.

Example: Calculate the percentage by volume if 2.5 mL of hydrogen peroxide is dissolved in enough water to make 50 mL of solution.

Solution: $v_{\text{solute}} = 2.5$ mL $\quad\quad v_{\text{solution}} = 50$ mL

$$\text{Concentration} = \left(\frac{2.5}{50}\right) \times 100 = 5.0\ \%(\text{v/v})$$

D Parts per million (ppm)
This means parts of a particular solute per million parts of mixture. This is often used in environmental contexts when the concentrations are very small, especially when considering pollutants. If mercury is present in a lake at a concentration of 3 ppm this means 3 g of mercury per million g of solution.

Example: If the maximum allowable concentration of DDT in beef is 1.5 ppm, would a 500 g sample of beef containing 0.075 g of DDT be acceptable?

Solution: $m_{DDT} = 0.075$ g $\qquad m_{beef} = 500$ g

Concentration = $(\frac{0.075}{500}) \times 1\,000\,000 = 150$ ppm

The beef is not acceptable as this is 100 times the allowable limit.

Problems

1 25 g of fertiliser is dissolved in 75 g of water. What is the concentration of this solution in:

a g/L

b % (w/w)

2 One teaspoon of sucrose (10 g) is dissolved in 240 g water. (Total volume 250 mL.) What is the concentration of this solution in:

a g/L

b % (w/w)

3 What mass of water would be needed to prepare 250 g of a 20% (w/w) solution of NaOH?

4 What mass of an 8% sodium chloride solution in water do you need in order to have 40 g NaCl?

5 An impurity is present at a concentration of 0.01%. How many ppm does this correspond to?

6 The concentration of Na^+ in blood plasma is 300 mg/100 mL. Express this in ppm. (Assume that blood plasma has a density of 1.00 g/mL.)

7 The water supply of many cities is fluoridated giving 1.00 ppm of F^-. One city's water supply was analysed and the results showed there were 0.04 g F^- in a 500 mL sample. Did this city's water have the correct level of fluoridation?

8 It is required to prepare 1 kg of 5% KI solution.

a What mass in g of KI is needed?

b What mass in g of KI must be added to 1 kg of 5% solution so that it becomes a 10% solution?

- **manipulate variables and solve problems to calculate concentration, mass or volume using:**
  - $c = \frac{n}{v}$ **(molarity formula) (ACSCH063)**
  - **dilutions (*number of moles before dilution = number of moles of sample after dilution)***

Read through Conquering Chemistry Prelim '8.9 Molarity' pg 217-221

Recall the equation for calculating molarity:

How can we rearrange the equation above to calculate the number of moles in a particular volume of solution?

When we combine this information with $n = \frac{m}{M}$ we can manipulate experimental data to calculate specific values.

Exercises:

Examine the Conquering Chemistry Prelim Examples on page 218-219 before attempting the questions on page 220-221

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MOLARITY CALCULATIONS INCLUDING DILUTIONS

Syllabus reference 8.4.4

The molarity of a solution is the number of moles of solute per litre of solution. It is expressed as mol/L or mol L^{-1}. The symbol M is often used as an abbreviation for molarity.

$$\text{Molarity } (c) = \frac{\text{number of moles of solute } (n)}{}$$

$$\text{Alternatively } c = \frac{n}{V}$$

Problems

1 How many moles of solute do you need to weigh out to make the following solutions?
   - **a** 2.00 L of 1.50 mol L^{-1} sodium chloride
   - **b** 3.5 L of 0.20 mol L^{-1} potassium hydroxide
   - **c** 250 mL of 0.115 mol L^{-1} sulfuric acid

2 Calculate the molarity of the solutions made by dissolving the following amounts of solute in water and making the volume up to the stated value.
   - **a** 5.0 mol nitric acid in 2.0 L
   - **b** 2.5 mol sodium hydroxide in 0.50 L
   - **c** 0.020 mol sulfuric acid in 100 mL

3 How many moles of solute are there in:
   - **a** 2.5 L of 1.27 mol L^{-1} hydrochloric acid
   - **b** 1.7 L of 0.047 mol L^{-1} sulfuric acid

c 50 mL of 0.116 mol L^{-1} sodium hydroxide

4 How many moles of **(i)** magnesium chloride, **(ii)** magnesium ions, **(iii)** chloride ions, are there in 16.4 mL 0.117 mol L^{-1} magnesium chloride solution?

5 How many moles of (i) sulfuric acid, (ii) sulfate ions, (iii) hydrogen ions, are there in 14.5 mL 0.056 mol L^{-1} sulfuric acid solution?

6 What mass of the indicated substance do you need to weigh out to make the following solutions?

a barium hydroxide to make 0.5 L of 0.060 mol L^{-1} hydroxide ion solution

b sulfuric acid to make 250 mL of 0.330 mol L^{-1} hydrogen ion solution

7 In the following table what volume of Solution A is needed to react completely with the stated amount of Solution B?

	SOLUTION A	SOLUTION B
a	0.215 mol L^{-1} nitric acid	25 mL 0.334 mol L^{-1} sodium hydroxide
b	0.450 mol L^{-1} hydrochloric acid	50 mL 0.202 mol L^{-1} sodium carbonate
c	0.104 mol L^{-1} hydrochloric acid	0.150 g sodium hydroxide
d	0.085 mol L^{-1} sulfuric acid	0.383 g sodium hydrogen carbonate

8 To determine the solubility of calcium hydroxide, $Ca(OH)_2$, a chemist took 25 mL of a saturated calcium hydroxide solution and found that it reacted completely with 8.13 mL of 0.102 mol L^{-1} hydrochloric acid.

a Calculate the molarity of the calcium hydroxide solution.

b Calculate the solubility of calcium hydroxide in g per litre.

9 A solution was made by dissolving 22.22 g of calcium chloride in enough water to make a 500 mL solution. This was then reacted with 0.5 mol L^{-1} sulfuric acid to produce insoluble calcium sulfate. Calculate:

a the volume of sulfuric acid needed to react

b the mass of calcium sulfate produced

10 What mass of insoluble lead bromide is formed when 30 mL of 0.050 mol L^{-1} potassium bromide is added to an excess of lead nitrate solution?

11 A 25 mL solution containing 0.033 mol L^{-1} sodium chloride was needed to precipitate all of the silver from 15 mL of a silver nitrate solution.

a Calculate the concentration of the silver nitrate solution.

b Calculate the mass of silver chloride produced.

DILUTIONS

Frequently we need to dilute a solution from one concentration to a lower one. This is done by measuring out a known volume of the original solution and adding water to make a new volume. In the process the number of moles of solute is unchanged, only the concentration has decreased.

The following formula can be used to calculate concentrations and volumes in dilutions where c_1 and V_1 are the original concentration and volume of solute and c_2 and V_2 are the new concentration and volume.

$$c_1V_1 = c_2V_2$$

12 The volume of solution in column A was diluted to the volume in Column B. Calculate the molarity of the diluted solution.

COLUMN A	COLUMN B
a 50 mL 0.242 mol L^{-1} hydrochloric acid	500 mL
b 25 mL 0.152 mol L^{-1} sulfuric acid	2.0 L
c 10 mL 0.114 mol L^{-1} sodium hydroxide	250 mL

13 What volume of solution in Column A in the following table is needed to prepare the solution in Column B?

	COLUMN A	COLUMN B
a	0.282 mol L^{-1} hydrochloric acid	250 mL 0.0113 mol L^{-1}
b	2.42 mol L^{-1} sulfuric acid	2.0 L 0.121 mol L^{-1}
c	0.318 mol L^{-1} sodium hydroxide	1.0 L 0.300 mol L^{-1}

14 50 mL 0.20 mol L^{-1} sodium hydroxide was mixed with 50 mL 0.30 mol L^{-1} magnesium chloride. A precipitate of magnesium hydroxide formed. Calculate the concentration of:

a sodium ions

b magnesium ions

c chloride ions

End of Section: Concentration and Molarity

Signed: ____________________ Resubmit? ☐ No ☐ Yes by due date:__________

Gas Laws

Inquiry question: How does the Ideal Gas Law relate to all other Gas Laws?

- **conduct investigations and solve problems to determine the relationship between the Ideal Gas Law and:**
  - **Gay-Lussac's Law (temperature)**
  - **Boyle's Law**
  - **Charles' Law**
  - **Avogadro's Law (ACSCH060)**

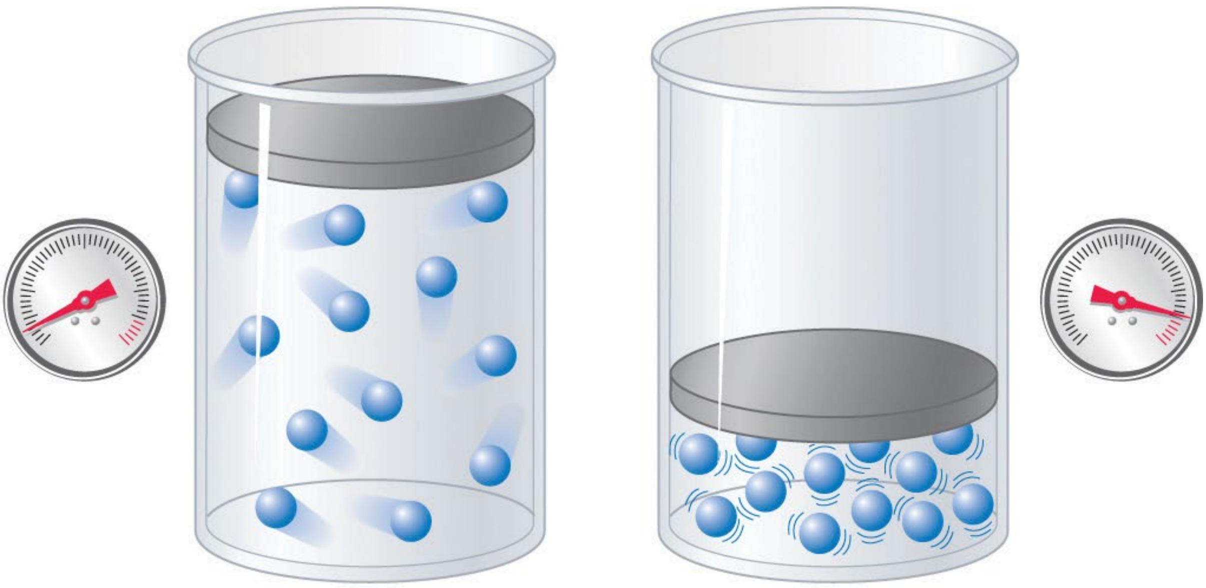
PROPERTIES OF GASES

Complete the table below from your prior knowledge:

Property of Gases	An Example of this Property
No definite shape	
Diffuse	
Can be poured	
Can be compressed	
Can exert pressure	

RELATING PRESSURE AND VOLUME

Examine the diagram below:



1. Summarise the trend in the diagram to explain the relationship between volume and pressure:

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2. Use a diagram to represent the particles making up a container of gas

a) describe the movement of these particles

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b) How would the particles change if you heated the container?

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c) If the container was a balloon, what would you observe upon heating it?

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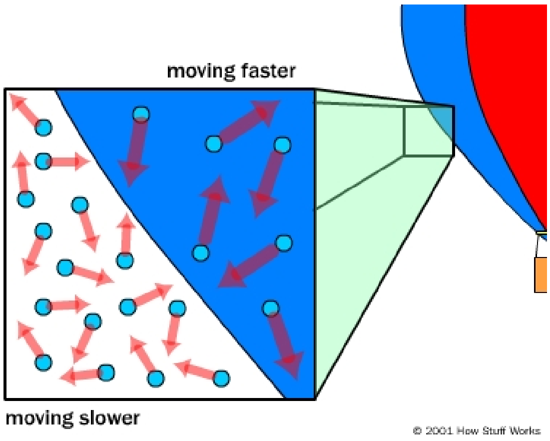
3. Explain why a hot air balloon rises upon heating the air inside it. Justify your response in terms of the behaviour of the gas particles.



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GAY-LUSSAC AND GAS STOICHIOMETRY

Read through Conquering Chemistry Prelim 'Gay-Lussac's Law of Combining Volumes and Avogadro's Hypothesis' pg 150-151

Recall the 'Law of Combining Volumes'

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Give an example to illustrate this law:

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How does this link back to Avogadro's findings?

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What did Avogadro propose as a result of Gay-Lussac's law?

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Sketch the Gay-Lussac experiment and Avogadro's theory to demonstrate the link:

Why can we compare volumes of gases combining, rather than having to convert them to moles?

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Is this an accurate process when considering reactions between other physical states? Why or why not?

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EXERCISES:

Complete the questions in Conquering Chemistry Prelim pg 152

Investigation 9.1 Verifying Gay-Lussac's Law

This resource is available on NelsonNet 'Chemistry in Focus' pg 141-142

AIM:

To process some results obtained from an experimental study of the reaction between hydrogen and oxygen to test the validity of Gay-Lussac's Law.

Hypothesis:

When measured at constant temperature and pressure the volumes of gases taking part in a chemical reaction show simple, whole-number ratios to one another.

Method:

Cylinders of compressed gas were used to fill two wash bottles, one with hydrogen and the other with oxygen. Immediately after filling, the wash bottles were places in a large container of water with their delivery tubes held below the water surface with weights. This was to stop air getting into the bottles and to allow water to enter the bottles as gas was used up.

By putting the hydrogen wash bottle tube under the modified burette and squeezing the wash bottle, hydrogen gas was injected into the burette. As the wash bottle was released, water was drawn through the wash bottle tube to replace the hydrogen gas.

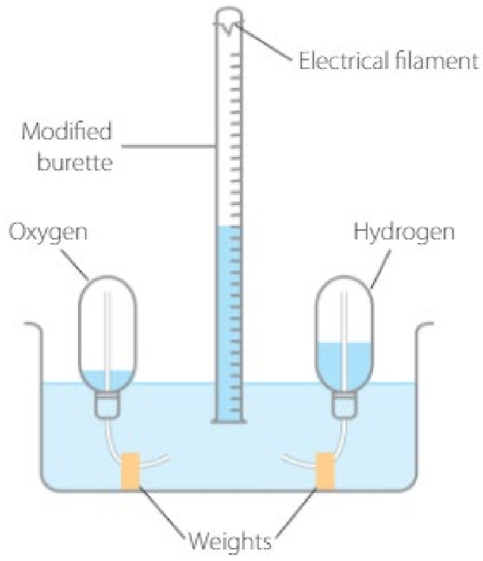

FIGURE 9.4 Apparatus for testing Gay-Lussac's law

Oxygen gas was then injected into the burette to make the total volume up to 20mL.

After a pause to allow the gases to mix, a battery was connected to the filament to make it glow, which caused the gas mixture to ignite. After another pause to allow the gas temperature and volume to stabilise, the final volume of gas was recorded.

The experiment was repeated several times.

1. Rewrite this method to meet the accepted conventions of scientific report writing.

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Results:

Description	Measurements					
Volume of hydrogen added to burette (mL)	2.2	3.9	7.2	9.8	14.9	17.9
Total Volume of Reactants (mL)	20.2	20.1	20.0	20.3	20.2	20.3
Final Volume of gases (mL)	16.8	14.2	9.4	5.5	4.6	13.2
Volume of Oxygen added to burette (mL)						
Reactant all used up						
Volume of hydrogen that reacted (mL)						
Volume of Oxygen that reacted (mL)						
Ratio of volumes of hydrogen to oxygen that reacted						

2. Calculate the 'volume of oxygen' provided in each experiment

3. The water formed in the reaction quickly condenses. So the volume remaining is unreacted Hydrogen or oxygen. Assume that the gas initially present in the lesser amount is the limiting reagent. Put a H or an O in the 'Reactant all used up' row. Enter its volume in the appropriate row and then calculate the volume of the other gas that reacted. This will be the initial volume of that gas, minus the volume of the gas left at the end of the experiment.

4. Calculate the ratio of the volumes of hydrogen and oxygen that reacted.

Discussion:

Watch YouTube "The Martian: Funny Explosion Scene' at https://youtu.be/kGfzTJlo-Ag

The Martian (2015) - Funny Explosion Scene [1440p]

1. Suggest reasons why you were not able to perform the actual experiment yourself

2. Identify sources of error in the experiment

Conclusion:

Decide on the validity of the original hypothesis:

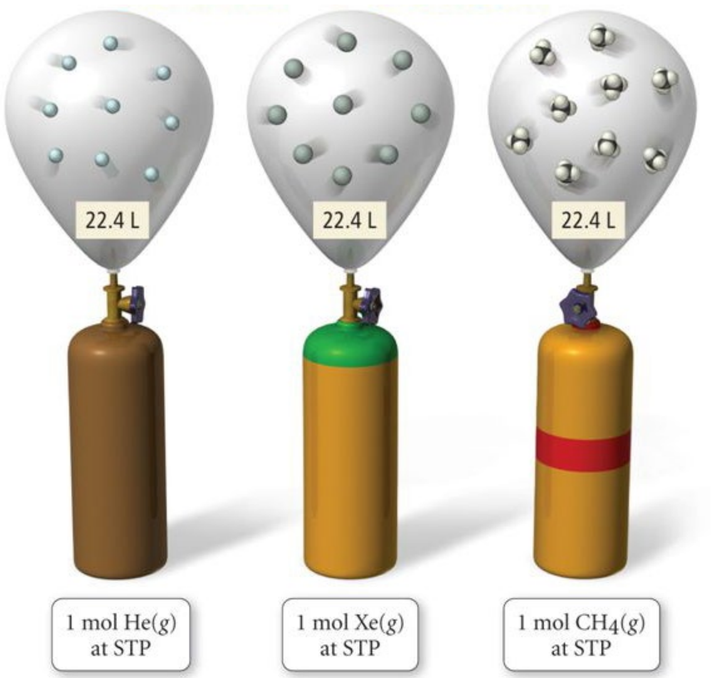
MOLAR VOLUME OF A GAS:

Avogadro's Law can be understood to state:

Equal numbers of molecules of different gases occupy the same volume at the same temperature and pressure.

At 0°C and exactly 100kPa, the molar volume of all gases is 22.71L

At 25°C and exactly 100kPa, the molar volume of all gases is 24.79L

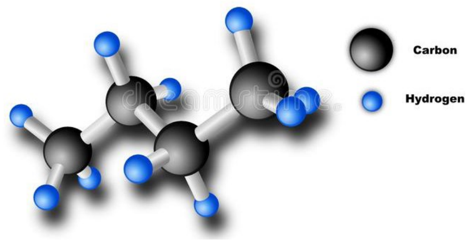


EXPERIMENT: Investigation 9.2 Measuring the Molar Mass of Butane

This resource is available on NelsonNet 'Chemistry in Focus' pg 141-142

In this investigation we shall determine the molar mass of the gas Butane, by measuring the volume occupied by a known mass of the gas. Butane will be obtained from a disposable cigarette lighter.

Butane (formula:_________________) is a liquid under pressure in a cigarette lighter, when the lighter valve is opened the butane vaporises and will bubble up into the measuring cylinder of water.



Aim:

To determine the molar mass of butane

Materials:

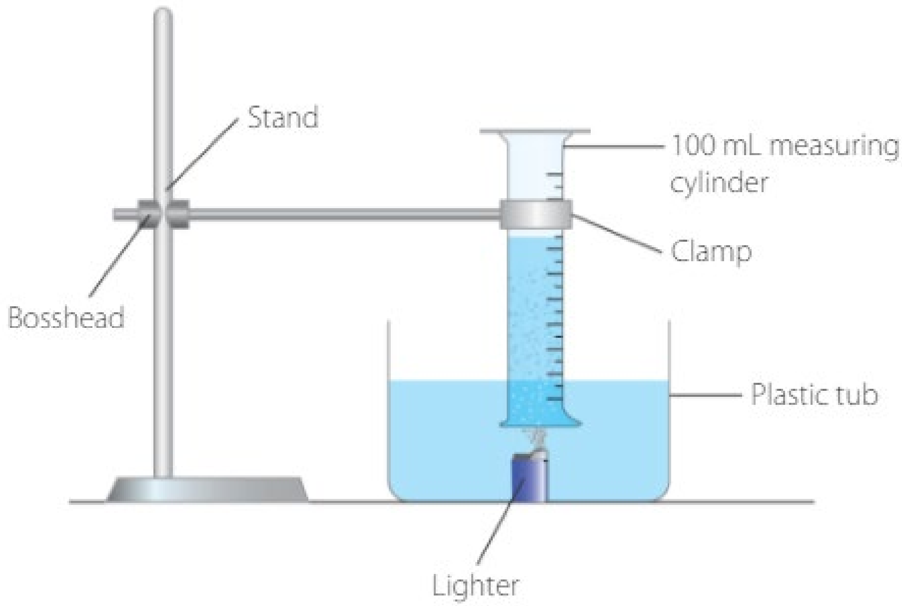
- Disposable butane lighter with flint removed
- Large plastic tub
- Water
- 100mL measuring cylinder
- Electronic balance
- Paper towel
- Thermometer
- Retort stand
- Boss head and clamp

Risk Assessment:

Risks	Precautions
Butane gas is flammable	Remove all ignition sources from the laboratory
Butane gas is a respiratory irritant	Work in a well-ventilated area. Do not inhale the gas.

Method:

1. Set up the following apparatus



2. Two thirds fill the large plastic tube with water

3. Weigh the lighter using the electronic balance

4. Completely fill the 100mL measuring cylinder with water

5. Hold your hand over the measuring cylinder and submerge it in the tub, clamping it in place

6. Press the button on the lighter to release the gas, ensuring it is bubbling into the measuring cylinder. Continue until the water levels inside and outside the measuring cylinder are level

7. Record the volume of gas in the measuring cylinder, and dry and weigh the lighter

8. Measure and record the temperature of the water

Results:

Description	Measurement
Initial mass of the lighter (g)	
Final mass of the lighter (g)	
Volume of butane (mL)	
Temperature of the water (°C)	
Molar Volume of Gas (100kPa & 25°C)	
Moles of Butane (mol)	
Mass of Butane used (g)	
Experimental Molecular Mass of Butane	
Actual Molecular Mass of Butane	

Discussion:

1. Identify any sources of error in both the experimental method and the calculations performed:

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2. Why is there a discrepancy between the experimental and standard molar mass values for butane?

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Conclusion:

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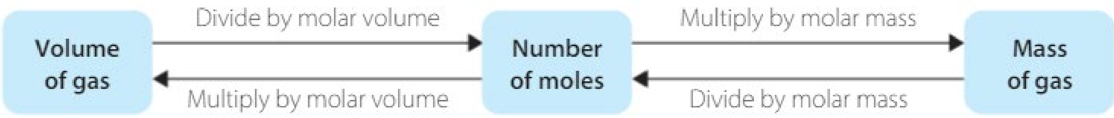
CALCULATIONS INVOLVING VOLUMES OF GASES

This resource is available on NelsonNet 'Chemistry in Focus' pg 199

The only difference for performing calculations involving gases from earlier calculations is the consideration of the volume of the gas. Many questions will state the experimental conditions regarding temperature and atmospheric pressure and give you a volume of gas to work with.

The method for doing these calculations can be shown schematically:



WORKED EXAMPLE 9.1

1 What mass of gas is present in a 515 mL flask containing carbon dioxide at 25°C and at a pressure of 100.0 kPa?

2 What volume is needed to store 3.11 g ethane, C_2H_6, at 100 kPa and 0°C?

ANSWER	LOGIC
1 Number of moles of CO_2 present $= \frac{\text{volume of gas in L}}{\text{volume of one mole in L/mol}}$ $= \frac{0.515}{24.79}$ $= 0.0208\ \text{mol}$	▪ Use the molar volume of a gas at 25°C and 100.0 kPa to calculate the number of moles of CO_2 present.
Molar mass of $CO_2 = 12.01 + 2 \times 16.00$ $= 44.01\ \text{g mol}^{-1}$ Mass of $CO_2 = 0.0208 \times 44.01$ $= 0.914\ \text{g}$	▪ Calculate the molar mass of CO_2 then use it to convert moles of CO_2 into mass.
2 Molar mass of ethane, $C_2H_6 = 2 \times 12.01 + 6 \times 1.008$ $= 30.07\ \text{g mol}^{-1}$ Number of moles of ethane $= \frac{3.11}{30.07}$ $= 0.103\ \text{mol}$	▪ Convert mass of ethane to moles by first calculating its molar mass.
Volume this will occupy $= 0.103 \times 22.71$ $= 2.35\ \text{L}$	▪ Calculate the volume this occupies by using the molar volume of a gas at 0°C.

EXERCISES:

1. What mass of methane (CH_4) is in a flask of internal volume of 2.33L at 0°C and 100kPa pressure?

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2. What volume is needed to store 23.0g of ammonia, NH_3, at 100kPa pressure and 25°C?

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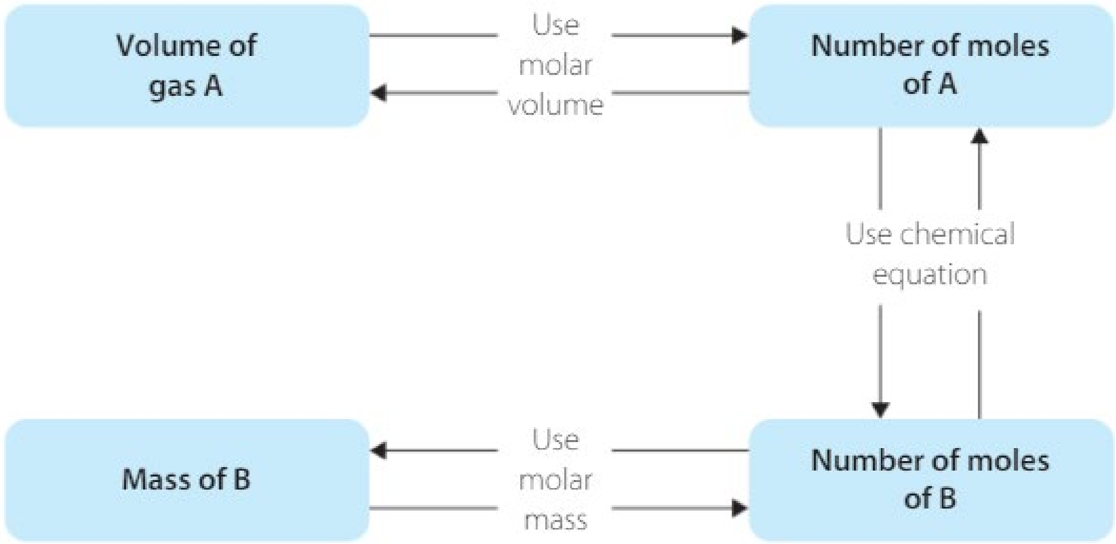
KEY CONCEPTS

- **Gay-Lussac's law** states that when measured at constant temperature and pressure, the volumes of gases taking part in a chemical reaction show simple whole-number ratios to one another.
- **Avogadro's law** states that when measured at the same temperature and pressure, equal volumes of different gases contain the same number of molecules.
- An alternative statement of Avogadro's law is that equal numbers of molecules of different gases occupy the same volume at the same temperature and pressure.
- The **molar volume of a gas** is the volume occupied by 1 mole of a gas. It is the same for all gases at a given temperature and pressure.

VOLUMES OF GASES IN CHEMICAL REACTIONS

When chemical reactions are involved, there is an additional step: When comparing the ratio of reactants vs gas products, we must do so in moles (This is very similar to previous calculations where we cannot compare masses of different substances, rather we must compare moles).



EXERCISES:

1. State Gay-Lussac's law of combining volumes

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2. Methane burns in oxygen according to the following equation:

$$CH_{4\,(g)} + 2O_{2\,(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

If the methane and oxygen are at the same temperature and pressure, calculate:

a) The volumes of carbon dioxide and water vapour produced by the combustion of 1.2L of methane in excess of oxygen

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b) The volume of gas produce be the combustion of 25mL of methane

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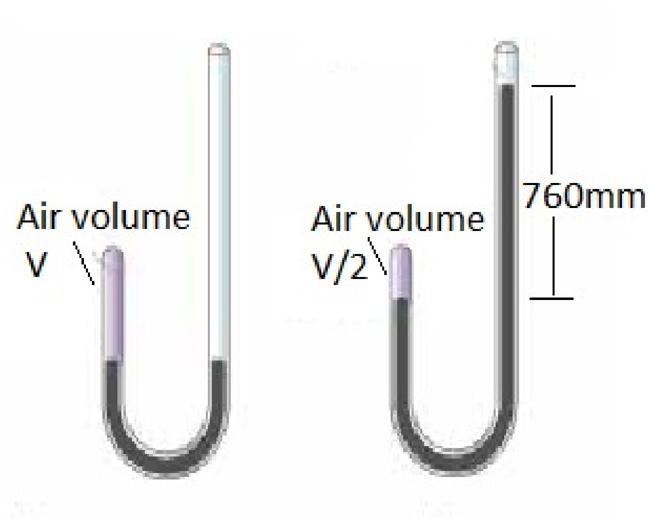
BOYLE'S LAW:

Boyles law states:

For a given mass of gas at constant temperature, its volume is inversely proportional to the pressure

$PV = k$ *(P = pressure, V = volume , k = a constant)*



This means that if the pressure on a gas is doubled, its volume will halve. This happens because the gas particles are compressed closer together by the increasing pressure.

In his experiments, Boyle trapped air in a curved glass tube, using liquid mercury, and measured the volume of the air. He then increased the pressure on the air by adding more mercury. He found that when the total pressure on the air was doubled, the volume halved.

A very useful form of Boyles Law is

$$P_2V_2 = P_1V_1$$

Each of the products, P_2V_2 and P_1V_1 is equal to the same constant, *k*. We can use any units we like, providing the units are the same for both sides of the equation.

WORKED EXAMPLE 9.3

A sample of gas originally had a volume of 1.50 L at 20°C and 100 kPa pressure. What volume would it have at 400 kPa pressure?

ANSWER	LOGIC
$P_2V_2 = P_1V_1$	▪ Use the alternative form of Boyle's law.
$P_1 = 100\,\text{kPa}$, $V_1 = 1.50\,\text{L}$, $P_2 = 400\,\text{kPa}$ and V_2 is what is wanted.	▪ Identify the values of the knowns in it.
$(400\,\text{kPa}) \times V_2 = (100\,\text{kPa}) \times 1.50\,\text{L}$	▪ Substitute the values into the equation and solve.
$V_2 = \dfrac{100 \times 1.50\,\text{kPa L}}{400\,\text{kPa}}$	▪ Put units along with the numbers into the equation to make sure they cancel out to give a sensible unit for the answer.
$= 0.38\,\text{L}$	▪ 100 and 400 are ambiguous regarding significant figures – could be 1, 2 or 3.
	▪ Using 2 is a safe guess.

EXERCISES:

1. A 2.3L sample of ammonia gas at 45°C and 101kPa was compressed until its pressure was 245kPa. What is its final volume (still at 45°C).

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2. 500mL of oxygen gas at 550kPa is allowed to expand into a volume of 3.60L at constant temperature. What is its final pressure?

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3. The pressure on 55.5L nitrogen gas is changed from 101.3kPa to 158.2kPa. Calculate its new volume

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4. The pressure on 2.5L of gas is increased to 180kPa. The volume of the gas decreases to 1.2L. What was the original pressure on the gas.

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5. Explain why carbonated drink 'fizz' when you open them

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EXPERIMENT: Investigation 9.3 Volume of a gas at different pressures

This resource is available on NelsonNet 'Chemistry in Focus' pg 141-142

Aim: To measure the change in volume of a gas when pressure is changed at constant temperature.

Materials:

60mL plastic syringe

Rubber stopper or blutack to seal the end of the syringe

Retort stand, clamp and boss head

Bathroom scales

Several 0.5kg, 1kg and 2kg weights to make a total of 6-8kg (bricks or pavers are ideal)

Method:

Design a method to test incremental weight increases against the volume of air within the syringe

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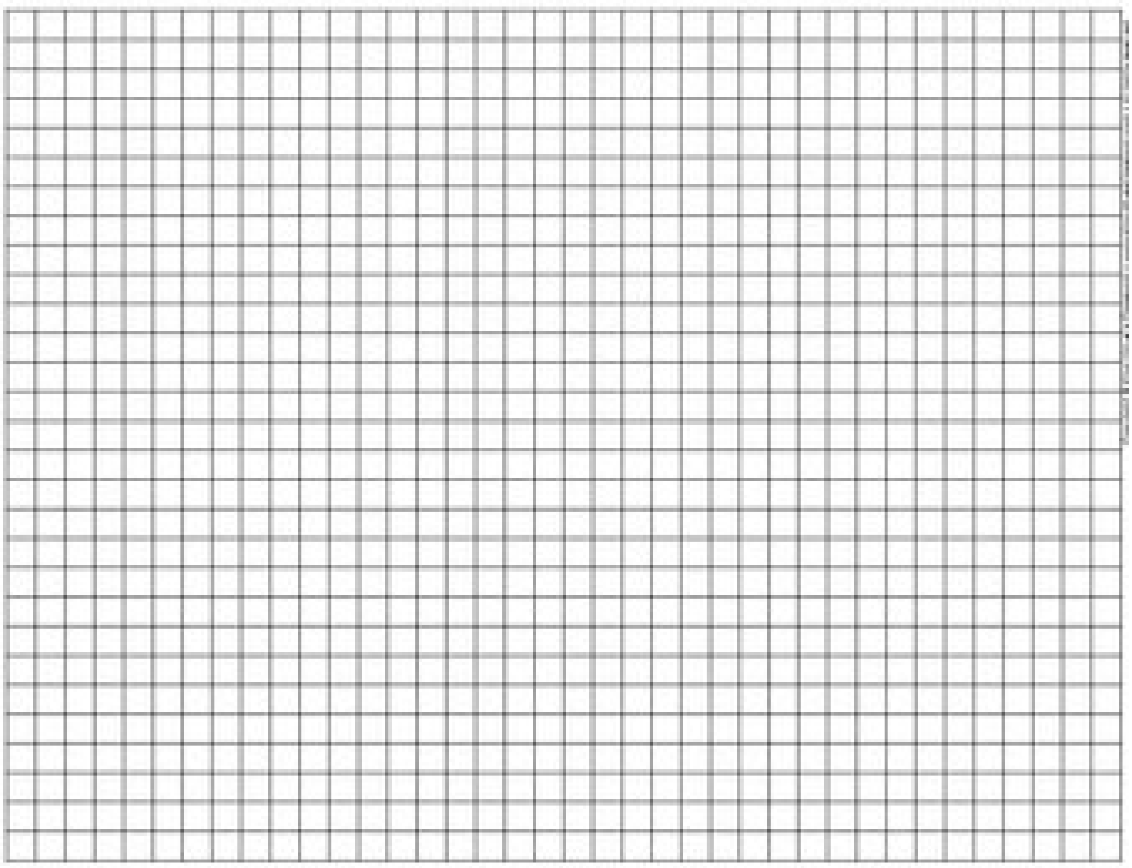
Results:

Total weight added to the top of the syringe (kg)					
Volume of gas in syringe (L)					

Graph the volume vs pressure data above



Conclusion:

Describe qualitatively the relationship between volume of gas and pressure at a constant temperature

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Can you devise an alternative plot of your data that may result in a straight line relationship?

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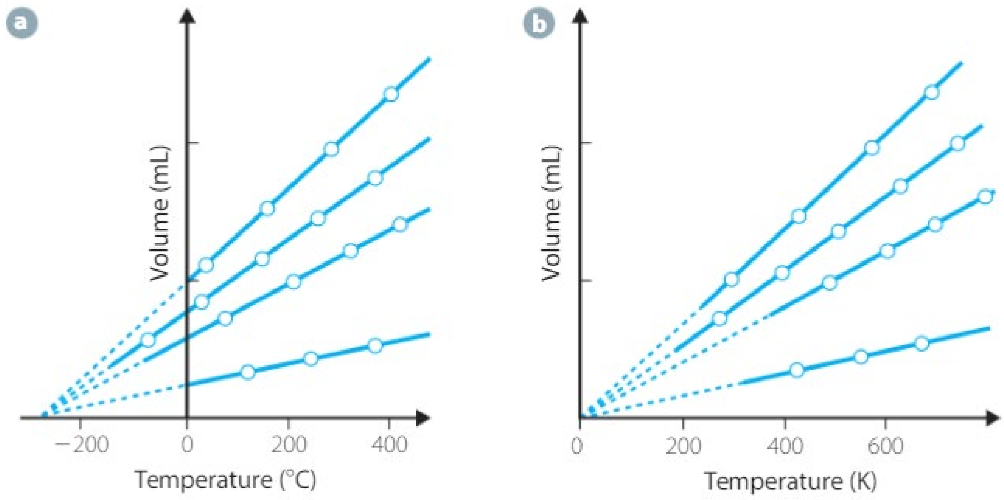
CHARLES' LAW

Charles' law states:

The volume of a fixed amount of a gas, held at a constant pressure, is directly proportional to its absolute temperature measured in kelvins

$\frac{V}{T} = k$ (v = volume, T = temperature, k = a constant value)



When various temperatures are tested and graphed it can be seen that, when extrapolated, all data plots intercept at -273.15°C. This is true of all gases studied.

Therefore, we introduced a new temperature scale, The Absolute Temperature Scale, which has the units of Kelvin (K). Unless otherwise stated, the units for symbol T are Kelvins.

We can move between Kelvins (K) and degrees Celsius (°C) according to:

T(K) = temperature in ℃ + 273.15

NOTE: 1°C is equal to 1K.

Similarly to Boyle's law, because of the presence of the constant, k, we can use Charles' law as:

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

WORKED EXAMPLE 9.4

A sample of gas at 101 kPa pressure had a volume of 2.5 L at 100°C. What would its volume be at 0°C at the same pressure?

ANSWER	LOGIC
$\frac{V_2}{T_2} = \frac{V_1}{T_1}$	▪ Use Charles's law.
$T_1 = 100 + 273.2 = 373.2\,\text{K}$ $T_2 = 0 + 273.2 = 273.2\,\text{K}$ $V_1 = 2.5\,\text{L}$	▪ Convert temperatures to kelvin because that is the temperature unit required for this equation.
$\frac{V_2}{273.2\,\text{K}} = \frac{2.5\,\text{L}}{373.2\,\text{K}}$ $V_2 = \frac{2.5 \times 273.2\,\text{L K}}{373.2\,\text{K}}$	▪ Substitute the values into the equation and solve it.
$= 1.8\,\text{L}$	▪ 2.5 L means 2 significant figures in the answer.

EXERCISES:

1. 250mL of nitrogen gas at 25°C was heated to 155°C at a constant pressure of 101kPa. What will be its new volume?

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2. A sample of chlorine gas held had a volume of 1.25L at 18°C What temperature must it be cooled to for its volume to decrease to 750mL at constant pressure?

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3. A sample of gas is kept at a constant pressure while its temperature is cooled from 100°C to 0°C. What was the original volume if its final volume is 1.83L.

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4. What is meant by the term 'absolute zero'?

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EXPERIMENT: Investigation 9.4 Volume of gas at different temperatures

Aim: To measure the change in volume of a gas when temperature is changes at constant pressure.

Materials:

- 20mL plastic syringe
- Rubber stopper or blutack to seal the end of the syringe
- Thermometer
- 2 rubber bands
- Ice
- Water
- 500mL beaker
- Tripod
- Wire gauze
- Bunsen Burner
- Matches
- Container for ice bath
- heat proof mat

Method:

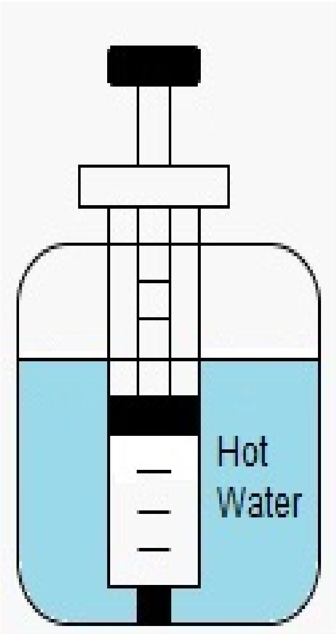
1. Add 10mL of air to the syringe and seal

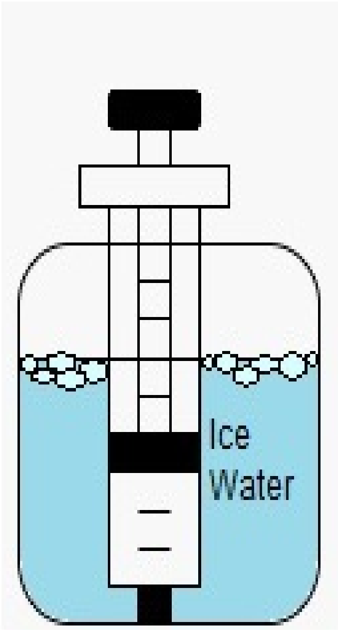
2. Attach a thermometer to the side of the syringe using rubber bands

3. Record the temperature of the air and the volume of gas

4. Place the syringe in an ice bath for 10minutes and record the temperature and volume of gas.

5. Repeat Step 4 using a water bath at 50°C, a water bath at 75°C and a water bath at boiling point.





Results:

Experimental Condition	Temperature (°C)	Volume (mL)
Ice Bath		
Room Temperature (25ish)		
water bath at 50°C		
water bath at 75°C		
water bath at boiling		

Graph your data and include a line of best fit:

Conclusion:

Describe the relationship between volume of gas and temperature at a constant pressure:

KEY CONCEPTS

- **Boyle's law** states that, for a given quantity of gas at a constant temperature, pressure is inversely proportional to volume or, stated differently, the product of volume, V, and pressure, P, is constant.
- The Kelvin or **absolute temperature scale** is defined as:

  temperature in kelvin (K) = temperature in °C + 273.15
- **Charles's law** states that, at constant pressure, the volume of a fixed quantity of gas is proportional to its absolute (or Kelvin) temperature.

THE IDEAL GAS LAW

The gas laws can be conveniently combined.

Boyle's law states that the pressure-volume product is constant:

$$PV = k$$

Charles's law shows that the volume is proportional to the absolute temperature:

$$\frac{V}{T} = k_2$$

Gay-Lussac's law says that the pressure is proportional to the absolute temperature:

$$P = k_3 T$$

Where:

P is the pressure in kilopascals (kPa)

V is the volume in litres (L)

T the absolute temperature of an ideal gas in kelvin (K).

By combining (1) and (2) or (3), a new equation is derived with P, V and T:

$$\frac{PV}{T} = k \qquad \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Units:
Kilopascals (kPa) for pressure
Litres (L) for volume
Kelvin (K) for temperature
Additionally;
Kilopascals x litres = Joules
kPa x L = J

Watch YouTube 'The Ideal Gas Law: Crash Course Chemistry' at [*https://youtu.be/BxUS1K7xu30*](https://youtu.be/BxUS1K7xu30)

The Ideal Gas Law: Crash Course Chemistry #12

WORKED EXAMPLE 9.6

1 Calculate the volume that 2.5 mol carbon dioxide occupies at 400 kPa pressure and 100°C.

2 Calculate the number of moles of gas needed to fill a 500 mL flask to a pressure of 220 kPa at 21.2°C.

ANSWER	LOGIC
1 $V=\frac{nRT}{P}$	• Use the ideal gas law in rearranged form with volume, *V*, as subject.
$P=400\,kPa$ (preferred unit) $T=100+273.2=373.2\,K$ (required unit) $n=2.50$ (no choice for unit here) $R=8.314\,J\,K^{-1}\,mol^{-1}$	• Identify the input data.
$V=\frac{2.50\times 8.314\times 373.2}{400}\ \frac{mol\,J\,K^{-1}\,mol^{-1}\,K}{kPa}$ $=19.4\,L$	• Substitute into the equation and solve. • (Note that $J\,kPa^{-1}=L$)
2 $n=\frac{PV}{RT}$	• Rearrange $PV=nRT$ to make the required quantity, *n*, the subject.
$P=220\,kPa$ $V=500\,mL=0.50\,L$ $T=21.2+273.2=294.4\,K$ $R=8.314\,J\,K^{-1}\,mol^{-1}$	• Identify the input data.
$n=\frac{220\times 0.50}{8.314\times 294.4}\ \frac{kPa\,L}{J\,K^{-1}\,mol^{-1}\,K}$ $=0.045\,mol$	• Substitute and solve.

QUESTIONS:

1. What volume is needed to store 0.62mol of gas at a pressure of 3.6×10^5Pa at a temperature of 295°C

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2. What is the combined gas law? How is it related to Boyle's, Gay-Lussac's and Charles's law? What units are needed?

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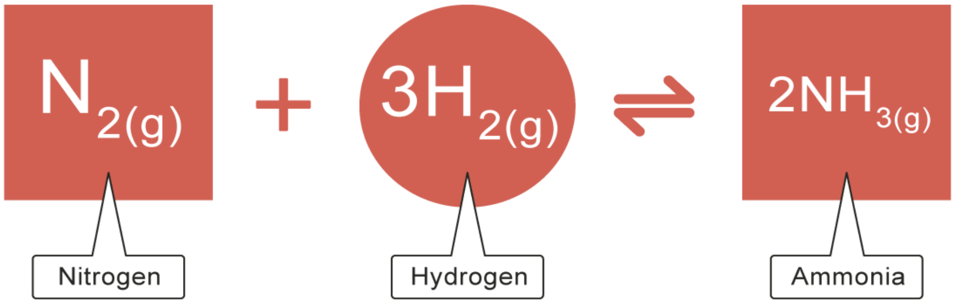
3. The Haber Process is used to combine hydrogen and nitrogen gas to form ammonia gas. The balanced equation is shown below:



a) Explain how it obeys Gay-Lussac's Law and the combined gas law.

b) If 1 mole of nitrogen reacts with excess hydrogen, how many moles of hydrogen will be used, and how many moles of ammonia will be produced?

c) If 2 litres of hydrogen is reacted, what will be the volume of nitrogen and ammonia involved in the reaction?

d) What volumes of hydrogen and ammonia are needed to produce 500mL of ammonia?

4. Each part of this question refers to a constant mass of gas at a fixed volume. Complete the following table:

Initial condition		Final condition	
Pressure	Temperature	Pressure	Temperature
5atm	300K		900K
760mmHg	500°C		250°C
693kPa	25°C	1582kPa	
300bar	300K	500bar	
200atm	300K		650K

5. Each part this question refers to a constant mass of gas. Complete the following table:

Initial condition			Final condition		
Pressure	Volume	Temperature	Pressure	Volume	Temperature
101.3kPa	10L	25°C	53.2kPa		-56°C
1atm	3.2L	3000K		1L	273K
5bar	116mL	-25°C		100mL	75°C
10atm	375mL	298K	1atm	7500mL	

6. A tyre is inflated to 30kPa pressure at 15°C. Assuming a constant tyre volume of 33L, what will be the pressure in the tyre at a temperature at 47°C?

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7. An upper atmosphere balloon is filled on the ground, where the temperature is 25°C and the pressure is 1atm, to a volume of 5L.
It rises through the atmosphere to a region where the temperature is -50°C and its volume increases to 10L. Calculate the pressure in the balloon.

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8. How many moles of helium are there in 3.45L cylinder of the gas at pressure of 235kPa at 35°C?

9. What pressure will 0.050mol argon exert in a 500mL flask at 18°C?

KEY CONCEPTS

- The **combined gas law** (equation 9.12) can be joined with Avogadro's law (equation 9.14) to give the ideal gas law.

$$PV = nRT$$

- The **ideal gas law** is also called the ideal gas equation or sometimes the general gas equation.
- The constant R is called the **universal gas constant**, because it applies to *all* gases.
- In using the ideal gas law, note that:

kilopascals × litres = joules *or* kPa × L = J

End of Section: Gas Laws

Signed: ______________________ Resubmit? ☐ No ☐ Yes by due date:__________