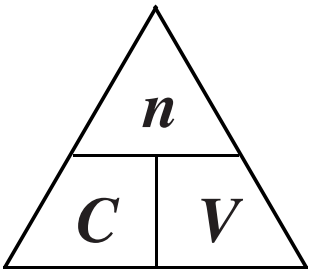
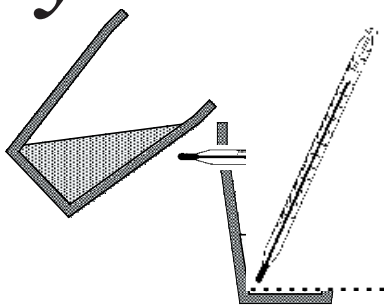
Higher Chemistry



Unit 1: Energy Matters



Section 4:

Enthalpy & Moles

Student:

Lesson	Activities			Done	Checked
4.1 ***Molar Relationships 1***	1.	Solids, liquids & gases			
	2.	Solutions			
	3.	Equations			
		Check Test	Score: /		
		Home Practice	Score: /		
4.2 ***Energy Changes***	1.	Energy out, energy in			
	2.	Where does the energy come from ?			
	3.	Important Enthalpies			
		Check Test	Score: /		
		Home Practice	Score: /		
4.3 ***Enthalpies from Experimental data***	1.	Specific Heat Capacity			
	2.	Enthalpy of Solution			
	3.	Enthalpy of Neutralisation			
		Check Test	Score: /		
		Home Practice	Score: /		
4.4 ***Enthalpy of Combustion***	1.	Burning Ethanol			
	2.	Processing Results			
	3.	Comparing Enthalpies			
		Check Test	Score: /		
		Home Practice	Score: /		
4.5 ***Molar Relationships 2***	1.	The Avogadro Constant			
	2.	Formula Units			
	3.	Empirical Formula			
		Check Test	Score: /		
		Home Practice	Score: /		
4.6 ***Molar Relationships 3***	1.	Density of a Gas			
	2.	Molar Volumes			
	3.	Reacting Volumes			
		Check Test	Score: /		
		Home Practice	Score: /		
4.7 ***Molar Relationships 4***	1.	Avogadro's Law			
	2.	Gases in Excess			
	3.	Other Chemicals in Excess			
		Check Test	Score: /		
		Home Practice	Score: /		
Consolidation Work		Consolidation A	Score: /		
		Consolidation B	Score: /		
		Consolidation C	Score: /		
		Consolidation D	Score: /		

End-of-Section Assessment	Score: %	Grade:	

	Notes	Revision Questions Questions	Text Book
4.1	**Molar Relationships**		
	Solids, liquids & gases	p16 Ex 1.3 Q1,2	p13 Q1
	Solutions	p16,17 Ex 1.3 Q3-7	p13 Q2,3
	Equations	p18,19 Ex 1.4 Q1-4	
4.2	**Energy Changes**		
	Energy out, Energy in	p23-25 Ex 1.7 various	p21 Q1-3
	Important Enthalpies	p26 Ex 1.8 Q1	
4.3	**Enthalpies from Experimental Data**		
	All lessons	p26-28 Ex 1.8 Q1-11(except 5)	p21 Q4-6
4.4	**Enthalpy of Combustion (PPA)**	p27 Ex 1.8 Q5	
4.5	**Molar Relationships 2**		
	The Avogadro Constant		
	Formula Units	p49,50 Ex 1.16 Q1-10	p56 Section A
	Empirical Formula		
4.6	**Molar Relationships 3**		
	Density of a Gas	p52 Ex 1.17 Q10,11	
	Molar Volumes	p51,52 Ex 1.17 Q1-9	p56,57 Section B
	Reacting Volumes	p53 Ex 1.17 Q12	p56,57 Section B
4.7	**Molar Relationships 4**		
	Avogadro's Law	p54 Ex 1.18 Q1-3	p56,57 Section B
	Gases in Excess	p54,55,56 Ex 1.18 Q4-12	p56,57 Section B
	Other Chemicals in Excess	p20 Ex 1.5 Q1-7	p13 Q6

4.1 Molar Relationships

This first lesson looks again at the work covered at Standard Grade which concentrated on Moles as Masses.

Solids, liquids & gases

*This activity looks at how the **masses** of substances can be used to calculate the **number of moles**.*

If you know the ***formula*** of a substance you can calculate its ***formula mass*** in amu (atomic mass units; mass of 1 proton = 1amu).

e.g. $CaCO_3$ = 1 x Ca = 1 x 40 = 40
1 x C = 1 x 12 = 12
3 x 0 = 3 x 16 = 48 =100 amu

By definition,

1 mole = ***formula mass*** in ***grammes*** or gramme formula mass ***gfm***

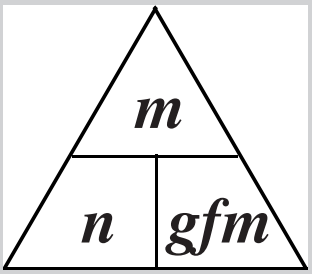
e.g. 1 mole of $CaCO_3$ = 100 g

Remember that ***1 mole*** of any substance contains the ***same amount*** of substance as ***1 mole*** of any other substance.

1 mole of $CaCO_3$	=	1 mole of $C_6H_{12}O_6$	=	1 mole of $(NH_4)_2SO_4$
100 g	=	180 g	=	128 g

moles *of a substance* = ***mass*** *of substance* ÷ ***gfm***

n = ***m*** ÷ ***gfm***



mass *of substance* = ***moles*** *of a substance* ***x*** ***gfm***

m = ***n*** ***x*** ***gfm***

e.g. 25g of $CaCO_3$

n = ***m*** ÷ ***gfm***
= 24 ÷ 100
= 0.25 moles

Solutions

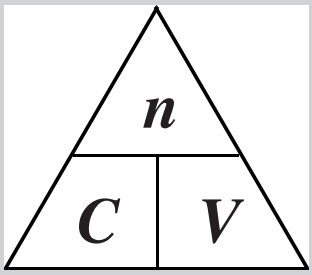
*This activity looks at how the **concentration** and **volume** of a solution can be used to calculate the **number of moles** of solute present.*

The ***concentration*** of a solution (***C***) is measured in ***moles per litre***, abbreviated to ***mol l^{-1}.***
The ***volume*** of a solution (***V***) is measured in cubic centimetres (cm^3) or in ***litres*** (*l*).

1000 cm^3 = 1 litre = 1000 m*l*

moles *in a solution* = ***concentration*** ***x*** ***volume***

$$n = C \times V$$



concentration = ***moles*** *dissolved* ÷ ***volume***

$$C = n \div V$$

e.g. 250 cm^3 of a 2 mol l^{-1} solution of hydrochloric acid.

$n = C \times V$
$= 2 \times 0.25$ (volumes must be in litres)
$= 0.5$ moles of HCl (= 0.5 moles of H^+ ions)
$= 18.25$g of HCl (*gfm* = 36.5g)

250 cm^3 of a 1 mol l^{-1} solution of sulphuric acid acid.

$n = C \times V$
$= 1 \times 0.25$ (volumes must be in litres)
$= 0.25$ moles of H_2SO_4 (= 0.5 moles of H^+ ions)
$= 24.5$g of H_2SO_4 (*gfm* = 98g)

e.g. 11g ofcarbon dioxide dissolved in 500 cm^3

gfm of CO_2 = 44g = 1mole
11g = 0.25 mole

$C = n \div V$
$= 0.25 \div 0.5$ (volumes must be in litres)
$= 0.5$ moles per litre
$= 0.5$ mole l^{-1} or 0.5 M

Equations

*This activity looks at how the **equation** for a reaction can be used to calculate reacting masses.*

A ***balanced chemical equation*** tells you the ***reactants*** and the ***products*** in a reaction, and the relative ***amounts*** involved.

e.g *methane* + *oxygen* → carbon dioxide + water

CH_4 + $2\ O_2$ → CO_2 + $2\ H_2O$

The ***number*** written in ***front*** of each ***formula*** tells you the ***number*** of each molecule involved in the **reaction**. This equation tells you that ***one*** *molecule of methane* reacts with ***two*** *molecules of oxygen* to form ***one*** *molecule of* ***carbon dioxide*** and ***two*** *molecules of* ***water***.

The ***molar relationships*** are exactly the same:

CH_4 ; *gfm* = 16 g	
O_2 ; *gfm* = g	
CO_2 ; *gfm* = g	
H_2O ; *gfm* = g	

CH_4 +	$2\ O_2$	→	CO_2 +	$2\ H_2O$
1*mole*	2 *moles*		1 *moles*	2 *moles*
16g	64 *g*		1 *g*	36 *g*

(mass must be conserved during chemical reactions)

Most of the time we only need to concentrate on two of the substances in the reaction.

e.g. What mass of *carbon dioxide* is produced when 64g of *methane* is burned in a plentiful supply of air ?

	CH_4 +	O_2	→	CO_2 +	H_2O
Step 1					
Step 2	***1****mole*			1 *moles*	
Step 3	*16g*			44 *g*	
Step 4	*64g*			*(64 / 16) x* 44 *g*	
				= 176 *g*	

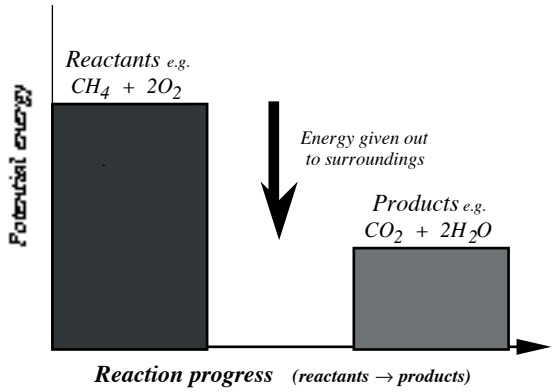
Step 4 involves the use of ***proportion***; which is an essential mathematical ability for anyone wishing to study Chemistry.

4.2 Energy Changes

This lesson considers the energy changes which are a characteristic feature of chemical reactions.

Energy out, energy in

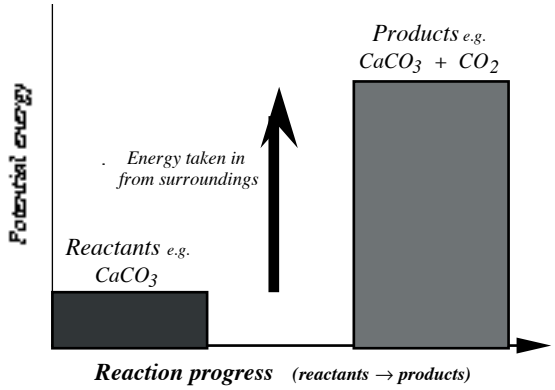
*This activity looks at **exothermic** and **endothermic** reactions.*



Many reactions ***give out energy*** and ***heat*** the surroundings. They are called ***exothermic*** reactions.

The ***products*** end up with ***less energy*** than the ***reactants*** had - but the ***surroundings*** end up with ***more***, and get ***hotter***.

We can *measure* the ***energy transferred*** to the ***surroundings*** and it is called the ***enthalpy change*** and given the symbol ***ΔH***



Very few reactions ***take in energy*** and ***cool down*** the surroundings. They are called ***endothermic*** reactions.

The ***products*** end up with ***more energy*** than the ***reactants*** had - but the ***surroundings*** end up with ***less***, and get ***cooler*** .

We can *measure* the ***energy transferred*** from the ***surroundings*** and it is called the ***enthalpy change*** and given the symbol ***ΔH***

We will soon be able to put ***numbers*** to the ***energy changes*** associated with chemical reactions ; to *quantitatively* measure ***enthalpy changes***, ***ΔH***.

By definition: $$\Delta H = H_{products} - H_{reactants}$$

This means that:

thermic reactions ΔH = - ve

thermic reactions ΔH = + ve

The units for ***enthalpy changes***, ***ΔH***, will be *kilojoules per mole*, ***kJ mol^{-1}***

Where does the energy come from ?

This activity looks at where the energy lost by the reactants or gained by the products comes from.

At the ***start*** of the reaction the ***reactant*** molecules have ***kinetic energy*** and ***potential energy***.

The ***potential energy*** is mainly energy ***stored*** within their ***bonds*** .

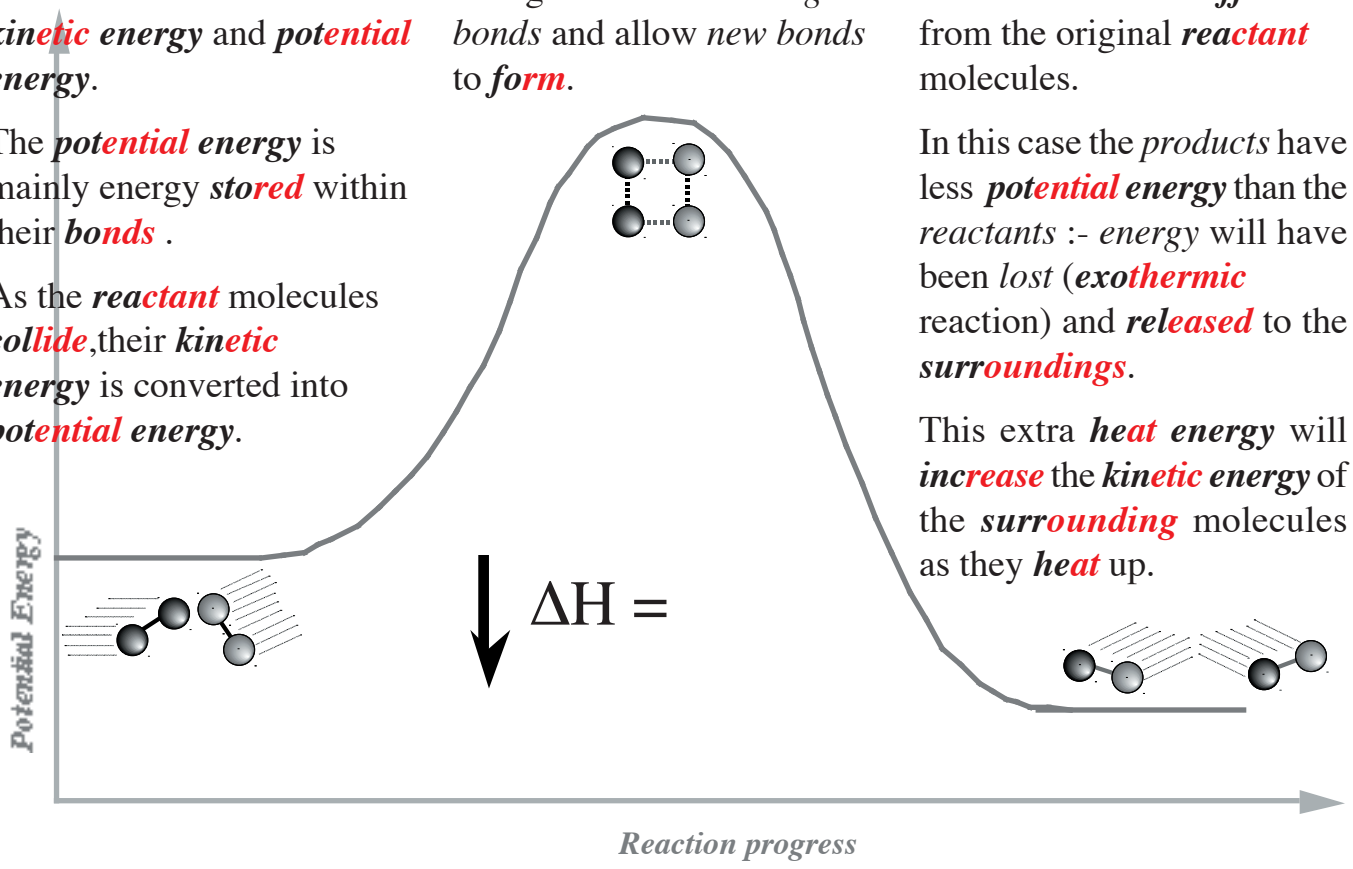
As the ***reactant*** molecules ***collide***,their ***kinetic energy*** is converted into ***potential energy***.

This ***extra energy*** can be enough to ***break*** *existing bonds* and allow *new bonds* to ***form***.

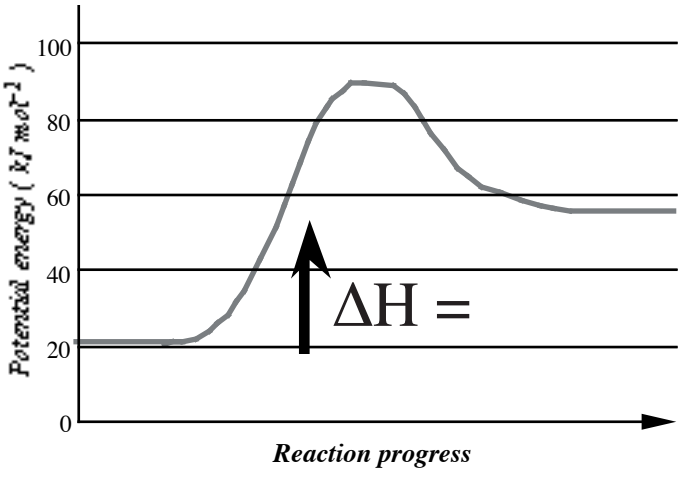
The *bonds* in the ***product*** molecules will be ***different*** from the original ***reactant*** molecules.

In this case the *products* have less ***potential energy*** than the *reactants* :- *energy* will have been *lost* (***exothermic*** reaction) and ***released*** to the ***surroundings***.

This extra ***heat energy*** will ***increase*** the ***kinetic energy*** of the ***surrounding*** molecules as they ***heat*** up.



In effect, ***kinetic energy*** is used to ***break*** *existing bonds* and then ***kinetic energy*** is ***released*** as *new bonds are* ***made***. During ***exothemic reactions*** more ***energy is released*** than was taken in. This energy was originally ***stored*** in the bonds of the ***reactants*** but is now providing extra ***kinetic energy*** to other ***reactant*** molecules, and the ***surroundings***, causing them to ***heat*** up.



During ***endothermic*** reactions, ***kinetic energy*** is again used in ***breaking*** *existing bonds*.

This time, though, the bonds in the ***product*** molecules will have ***more energy stored*** in them than the reactant molecules had.

Not all of the ***kinetic energy*** that was available at the start of the reaction is now available to the ***product*** molecules.

The reaction mixture will now have ***less kinetic energy*** (be colder) than the ***surroundings*** so ***energy will flow from the surroundings to the reaction mixture***.

Important Enthalpies

*This activity introduces the enthalpies of **combustion**, **solution**, and **neutralisation**.*

Any ***chemical*** process can have an ***enthalpy of reaction***, but certain types of reactions are deemed important enough to be ***named*** and defined carefully. ***Values*** for many of these processes are to be found in *Data Books*.

Along with the written definition, it is important that ***balanced equations*** can also be written that describe *exactly* the reaction to which the *numerical* value of ΔH corresponds.

*The **enthalpy of combustion** of a substance is the enthalpy change when **one mole** of the substance **burns completely** in oxygen*

e.g propane

$\Delta H =$ $C_3H_{8\,(g)} + 5\,O_{2\,(g)} \rightarrow 3\,CO_{2\,(g)} + 4\,H_2O_{(l)}$

methanol

$\Delta H =$ $CH_3OH_{(l)} + {}^3/_2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{(l)}$

Notice that for the ***equation*** to match the ΔH value in the Data Book it must be ***balanced*** using *'fractions'* of moles so that the ***quantity*** of fuel remains as ***one mole***.

Notice that ***enthalpies of combustion*** are *always* ***exothermic***, ie. their ΔH values are *always* ***negative***.

*The **enthalpy of solution** of a substance is the enthalpy change when **one mole** of the substance **dissolves completely** in water*

e.g ammonium nitrate

$\Delta H = +26.8\ kJmol^{-1}$ $NH_4NO_{3\,(s)} + aq \rightarrow NH_4NO_{3\,(aq)}$

sodium hydroxidel

$\Delta H = -41{,}8\ kJmol^{-1}$ $NaOH_{(s)} + aq \rightarrow NaOH_{(aq)}$

Notice that there is no ***quantity*** for the ***water*** involved; this is why ***aq*** is used rather than the ***formula*** H_2O, which would imply *one mole* of water.

Notice that ***enthalpies of solution*** can be *either* ***exothermic***, ΔH ***negative***, *or* ***endothermic***, ΔH ***positive***.

Once ***dissolved*** the ***ionic network***, *crystal lattice*, is completely broken so the *individual ions* are now only ***bonded*** to ***water*** molecules. For this reason, these ***equations*** are more often written as:-

$NH_4NO_{3\,(s)} + aq \rightarrow NH_4^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)}$

$NaOH_{(s)} + aq \rightarrow Na^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$

*The **enthalpy of neutralisation** of an acid is the enthalpy change when the acid is neutralised to form **one mole** of **water**.*

e.g *hydrochloric acid* / NaOH

$\Delta H = -57.3\ kJmol^{-1}$ $NH_4NO_{3\,(s)} + Na^+OH^-_{(aq)} \rightarrow {}_{(aq)} + H_2O_{(l)}$

sulphuric acid / KOH

$\Delta H = -57.3\ kJmol^{-1}$ $NH_4NO_{3\,(s)} + K^+OH^-_{(aq)} \rightarrow {}_{(aq)} + H_2O_{(l)}$

Notice, again, that there is often a need to use '***fractions***' of ***moles*** in order to ***balance*** these equations in terms of ***one mole of water*** ,as defined.

Notice that ***enthalpies of neutralisation*** are *always* ***exothermic***, ΔH ***negative***, and that the value is the ***same*** for many combinations of ***acids*** / ***alkalis***.

This is because the actual reaction is ***identical*** in most cases;

$H^+_{(s)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$ (omitting ***spectator ions***)

Later in the course, you will be introduced to ***weak acids*** like *ethanoic acid*,CH_3COOH, and ***weak alkalis*** like *ammonium hydroxide*, NH_4OH. Acids like these have ***different*** enthalpies of neutralisation.

For any combination of ***strong acid*** / ***strong alkali***, however, the enthalpy of neutralisation is always very close to $-57.3\ kJmol^{-1}$.

4.3 Enthalpies from Experimental data

This lesson looks at how enthalpies for reactions can be measured experimentally.

Specific Heat Capacity

ΔH	=	*energy gained or lost by the water*
c	=	*specific heat capacity of the water*
m	=	*mass of the water*
ΔT	=	*rise or fall in temperature*

$$\Delta H = c \times m \times \Delta T$$

*This activity looks at how the **specific heat capacity** of water can be used to experimentally determine the enthalpy of a reaction.*

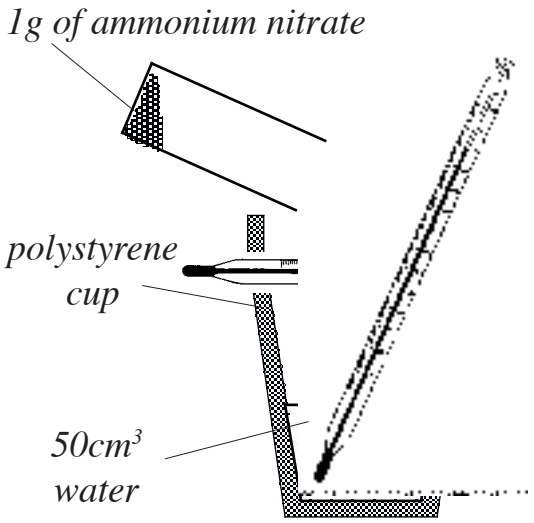
During ***chemical*** reactions ***energy*** is ***released*** *to the* ***surroundings***, ***exothermic***, or ***taken in*** *from the* ***surroundings***, ***endothermic***.

Often, the ***surroundings*** are the ***water*** that the chemicals are ***dissolved*** in.

We know exactly ***how much energy*** it takes to heat, or cool down, ***1kg*** of water by exactly ***1°C***. This is the ***specific heat capacity***, ***c = 4.18 kJ kg^{-1} °C^{-1}***

Enthalpy of Solution

*This activity looks at how the **enthalpy of solution** can be measured experimentally.*



This is an ***endothermic*** process and ***energy*** will be ***taken from*** the *water*, causing the *water* to ***cool down.***

typical results

There was a ***temperature drop*** of **1.6** °C.

Energy absorbed from water = ***c*** *x* ***m*** *x* ***ΔT***

= ***4.18*** *x* ***0.05*** *x* ***1.6***

= ***kJ***

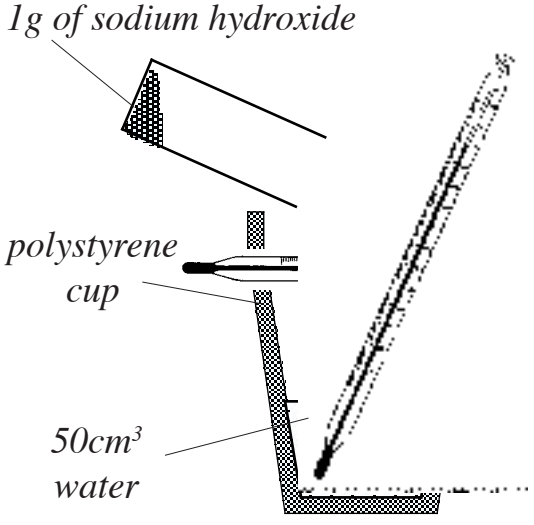
This tells us how much ***energy*** was ***absorbed*** when ***1g*** of *ammonium nitrate* ***dissolved*** in water. We need to ***calculate*** how much ***energy*** would have been absorbed if ***1 mole*** of *ammonium nitrate* had ***dissolved***.

Formula = gram formula mass (1 mole) =

1 g of *ammonium nitrate* ⟶ ***kJ*** of energy

1 mole, g of *ammonium nitrate* ⟶ *x*

= ***kJ*** of energy



This is an ***exothermic*** process and energy will be ***released into*** the *water*, causing the *water* to ***heat up***.

typical results

There was a ***temperature rise*** of °C.

Energy released into the water = ***c*** *x* ***m*** *x* ***ΔT***

= *x* *x*

= ***kJ***

This tells us how much ***energy*** was ***released*** when ***1g*** of *sodium hydroxide* ***dissolved*** in water. We need to ***calculate*** how much ***energy*** would have been released if ***1 mole*** of *sodium hydroxide* had ***dissolved***.

Formula = *gram formula mass* (1 mole) =

1 g of *sodium hydroxide* ⟶ ***kJ*** of energy

1 mole, g of *sodium hydroxide* ⟶ *x*

= ***kJ*** of energy

These ***experimental*** values are always *slightly* ***inaccurate***, because,

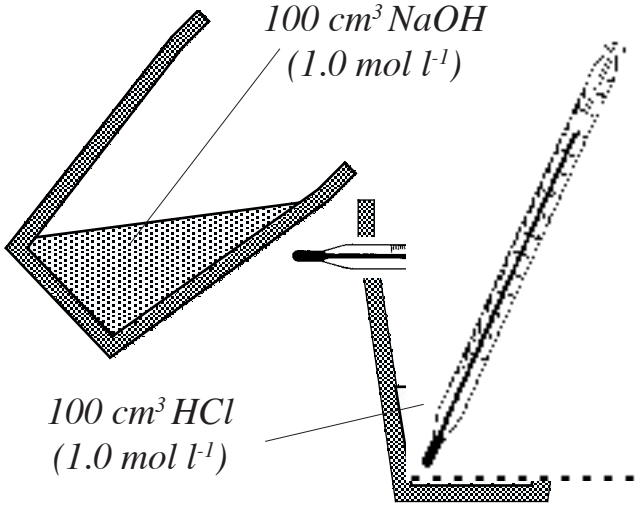
- we assume that the ***specific heat capacity*** for the ***solution*** is the *same* as ***pure water***, 4.18 kJ kg^{-1} $°C^{-1}$
- we assume that the ***volume*** of the ***solution*** is the *same* as the ***water*** we started with, 50 cm^3, and that ***1cm³ = 1g*** as for ***pure*** water.
- we assume that all the ***heat released*** stays in the ***water***; that none ***escapes*** to the air etc.

Enthalpy of Neutralisation

*This activity looks at how the **enthalpy of neutralisation** can be measured experimentally.*

This is an ***exothermic*** process and ***energy*** will be ***released into*** the *water*, causing the *water* to ***heat up***.

typical results There was a ***temperature rise*** of **6.5** °C.



$NaOH + HCl \rightarrow NaCl + H_2O$

Energy released into water = ***c*** x ***m*** x ***ΔT***

= ***4.18*** x ***0.2*** x ***6.5***

= ***kJ***

These ***volumes*** and ***concentrations*** mean that 0.1 ***moles*** of *NaOH* react with 0.1 ***moles*** of *HCl* to produce 0.1 ***moles*** of H_2O.

We simply need to multiply by 10 to ***calculate*** the ***energy released*** when ***1 mole*** of H_2O is formed.

$\Delta H_{neutralisation}$ = *kJ* mol^{-1}

4.4 Enthalpy of Combustion

PPA

This lesson looks at how an enthalpy of combustion can be measured experimentally.

Burning ethanol

*This activity looks at how the **enthalpy of combustion** of ethanol can be measured experimentally.*

* *State the aim of the experiment*

Procedure

Ethanol can be ***burned*** in a ***spirit burner***. The ***heat released*** (***exothermic***) must be ***absorbed*** by ***water*** placed in a beaker, or, if possible, a ***copper can***, just above the ***flame***. (***Copper*** is an excellent ***conductor***).

Heatproof mats can be used to provide a ***draught shield*** around the apparatus; to keep the ***flame*** steady and cut down ***heat loss*** from the experiment.

A ***thermometer*** is used to measure the ***temperature change*** as the ***water heats up***.

* *Draw a labelled diagram of the assembled apparatus*

Enough ***ethanol*** is burnt to produce a ***10°C temperature rise*** (approximately), and the ***amount*** of ***ethanol*** used is measured. The following measurements should be taken:

* *List the five measurements that were made during the experiment*

1

2

3

4

5

Processing Results

*This activity looks at how the **enthalpy of combustion** of ethanol can be calculated from experimental data.*

Results

* *Present your results in an appropriate manner*

Calculation / Conclusion

* *Carry out a calculation to determine the enthalpy of combustion of ethanol.*

Look up your Data Book, and find the value for the enthalpy of combustion of ethanol.

Data Book value =

This is very much greater than the experimental value.

* *Suggest sources of error which could account for this difference.*

Comparing Enthalpies

This activity compares the enthalpies of combustion of a family of alkanols.

Use your Data Book to complete the following table:

alcohol	*structural formula*	*difference*	*enthalpy of combustion kJmol^{-1}*	*difference*
methanol	CH_3OH	—		—
ethanol	CH_3CH_2OH	CH_2		
propan-1-ol	$CH_3CH_2CH_2OH$			
butan-1-ol	$CH_3CH_2CH_2CH_2OH$		- 2673	

The ***alkanol*** family is another example of a ***homologous series*** - each ***member*** differs from the previous ***member*** by the ***same amount*** - by a CH_2.

There is also a fairly ***constant difference*** in their ***enthalpy values*** . This supports the idea that it is energy ***stored in bonds*** that is ***released*** during ***combustion*** reactions.

4.5 Molar Relationships 2

*This lesson looks at the mole as a **number**.*

The Avogadro Constant

Sad, but true.....

6.02×10^{23} grains of pollen would cover the city of London to a depth of 1 mile !

6.02×10^{23} marshmallows would cover the United States of America to a depth of 600 miles !

6.02×10^{23} marbles would coverGreat Britain to a depth of 1000 miles !

It would take a computer, counting at the rate of 10 million numbers per second, 2 billion years to count to 6.02×10^{23} !

All the grains of sand on all the beaches of the World do not number as many as 6.02×10^{23} !

If you laid 6.02×10^{23} Rangers supporters head to foot across the Atlantic Ocean, a lot of them would drown !

This activity introduces the number of 'entities' contained in a mole of a substance.

Everyday objects are often counted in group amounts; a ***pair*** of *socks*, a ***dozen*** *eggs*, a ***gross*** (144) of *nails* and a ***ream*** (500) of *paper*.

The ***smaller*** the object, the ***larger*** the number of objects in the group.

Since ***atoms*** are so ***small*** we use a very large number, 602252000000000000000000, in our group called a ***mole***.

This number is called ***The Avogadro Constant***, and given the symbol **L**.

Avogadro Constant, L = 6.02×10^{23}

One of the ***quantities*** available to early ***Chemists***, including the Italian ***Amadeo Avogadro*** (1776 - 1856), were ***atomic masses***.

The element ***Carbon*** was thought of as 'special' because its ***atomic mass*** was very close to a ***whole number***; 12.0.

The mass of a ^{12}C atom = 1.99252×10^{-23} g
(12 x mass of 1 proton)

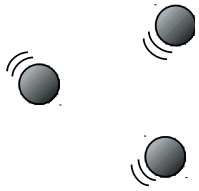
The number of C atoms in 12g of carbon =
$12 / 1.99252 \times 10^{-23} = 6.02252 \times 10^{23}$

1 mole *of any substance* ***= gram formula mass*** *of that substance*
1 mole *of any substance* ***= 6.02 x 10²³*** *entities*

Formula Units

This activity explains which 'entities' are contained in a mole of different substances.

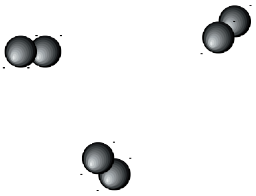
- ***Monatomic gases*** The basic entity is the ***atom***.



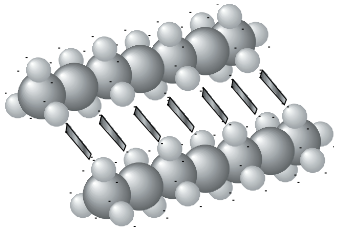
Formula = Ne ; 10g = ***1mole*** = 6.02×10^{23} ***atoms***

Formula = Ar ; 40g = ***1mole*** = 6.02×10^{23} ***atoms***

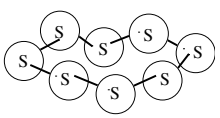
- ***Covalent molecules*** The basic entity is usually the ***molecule***.



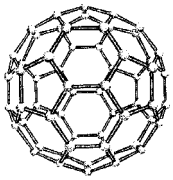
Formula = O_2 ; 32g = ***1mole*** = 6.02×10^{23} ***molecules***
= $2 \times 6.02 \times 10^{23}$ ***atoms***
= 1.204×10^{24} ***atoms***



Formula = C_7H_{16} ; 100g = ***1mole*** = 6.02×10^{23} ***molecules***
= $23 \times 6.02 \times 10^{23}$ ***atoms***
= $7 \times 6.02 \times 10^{24}$ ***C atoms***
= $16 \times 6.02 \times 10^{24}$ ***H atoms***

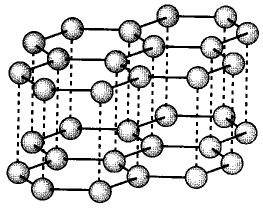


Formula = S ; 32g = ***1mole*** = 6.02×10^{23} ***atoms***



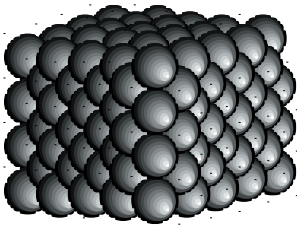
Formula = C_{60} ; 720g = ***1mole*** = 6.02×10^{23} ***molecules***
= $60 \times 6.02 \times 10^{23}$ ***C atoms***

- ***Covalent network*** The basic entity is the ***atom***.



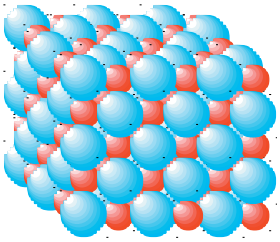
Formula = C; 12g = ***1mole*** = 6.02 x 10^{23} ***atoms***

- ***Metallic Networks*** The basic entity is the ***atom***



Formula = Cu; 63.5g = ***1mole*** = 6.02 x 10^{23} ***atoms***

- ***Ionic Networks*** The basic entity is the ***ionic ratio (formula unit)***



Formula = NaCl ; 58.5g = ***1mole*** = 6.02 x 10^{23} ***Na^+Cl^- units***

= 2 x 6.02 x 10^{23} ***ions***

= 6.02 x 10^{23} ***Na^+ ions***

= 6.02 x 10^{23} ***Cl^- ions***

Formula = $CaCl_2$; 111g = ***1mole*** = 6.02 x 10^{23} ***$Ca^{2+}(Cl^-)_2$ units***

= 3 x 6.02 x 10^{23} ***ions***

= 6.02 x 10^{23} ***Ca^{2+} ions***

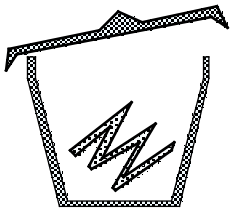
= 2 x 6.02 x 10^{23} ***Cl^- ions***

From these examples you can see that, since it is the ***formula*** of a substance that determines how you ***calculate*** the ***mass*** of ***1 mole***; it is the ***formula unit*** that there will be ***6.02 x 10^{23}*** of in ***1 mole***.

1 mole *of any substance* **=**	***gram formula mass*** *of that substance*
1 mole *of any substance* **=**	***6.02 x 10^{23} formula units***

Empirical Formula

This activity reminds you how to calculate empirical formulae.



mass of Mg = 0.84g

mass of oxide = 1.40g

mass of O = 0.56g

One of the first things that Chemists were able to do with the ***Avogadro Constant*** was work out the ***formulae*** of ***compounds*** by ***experimentation***.

For example, a *weighed* piece of ***magnesium ribbon*** could be heated in a ***crucible***, with the lid raised occasionally to allow ***oxygen*** in.

The mass of ***magnesium oxide*** produced would allow the mass of ***oxygen*** to be calculated.

Mg	***O***
mass of magnesium = 0.84 g	*mass of oxygen* = 0.56 g
24 g ⟶ 6.02 x 10^{23} Mg atoms	16 g ⟶ 6.02 x 10^{23} O atoms
0.84 g ⟶ $\frac{0.84}{24}$ x 6.02 x 10^{23} atoms	0.56 g ⟶ $\frac{0.56}{16}$ x 6.02 x 10^{23} atoms
= atoms	= atoms

The ***number ratio*** of the atoms show that the ***formula*** for ***magnesium oxide*** must be :- MgO

6. Empirical Formula

The formula of a compound was originally determined by experiment, (and the Valency rules were established later). The steps involved are:

① *determine **mass** of each element in compound*
② *use **molar masses** (atomic mass in g) to*
③ *calculate the **number** of moles of each element and*
④ *express in the **simplest ratio** possible to*
⑤ *write the **empirical formula** (simplest possible)*

The best way of setting out these five steps is in the form of a table.

Worked Example: A sample of titanium chloride is found to contain 4g of titanium and 12g of chlorine.

	Ti	***Cl***
① ***mass** of each element*	*4g*	*12g*
② ***molar mass** (atomic mass in g)*	*48g*	*35.5g*
③ ***number** of moles*	*4 / 48 = 0.083*	*12 / 35.5 =0.338*
④ ***ratio** (divide both numbers by smallest)*	*0.083 / 0.083 = 1*	*0.338 / 0.083 = 4.07*

⑤ *rounding to the nearest whole number,* ***empirical formula*** = $TiCl_4$

Most of the time the ratios will be very close to whole numbers.

*Sometimes a ratio of 1: 1.5 is obtained. This is the **simplest** ratio and should not be rounded. Instead, you should realise that the true ratio is 2:3 eg as in Al_2O_3.*

In fact, since

1 mole of any element = 6.02 x 10^{23} atoms

we don't really need to work with such large numbers at all.

In Standard Grade you were taught to calculate the ***mole ratio*** by dividing the ***mass of each element*** by the ***mass of 1 mole*** (***gfm***).

This method can be used for any number of elements, but remember that it produces the ***simplest*** (*empirical*) formula.

Glucose would, for example, come out as CH_2O. The *molecular weight* of 180 tells you that the actual molecule is 6 times bigger,ie. $C_6H_{12}O_6$.

Remember that ***percentages***, e.g. 80% Cu, 20% O are used as ***weights***; 80g Cu, 20g O

4.6 Molar Relationships 3

*This lesson looks at how to do calculations involving **volumes** of gases.*

Density of a Gas

*This activity looks at how the **density** of a gas can be used to calculate the **volume** occupied by one mole of the gas*

$$\textit{Density} = \frac{\textit{mass}}{\textit{volume}}$$

$$= \textit{g / cm}^3$$

O_2	=	0.0014 g / cm³
N_2	=	0.0013 g / cm³
F_2	=	0.0017 g / cm³
Cl_2	=	0.0032 g / cm³
Ne	=	0.0009 g / cm³
Ar	=	0.0018 g / cm³

All these densities were measured at 1 atm pressure and at 0°C

Density values for many ***gases*** are readily available in *Data Books*.

They tell us the ***mass*** of ***1 cm³*** of the gas , eg the density of ***oxygen*** is 0.00143 g / cm³

We now have a link between ***mass*** and ***volume***;

0.00143 g ⟶ 1 cm³

We also know the ***mass*** of ***1 mole*** of ***oxygen*** and so can easily calculate the ***volume*** occupied by ***a mole*** of ***oxygen*** - the ***Molar Volume***.

0.00143 g ⟶ 1 cm³

32 g ⟶ (32 / 0.00143) *x* 1 cm³

= 22 400 cm³

= 22.400 *l*

Similar ***calculations*** can be done for some of the other ***gases*** whose ***densities*** are in the *Data Book*:

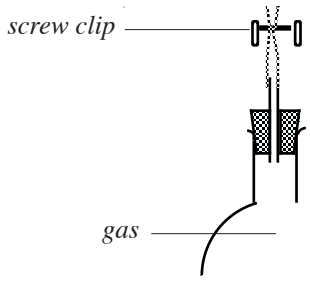
Element	***N_2***	***O_2***	***F_2***	***Cl_2***	***Ne***	***Ar***
Density (g / cm³)		0.00143				
Mass of 1 mole (g)		32				
Molar Volume (l)		***22.4***				

The most striking thing about these ***Molar Volumes*** is how ***similar*** they are. Despite quite big ***differences*** in the ***size*** and ***mass*** of their molecules/atoms, ***1 mole of each of these gases occupies the same volume as 1 mole of any other gas.***

The molecules in a gas are very ***far apart***, so most of the ***volume*** of a gas is ***empty space*** (less than 0.1% is due to the actual molecules). The ***size*** *of the molecule* has a negligible effect. More important is the ***number*** *of molecules* and since ***1 mole*** of every gas contains 6 x 10^{23} molecules it is not too difficult to see why their ***Molar Volumes are so similar***.

Molar Volumes

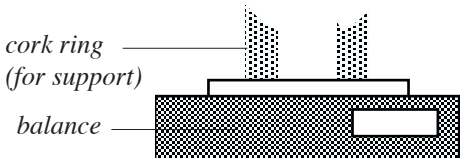
*This activity looks at how the **molar volume** of a gas can be obtained experimentally*



Any ***experiment*** that can be used to link the ***mass*** *of a gas* to the ***volume of the gas*** can be used to ***calculate*** the ***Molar Volume***.

mass ⟶ *volume*

*mass of **1 mole*** ⟶ *volume*



With a good pump it should be possible to ***remove*** all the air from a ***flask*** and ***weigh*** it ***empty***.

e.g *mass of empty flask =* ***107.49*** g

The ***flask*** could then be filled with, *for example*, ***carbon dioxide*** and reweighed:

mass of flask + carbon dioxide = ***108.37*** g

mass of carbon dioxide = 108.37 - 107.49 = ***0.88*** g

The ***flask*** would then be filled with ***water*** and the water poured into a ***measuring*** *cylinder* to determine the ***volume*** *of the flask*:

volume of flask = ***480*** cm^3

The ***Molar Volume*** can now be calculated:

	mass	⟶	*volume*
	0.88 g of CO_2	⟶	480 cm^3
*(scale up to mass of **1 mole**)*	44 g of CO_2	⟶	44 / 0.88 x 480 cm^3
		⇒	24 000 cm^3

This ***Molar Volume*** is ***higher*** than those calculated from ***densities*** because the ***densities*** are measured at ***0 °C***, wheras these results would have been obtained at ***room*** *temperature* (about ***20°C***).

Molar Volumes are dependant on ***temperature*** and ***pressure***; it is not a *constant* value like the *Avogadro Constant*.

However, for ***any*** gases at the ***same temperature*** and the ***same pressure***;

1 mole of any gas occupies the same volume as 1 mole of any other gas.

This is an extremely useful relationship when dealing with ***reactions involving gases***.

Reacting Volumes

*This activity looks at how the **molar volume** of a gas can be used in calculations for reactions involving gases.*

Up to now, we have treated ***gases*** the same as other substances and ***calculated*** their ***mass***.

e.g. What ***mass*** of *carbon dioxide* is produced when 64g of *methane* is burned in a plentiful supply of air ?

Step 1	CH_4 +	2 O_2	→	CO_2	+	2 H_2O
Step 2	***1****mole*	⟶		1 *moles*		
Step 3	*16g*	⟶		44 *g*		
Step 4	*64g*	⟶		*(64 / 16) x* 44		*g*
				= 176		*g*

The ***volume of a gas*** is a much more ***useful quantity*** to deal with, so

e.g. What ***volume*** of *carbon dioxide* is produced when 64g of *methane* is burned in a plentiful supply of air ? (Take the molar volume to be 23.2 *l*).

Step 1	CH_4 +	2 O_2	→	CO_2	+	2 H_2O
Step 2	***1****mole*	⟶		1 *moles*		
Step 3	*16g*	⟶		23.2 *l*		
Step 4	*64g*	⟶		*(64 / 16) x* 23.2		*l*
				= 92.8		*l*

4.7 Molar Relationships 4

*This lesson looks at how to do calculations when most of the reactants/products are gases, and intoducesthe idea of **excess** and calculating excess.*

Avogadro's Law

*This activity looks at how the **molar volumes** of gases can be used to do calculations involving **volumes only**.*

Avogadro's Law - ***Equal volumes*** of all ***gases*** at the ***same temperature*** and ***same pressure*** ,will contain an ***equal number*** of molecules.

1 mole of any gas occupies the same volume as 1 mole of any other gas.

In practice, this means that we can ***replace moles*** with ***volumes*** in any relationship involving gases, as long as they are at the ***same temperature*** and the ***same pressure***.

For example, consider the manufacture of *ammonia* from ***nitrogen*** and ***hydrogen***.

N_2	+	$3\ H_2$	→	$2\ NH_3$
1 *moles*		3 *moles*		2 *moles*
1 *volumes*		3 *volumes*		2 *volumes*

For any volume of N_2, e.g. ***100cm³*** , we would need 300 ***cm³*** of H_2 and would be able to produce 200 ***cm³*** of ammonia.

e.g. What ***volume*** of *carbon dioxide* is produced when 100cm³ of *methane* is burned in a plentiful supply of air ?

	CH_4 +	$2\ O_2$ →	CO_2 +	$2\ H_2O$
Step 2	***1****mole*	⟶	2 *moles*	
Step 3	***1****volume*	⟶	2 *volumes*	
Step 4	*100 cm³*	⟶	200 *cm³*	

Gases In Excess

*This activity looks at how to calculate the **limiting reactant** and the **reactant in excess**.*

Up to know you have been allowed to ***assume*** that ***reactants*** were always in the ***correct proportions*** to ***react completely***, or phrases like '*is burned in a plentiful supply of air*' have reassured you that the reactant you are interested in will ***react completely***.

Often there will be ***less of one reactant*** than is ***needed*** to allow ***all*** the reactants to ***react completely***:- *one of the reactants will run out first and stop the reaction.* This reactant is called the ***limiting reactant*** because it ***limits*** the amount of product that can be made.

Any reactant ***left over*** at the end is said to be ***in excess***.

e.g. A mixture of 20 cm³ of ethane and 120 cm³ oxygen was ignited and than allowed to cool. Calculate the volume and composition of the gas mixture that would be left at the end. All volumes are measured at room temperature and pressure.

Step 1 Write the balanced equation and establish the mole relationships

C_2H_6	+	$^7/_2\ O_2$	→	$2\ CO_2$	+	$3\ H_2O$
1 *moles*		$^7/_2$ *moles*		2 *moles*		3 *moles*

Step 2 Using Avogadro's Law we can rewrite the relationships as volumes

1 *volumes*	$^7/_2$ *volumes*	2 *volumes*

Step 3 By inspection, or calculation if necessary, identify the limiting reactant and the reactant in excess.

$$C_2H_6 + {}^7/_2\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$$

1 *volumes* ⟶ $^7/_2$ *volumes*

20 cm^3 ⟶ 70 cm^3

limiting reactant = *ethane*
reactant in excess = *oxygen*
amount in excess = 120 - 70 = 50 cm^3

Step 4 Use the volume of the limiting reactant to find the volume of CO_2 produced

$$C_2H_6 + {}^7/_2\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$$

1 *volumes* ⟶ 2 *volumes*

20 cm^3 ⟶ 40 cm^3

The resulting mixture will contain 50 cm^3 *of* O_2 *and* 40 cm^3 *of* CO_2

Variations include; *calculating volumes at, for example, 120°C (water is now also a gas) and then re-calculating at room temperature (volume of water 'disappears')*

passing mixture through $NaOH_{(aq)}$,(or some other alkali), removes the acidic CO_2.

Other Chemicals in Excess

*This activity looks at how to calculate the **limiting reactant** and the **reactant in excess** for non-gases.*

e.g. Calculate the mass of sodium chloride produced when 25 cm^3 of 0.1 mol l^{-1} sodium hydroxide reacts with 10 cm^3 of 0.2 mol l^{-1} hydrochloric acid.

Step 1 Write the balanced equation and establish the mole relationships

$$NaOH + HCl \rightarrow NaCl + H_2O$$
1 *moles* 1 *moles* 1 *moles* 1 *mole*

Step 2 Calculate the number of moles of each reactant

Step 3 By inspection, or calculation if necessary, identify the limiting reactant and the reactant in excess.

$$NaOH + HCl \rightarrow NaCl + H_2O$$
0.0025 *moles* 0.0020 *moles*

limiting reactant = *HCl*

Step 4 Use the moles of the limiting reactant to find the mass of NaCl produced

$$NaOH + HCl \rightarrow NaCl + H_2O$$

1 *moles* → 1 *moles* = 58.5 *g*

0.0020 *moles* ⟶ *g*

UNIT 1. ENERGY MATTERS

Section 4: Enthalpy / Moles

Potential energy diagrams

1. ***Exothermic*** *changes cause heat to be released to the surroundings*

2. ***Endothermic*** *changes cause absorption of heat from the surroundings*

3. *The* ***enthalpy change*** *is the energy difference between products and reactants*

   $\Delta H = H(products) - H(reactants)$

4. *The enthalpy change can be calculated from potential energy diagrams*

5. *The enthalpy change has a* ***negative*** *value for exothermic reactions*

   $\Delta H = H(products) - H(reactants)$

6. *The enthalpy change has a* ***positive*** *value for endothermic reactions*

   $\Delta H = H(products) - H(reactants)$

Enthalpy Changes

7. *The* ***enthalpy of combustion*** *of a substance is the enthalpy change when* ***one mole*** *of the substance burns completely in oxygen*

   e.g. $CH_3OH + {}^3/_2O_2 \rightarrow CO_2 + 2H_2O$

   ΔH = -ve, always exothermic

8. *The* ***enthalpy of solution*** *of a substance is the enthalpy change when* ***one mole*** *of the substance dissolves in water*

   e.g. $NaOH + aq \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$

   ΔH = -ve or +ve, exothermic or endothermic

9. *The* ***enthalpy of neutralisation*** *of an acid is the enthalpy change when the acid is neutralised to form one mole of water*

   e.g. ${}^1/_2H_2SO_4 + LiOH \rightarrow {}^1/_2Li_2SO_4 + H_2O$

   ΔH = -ve, always exothermic

10. *The enthalpy changes can be calculated using* $\Delta H = c\ m\ \Delta T$

    If water is used; $c = 4.18\ kJ\ kg^{-1}\ °C^{-1}$

The Avogadro Constant

11. *One mole of any substance contains* 6.02×10^{23} *formula units*

12. *Equimolar amounts of substances contain equal numbers of formula units*

Molar Volume

13. *The molar volume (in units of* $dm^3\ mol^{-1}$*) is the same for all gases at the same temperature and pressure*

14. *The volume of a gas can be calculated from the number of moles and vice versa.*

Reacting Volumes

15. *The volumes of reactant and product gases can be calculated from the number of moles of each reactant and product*

The idea of Excess

16. *The reactant that is in excess can be calculated*