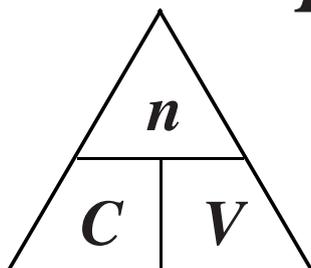
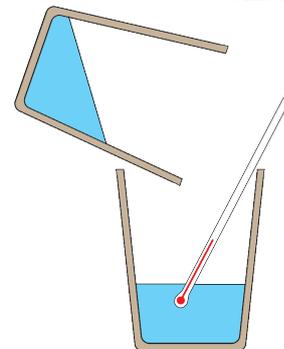


Higher Chemistry



Unit 3:

Principles to Production - Part 2

Student: _____

Lesson	Activities	Done	Checked
3.8 Molar Relationships 1	1. Solids, liquids & gases		
	2. Solutions		
	3. Equations		
	Check Test	Score: / 8	
	Home Practice	Score: / 10	
3.9 Energy Changes	1. Energy out, energy in		
	2. Where does the energy come from ?		
	3. Important Enthalpies		
	Check Test	Score: / 8	
	Home Practice	Score: / 10	
3.10 Enthalpies from Experimental data	1. Specific Heat Capacity		
	2. Enthalpy of Solution		
	3. Enthalpy of Neutralisation		
	Check Test	Score: / 8	
	Home Practice	Score: / 10	
3.11 Enthalpy of Combustion	1. Burning Ethanol		
	2. Processing Results		
	3. Comparing Enthalpies		
	4. Bond Enthalpies		
	Check Test	Score: / 8	
Home Practice	Score: / 10		
3.12 Hess's Law Calculations	1. Using Diagrams		
	2. Using Algebra		
	Check Test	Score: / 8	
	Home Practice	Score: / 10	
	3.13 Molar Relationships 2	1. The Avogadro Constant	
2. Formula Units			
3. Empirical Formula			
Check Test		Score: / 8	
Home Practice		Score: / 10	
3.14 Molar Relationships 3	1. Density of a Gas		
	2. Molar Volumes		
	3. Reacting Volumes		
	Check Test	Score: / 8	
	Home Practice	Score: / 10	
3.15 Molar Relationships 4	1. Avogadro's Law		
	2. Gases in Excess		
	3. Other Chemicals in Excess		
	Check Test	Score: / 8	
	Home Practice	Score: / 10	
Consolidation Work	Consolidation A	Score: / 10	
	Consolidation B	Score: / 10	
	Consolidation C	Score: / 10	
	Consolidation D	Score: / 10	
	End-of-Section Assessment	Score: _____ %	Grade: _____

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

3.8 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 *for* of a substance you can calculate its *for ma* in amu (atomic mass units; mass of 1 proton = 1amu).

$$\begin{array}{rclclclclclcl} \text{e.g.} & \text{CaCO}_3 & = & 1 \times \text{Ca} & = & 1 \times & = & & & & \\ & & & 1 \times \text{C} & = & 1 \times & = & & & & \\ & & & 3 \times \text{O} & = & 3 \times & = & & & & \\ & & & & & & & & & & = \text{amu} \end{array}$$

By definition,

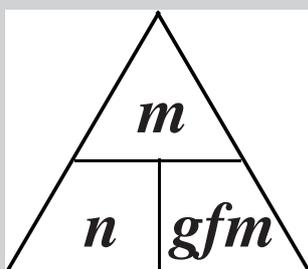
$$1 \text{ mole} = \text{for ma in gra} \quad \text{or gramme formula mass } \mathbf{gfm}$$

$$\text{e.g.} \quad 1 \text{ mole of CaCO}_3 = \text{g}$$

Remember that *1 mo* of any substance contains the *sa amount* of substance as *1 mo* of any other substance.

$$\begin{array}{rclclclclcl} 1 \text{ mole of CaCO}_3 & = & 1 \text{ mole of C}_6\text{H}_{12}\text{O}_6 & = & 1 \text{ mole of (NH}_4)_2\text{SO}_4 \\ \text{g} & = & \text{g} & = & \text{g} \end{array}$$

$$\begin{array}{rclclcl} \mathbf{moles \text{ of a substance}} & = & \mathbf{mass \text{ of substance}} & \div & \mathbf{gfm} \\ \mathbf{n} & = & \mathbf{m} & \div & \mathbf{gfm} \end{array}$$



$$\begin{array}{rclcl} \mathbf{mass \text{ of substance}} & = & \mathbf{moles \text{ of a substance}} & \times & \mathbf{gfm} \\ \mathbf{m} & = & \mathbf{n} & \times & \mathbf{gfm} \end{array}$$

$$\begin{array}{rclcl} \text{e.g.} & 25\text{g of CaCO}_3 & \mathbf{n} & = & \mathbf{m} \div \mathbf{gfm} \\ & & & = & 24 \div 100 \\ & & & = & 0.25 \text{ moles} \end{array}$$

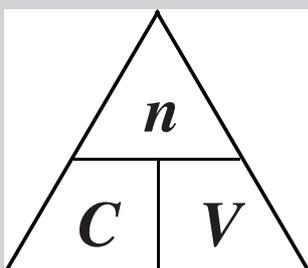
Solutions

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 **conc** of a solution (**C**) is measured in **mo per litre**, abbreviated to **mol l⁻¹**.
The **volume** of a solution (**V**) is measured in cubic centimetres (cm³) or in **litres (l)**.

$$1000 \text{ cm}^3 = 1 \text{ litre} = 1000 \text{ ml}$$

$$\begin{array}{l} \text{moles in a solution} = \text{concentration} \times \text{volume} \\ n = C \times V \end{array}$$



$$\begin{array}{l} \text{concentration} = \text{moles dissolved} \div \text{volume} \\ C = n \div V \end{array}$$

e.g. 250 cm³ of a 2 mol l⁻¹ solution of hydrochloric acid.

$$\begin{array}{l} n = C \times V \\ = 2 \times \quad \quad \quad \text{(volumes must be in litres)} \\ = \quad \quad \quad \text{moles of HCl} \quad (= \quad \quad \quad \text{moles of H}^+ \text{ ions)} \\ = \quad \quad \quad \text{g of HCl} \quad \quad \quad \text{(gfm} = 36.5\text{g)} \end{array}$$

250 cm³ of a 1 mol l⁻¹ solution of sulphuric acid acid.

$$\begin{array}{l} n = C \times V \\ = 1 \times \quad \quad \quad \text{(volumes must be in litres)} \\ = \quad \quad \quad \text{moles of H}_2\text{SO}_4 \quad (= \quad \quad \quad \text{moles of H}^+ \text{ ions)} \\ = \quad \quad \quad \text{g of H}_2\text{SO}_4 \quad \quad \quad \text{(gfm} = 98\text{g)} \end{array}$$

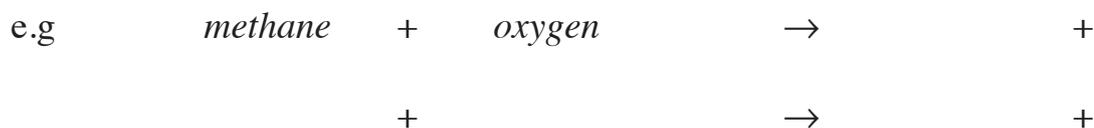
e.g. 11g of carbon dioxide dissolved in 500 cm³

$$\begin{array}{l} \text{gfm of CO}_2 = 44\text{g} = 1 \text{ mole} \\ \quad \quad \quad 11\text{g} = \quad \quad \quad \text{mole} \\ \\ C = n \div V \\ = \quad \quad \quad \div 0.5 \quad \quad \quad \text{(volumes must be in litres)} \\ = \quad \quad \quad \text{moles per litre} \\ = \quad \quad \quad \text{mole l}^{-1} \text{ or} \quad \quad \quad \text{M} \end{array}$$

Equations

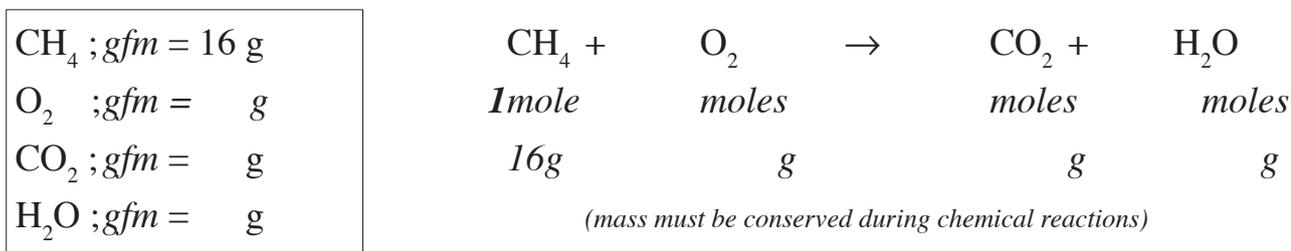
This activity looks at how the **equation** for a reaction can be used to calculate reacting masses.

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



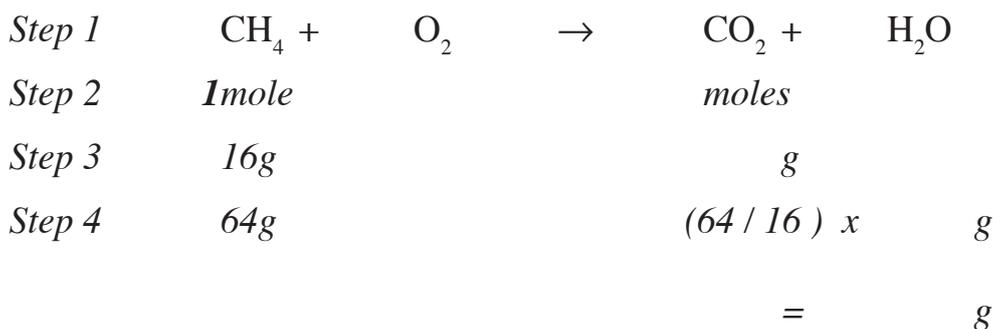
The **number** written in **front** of each **formula** tells you the **number** of each molecule involved in the **reaction**. This equation tells you that **one molecule of methane** reacts with **two molecules of oxygen** to form **one molecule of carbon dioxide** and **two molecules of water**.

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



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 ?



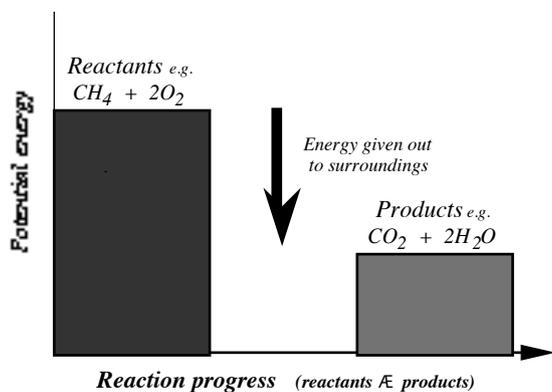
Step 4 involves the use of **proportion**; which is an essential mathematical ability for anyone wishing to study Chemistry.

3.9 Energy Changes

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

Energy out, energy in

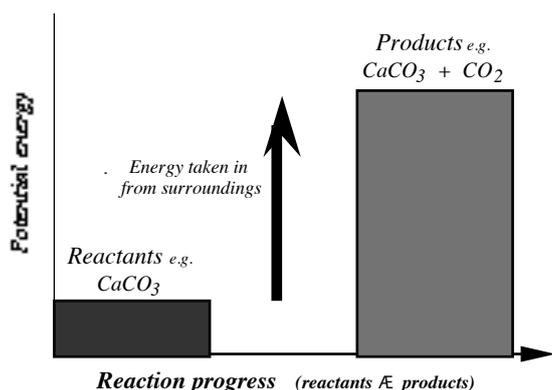
This activity looks at *exothermic* and *endothermic* reactions.



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 energy*, 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 energy*, and get *colder*.

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 energy changes*, ΔH .

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

This means that:

exothermic reactions $\Delta H = -ve$

endothermic reactions $\Delta H = +ve$

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 *rea* molecules have *kin* energy and *pot* energy.

The *pot* energy is mainly energy *sto* within their *bo*.

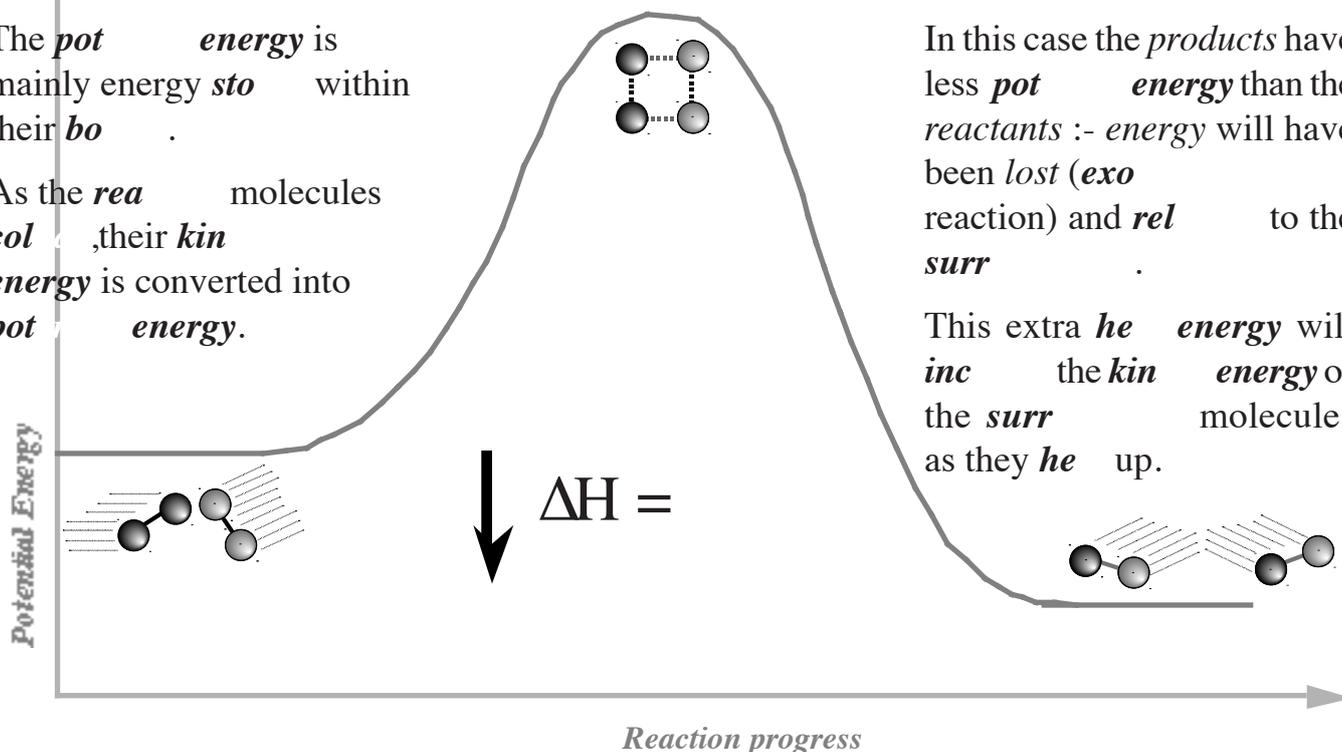
As the *rea* molecules *col*, their *kin* energy is converted into *pot* energy.

This *extra* energy can be enough to *br* existing bonds and allow *new* bonds to *fo*.

The *bonds* in the *pro* molecules will be *diff* from the original *rea* molecules.

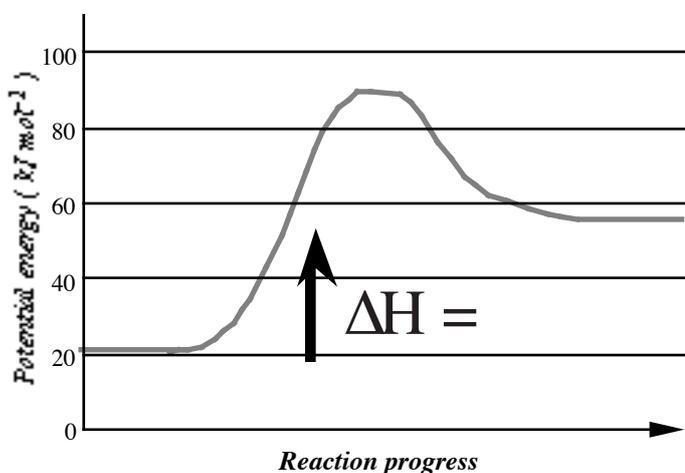
In this case the *products* have less *pot* energy than the *reactants* :- energy will have been *lost* (*exo* reaction) and *rel* to the *surr*.

This extra *he* energy will *inc* the *kin* energy of the *surr* molecules as they *he* up.



In effect, *kin* energy is used to *br* as *new* bonds are *ma*. During *exo* taken in. This energy was originally *st* ing extra *kin* energy to other *reac* to *he* up.

existing bonds and then *kin* energy is *rel* *reactions* more *en* is *rel* than was in the bonds of the *reac* but is now provid- molecules, and the *surr*, causing them



During *endo* reactions, *kin* energy is again used in *br* existing bonds.

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

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

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

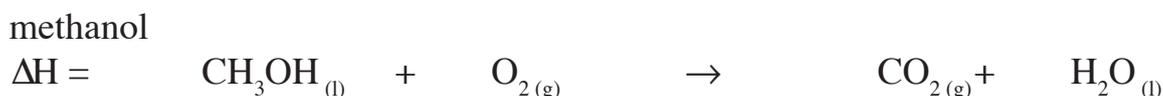
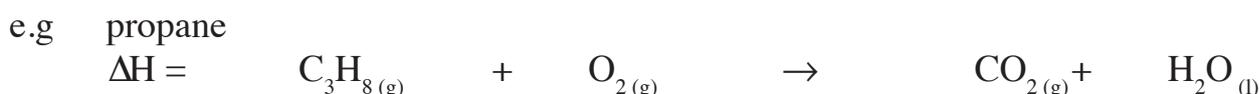
Important Enthalpies

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

Any **che** process can have an **enth of reaction**, but certain types of reactions are deemed important enough to be **na** and defined carefully. **Val** for many of these processes are to be found in *Data Books*.

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

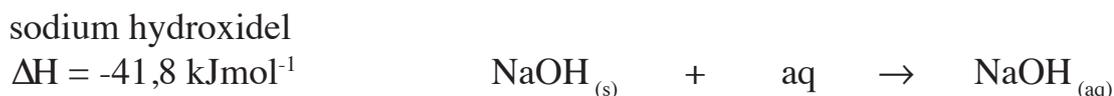
The **enthalpy of combustion** of a substance is the enthalpy change when **one mole** of the substance **burns completely** in oxygen



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

Notice that **enth of comb** are *always exo*, ie. their ΔH values are *always neg*.

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



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

Notice that **enth of sol** can be *either exo*, ΔH **neg**, or **endo**, ΔH **pos**.

Once **diss** the **ionic net**, *crystal lattice*, is completely broken so the *individual ions* are now only **bon** to **wa** molecules. For this reason, these **equ** are more often written as:-

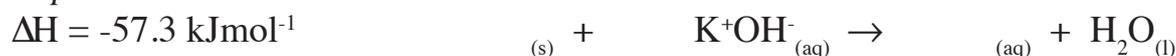


The **enthalpy of neutralisation** of an acid is the enthalpy change when the acid is neutralised to form **one mole of water**.

e.g. hydrochloric acid / NaOH



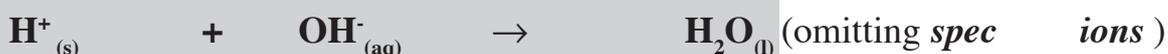
sulphuric acid / KOH



Notice, again, that there is often a need to use '**fra**' of **mo** in order to **bal** these equations in terms of **one mole of wa**, as defined.

Notice that **enth** of **neut** are always **exo**, ΔH **neg**, and that the value is the **sa** for many combinations of **ac** / **alk**.

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



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} .

3.10 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

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

During **chem** reactions **en** is **rel** to the **surr**, **thermic**, or **tak in** from the **surr**, **thermic**.

Often, the **surr** are the **wa** that the chemicals are **diss** 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**, $c = 4.18 \text{ kJ kg}^{-1} \text{ °C}^{-1}$

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

This tells us how much *en* was *rel* when **1g** of *sodium hydroxide* *diss* in water. We need to *calc* how much *en* would have been released if **1 mole** of *sodium hydroxide* had *diss*.

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 *exper* values are always *slightly inacc*, because,

- we assume that the *specific heat capacity* for the *sol* is the *same* as *pu water*, 4.18 kJ kg⁻¹ °C⁻¹
- we assume that the *vol* of the *sol* is the *same* as the *wa* we started with, 50 cm³, and that **1cm³ = 1g** as for *pu water*.
- we assume that all the *heat rel* stays in the *wa*; that none *esc* to the air etc.

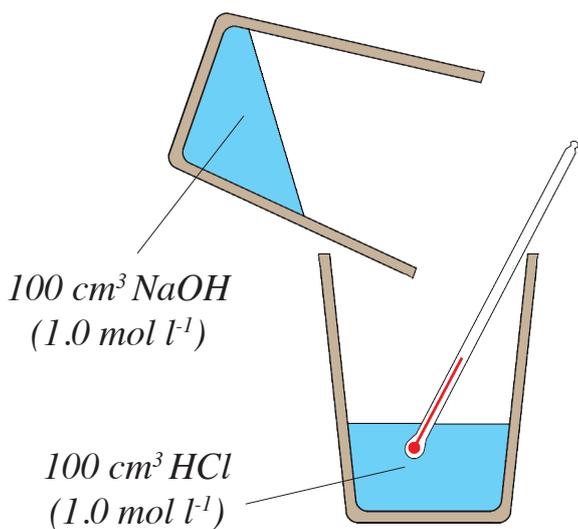
Enthalpy of Neutralisation

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

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

typical results

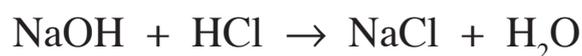
There was a *temp* rise of **6.5 °C**.



$$\begin{aligned} \text{Energy released} &= \mathbf{c \times m \times \Delta T} \\ \text{into water} & \\ &= \mathbf{x \times x} \\ &= \mathbf{kJ} \end{aligned}$$

These *vol* and *conc* mean that **moles** of *NaOH* react with **moles** of *HCl* to produce **moles** of *H₂O*.

We simply need to multiply by **to** *calc* the *en* **released** when **1 mole** of *H₂O* is formed.



$\Delta H_{\text{neutralisation}}$ = ***kJ mol⁻¹***

3.11 Enthalpy of Combustion

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

Eth can be **bur** in a **sp burner**.
The **he released** (**thermic**) must be
abs by **wa** placed in a beaker, or, if
possible, a **co can**, just above the **fl**.
(**Co** is an excellent **con**).

H proof mats can be used to provide
a **dr shield** around the apparatus; to
keep the **fl** steady and cut down **he**
loss from the experiment.

A **ther** is used to measure the
temp change as the **wa heats up**.

* Draw a labelled diagram of the assembled apparatus

Enough **eth** is burnt to produce a **°C temp rise** (approximately), and the **am**
of **eth** 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	—		—
anol		CH_2		
an-1-ol				
an-1-ol	$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$		- 2673	

The *ol* family is another example of a *homo series* - each *mem* differs from the previous *mem* by the *sa amount* - by a CH_2 .

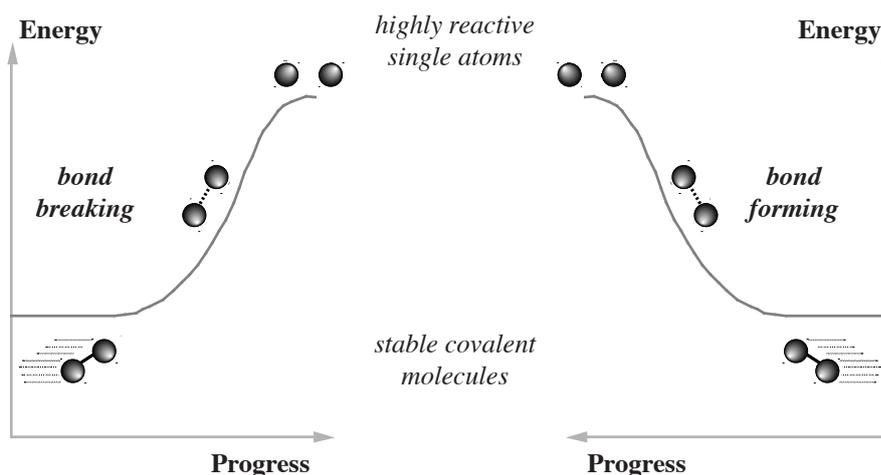
There is also a fairly *con difference* in their *enth values*. This supports the idea that it is energy *sto in bonds* that is *rel* during *com* reactions.

Bond Enthalpies

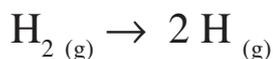
This activity looks at how Bond Enthalpies can be used to calculate the Enthalpies of certain reactions.

In your data book there are tables of *Bond Enthalpies*.

Bond	Enthalpy kJ mol^{-1}
H - H	432
O = O	497
N = N	941
F - F	155
Cl - Cl	243
Br - Br	194
I - I	149
H - F	569
H - Cl	428
H - Br	362
H - I	295

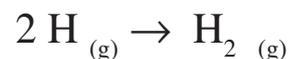


e.g. to *break* the H—H bond.



$$\Delta H = + 432 \text{ kJmol}^{-1}$$

e.g. to *form* the H—H bond.



$$\Delta H = - 432 \text{ kJmol}^{-1}$$

Bond Breaking - Endothermic

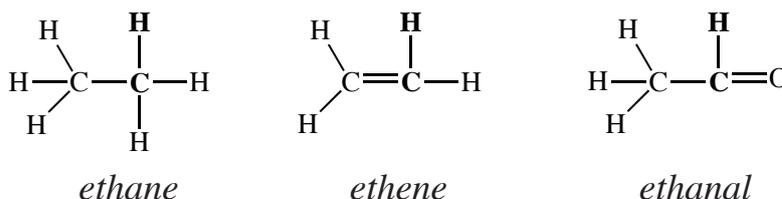
Bond Forming - Exothermic

Bond	Mean enthalpy /kJ mol ⁻¹
Si – Si	222
C – C	346
C = C	602
C ≡ C	835
$\left. \begin{array}{l} \text{C} \equiv \text{C} \\ \text{(aromatic)} \end{array} \right\}$	519
H – O	458
H – N	387
C – H	414
C – O	358
C = O	798
C – F	486
C – Cl	326
C – Br	285
C – I	213

In the previous table, all the substances were complete *covalent molecules* and each bond listed would be *unique* to that particular molecule.

In more complex covalent molecules, other types of covalent bonds will be found. e.g. C—H and O—H.

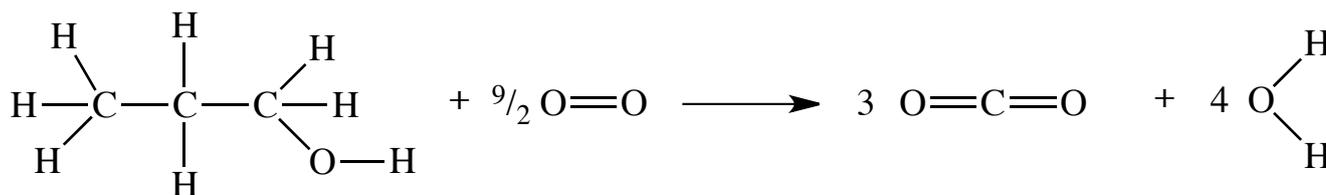
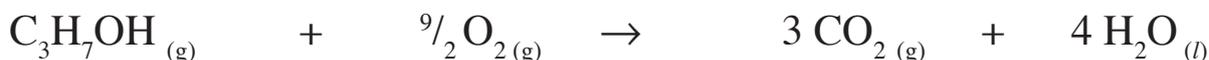
However, there are many slightly different C—H bonds, depending on their exact position in a molecule, e.g.



Notice, therefore, that the values in this table are *mean bond enthalpies*.

With reactions involving *only covalent molecules*, it will be possible to calculate the *enthalpy of reaction* using *bond enthalpies* only. It can be useful to write (*or think*) the equation for the reaction with *full structural formulae*.

For example, to calculate the *enthalpy of combustion* of *propan-1-ol*.



BOND BREAKING

$$2 \times \text{C}-\text{C} = 2 \times \quad =$$

$$7 \times \text{C}-\text{H} = 7 \times \quad =$$

$$1 \times \text{C}-\text{O} = 1 \times \quad =$$

$$1 \times \text{O}-\text{H} = 1 \times \quad =$$

$$\frac{9}{2} \times \text{O}=\text{O} = \frac{9}{2} \times \quad =$$

$$\text{Total} = + \quad \text{kJ}$$

BOND FORMING

$$6 \times \text{C}=\text{O} = 6 \times \quad =$$

$$8 \times \text{O}-\text{H} = 8 \times \quad =$$

$$\text{Total} = - \quad \text{kJ}$$

Bond Breaking - Endothermic
Bond Forming - Exothermic

enthalpy of combustion of *propan-1-ol* =

kJ mol⁻¹

3.12 Hess's Law Calculations

Using Diagrams

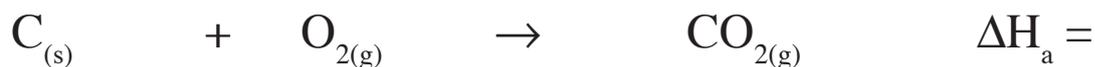
This activity shows how enthalpies of reactions can be calculated using other experimentally determined enthalpies.

Only certain enthalpies can be *measured* directly. Hess's Law allows us to *calculate* enthalpy changes that are very difficult or even impossible to measure.

For example, carbon burns to produce carbon monoxide but it is impossible to burn carbon in such a way that the *only* product is carbon monoxide. ie. we *cannot measure* the ΔH for this reaction:



We can, however, *measure* the enthalpies of these two reactions (values in Data Book).



There are two routes that we can use to convert $\text{C}_{(s)} \rightarrow \text{CO}_{2(g)}$

A diagram can be drawn to show these two routes:

Using the diagram, and Hess's Law, it is clear that:

$$\Delta H_a = \Delta H_c + \Delta H_b$$

Putting in the values we know, will allow us to *calculate* ΔH_c

An extremely useful enthalpy to know is called the enthalpy of formation of a substance: by comparing enthalpies of formation chemists learn a lot about relative stabilities and the energy changes needed to change one substance into another,

*The **enthalpy of formation** of a substance is the enthalpy change when **one mole** of the substance **is formed** from its elements in their natural state*



Most of these enthalpies, particularly of organic molecules, cannot be measured experimentally; either the reactions are impossible or there will be a mixture of products.

Many of the enthalpies can be measured. Notice that the enthalpy of formation of carbon dioxide is also the enthalpy of combustion of carbon.

Again a diagram can be constructed that links a reaction we cannot measure;



to reactions that can be measured;



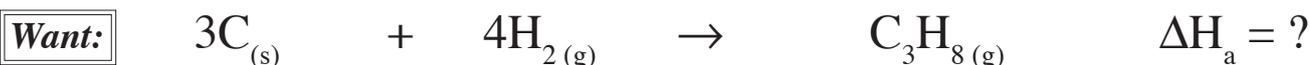
Using Algebra

This activity shows an alternative method for calculating enthalpies of reactions.

Sometimes, it can be difficult to see immediately a diagram that links all the reactions together to form two *different routes*.

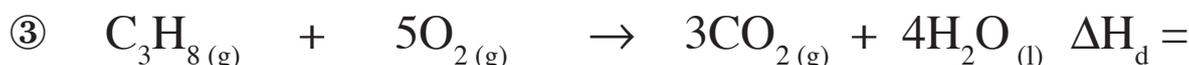
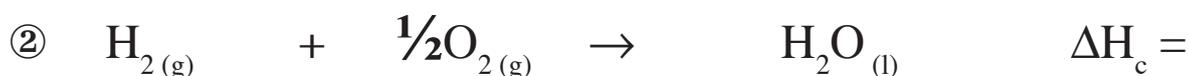
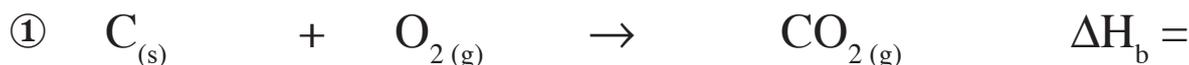
An alternative (more *algebraic*) approach has the advantage that you can follow a series of simple steps and the correct answer will '*pop out*' at the end. The disadvantage is that it can be slower.

Step 1: Write (if not given) an equation that correctly represents the reaction whose enthalpy you *want to calculate*:



Step 2: Write (if not given) equations that correctly represent the reactions whose enthalpies you have been *directed to use* (often in *Data Book*). It can be useful to number these equations:

Given:



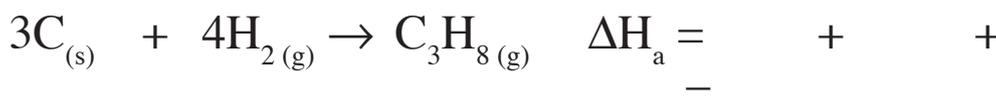
Step 3: Taking each equation in turn (no need to visualise the overall picture) decide how it needs to be *rearranged* in order to match up exactly with the '*Want*' equation :

Rearranging:



Step 4: After *cancelling* identical substances, *adding* produces the '*Want*' equation :

Adding:



3.13 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 cover Great 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 **pa** of socks, a **do** eggs, a **gr** (144) of nails and a **re** (500) of paper.

The **sm** the object, the **lar** the number of objects in the group.

Since **at** are so **sm** we use a very large number, 602252000000000000000000, in our group called a **mo**.

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

$$\text{Avogadro Constant, } L = 6.02 \times 10^{23}$$

One of the **qua** available to early **Che**, including the Italian **Amadeo Avogadro** (1776 - 1856), were **atomic ma**.

The element **Car** was thought of as 'special' because its **ato mass** was very close to a **who 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 = 6.02×10^{23} 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**

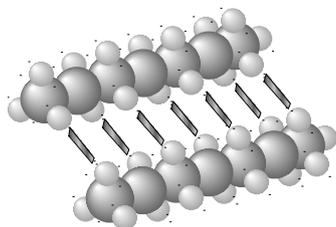


Formula = Ar ; 40g = **1mole** = 6.02×10^{23} **atoms**

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

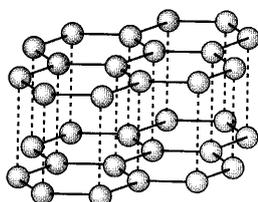


$$\begin{aligned} \text{Formula} = \text{O}_2; \quad 32\text{g} = 1\text{mole} &= 6.02 \times 10^{23} \text{ molecules} \\ &= 2 \times 6.02 \times 10^{23} \text{ atoms} \\ &= 1.204 \times 10^{24} \text{ atoms} \end{aligned}$$



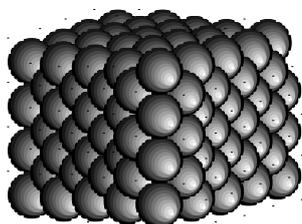
$$\begin{aligned} \text{Formula} = \text{C}_7\text{H}_{16}; \quad 100\text{g} = 1\text{mole} &= 6.02 \times 10^{23} \text{ molecules} \\ &= 23 \times 6.02 \times 10^{23} \text{ atoms} \\ &= 7 \times 6.02 \times 10^{24} \text{ C atoms} \\ &= 16 \times 6.02 \times 10^{24} \text{ H atoms} \end{aligned}$$

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



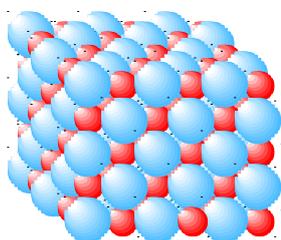
$$\text{Formula} = \text{C}; \quad 12\text{g} = 1\text{mole} = 6.02 \times 10^{23} \text{ atoms}$$

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



$$\text{Formula} = \text{Cu}; \quad 63.5\text{g} = 1\text{mole} = 6.02 \times 10^{23} \text{ atoms}$$

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



$$\begin{aligned} \text{Formula} = \text{NaCl}; \quad 58.5\text{g} = 1\text{mole} &= 6.02 \times 10^{23} \text{ Na}^+\text{Cl}^- \text{ units} \\ &= 2 \times 6.02 \times 10^{23} \text{ ions} \\ &= 6.02 \times 10^{23} \text{ Na}^+ \text{ ions} \\ &= 6.02 \times 10^{23} \text{ Cl}^- \text{ ions} \end{aligned}$$

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

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

3.14 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

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

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

$$O_2 = 0.0014 \text{ g / cm}^3$$

$$N_2 = 0.0013 \text{ g / cm}^3$$

$$F_2 = 0.0017 \text{ g / cm}^3$$

$$Cl_2 = 0.0032 \text{ g / cm}^3$$

$$Ne = 0.0009 \text{ g / cm}^3$$

$$Ar = 0.0018 \text{ g / cm}^3$$

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

Den values for many **gas** are readily available in *Data Books*.

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

We now have a link between **ma** and **vol** ;

$$0.00143 \text{ g} \longrightarrow 1 \text{ cm}^3$$

We also know the **ma** of **1 mole** of **ox** and so can easily calculate the **vol** occupied by **a mole** of **ox** - the **Molar Volume**.

$$\begin{aligned} 0.00143 \text{ g} &\longrightarrow 1 \text{ cm}^3 \\ 32 \text{ g} &\longrightarrow \left(\frac{32}{0.00143} \right) \times 1 \text{ cm}^3 \\ &= \text{cm}^3 \\ &= \text{l} \end{aligned}$$

Similar **cal** can be done for some of the other **ga** whose **den** are in the *Data Book*:

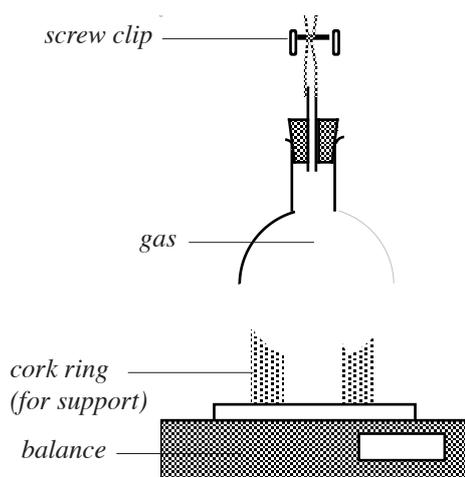
Element	N_2	O_2	F_2	Cl_2	Ne	Ar
Density (g / cm ³)		0.00143				
Mass of 1 mole (g)		32				
Molar Volume (l)		22.4				

The most striking thing about these **Mo Vol** is how **sim** they are. Despite quite big **diff** 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 ap**, so most of the **vol** of a gas is **em space** (less than 0.1% is due to the actual molecules). The **size of the molecule** has a negligible effect. More important is the **num of molecules** and since **1 mole** of every gas contains \times molecules it is not too difficult to see why their **Mo Vol are so similar**.

Molar Volumes

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



Any *exp* that can be used to link the *ma* of a gas to the *vol* of the gas can be used to *cal* the *Mo Volume*.
 $mass \longrightarrow volume$
 $mass \longrightarrow volume$
 of 1 mole

With a good pump it should be possible to *rem* all the air from a *fl* and *weigh* it *emp*.
 e.g. $mass\ of\ empty\ flask = 107.49\ g$

The *fl* 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 *fl* would then be filled with *wa* and the water poured into a *mea* cylinder to determine the *vol* of the flask:

$$volume\ of\ flask = 480\ cm^3$$

The *Mo Volume* can now be calculated:

	$mass \longrightarrow$	$volume$
	$0.88\ g\ of\ CO_2 \longrightarrow$	$480\ cm^3$
<i>(scale up to mass of 1 mole)</i>	$44\ g\ of\ CO_2 \longrightarrow$	$44 / 0.88 \times 480\ cm^3$
		$\Rightarrow 24\ 000\ cm^3$

This *Mo Volume* is *hi* than those calculated from *den* because the *den* are measured at $0\ ^\circ C$, whereas these results would have been obtained at *ro* temperature (about $20\ ^\circ C$).

Mo Volumes are dependant on *temp* and *pre*; it is not a *constant* value like the *Avogadro Constant*.

However, for *any* gases at the *same temp* and the *same pre*;

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

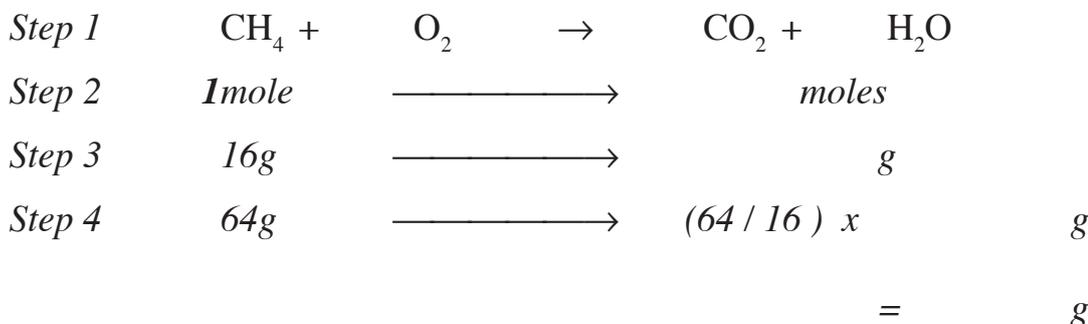
This is an extremely useful relationship when dealing with *reactions involving gases*.

Reacting Volumes

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

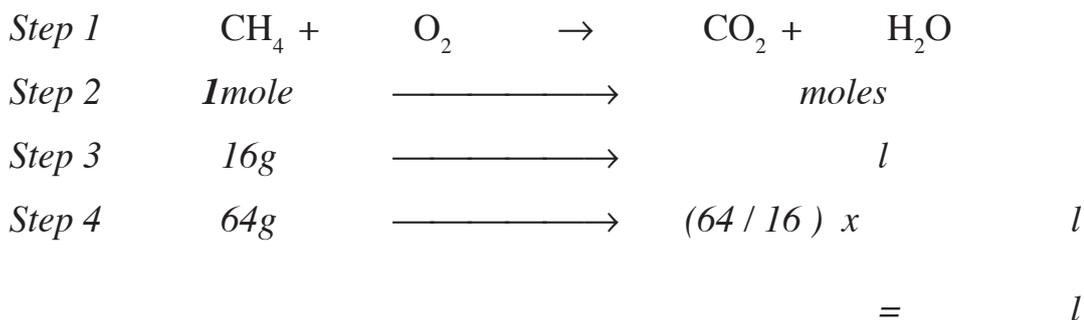
Up to now, we have treated **gas** the same as other substances and **calculate** their **mass**.

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



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).

**3.15 Molar Relationships 4**

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

Avogadro's Law

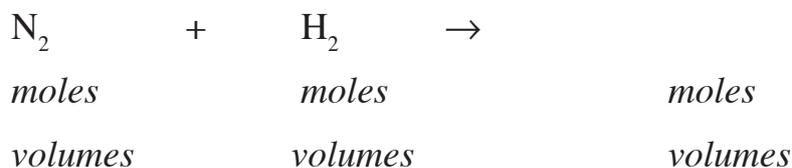
This activity looks at how the **molar volumes** of gases can be used to do calculations involving **volumes only**.

Avogadro's Law - **Equal volume** 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.

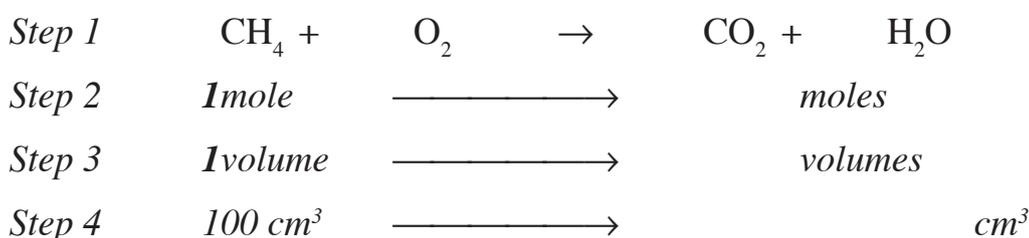
In practice, this means that we can **replace moles** with **vol** in any relationship involving gases, as long as they are at the **same temp** and the **same pre**.

For example, consider the manufacture of *ammonia* from **nit** and **hyd**.



For any volume of N_2 , e.g. 100cm^3 , we would need cm^3 of H_2 and would be able to produce cm^3 of ammonia.

e.g. What **volume** of *carbon dioxide* is produced when 100cm^3 of *methane* is burned in a plentiful supply of air?



Gases In Excess

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

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

Often there will be **le of one reactant** than is **nee** to allow **all** the reactants to **react comp** :- *one of the reactants will run out first and stop the reaction*. This reactant is called the **lim reactant** because it **lim** the amount of product that can be made.

Any reactant **le over** at the end is said to be **in ex**.

e.g. A mixture of 20 cm^3 of ethane and 120 cm^3 oxygen was ignited and then 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.

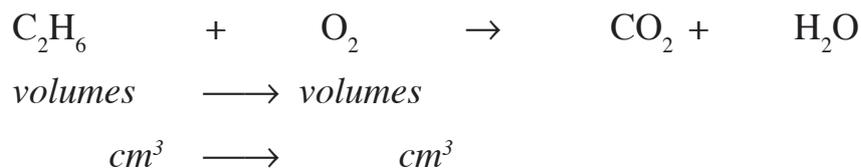
Step 1 Write the balanced equation and establish the mole relationships



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

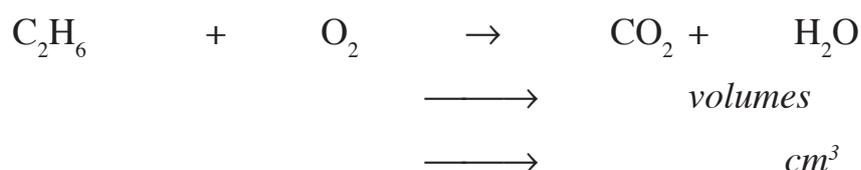


Step 3 By inspection, or calculation if necessary, identify the limiting reactant and the reactant in excess.



limiting reactant =
 reactant in excess =
 amount in excess = - = cm^3

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



The resulting mixture will contain cm^3 of and cm^3 of

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')

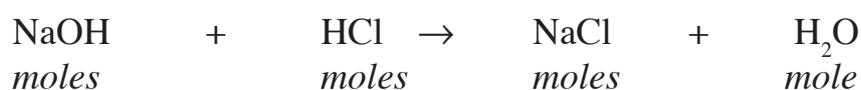
passing mixture through $\text{NaOH}_{(aq)}$, (or some other alkali), removes the acidic CO_2 .

Other Chemicals in Excess

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

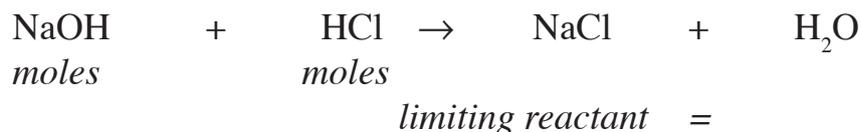
e.g. Calculate the mass of sodium chloride produced when 25 cm^3 of 0.1 mol l^{-1} sodium hydroxide reacts with 10 cm^3 of 0.2 mol l^{-1} hydrochloric acid.

Step 1 Write the balanced equation and establish the mole relationships

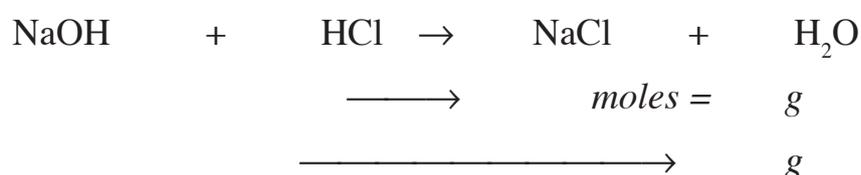


Step 2 Calculate the number of moles of each reactant

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



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



UNIT 3. Principles to Production

Topic 2: 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 $\Delta H = H(\text{products}) - H(\text{reactants})$
4. The enthalpy change can be calculated from potential energy diagrams
5. The enthalpy change has a **negative** value for exothermic reactions $\Delta H = H(\text{products}) - H(\text{reactants})$
6. The enthalpy change has a **positive** value for endothermic reactions $\Delta H = H(\text{products}) - H(\text{reactants})$

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. $\text{CH}_3\text{OH} + \frac{3}{2}\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
 $\Delta H = -\text{ve}$, 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. $\text{NaOH} + \text{aq} \rightarrow \text{Na}^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$
 $\Delta H = -\text{ve}$ or $+\text{ve}$, 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. $\frac{1}{2}\text{H}_2\text{SO}_4 + \text{LiOH} \rightarrow \frac{1}{2}\text{Li}_2\text{SO}_4 + \text{H}_2\text{O}$
 $\Delta H = -\text{ve}$, always exothermic
10. The enthalpy changes can be calculated using $\Delta H = c m \Delta T$ If water is used; $c = 4.18 \text{ kJ kg}^{-1} \text{ }^\circ\text{C}^{-1}$
11. The **bond enthalpy** is the enthalpy change associated with the breaking or forming of one mole of a particular bond
12. **Bond breaking** is **endothermic** while **bond forming** is **exothermic**.

Hess's Law

13. **Hess's law** states that the enthalpy change for a chemical reaction is independent of the route taken.

14. *Enthalpy changes can be calculated by application of Hess's law*

The Avogadro Constant

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

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

Molar Volume

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

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

Reacting Volumes

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

The idea of Excess

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