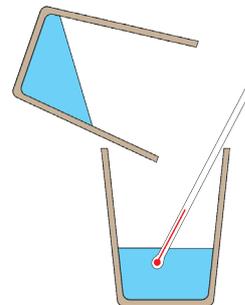


Higher Chemistry



Topic 4:

Enthalpy, Moles & RedOx

Student:

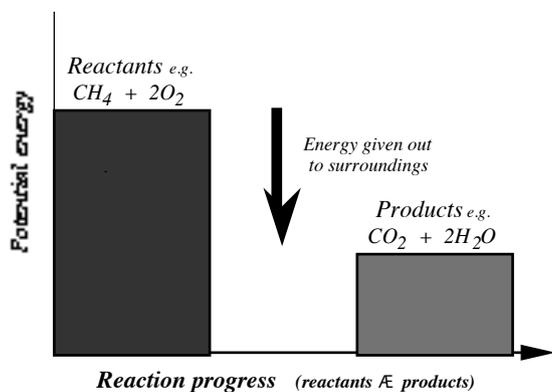
Lesson	Activities	Done	Checked
4.1 Energy Changes	1. Energy out, energy in		
	2. Where does the energy come from ?		
	3. Important Enthalpies		
	4.1 Check Test / Home Practice		
4.2 Enthalpies from Experimental data	1. Specific Heat Capacity		
	2. Enthalpy of Solution		
	3. Enthalpy of Neutralisation		
4.3 Enthalpy of Combustion	1. Burning Ethanol		
	2. Processing Results		
	3. Comparing Enthalpies		
	4. Bond Enthalpies		
	4.2 / 4.3 Check Test / Home Practice		
4.4 Hess's Law Calculations	1. Using Diagrams		
	2. Using Algebra		
		4.4 Check Test / Home Practice	
4.5 Molar Relationships 2	1. The Avogadro Constant		
	2. Formula Units		
	3. Empirical Formula		
4.6 Molar Relationships 3	1. Density of a Gas		
	2. Molar Volumes		
	3. Reacting Volumes		
		4.5 / 4.6 Check Test / Home Practice	
4.7 Molar Relationships 4	1. Avogadro's Law		
	2. Gases in Excess		
	3. Other Chemicals in Excess		
		4.7 Check Test / Home Practice	
4.8 Redox Agents	1. Definitions		
	2. Oxidising & Reducing Agents		
	3. Donating & Accepting Electrons		
	4. Electrochemical Series		
	4.8 Check Test / Home Practice		
4.9 Redox Equations	1. Combining Simple Ion-Electron Equations		
	2. More Complex Examples		
	3. Ion-Electron Equations From Scratch		
4.10 Redox Titration	1. The Redox Reaction		
	2. Standardisation		
	3. Determination of Iron		
		4.9 / 4.10 Check Test / Home Practice	
End-of-Section Assessment	Score:		Grade:
		%	

4.1 Energy Changes

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

Energy out, energy in

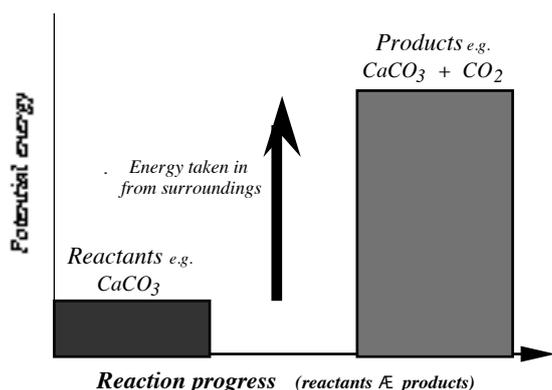
This activity looks at *exothermic* and *endothermic* reactions.



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

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 **reactant** molecules have **kinetic energy** and **potential energy**.

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

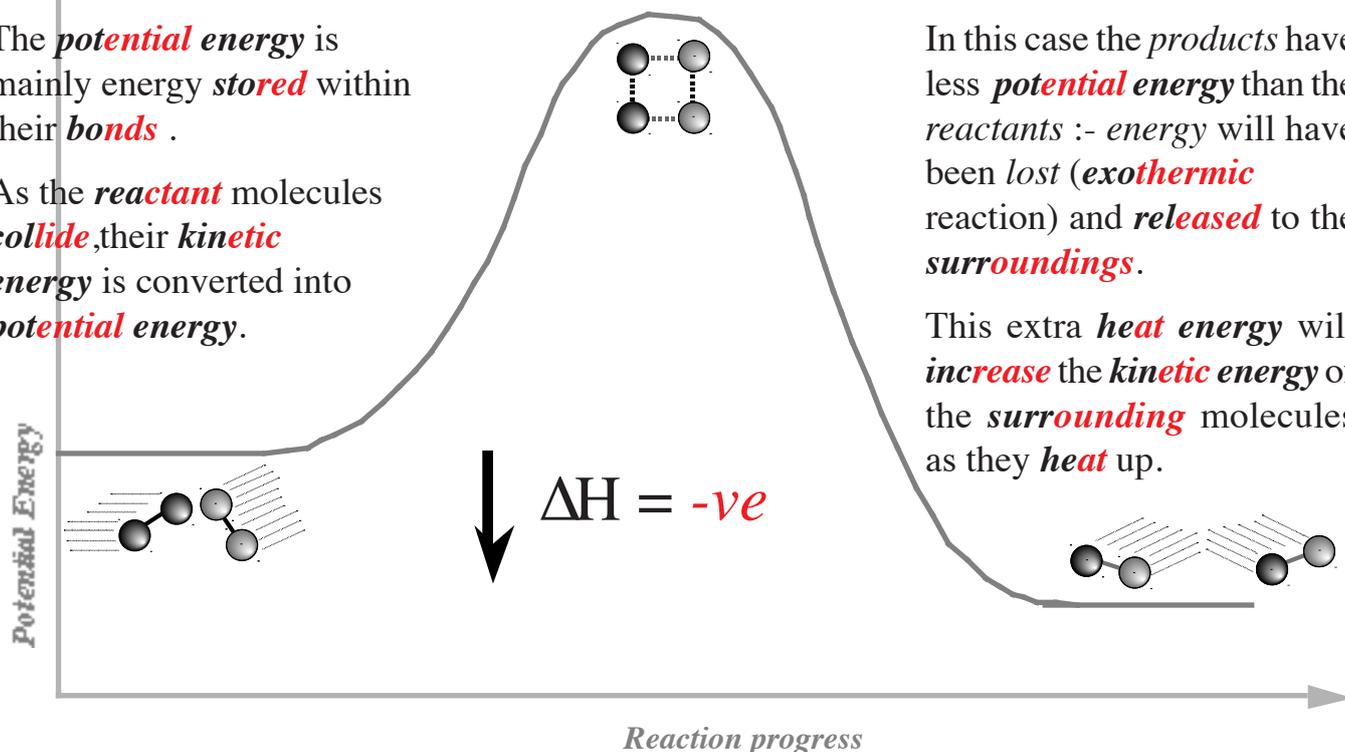
As the **reactant** molecules **collide**, their **kinetic energy** is converted into **potential energy**.

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

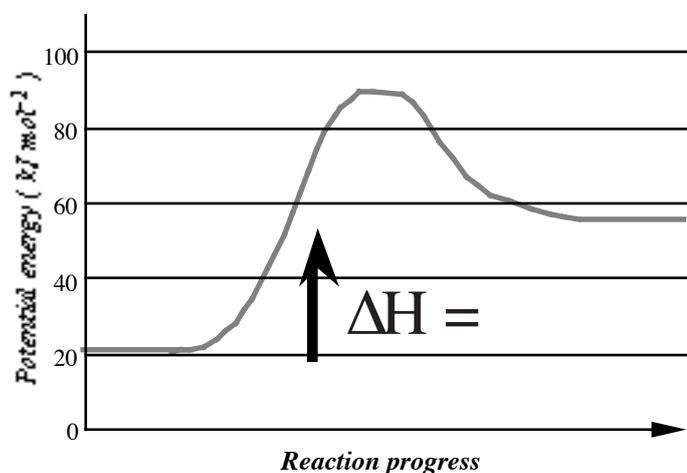
The **bonds** in the **product** molecules will be **different** from the original **reactant** molecules.

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 **exothermic 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

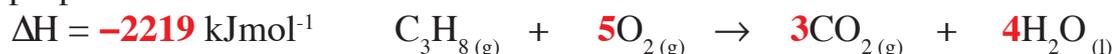
This activity introduces the enthalpies of **combustion**, **solution**, and **neutralisation**.

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*.

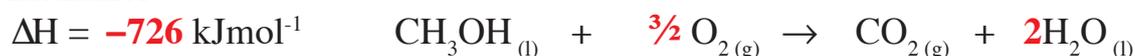
Along with the written definition, it is important that **balanced equations** can also be written that describe *exactly* the reaction to which the **numerical** value of ΔH corresponds.

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

e.g propane



methanol



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



sodium hydroxide



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.

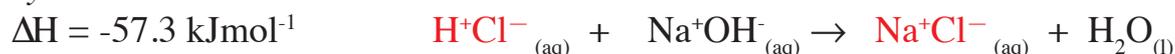
Notice that **enthalpies of solution** can be *either exothermic*, ΔH **negative**, or *endothermic*, ΔH **positive**.

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:-

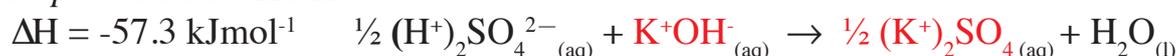


The **enthalpy of neutralisation** of an acid is the enthalpy change when the acid is neutralised to form **one mole of water**.

e.g hydrochloric acid / NaOH



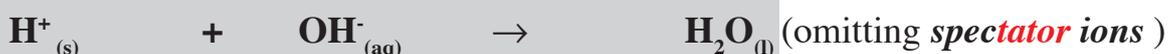
sulphuric acid / KOH



Notice, again, that there is often a need to use '**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;



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} .

4.2 Enthalpies from Experimental data

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

Specific Heat Capacity

$\Delta H =$ energy gained or lost by the water (E_h)

$c =$ specific heat capacity of the water

$m =$ mass of the water

$\Delta T =$ 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.

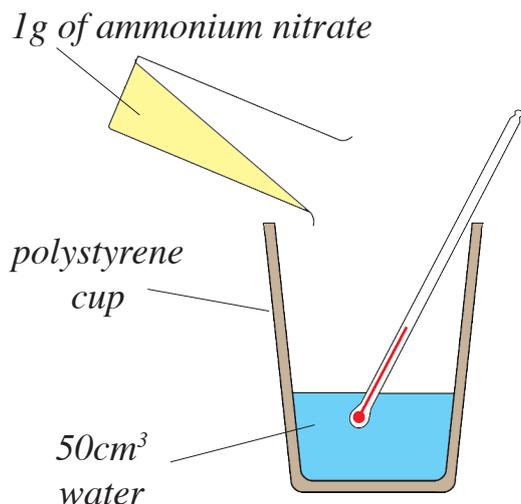
During **chemical** reactions **energy** is **released** to the **surroundings**, **exothermic**, or **taken in** from the **surroundings**, **endothermic**.

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**, $c = 4.18 \text{ kJ kg}^{-1} \text{ °C}^{-1}$

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.

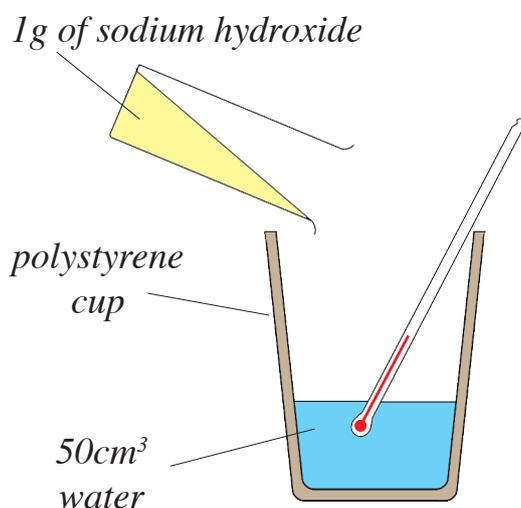
$$\begin{aligned}
 \text{Energy absorbed} &= c \times m \times \Delta T \\
 \text{from water} &= 4.18 \times 0.050 \times -1.6 \\
 &= -0.3344 \text{ kJ}
 \end{aligned}$$

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 = NH_4NO_3

gram formula mass (1 mole) = 80 g

$$\begin{aligned}
 1 \text{ g of ammonium nitrate} &\longrightarrow 0.3344 \text{ kJ of energy} \\
 1 \text{ mole, } 80 \text{ g of ammonium nitrate} &\longrightarrow 0.3344 \times 80 / 1 \\
 &= -26.75 \text{ kJ of energy}
 \end{aligned}$$



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

experimental results

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

$$\begin{aligned}
 \text{Energy released} &= c \times m \times \Delta T \\
 \text{into the water} &= 4.18 \times 0.050 \times \\
 &= \text{kJ}
 \end{aligned}$$

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 = **NaOH**

gram formula mass (1 mole) = **40** g

$$\begin{array}{rcl}
 1 \text{ g of sodium hydroxide} & \longrightarrow & \text{kJ of energy} \\
 1 \text{ mole, } \mathbf{40} \text{ g of sodium hydroxide} & \longrightarrow & x / 1 \\
 & = & \text{kJ of energy}
 \end{array}$$

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

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

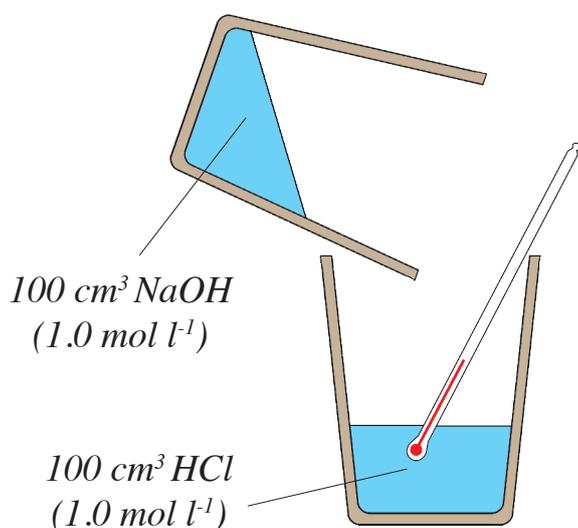
Enthalpy of Neutralisation

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

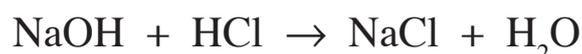
There was a **temperature rise** of **6.5** °C.



$$\begin{aligned}
 \text{Energy released} &= c \times m \times \Delta T \\
 \text{into water} &= \mathbf{4.18} \times \mathbf{0.2} \times \mathbf{6.5} \\
 &= \mathbf{5.434} \text{ kJ}
 \end{aligned}$$

These **volumes** and **concentrations** mean that **0.1 moles** of **NaOH** react with **0.1 moles** of **HCl** to produce **0.1 moles** of **H₂O**.

We simply need to multiply by **10** to **calculate** the **energy released** when **1 mole** of **H₂O** is formed.



$$\Delta H_{\text{neutralisation}} = \mathbf{54.34} \text{ kJ mol}^{-1}$$

4.3 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	—	- 726	—
ethanol	$\text{CH}_3\text{CH}_2\text{OH}$	CH_2	- 1367	- 641
propan-1-ol	$\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$	CH_2	- 2021	- 654
butan-1-ol	$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$	CH_2	- 2673	- 652

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 .

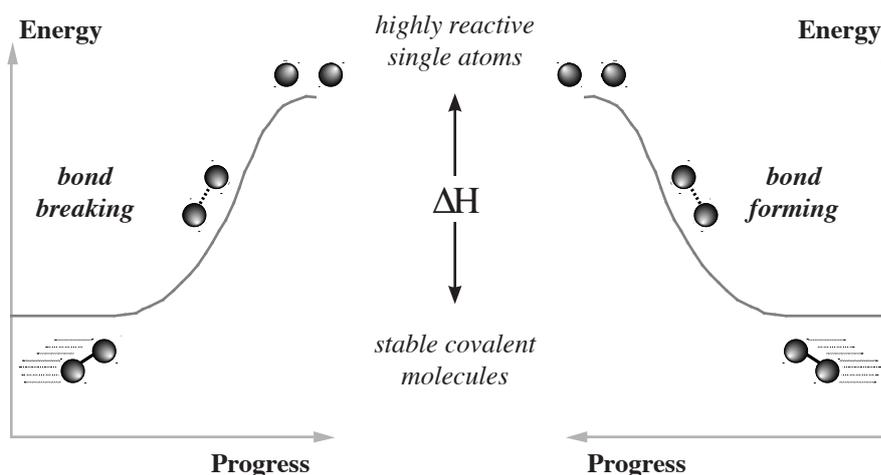
There is also a fairly **constant difference** in their **enthalpy values**. This supports the idea that it is energy **stored in bonds** that is **released** during **combustion** reactions.

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**.

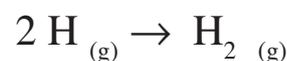
Bond	Enthalpy kJ mol^{-1}
H - H	432
O = O	497
N = N	941
F - F	155
Cl - Cl	243
Br - Br	194
I - I	149
H - F	569
H - Cl	428
H - Br	362
H - I	295



e.g. to **break** the H—H bond.



e.g. to **form** the H—H bond.



Bond Breaking - Endothermic

Bond Forming - Exothermic

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

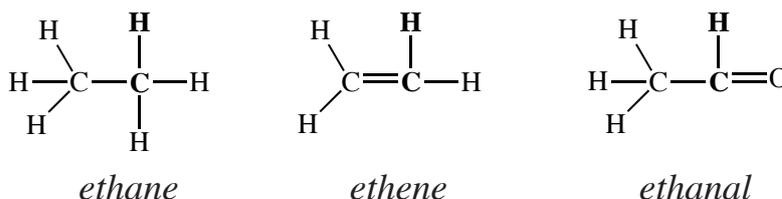
$$\Delta H = - 432 \text{ kJmol}^{-1}$$

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.

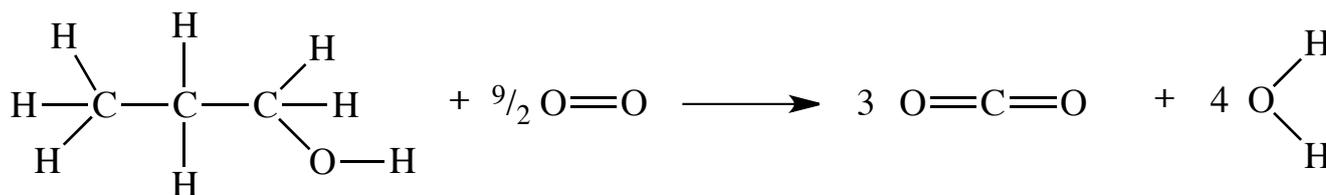
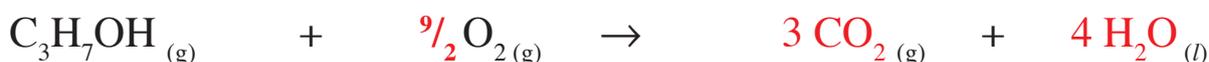
However, there are many slightly different C—H bonds, depending on their exact position in a molecule, e.g.



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

$$2 \times \text{C}-\text{C} = 2 \times 348 = 696$$

$$7 \times \text{C}-\text{H} = 7 \times 412 = 2884$$

$$1 \times \text{C}-\text{O} = 1 \times 360 = 360$$

$$1 \times \text{O}-\text{H} = 1 \times 463 = 463$$

$$\frac{9}{2} \times \text{O}=\text{O} = \frac{9}{2} \times 498 = 2241$$

$$\text{Total} = + 6644 \text{ kJ}$$

BOND FORMING

$$6 \times \text{C}=\text{O} = 6 \times 743 = 4458$$

$$8 \times \text{O}-\text{H} = 8 \times 463 = 3704$$

$$\text{Total} = - 8162 \text{ kJ}$$

Bond Breaking - Endothermic

Bond Forming - Exothermic

$$\text{enthalpy of combustion of propan-1-ol} = 6644 + (-8162) = -1518 \text{ kJ mol}^{-1}$$

4.4 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:

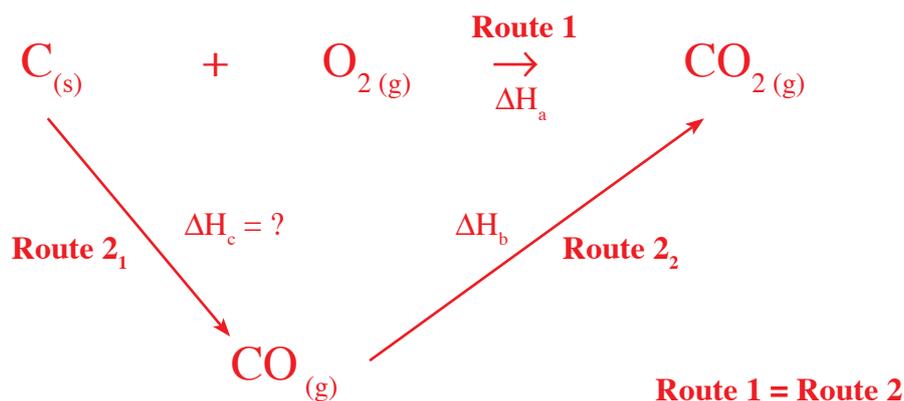


We can, however, *measure* the enthalpies of these two reactions (values in Data Books).



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:

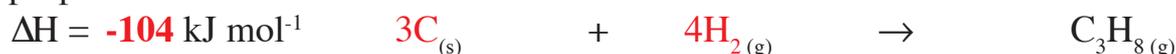
$$\Delta H_a = \Delta H_c + \Delta H_b$$

$$\begin{aligned} \text{Putting in the values we know, will allow us to calculate } \Delta H_c &= \Delta H_a - \Delta H_b \\ &= -394 - (-283) \\ &= -111 \text{ kJ mol}^{-1} \end{aligned}$$

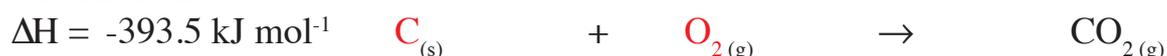
An extremely useful enthalpy to know is called the enthalpy of formation of a substance: by comparing enthalpies of formation chemists learn a lot about relative stabilities and the energy changes needed to change one substance into another,

*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*

e.g propane



carbon dioxide



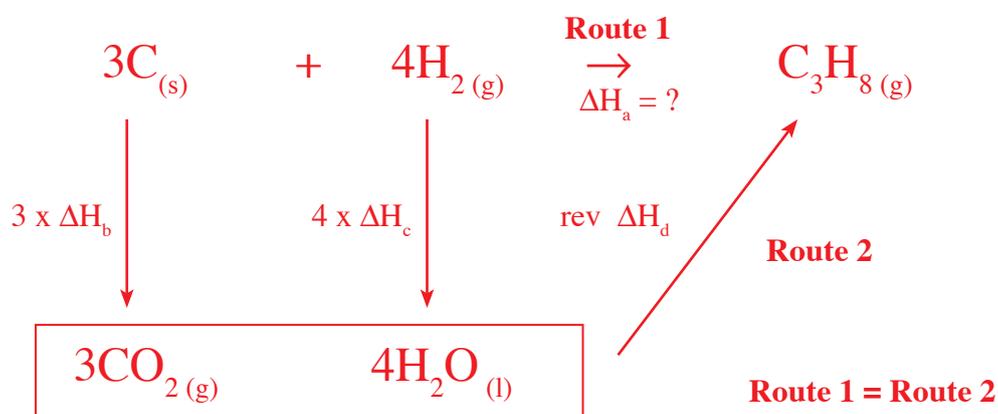
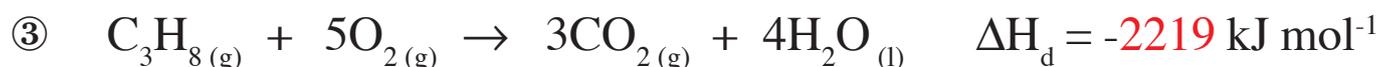
Most of these enthalpies of formation, particularly of organic molecules, cannot be measured experimentally; either the reactions are impossible or there will be a mixture of products.

Some 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 (like combustions) and whose values are in Data Books;



$$\text{Route 2} = (3 \times \Delta H_b) + (4 \times \Delta H_c) + (\text{rev } \Delta H_d) = (-1182) + (-1144) + (+2219) = -107 \text{ kJ} = \text{Route 1}$$

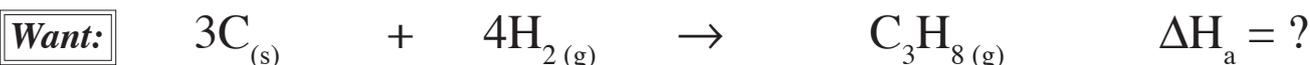
Using Algebra

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*.

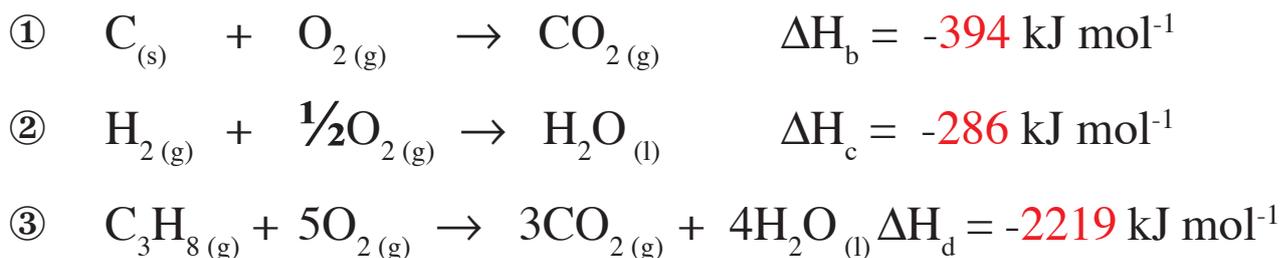
An alternative (more *algebraic*) approach has the advantage that you can follow a series of simple steps and the correct answer will '*pop out*' at the end. The disadvantage is that it can be slower.

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



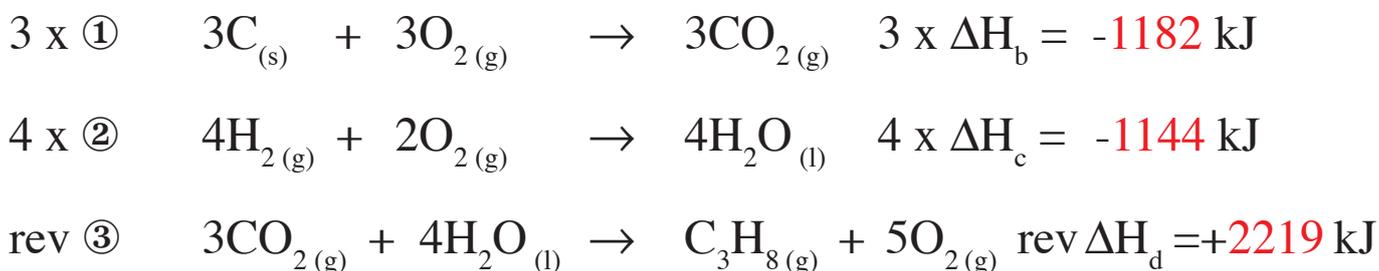
Step 2: Write (if not given) equations that correctly represent the reactions whose enthalpies you have been *directed to use* (often in *Data Book*). It can be useful to number these equations:

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:



4.5 Molar Relationships 2 - Numbers

This lesson looks at the mole in its purest form - 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} !

An eastern mole (*Scalopus aquaticus*) weighs about 75 grams, which means a mole of moles weighs $(6.022 \times 10^{23}) \times 75 \text{g} \approx 4.52 \times 10^{22} \text{kg}$. About half the mass of the moon.

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**.

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

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 = 6.02×10^{23} entities

Formula Units

This activity explains which 'entities' are contained in a mole of different substances - revision of Bonding Structures.

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



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

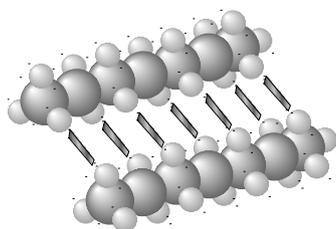


Formula = Ar ; 40g = **1mole** = 6.02×10^{23} **atoms**

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

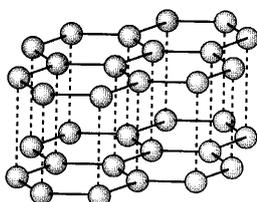


$$\begin{aligned} \text{Formula} = \text{O}_2; \quad 32\text{g} = 1\text{mole} &= 6.02 \times 10^{23} \text{ molecules} \\ &= 2 \times 6.02 \times 10^{23} \text{ atoms} \\ &= 1.204 \times 10^{24} \text{ atoms} \end{aligned}$$



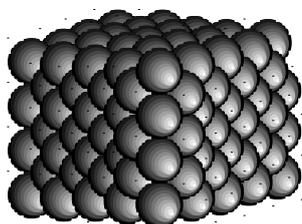
$$\begin{aligned} \text{Formula} = \text{C}_7\text{H}_{16}; \quad 100\text{g} = 1\text{mole} &= 6.02 \times 10^{23} \text{ molecules} \\ &= 23 \times 6.02 \times 10^{23} \text{ atoms} \\ &= 7 \times 6.02 \times 10^{23} \text{ C atoms} \\ &= 16 \times 6.02 \times 10^{23} \text{ H atoms} \end{aligned}$$

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



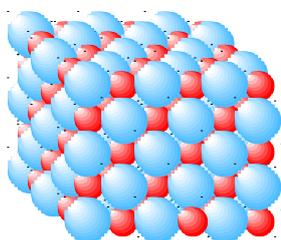
$$\text{Formula} = \text{C}; \quad 12\text{g} = 1\text{mole} = 6.02 \times 10^{23} \text{ atoms}$$

- **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* 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*.

$$\begin{aligned} 1 \text{ mole of any substance} &= \text{gram formula mass of that substance} \\ 1 \text{ mole of any substance} &= 6.02 \times 10^{23} \text{ formula units} \end{aligned}$$

4.6 Molar Relationships 3 - Gases

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

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

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

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

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

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

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

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

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

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**;

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

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**.

$$\begin{aligned} 0.00143 \text{ g} &\longrightarrow 1 \text{ cm}^3 \\ 32 \text{ g} &\longrightarrow (32 / 0.00143) \times 1 \text{ cm}^3 \\ &= 22\,400 \text{ cm}^3 \\ &= 22.400 \text{ l} \end{aligned}$$

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

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

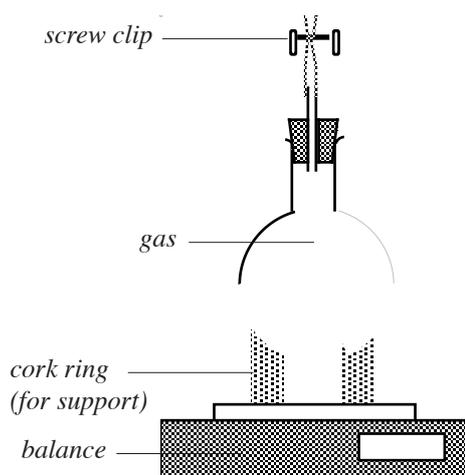
Element	N ₂	O ₂	F ₂	Cl ₂	Ne	Ar
Density (g / cm ³)	0.0013	0.00143	0.0017	0.0032	0.0009	0.0018
Mass of 1 mole (g)	28	32	38	71	20.2	39.9
Molar Volume (l)	21.5	22.4	22.4	22.2	22.4	22.2

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 10²³** 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*.

<i>mass</i>	→	<i>volume</i>
<i>mass</i>	→	<i>volume</i>
<i>of 1 mole</i>		

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

e.g. $\text{mass of empty flask} = 107.49 \text{ g}$

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

$\text{mass of flask} + \text{carbon dioxide} = 108.37 \text{ g}$

$\text{mass of carbon dioxide} = 108.37 - 107.49 = 0.88 \text{ g}$

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

$\text{volume of flask} = 480 \text{ cm}^3$

The *Molar Volume* can now be calculated:

<i>mass</i>	→	<i>volume</i>
0.88 g of CO ₂	→	480 cm³
(scale up to mass of 1 mole) 44 g of CO ₂	→	44 / 0.88 x 480 cm³
		⇒ 24 000 cm³

This *Molar Volume* is *higher* than those calculated from *densities* because the *densities* are measured at **0 °C**, whereas 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.

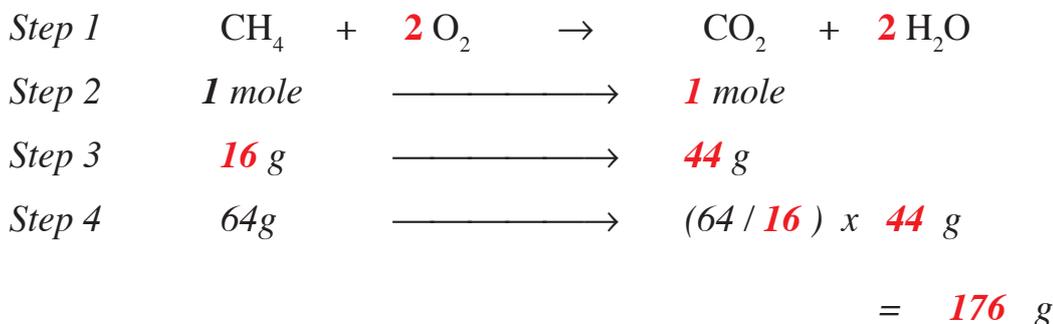
This is an extremely useful relationship when dealing with *reactions involving gases*.

Reacting Volumes

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

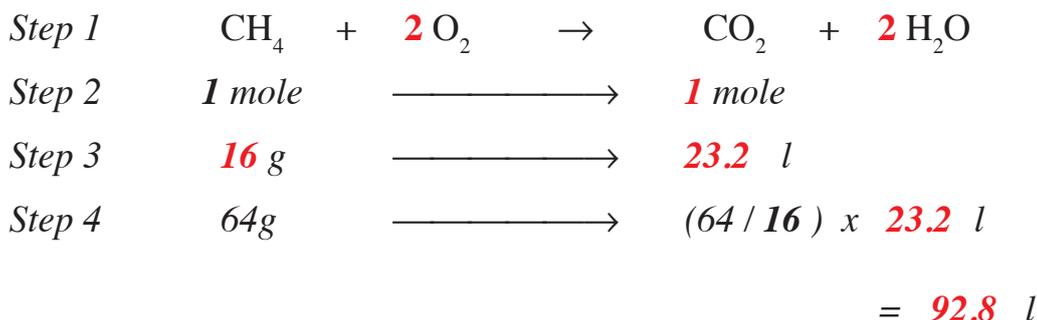
Up to now, we have treated **gases** the same as other substances and **calculated** their **mass**.

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



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).

**4.7 Molar Relationships 4**

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

Avogadro's Law

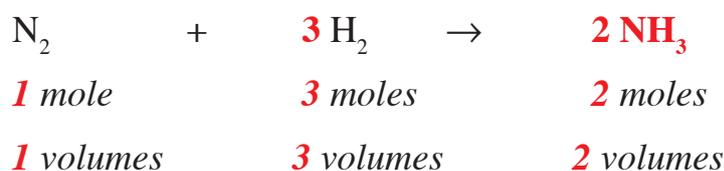
This activity looks at how the **molar volumes** of gases can be used to do calculations involving **volumes only**.

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

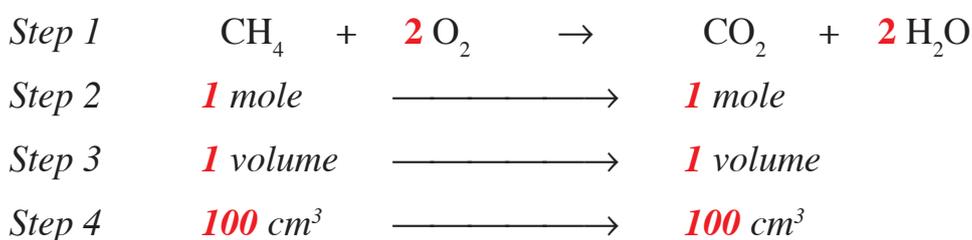
In practice, this means that we can **replace moles** with **volumes** in any relationship involving gases, as long as they are at the **same temperature** and the **same pressure**.

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



For any volume of N_2 , e.g. 100 cm^3 , we would need 300 cm^3 of H_2 and would be able to produce 200 cm^3 of ammonia.

e.g. What **volume** of *carbon dioxide* is produced when 100 cm^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 **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.

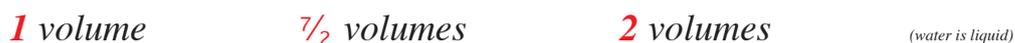
Any reactant **left over** at the end is said to be **in excess**.

e.g. A mixture of 20 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.

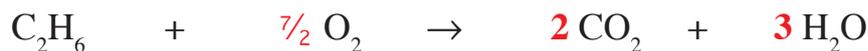
Step 1 Write the balanced equation and establish the mole relationships



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.



1 volume \longrightarrow **$\frac{7}{2}$ volumes**

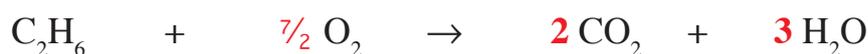
20 cm³ \longrightarrow **$\frac{7}{2} \times 20 = 70 \text{ cm}^3$**

limiting reactant = **C_2H_6**

reactant in excess = **O_2**

amount in excess = **120 - 70 = 50 cm³**

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



1 volume \longrightarrow **2 volumes**

20 cm³ \longrightarrow **$2 \times 20 = 40 \text{ cm}^3$**

The resulting mixture will contain **40 cm³ of CO_2** and **50 cm³ of O_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')

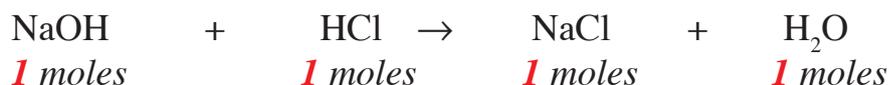
passing mixture through $\text{NaOH}_{(aq)}$, (or some other alkali), removes the acidic CO_2 .

Other Chemicals in Excess

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

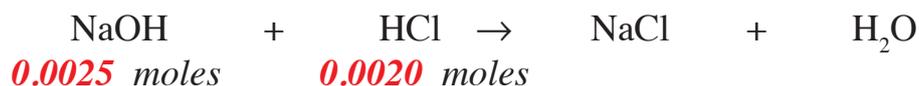
e.g. Calculate the mass of sodium chloride produced when 25 cm³ of 0.1 mol l⁻¹ sodium hydroxide reacts with 10 cm³ of 0.2 mol l⁻¹ hydrochloric acid.

Step 1 Write the balanced equation and establish the mole relationships



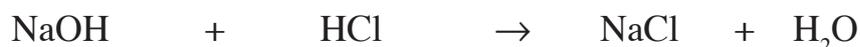
Step 2 Calculate the number of moles of each reactant

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



limiting reactant = **HCl**

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



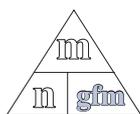
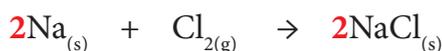
1 mole \longrightarrow **1 mole = 58.5 g**

0.0020 moles \longrightarrow **$0.0020 \times 58.5 \text{ g}$**

= 0.117 g

Example 2

What mass of sodium chloride is formed when 4.6g of sodium burns in 5g of chlorine.



$$\text{moles of Na} = 4.6 / 23 = 0.2 \text{ mol} \quad \text{mole ratio} = 0.2 / 2 = 0.1$$

$$\text{moles of Cl}_2 = 5 / 71 = 0.07 \text{ mol} \quad \text{mole ratio} = 0.07 / 1 = 0.07$$

So Cl_2 is the limiting reagent.

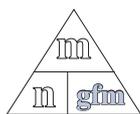


$$71 \text{ g} \rightarrow 117 \text{ g}$$

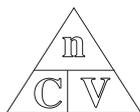
$$5 \text{ g} \rightarrow 117 \times 5 / 71 = 8.24 \text{ g}$$

Example 3

What volume of carbon dioxide is formed when 20g of calcium carbonate reacts with 500 cm³ of 1 mol l⁻¹ hydrochloric acid. (Molar volume = 25 l).



$$\text{moles of CaCO}_3 = 20 / 100 = 0.2 \text{ mol} \quad \text{mole ratio} = 0.2 / 1 = 0.2$$



$$\text{moles of HCl} = 1 \times 0.500 = 0.5 \text{ mol} \quad \text{mole ratio} = 0.5 / 2 = 0.25$$

So CaCO_3 is the limiting reagent.



$$100 \text{ g} \rightarrow 25 \text{ l}$$

$$20 \text{ g} \rightarrow 25 \times 20 / 100 = 5 \text{ l}$$

Example 4 Zinc chloride can be prepared in the laboratory by the reaction between zinc oxide and hydrochloric acid. A 6.75g sample of pure zinc oxide was added to 100 cm³ of 1.20 mol l⁻¹ hydrochloric acid.

Calculate the maximum mass of zinc chloride that could be obtained from this reaction.



$$\text{moles of ZnO} = 6.75 / 81.4 = 0.083 \text{ mol} \quad \text{mole ratio} = 0.083 / 1 = 0.083$$

$$\text{moles of HCl} = 1.2 \times 0.100 = 0.12 \text{ mol} \quad \text{mole ratio} = 0.12 / 2 = 0.06$$

So HCl is the limiting reagent.



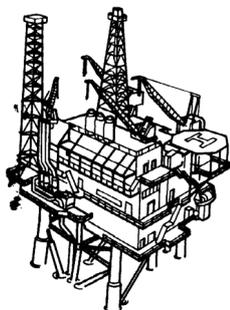
$$2 \text{ mol} \rightarrow 136.4 \text{ g}$$

$$0.12 \text{ mol} \rightarrow 136.4 \times 0.12 / 2 = 8.184 \text{ g}$$

4.8 Redox Agents

This lesson is about reactions involving reduction and oxidation and builds on the concepts learnt in earlier Chemistry Courses.

Definitions



OIL RIG

Oxidation Is Loss of electrons by a reactant.

Gaining **oxygen**/ losing **hydrogen** (increasing O:H **ratio**)
can also be described as **Oxidation**:-

e.g. aldehyde (—CHO) → acid (—COOH)

e.g. alcohol (—CH₂OH) → aldehyde (—CHO)

Reduction Is Gain of electrons by a reactant.

Losing **oxygen**/gaining **hydrogen** (decreasing O:H **ratio**)
can also be described as **Reduction**:-

e.g. ketone (—CO—) → alcohol (—CHOH—)

Oxidation and **Reduction** always go together so we call these reactions **REDOX**.

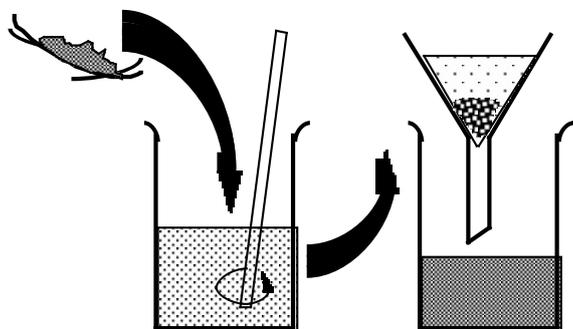
Oxidising & Reducing Agents

This activity considers how the term agent can be applied to oxidation & reduction reactions

An **oxidising agent** is a chemical which acts on another substance to oxidise it

A **reducing agent** is a chemical which acts on another substance to reduce it

For example, during the **displacement** of copper by zinc, the following reaction occurs.



By losing electrons **zinc atoms** are causing the reduction of Cu²⁺ ions so the **zinc atoms** are the **reducing agents**

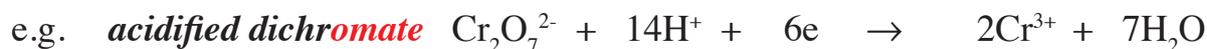


By gaining electrons **copper ions** are causing the oxidation of Zn atoms so the **copper ions** are the **oxidising agent**

During the Organic Unit you will be asked to learn the following **oxidation** reactions:



A number of **oxidising agents** can be used; each time the **agent** will itself be **reduced** as they **accept electrons**



An oxidising agent will be **reduced** when it reacts

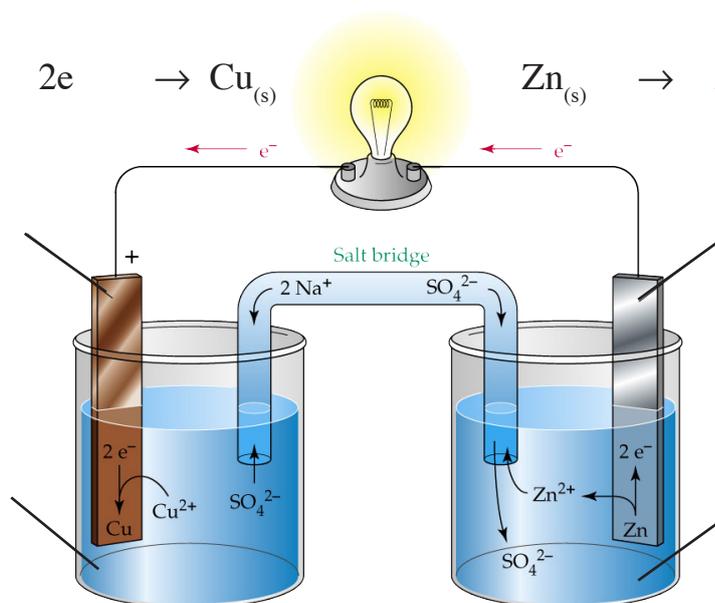
A reducing agent will be **oxidised** when it reacts

Donating & Accepting Electrons

This activity demonstrates electron transfer in a redox reaction in an electrochemical cell

Reduction

Oxidation



By gaining electrons **copper ions** are causing the oxidation of Zn atoms so the **copper ions** are the **oxidising agents**

By losing electrons **zinc atoms** are causing the reduction of Cu^{2+} ions so the **zinc atoms** are the **reducing agents**

By gaining electrons **copper ions** are accepting electrons from the Zn atoms.

By losing electrons **zinc atoms** are donating electrons to the Cu^{2+} ions.

In an electrochemical cell electrons flow from the reducing agent to the oxidising agent.

An oxidising agent will **accept electrons** when it reacts

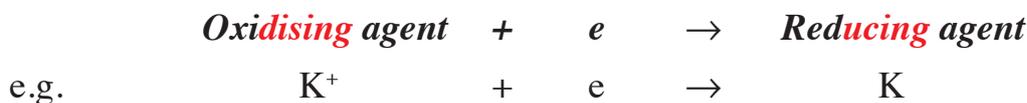
A reducing agent will **donate electrons** when it reacts

Electrochemical Series

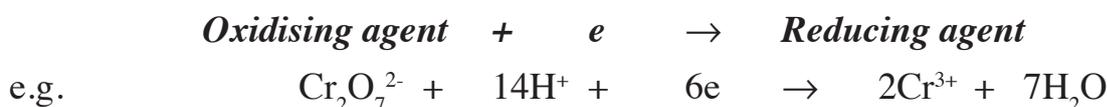
This activity is about how the electrochemical series can be used to write both reduction and oxidation equations

Help is available in the Data Booklet in the form of the *Electrochemical Series*.

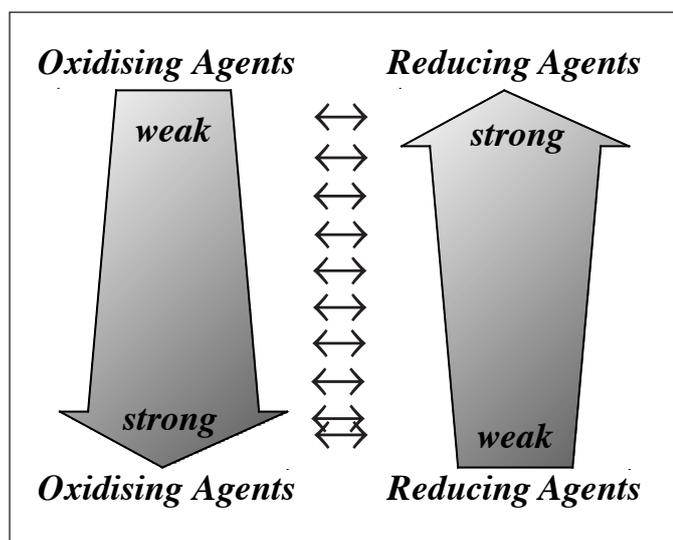
Equations are always written as **Reduction** reactions, but can be **reversed** for **Oxidation** reactions. Generally speaking, each equation represents the following change:



This reaction, from near the **top** of the *electrochemical series*, is **very unlikely** to go **as written** so we can describe **potassium ions** as **very weak oxidising agents** while **potassium atoms** would make **very strong reducing agents**.



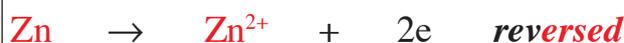
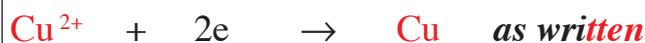
This reaction, from near the **bottom** of the *electrochemical series*, is **very likely** to go **as written** so we can describe **dichromate ions** as **very strong oxidising agents** while **chromium (III) ions** would make **very weak reducing agents**.



A useful 'rule-of-thumb' that is worth remembering is that the

reaction lower in the series goes as written while the **reaction further up will be reversed**.

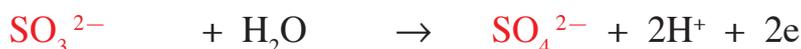
e.g. a simple **displacement** reaction, eg Zn/Cu



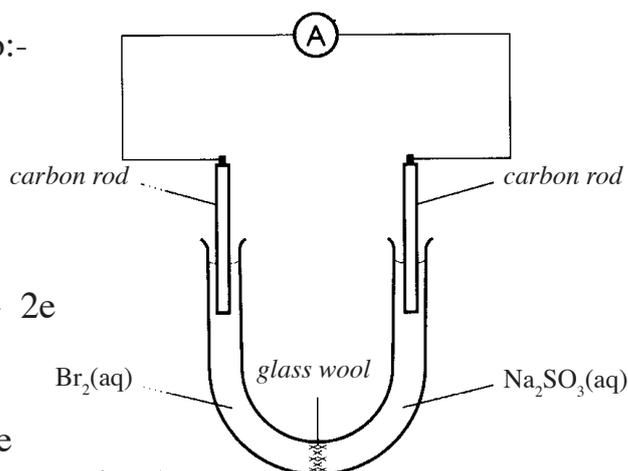
In this example, **bromine** is **lower** in the series so:-



While **sulphite ions** are found **higher** in the series so:-



(Sodium ions are so far up the series that it is safe to assume that they will never react; they are **spectator ions**)



4.9 Redox Equations

This lesson looks at how the overall equation for a redox reaction can be worked out from the ion-electron equations for the oxidation and reduction changes involved.

Combining Simple Ion-Electron Equations

This activity looks at the conditions for combining ion-electron equations and how this can lead to an overall balanced equation for a redox reaction.

Before ion-electron equations can be combined, the following two conditions must be satisfied :-

One equation must be a **reduction** (as written in electrochemical series) ,
One equation must be an **oxidation** (reverse of equation in series)

and

The **number of electrons** in each equation must be the **same**.

So

1. Identify and copy the two ion-electron equations from the Electrochemical series.
(**lower in series goes as written, higher in series is reversed**)
2. Multiply each ion-electron equation as necessary to ensure
number of electrons lost = number of electrons gained
3. Combine by writing everything on the left of both ion-electron equations on the left of RedOx equation, similarly for the right.
(**electrons should be equal and cancel each other out**)

Example 1: Aluminium metal and silver ions (displacement).



Example 2: oxidation of iron(II) ions by chlorine (a strong oxidising agent).



More Complex Examples

This activity is about combining ion-electron equations which have ions containing oxygen

These examples appear more complicated because they also involve ***H⁺ ions*** and ***H₂O molecules*** - so they are usually ***only able to react when in solution*** .

(Remember that, from your work on acids, water always contains a small proportion of H⁺ ions, though we often add some acid to help these reactions go faster).

Many of these ***ion-electron equations*** are available in the ***Electrochemical Series*** in your ***Data Booklet*** and can be dealt with in exactly the same way.

Example 3: *oxidation of sulphite ions, SO₃²⁻, by acidified permanganate solution, MnO₄⁻, (an extremely strong oxidising agent).*



Oxidising Agent: MnO₄⁻ is a powerful *oxidising agent* and can oxidise your skin, turning it brown so care must be taken when using permanganate solutions.

Reducing Agent: SO₃²⁻ is a useful *reducing agent* and can be seen on the labels of many foodstuffs - it slows down the oxidation of food and is labelled as a '*preservative*'.

Ion-Electron Equations From Scratch

This activity is about how ion-electron equations can be written without using a data booklet

As you have seen, many ***oxidation/reduction*** reactions need more than just ***electrons*** to keep them ***balanced***. We must also introduce and balance, as necessary, ***H⁺ ions*** and ***H₂O molecules***.

However, every balanced equation (even an ion-electron equation) must still have the ***same number of each type of atom*** on both sides and this must be done ***before*** adding ***H⁺ ions*** and ***H₂O molecules*** and ***before*** we add in ***electrons*** to ensure that the ***overall charge*** on both sides is ***the same***.

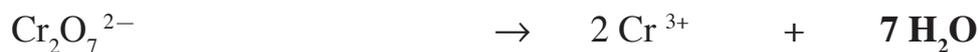
e.g. **dichromate ions** (orange) are strong **oxidising agents** because they are **easily reduced** to form **chromium (III) ions** (green).



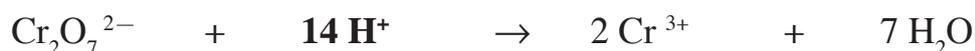
Step 2 there must be an *equal number of chromiums* on both sides so,



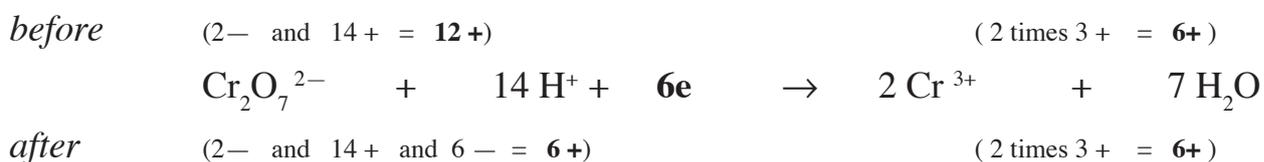
Step 3 there must be an *equal number of oxygens* on both sides so **water is added**,



Step 4 there must be an *equal number of hydrogens* on both sides so **H⁺ ions are added**,

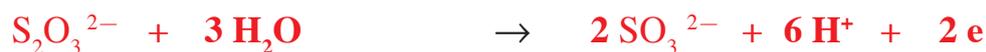


Step 5 there must be an *equal amount of charge* on both sides so **electrons are added**,



(Most **redox** reactions will only take place in **solution**, so there will be plenty of H_2O and H^+ available when needed. The pH of a solution often **changes** during a Redox reaction. If a large number of H^+ ions are needed then **acid** may be added e.g. 'acidified dichromate'.)

Example 4: Thiosulphate ion, $\text{S}_2\text{O}_3^{2-}$, being oxidised to the sulphite ion, SO_3^{2-}

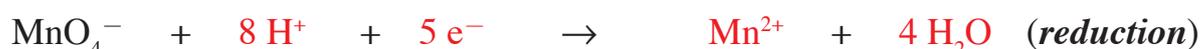


4.10 Redox Titrations

This lesson considers how titration could be used to determine the amount of iron present in a commercial iron tablet. The iron is present in the form of iron (II) compounds.

The Redox Reaction

This activity considers the overall equation for the oxidation of iron (II) by permanganate ions.



As shown by the overall equation, a **large number** of H^{+} ions are used up during the **RedOx** reaction. Though there are always some H^{+} ions present in water, the **concentration** is **very low** and the **rate of the reaction** would be **very slow** - too slow for the **rapid colour change** needed during a **titration**. Therefore, **acid** is always added to ensure a **rapid reaction** - the **solution is acidified**.

Standardisation

In this activity, the making of a standard solution of iron (II) sulphate is described along with the titration with potassium permanganate to determine KMnO_4 concentration.

Like many chemicals, **potassium permanganate** is either **not available in a pure enough form** or is **too unstable** to simply **weigh** out and **dissolve** to make a solution of **accurately known concentration** - a **standard** solution.

Instead another chemical - in this case **ammonium iron(II) sulphate hexahydrate** - **AFS** - $(\text{NH}_4)_2\text{Fe}(\text{SO}_4)_2 \cdot 6\text{H}_2\text{O}$ - is **weighed accurately** and then carefully **dissolved** in **deionised water** and transferred into a **volumetric flask**. With great care more **deionised water** is added to the 250 cm^3 flask (eventually **drop by drop**) until the **bottom of the meniscus** is sitting exactly on the **calibration line**.



$$\text{Initial mass of AFS} = 25.367\text{g}$$

$$\text{Molar Mass} = 392.14 \text{ g}$$

$$\text{Final mass of AFS} = 15.682\text{g}$$

$$392.14\text{g} \longrightarrow 1 \text{ mole}$$

$$\text{Mass of AFS used} = \mathbf{9.685 \text{ g}}$$

$$\mathbf{9.685 \text{ g}} \longrightarrow 1 \times \mathbf{9.685} / 392.14 = \mathbf{0.0247}$$



Moles of AFS, $n = 0.0247$

$$C = n / V = 0.0247 / 0.250$$

Volume of flask, $V = 0.250 \text{ l}$

$$= 0.0988 \text{ mol l}^{-1}$$

Deionised water is used instead of tap water to avoid possible side reactions such as precipitation caused by, in particular, the presence of Ca^{2+} and Mg^{2+} ions which form many insoluble compounds. The flask can be washed thoroughly with tap water but will then be rinsed with deionised water. Water remaining in the flask is not a problem as it simply becomes part of the water needed to make up to the 250 cm^3 mark.

*Accuracy is achieved by using the best balance available that reads to (at least) 2 decimal places but ideally 3 or 4. Accuracy is achieved by using (at least) a grade B volumetric flask (± 0.30) and ideally a grade A volumetric flask capable of $250 \text{ cm}^3 \pm 0.15 \text{ cm}^3$. Precision is achieved by working carefully to ensure that **all** of the solid AFS is transferred to the flask and the bottom of the meniscus rests exactly on the mark.*

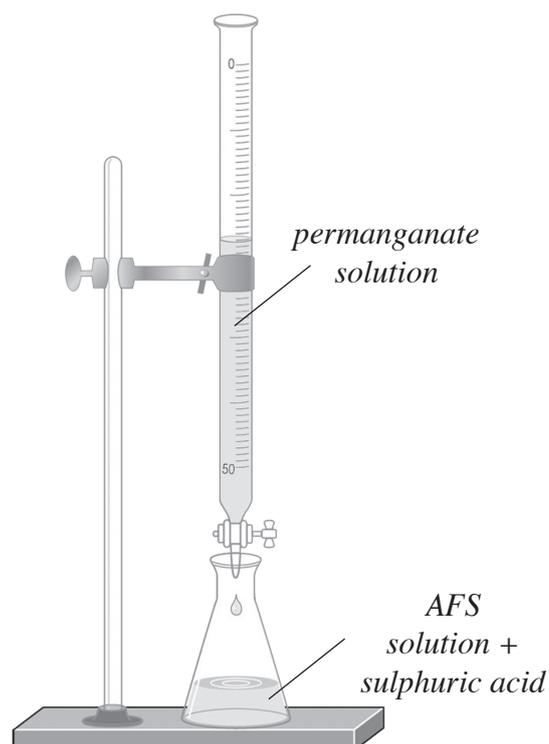
The ammonium iron(II) sulphate (AFS) solution can now be used to **determine accurately** the **concentration** of a **solution** of **potassium permanganate** (standardise the solution).

Exactly 25 cm^3 of AFS solution was **pipetted** into the flask, and **acidified** using *sulphuric acid*.

Permanganate solution (**purple**), of **unknown concentration**, was added from the **burette**. As the MnO_4^- ions reacted with the Fe^{2+} ions, the **purple** colour disappears.

More *permanganate* solution was added, eventually drop by drop, until a **faint pink/purple colour persisted**.

The strength of the MnO_4^- **colour** means that even a single *excess* drop produces a visible colour so no need for a separate **indicator** to detect end-point.



The **pipette** can be washed thoroughly with tap water but will then be rinsed with some of the AFS solution to remove the water and prevent any dilution of the AFS solution occurring. Similarly the **burette** will have a final rinsing with potassium permanganate solution. The **flask** can be rinsed with deionised water as any water left in the flask will have no effect on the **amount** of AFS present and the volume of KMnO_4 needed.

Accuracy is achieved by using the best pipette and best burette available. **Precision** is achieved by working carefully to ensure that the bottom of the meniscus rests exactly on the mark in the pipette and the pipette drains correctly. Similarly, all readings of the burette are made carefully. **Reliability** of the results is ensured by **repeating** several times until **concordant** results are achieved and then **averaging** if necessary.

Chemistry Reliability Accuracy Precision

	Rough	1	2	3
initial volume (cm^3)	0.10	0.10	0.05	0.10
final volume (cm^3)	25.50	25.00	24.75	24.90
volume added (cm^3)	25.40	24.90	24.70	24.80

The results obtained are shown in the table. Results 1, 2 & 3 are all **concordant** (within 0.1 cm^3 of another result) so all 3 results are used to calculate the **average** volume of KMnO_4 used.

$$\text{Average volume of } \text{KMnO}_4 \text{ required} = (24.90 + 24.70 + 24.80) / 3 = 24.80 \text{ cm}^3$$

The concentration of the KMnO_4 solution can now be calculated.

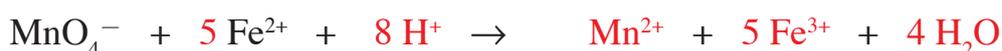
Chemistry calculations can be done in a variety of ways - the 'best' method is often a matter of personal choice. However, the formula below is provided in your Data Booklet so it is likely that most questions will be set up to use this formula.

Method 1 **provided:** overall RedOx equation concentration of AFS solution
required: concentration of potassium permanganate

n_1 = number of moles
of reactant 1
(AFS)

$$\frac{V_1 C_1}{n_1} = \frac{V_2 C_2}{n_2}$$

n_2 = number of moles
of reactant 2
 KMnO_4



$$n_2 = 1 \quad n_1 = 5$$

$$C_1 = 0.0988 \text{ mol l}^{-1}, V_1 = 25 \text{ cm}^3$$

$$V_2 = 24.80 \text{ cm}^3, C_2 = \text{calculate}$$

$$\frac{25 \times 0.0988}{5} = \frac{24.80 \times C_2}{1}$$

$$C_2 = \frac{25 \times 0.0988 \times 1}{5 \times 24.80}$$

$$C_2 = 2.47 / 124 = 0.0199 \text{ mol l}^{-1}$$

Determination of Iron

In this activity, the method used to determine the iron content in iron tablets is described.



Five iron tablets were weighed accurately.

$$\text{mass of iron tablets} = 1.067 \text{ g}$$

After being ground to a fine with a mortar and pestle, they were dissolved in about 25 cm³ of 2M sulphuric acid.

They were then carefully transferred to a 250 cm³ volumetric flask and filled to the mark with some 1M sulphuric acid.

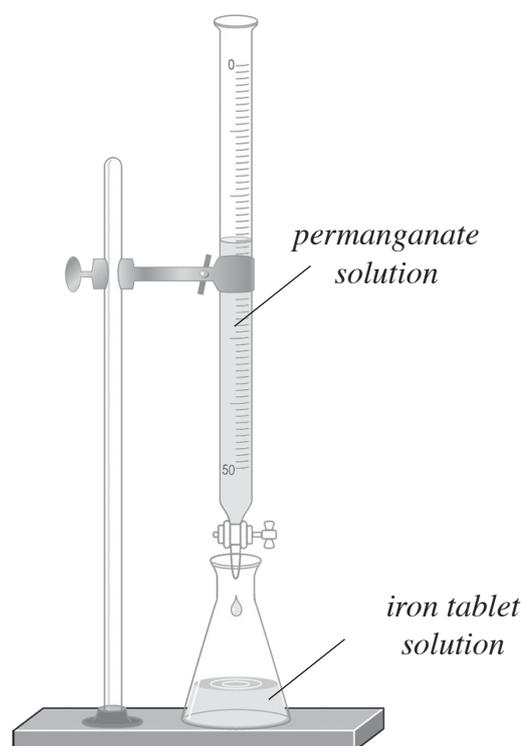


Exactly 25 cm³ of AFS solution was **pipetted** into the flask (already **acidified** using sulphuric acid).

Permanganate solution (**purple**), of **known concentration**, was added from the **burette**. As the MnO₄⁻ ions reacted with the Fe²⁺ ions, the **purple** colour disappears.

More permanganate solution was added, eventually drop by drop, until a **faint pink/purple colour persisted**.

The strength of the MnO₄⁻ **colour** means that even a single **excess** drop produces a visible colour so no need for a separate **indicator** to detect end-point.



Glassware used	Rinse with water	Rinse with iron tablet sol.	Rinse with KMnO _{4(aq)}
pipette		✓	
burette			✓
conical flask	✓		

	Rough	1	2	3
initial volume (cm ³)	0.00	3.00	5.70	8.40
final volume (cm ³)	3.00	5.70	8.40	11.20
volume added (cm ³)	3.00	2.70	2.70	2.80

$$\begin{aligned} \text{Average volume of KMnO}_4 \text{ required} &= (2.70 + 2.70 + 2.80) / 3 \\ &= 2.73 \text{ cm}^3 \end{aligned}$$

Evaluation: *The volumes in the previous titration are too small. There are rounding errors involved in reading burettes (typically $\pm 0.05 \text{ cm}^3$ for initial volume and $\pm 0.05 \text{ cm}^3$ for final volume = $\pm 0.10 \text{ cm}^3$ in total).*

Precision

$$\text{Average volume} = 2.5 \pm 0.10 \text{ cm}^3 \quad \% \text{ error} = 0.1 / 2.5 \times 100 = 4 \%$$

Accuracy

$$\text{Average volume} = 25 \pm 0.10 \text{ cm}^3 \quad \% \text{ error} = 0.1 / 25 \times 100 = 0.4 \%$$

An increased volume could have been achieved by either using more iron tablets (wasteful) or using a more dilute KMnO_4 solution (better).

Reliability

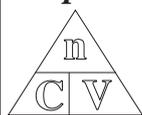
The determination could have been repeated using 3 sets of iron tablets and an average value calculated.

As mentioned earlier, chemistry calculations can be done in a variety of ways - the 'best' method is often a matter of personal choice. Whilst it is possible to use the titration formula provided in the data booklet, the need to calculate a **mass** of iron rather than a **concentration** could tip the balance towards a more 'basic' approach to this calculation.

Method 2 *provided:* overall RedOx equation concentration of KMnO_4
required: mass of iron in a tablet

Known: $C_1 (\text{KMnO}_4) = 0.0199 \text{ mol l}^{-1}$ $V_1 = 2.73 \text{ cm}^3$ $V_2 = 25 \text{ cm}^3$

Step 1: Start with 'known' chemical and calculate number of moles added from burette.



$$n_1 = C_1 \times V_1 \quad n_1 = 0.0199 \times 0.00273 \quad n_1 = 5.437 \times 10^{-5} \text{ moles}$$

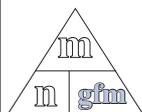
Step 2: Use balanced equation to convert from n_1 to n_2



$$n_2 = n_1 \times 5 = 5.437 \times 10^{-5} \times 5 = 2.719 \times 10^{-4} \text{ moles}$$

This is number of moles of Fe^{2+} present in the 25 cm^3 pipette. Total Fe^{2+} originally in the 250 cm^3 flask = $10 \times 2.719 \times 10^{-4} = 2.719 \times 10^{-3}$ moles

Step 3: Use gram formula mass of iron (55.8) to calculate mass of iron present.



$$\text{mass} = n_2 \times 55.8 \text{ g} = 2.719 \times 10^{-3} \times 55.8 \text{ g} = 0.152 \text{ g} \quad \text{or} \quad 152 \text{ mg}$$

This is mass of Fe^{2+} present in 5 tablets.

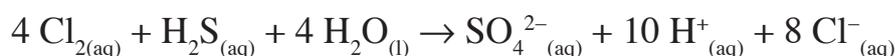
$$\text{Fe}^{2+} \text{ per tablet} = 152 / 5 = 30.4 \text{ mg}$$

* The recommended daily allowance (RDA) of iron is 7 mg (young children) - 30 mg (pregnant women)

Example 1

Hydrogen sulfide, H_2S , can cause an unpleasant smell in water supplies.

The concentration of hydrogen sulfide can be measured by titrating with a chlorine standard solution. The equation for the reaction taking place is



50.0 cm³ samples of water were titrated using a 0.010 mol l⁻¹ chlorine solution. An average of 29.4 cm³ of 0.010 mol l⁻¹ chlorine solution was required to react completely with a 50.0 cm³ sample of water.

Calculate the hydrogen sulfide concentration, in mol l⁻¹, present in the water sample.

(method 1 would be fine)

$$\frac{V_{\text{ox}} C_{\text{ox}}}{n_{\text{ox}}} = \frac{V_{\text{red}} C_{\text{red}}}{n_{\text{red}}}$$

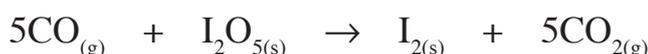
oxidising agent is Cl_2 but symmetry of equation means that it is not crucial to have correct way round.

$$\frac{29.4 \times 0.010}{4} = \frac{50 \times C}{1} \quad \text{so} \quad C = \frac{29.4 \times 0.0100}{4 \times 50} = 0.00147 \text{ mol l}^{-1}$$

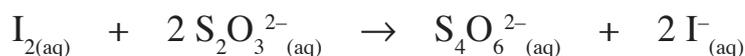
Example 2

The number of moles of carbon monoxide in a sample of air can be measured as follows.

Step 1 The carbon monoxide reacts with iodine(V) oxide, producing iodine.



Step 2 The iodine is then dissolved in potassium iodide solution and titrated against sodium thiosulphate solution.



If 50.4 cm³ of 0.10 mol l⁻¹ sodium thiosulphate solution was used in a titration, calculate the number of moles of carbon monoxide in the sample of air.

*(given the two equations involved and the fact that we are not calculating a final concentration, a step-wise **method 2** would be better)*

Step 1: Start with 'known' chemical and calculate n_1 (moles of S_2O_3) added from burette.

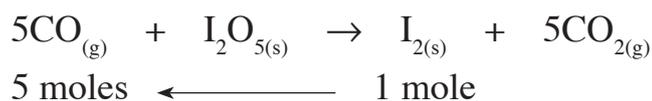
$$n_1 = C_1 \times V_1 \quad n_1 = 0.10 \times 0.0504 \quad n_1 = 0.00504 \text{ moles}$$

Step 2: Use balanced equation to convert from n_1 (moles of S_2O_3) to n_2 (moles of I_2)



$$n_2 = n_1 \div 2 = 0.00504 \div 2 = 0.00252 \text{ moles}$$

Step 3: Use second balanced equation to convert from n_2 (moles of I_2) to moles of CO .

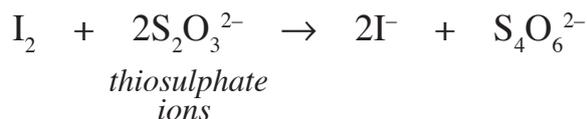


$$\text{moles of } CO = 5 \times \text{moles of } I_2 = 5 \times 0.0252 \text{ moles} = 0.126 \text{ moles}$$

Example 3

Solutions containing iodine are used to treat foot rot in sheep.

The concentration of iodine in a solution can be determined by titrating with a solution of thiosulphate ions.



20.0 cm³ samples of a sheep treatment solution were titrated with 0.10 mol l⁻¹ thiosulphate solution and an average titre of 18.15 cm³ obtained.

Calculate the concentration of iodine, in mol l⁻¹, in the foot rot treatment.

(balanced equation provided and concentration wanted so **method 1** would be fine)

$\frac{V_{\text{ox}} C_{\text{ox}}}{n_{\text{ox}}} = \frac{V_{\text{red}} C_{\text{red}}}{n_{\text{red}}}$
--

oxidising agent is I_2 but symmetry of equation means that it is not crucial to have correct way round.

$$\frac{20}{1} \times C = \frac{18.15 \times 0.10}{2} \quad \text{so} \quad C = \frac{18.15 \times 0.10}{2 \times 20} = 0.045 \text{ mol l}^{-1}$$

Outcome 1: It is intended to use a Redox titration involving determination of Vitamin C in a Vitamin C tablet to give an opportunity to pass this course requirement.

Planning an experiment

- The plan should include:
- ◆ a clear statement of the aim
 - ◆ a hypothesis
 - ◆ a dependent and independent variable
 - ◆ variables to be kept constant
 - ◆ measurements/observations to be made
 - ◆ the equipment/materials
 - ◆ a clear and detailed description of how the experiment should be carried out, including safety considerations

Presenting results in an appropriate format

One format from: table, line graph, chart, key, diagram, flow chart, summaries or extended text or other appropriate formats

Draw a valid conclusion

Include reference to the aim

Evaluating experimental procedures

Suggest two improvements with justification

UNIT 3. Chemistry In Society

Topic 4: Enthalpy , Moles & Redox

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

Not formally included in CfE Chemistry but such a fundamental part of understanding the concept of a mole that you should at least be aware of it. Will not feature as a 'recall' question in exams but may be introduced as part of an application of knowledge.

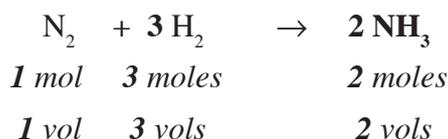
Molar Volume

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.

1 mol \longleftrightarrow molar volume

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

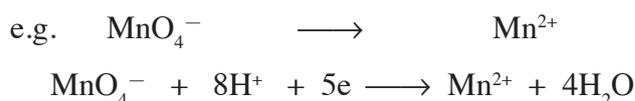
Redox

21. An **oxidising agent** is a substance which accepts electrons; a **reducing agent** is a substance which donates electrons
22. Oxidising and reducing agents can be identified in redox reactions
23. **Ion-electron equations** can be written for oxidation and reduction reactions.
24. Ion-electron equations can be combined to produce **redox** reactions
25. Given reactant and product species, ion-electron equations can be written which include H_2O and H^+ .
26. The concentration of a reactant can be calculated from the results of **volumetric titrations**

Oxidising agents gain electrons - are reduced
Reducing agents lose electrons - are oxidised

Data Booklet contains most equations you will ever need

Balance electrons lost with electrons gained



Method 1:

$$\frac{V_{\text{ox}} C_{\text{ox}}}{n_{\text{ox}}} = \frac{V_{\text{red}} C_{\text{red}}}{n_{\text{red}}}$$

Method 2:

$$C_{\text{red}} \times V_{\text{red}} = n_{\text{red}} \rightarrow \text{equation} \rightarrow n_{\text{ox}} \rightarrow n_{\text{ox}} / V_{\text{ox}} = C_{\text{ox}}$$

(known) (known)