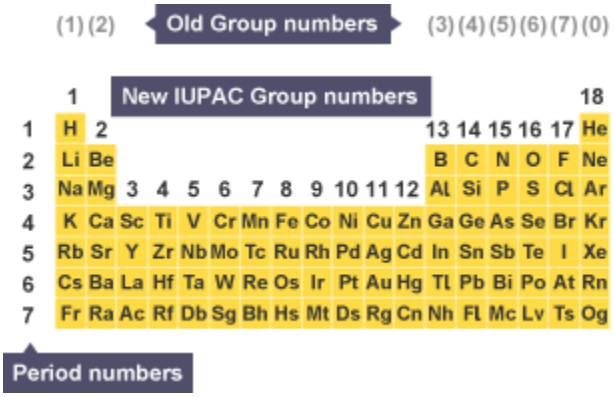
SS2 FIRST TERM CHEMISTRY, 2026/2027 NOTES

WEEK 1 & 2

Periodic Table: is defined as the systematic arrangement of elements in order of their increase in atomic number.

In the modern *periodic table*, the elements are arranged according to their atomic number - not their *relative atomic mass*.



In the periodic table the elements are arranged into:

- horizontal rows, called *periods*, in order of increasing atomic number
- vertical columns, called *groups*, where the elements have similar *properties*

Group numbers

Group numbering has changed over time:

- in older periodic tables, group numbers 1-7 and 0 are given for the main groups
- in newer periodic tables, 'IUPAC' group numbers 1-18 are given for all the groups

Hydrogen, H, is placed above the top of group 1 - it is not a member of group 1.

Metals and non-metals in the table

The *metal* elements are found on the left-hand side of the periodic table and the *non-metal* elements are found on the right. You can imagine a zig-zag line, starting at B-Al-Si, separating metals from non-metals.

Resolving pair reversals

Mendeleev did not know about *isotopes*, but their existence is an explanation for pair reversals in his table. The positions of iodine and tellurium were reversed in Mendeleev's table because:

- iodine has one naturally occurring isotope, ^{127}I
- the most *abundant* tellurium isotopes are ^{128}Te and ^{130}Te

The high relative abundance of these tellurium isotopes gives tellurium the greater relative atomic mass. The atomic number of tellurium is 52 and the atomic number of iodine is 53, so these elements are in the correct order in the modern periodic table.

Electron arrangements

An *electron arrangement* is the way in which *electrons* are arranged in an *atom*.

Electrons in shells

Different shells can hold different maximum numbers of electrons. Electrons occupy shells starting with the innermost one. They begin to occupy the next shell when a shell becomes full.

For elements with *atomic number* 1 to 20:

Electron shell	Maximum number of electrons
First	2
Second	8
Third	8

Predicting an electron arrangement

The electron arrangement of an atom can be predicted from its atomic number. For example, the atomic number of sodium is 11. Sodium atoms have 11 *protons* and so 11 electrons:

- 2 electrons occupy the first shell
- 8 electrons occupy the second shell
- 1 electron occupies the third shell

This electron arrangement can be written as 2.8.1 (each dot separates one shell from the next). This electron arrangement can also be shown as a diagram. In these diagrams:

- each shell is modelled as a circle
- each electron is modelled as a dot or a cross

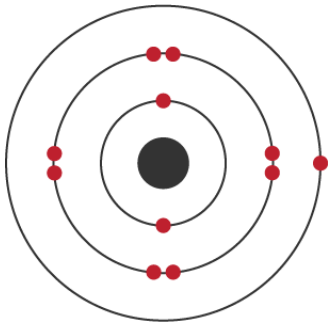


Fig. The electron arrangement of sodium as a diagram

Electron arrangements and the periodic table

The electron arrangement of an element is related to its position on the periodic table.

Electron arrangement feature	Link to the periodic table
Number or numbers of circles	Period number
Number of electrons in outermost shell	Old group number
Total number of electrons in all shells	Atomic number

Note that:

- helium and the other elements in group 0 (IUPAC group 18) have full outer shells
- hydrogen has 1 electron, so it is placed above the top of group 1 (but it is not in group 1)

Example

The electron arrangement of sodium is 2.8.1. This shows that sodium:

- is in period 3
- is in old group 1 (IUPAC group 1)
- has an atomic number of (2 + 8 + 1) = 11

Question

The electron arrangement of nitrogen is 2.5. Explain what this shows about the position of nitrogen in the periodic table.

Answer

2.5 shows that nitrogen:

- is in period 2
- is in group 5 (IUPAC group 15)
- has an atomic number of (2 + 5) = 7

Physical properties

The table summarises some typical *physical properties* of metals and non-metals.

Metals	**Non-metals**
Shiny	Dull
High melting points	Low boiling points
Good conductors of electricity	Poor conductors of electricity
Good conductors of heat	Poor conductors of heat
High density	Low density
Malleable	Brittle

Some *elements* have properties that are not typical. For example:

- mercury (a metal) has a low *melting point* and exists as a liquid at room temperature
- graphite, a form of carbon (a non-metal), has a high *boiling point* and is also a good *conductor* of electricity

A substance with a high *density* means it has a high mass for its size.

Malleable substances can be bent or hammered into shape without shattering, while *brittle* substances shatter when bent or hit.

Ductile means that a substance can be drawn out into a long wire without snapping or breaking.

Chemical properties

The reactions of elements are related to the *electron arrangement* of their atoms (and so to their *atomic number*). Electrons in the outer shells are lost, gained or shared in reactions. Elements in a group have similar chemical properties because they have the same number of electrons in their outer shells.

Reaction	Electrons in outer shells	Type of bonding formed
metal + non-metal	Lost from metal atoms, gained by non-metal atoms	Ionic
non-metal + non-metal	Pairs of electrons shared	Covalent

In general:

- metals form positively charged *ions* but non-metals form negatively charged ions
- metals form *basic* oxides, some of which dissolve to form *alkaline* solutions, but non-metals form *acidic* oxides

ELECTRONIC CONFIGURATION OF THE FIRST THIRTY ELEMENTS

The Electronic configuration of an atom is the representation of the arrangement of the electrons distributed among the orbital shells and subshells. Commonly, the electronic configuration is used to describe the orbitals of an atom in its ground state

MEANING OF ATOMIC ORBITAL

Orbital is the region of space around the nucleus where there is a high probability of finding electron. The four different types of orbitals are *s, p, d*, and *f*. These orbitals have different shapes and one orbital can hold a maximum of two electrons. The p-orbital has three degenerate orbitals, with a maximum of six electrons, d has five sub orbitals with a maximum of ten electrons and the f-orbital has seven sub-orbitals with maximum of fourteen electrons.

Orbitals	Number of sub-ortals	Max.no. electrons	Shape
S	-	2	Spherical
P	3	6	Dumb-bell
D	5	10	Double dumb-bell
F	7	14	Complex

RULES AND PRINCIPLES FOR FILLING IN ELECTRONS

1. Aufbau Principle: states that electrons are filled in their orbital in order of increasing energy level.
The order is as follows:

1s < 2s < 2p< 3s < 3p <4s < 3d <4p etc.

PERIOD 1 AND 2:

PERIODIC PROPERTIES. Some properties of the atom change along a group or across a period on the periodic table. Atomic radius which is measured of the size is one of such properties. The orbiting electrons in an atom are best represented by an electron cloud which has no distinct limit as the size of an action cannot be defined easily.

1. **Atomic radius:** This has been defined as the distance of closest approach to another identical atom in a given bonding situation. There are two types of atomic radii. Covalent radius and Van der Waals radius. Covalent radius is half the distance between two identical atoms which are not chemically bonded.
For the two types of atomic radius two variations are noticeable:

(i) The atom radius increases down a group

(ii) The atomic radius decreases along a period.

This is because going down any group on the periodic table the number of valence electrons remains constant while the shells increase in size (radius) despite increase in nuclear charge. The atomic radius of potassium is greater than that of Sodium. The atomic radius of caesium is greater than that of rubidium.

Across a period, electrons are added to orbitals in the same shell, all the valence electrons are therefore at the same energy level. As atomic number increases, the positive charge of the nucleus increases giving rise to greater attraction between the positive nucleus and negative electrons. This in turn result in contraction of the electrons cloud resulting in a smaller atom. Atomic radii therefore decrease across a given period on the periodic table.

2. **Ionic Radius**: Ions are formed by a loss or gain of electrons by an atom. A positive ion (cation) is smaller than the original metal atom because electrons are pulled in due to increase in effective nuclear charge.

A negative ion (anion) is bigger than the corresponding non- metal atom because the effective nuclear charge is reduced.

As we move across the second short period, the cationic radii decrease from Sodium to Aluminium while the anionic radii increase from phosphorous to chlorine.

3. **Ionization energy**: Ionization occurs when gaseous atom loses electrons from its outer most shell to become positively charged

K ——— e^- K^+ ———

The energy required to do this is called ionization energy or ionization potential. **First ionization energy** of an element is the energy needed to remove one mole of electron(s) from one mole of atoms in the gaseous state. It is expressed in kilo- joules per mole of atoms ionized.

First ionization energy increase across the period with noble gases having the highest. As we go down he the group, the value of first ionization energy decreases.

FIRST IONIZATION ENERGIES OF ALKALI METALS

Element	LI	Na	K	Rb	Cs
First ionization energy $KJMOL^{-1}$	520	500	420	400	380

FIRST IONIZATION ENERGIES OF THE ELEMENTS IN THE THIRD PERIOD OF THE PERIODIC TABLE

Element	Na	Mg	Al	Si	P	S	Cl	Ar
First ionization energy $KJMOL^{-1}$	496	737	577	786	1012	999	1255	1521

Three factors that affect the ionization potential of an atom are:

(i) Distance of the outer most electrons from the nucleus.

(ii) Size of the positive or effective nuclear change.

(iii) Screening effect of the inner electrons.

Moving from left to right across a period, there is a general rise in the first ionization energy. This is due to the fact that the nuclear charge is increasing across the period. This in turn causes a decrease in atomic radius that is a decrease in the distance of the outermost electrons from the nucleus. The screening effect is almost the same across the period. Down a group of the periodic table, ionization energy decreases because the nuclear charge on the outermost electron is reduced. The outermost electrons are properly shielded from the effect of nuclear charge

4. **Electron Affinity:** is the energy released when an electron is added to gaseous atom in its lowest energy state. Its unit is $kJmol^{-1}$ or electron volts (eV). Electron affinity increases across a period from left to right and decrease down the group on the periodic table.

Group 1 elements, alkali metals have the least tendency to add electrons to their neutral atoms.

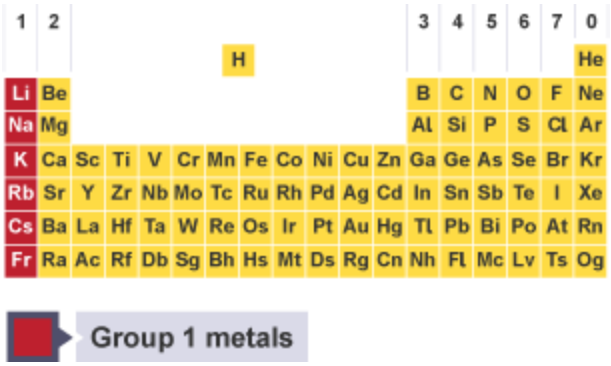
Elements in groups VI and VII have greatest tendency to accept electron. Noble gases (group 8 or 0) have stable electronic configuration

5. **Electronegativity:** Electronegativity is the ability or power of that atom in a molecule to attract shared pair of electrons. It is more pronounced in heteronuclear molecules where two dissimilar atoms share one or more pairs of electrons.

Electronegativity increases across the period, i.e. going from left to right of the Periodic Table but decreases down the group i.e. going down the Periodic Table. The steady increase as one goes across the period is due to a steady increase in nuclear charge and decrease in atomic size. Consequently, the halogen atom, Fluorine, has the highest electronegativity in the period, due to the strong affinity for electrons. But down the group, the increase in atomic size due to screening effect of the inner shells of electrons decreases the nuclear attraction for shared electrons. The noble gases of group O are not assigned electronegativity values since they have completed shells of electrons.

Group 1 - physical properties

Group 1 contains *elements* placed in a vertical column on the far left of the *periodic table*. The elements in group 1 are called the *alkali metals*.



Group 1 is on the left-hand side of the periodic table

The alkali metals share similar *physical properties*. For example, they:

- are soft (they can be cut with a knife)
- have relatively low *melting points*
- have low densities

Question

The table shows the melting points of five alkali metals.

Use this information to describe how melting point changes in group 1.

Element	**Melting point (°C)**
Lithium, Li	180
Sodium, Na	98
Potassium, K	63
Rubidium, Rb	39
Caesium, Cs	28

Answer

Going down group 1, the melting point decreases.

Group 1 - Chemical Reactions with Water

Atoms of *group 1 elements* all have one *electron* in their outer shell. This means that the alkali metals all have similar *chemical properties*.

When a group 1 element takes part in a reaction, its atoms each lose one electron. This forms positively charged *ions*. The ions have a stable arrangement of electrons, with a complete outer shell.

Reactions with water

The alkali metals react with water to produce a metal hydroxide and hydrogen. For example, sodium reacts with water:

sodium + water → sodium hydroxide + hydrogen

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Sodium hydroxide is an *alkali*. It is a *base* that *dissolves* in water to form an *alkaline solution*. This solution:

- has a *pH* greater than 7
- turns *universal indicator* solution blue or purple

Question

Write the word equation and balanced symbol equation for the reaction of potassium with water.

Answer

potassium + water → potassium hydroxide + hydrogen

$2K(s) + 2H_2O(l) \rightarrow 2KOH\ (aq) + H_2(g)$

Group 1 - chemical reactions with oxygen and chlorine

Reactions with oxygen

The group 1 elements react with oxygen from the air to make metal *oxides*.

At room temperature, oxygen reacts with the surface of the metal. This forms a white oxide, which covers the surface. The metal below the surface does not react.

The *alkali metals* burn vigorously when heated and placed in a gas jar of oxygen. The oxide forms as white smoke.

For example:

sodium + oxygen → sodium oxide

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

The *reactivity* of the group 1 elements increases down the group, so their reactions with oxygen get more vigorous.

Question

Predict which becomes white more quickly on exposure to air - a piece of rubidium, or a piece of lithium. Explain your answer.

Answer

The rubidium becomes white more quickly. This is because rubidium is below lithium in group 1, so rubidium is more *reactive*.

Reactions with chlorine

The group 1 elements react vigorously with chlorine. The products of the reactions are chlorides. At room temperature the chlorides are white solids. They dissolve in water to form colourless solutions. For example:

sodium + chlorine → sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

The reactions with chlorine get more vigorous going down the group.

Question

Write a balanced equation for the reaction of potassium with chlorine.

Answer

$2K(s) + Cl_2(g) \rightarrow 2KCl(s)$

Explaining the trend in reactivity

When a group 1 element takes part in a reaction, each of its *atoms* loses its outer electron to form a positively charged *ion*. The more easily the ions form, the more reactive the metal.

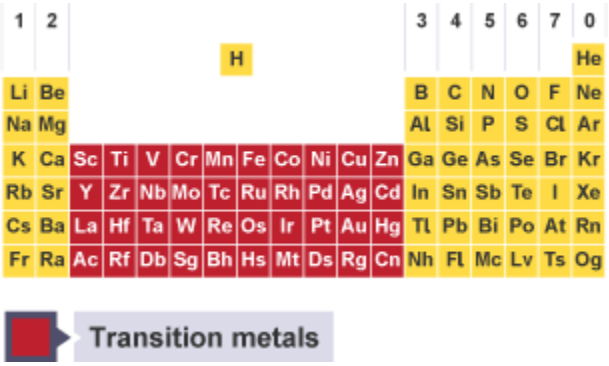
Going down group 1:

- the atoms become larger
- the outer electron becomes further from the *nucleus*
- the force of attraction between the nucleus and the outer electron decreases
- the outer electron is lost more easily

Atomic number	Name	Electronic structure	Diagram of atom
3	Lithium	2.1	
11	Sodium	2.8.1	
19	Potassium	2.8.8.1	

Physical properties of transition elements

Most *metals* are *transition metals*. They include iron, copper and chromium. The transition elements are in the central part of the *periodic table*.



The transition elements are placed between groups 2 and 3 in the periodic table

Physical properties

The transition elements share some *physical properties* with all metals:

- they conduct electricity in the solid and liquid states
- they are shiny when freshly cut

Some properties of transition elements are different from those of the metals in *group 1*. Compared to other metals, most transition metals have:

- higher *melting points*
- higher *densities*

- greater strength
- greater hardness

Remember that these are typical properties – some transition metals may not show one or more of them. For example, mercury *melts* at just -39°C, so it is a liquid at room temperature.

Typical transition elements

The elements below have properties that are typical of transition elements:

- chromium, Cr
- manganese, Mn
- iron, Fe
- cobalt, Co
- nickel, Ni
- copper, Cu

The table shows the melting point and density of some transition elements, compared to three metals in the *periodic table* that are not transition elements.

Metal	Position	Melting point	Density
Sodium	Group 1	98°C	0.97 g/cm^3
Magnesium	Group 2	650°C	1.74 g/cm^3
Aluminium	Group 3	660°C	2.70 g/cm^3
Chromium	Transition elements	1890°C	7.19 g/cm^3
Manganese	Transition elements	1240°C	7.20 g/cm^3
Iron	Transition elements	1538°C	7.87 g/cm^3
Cobalt	Transition elements	1492°C	8.90 g/cm^3
Nickel	Transition elements	1453°C	8.90 g/cm^3

Metal	Position	Melting point	Density
Copper	Transition elements	1083°C	8.92 g/cm^3

Worked example

Compare the properties of sodium and chromium. How does the data show that chromium is a typical transition element, and sodium is a typical group 1 metal?

Answer

Chromium has a much higher *melting point* than sodium. The *density* of chromium is also greater than sodium. Most transition elements have high melting points and densities, so chromium is a typical transition element. The group 1 elements have low melting points and densities, so sodium is a typical group 1 element.

Question

Predict the differences in the hardness and strength of nickel and potassium. Give a reason for your prediction.

Answer

Nickel is harder and stronger than potassium. The reason for this prediction is that nickel is a transition element and potassium is in group 1.

Chemical properties of transition elements

There are differences in the *chemical properties* of metals in *group 1* and the *transition elements*.

Chemical reactions

Reactions with oxygen

The group 1 elements react quickly with oxygen in the air at room temperature. Most transition elements react slowly, or not at all, with oxygen at room temperature. Some transition metals react with oxygen on heating, for example:

copper + oxygen → copper oxide

$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$

Reactions with water

The group 1 elements react vigorously with cold water. Most transition elements react slowly with cold water, or not at all.

Iron reacts with water and oxygen at room temperature to form hydrated iron(III) oxide, or rust.

Reactions with halogens

The group 1 elements react vigorously with the halogens. Some transition elements also react with halogens, for example:

iron + chlorine → iron(III) chloride

$Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$

Ions with different charges

Transition elements form *ions* with different charges. For example:

- manganese forms Mn^{2+} and Mn^{3+} ions
- copper forms Cu^{+} and Cu^{2+} ions

Coloured compounds

Metals that are not transition elements usually form white *compounds*. Transition elements form coloured compounds.

The table shows the colours of some iron compounds.

Compound	**Colour**
Iron(II) hydroxide, $Fe(OH)_2$	Pale green
Iron(III) hydroxide, $Fe(OH)_3$	Orange-brown
Iron(III) oxide, Fe_2O_3	Red-brown

Catalytic activity

Catalysts are substances that increase the *rate of reaction* without being used up in the reaction. For example:

- iron is the catalyst in the *Haber process*, which makes ammonia
- manganese(IV) oxide increases the decomposition of hydrogen peroxide to oxygen and water

WEEK 3

Alkanols / Alcohols

The alkanols/*alcohols* form a *homologous series*. Like all homologous series, the alcohols:

- have the same *general formula*

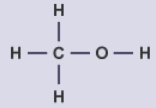
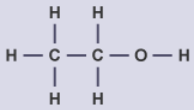
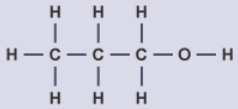
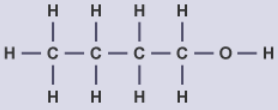
- differ by CH_2 in *molecular formulae* from neighbouring *compounds*
- show a gradual variation in physical *properties*, such as their *boiling points*
- have similar chemical properties

Functional group

The *functional group* in the alcohols is the *hydroxyl* group, -OH. It is responsible for the typical reactions of alcohols. Take care not to confuse the -OH group with the hydroxide ion, OH^-.

Structures

The table shows four alcohols, their molecular formulae and their structures.

Alcohol	Formula
Methanol	CH_3OH
Structure (showing all the covalent bonds)	
	
Alcohol	**Formula**
Ethanol	C_2H_5OH
Structure (showing all the covalent bonds)	
	
Alcohol	**Formula**
Propanol	C_3H_7OH
Structure (showing all the covalent bonds)	
	
Alcohol	**Formula**
Butanol	C_4H_9OH
Structure (showing all the covalent bonds)	
	

Making ethanol by fermentation

Ethanol is the *alcohol* found in beer, wine and other alcoholic drinks. It is also used as a *fuel* for vehicles, either on its own or mixed with *petrol*. Ethanol can be produced by *fermentation* and concentrated using *fractional distillation*.

Fermentation

Fermentation is an *anaerobic* process:

glucose → ethanol + carbon dioxide

Yeast, a type of single-celled *fungus*, provides the *enzymes* needed for fermentation. If the yeast *cells* become too cold, fermentation happens very slowly, or may not happen at all. If the yeast cells become too hot, their enzymes become *denatured* and fermentation stops.

The typical conditions needed for fermentation include:

- sugars *dissolved* in water, and mixed with yeast
- an air lock to allow carbon dioxide out, while stopping air getting in
- warm *temperature*, 25-35°C

The yeast dies when the ethanol concentration reaches about 15%. Fermentation is a slow reaction and takes several days or weeks to finish. If air is present, the oxygen causes the ethanol to oxidise to ethanoic acid, so the drink tastes of vinegar.

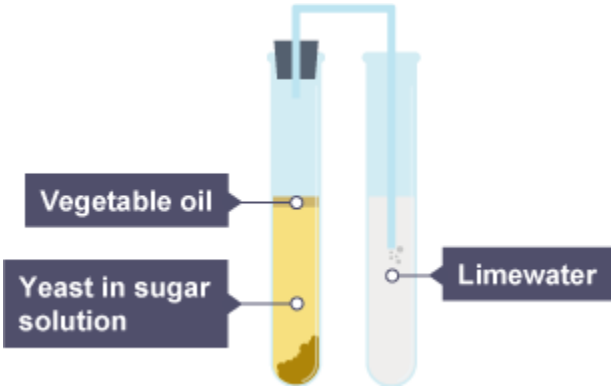


Fig. Apparatus used to investigate fermentation in the lab.

Question

Describe and explain what happens to the limewater during the experiment shown above.

Answer

The limewater turns milky. This is because carbon dioxide is produced during fermentation, and passes through the limewater.

Reactions of alkanols

Uses of the first four alkanols

Methanol is used as a chemical *feedstock*. It is *toxic*, so it is deliberately added to industrial ethanol (methylated spirits) to prevent people from drinking it.

Ethanol is the alcohol present in alcoholic drinks. It is also used as a *fuel* and a *solvent*.

Propanol and butanol are also used as solvents and fuels.

Combustion

The alcohols undergo *complete combustion* to form carbon dioxide and water. For example, ethanol is used as a fuel:

ethanol + oxygen → carbon dioxide + water

$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$

When less oxygen is present, *incomplete combustion* will occur, producing water and either carbon monoxide or carbon.

Reactions with sodium

If a small piece of sodium is dropped into ethanol, bubbles of hydrogen gas are produced and the liquid contains sodium ethoxide. The reaction is:

sodium + ethanol → sodium ethoxide + hydrogen

$2Na + 2C_2H_5OH \rightarrow 2C_2H_5ONa + H_2$

Methanol, propanol and butanol undergo similar reactions.

Solubility in water

When the alcohols with the shortest hydrocarbon chains, eg methanol, ethanol or propanol, are added to water, they mix easily to produce a solution. However, the solubility decreases as the length of the alcohol molecule gets longer, so butanol is less soluble than propanol. It may not mix easily, and two distinct layers might be left in the container.

Oxidation of alcohols

The alcohols can also be *oxidised* without *combustion* to produce *carboxylic acids*. For example, ethanol can be oxidised to ethanoic acid using an *oxidising agent*.

It is easier to understand what happens if ethanol is shown as CH_3CH_2OH in the balanced equation:

ethanol + oxidising agent → ethanoic acid + water

$CH_3CH_2OH + 2[O] \rightarrow CH_3COOH + H_2O$

Each of the two oxygen atoms provided by the oxidising agent are shown as [O]. Notice that the left-hand side of the ethanol molecule is unchanged. The reaction involves the -OH group on the right-hand side.

Question

Propanol is oxidised by heating with an oxidising agent. Name the carboxylic acid formed in the reaction.

Answer

Propanoic acid.

WEEK 4

CHEMICAL REACTIONS

Energy level diagrams

Energy level diagrams are used to model *energy* changes during reactions. They show the relative energy levels of the *products* and *reactants*.

Exothermic reaction

The energy level decreases in an *exothermic* reaction. This is because energy is given out to the surroundings.

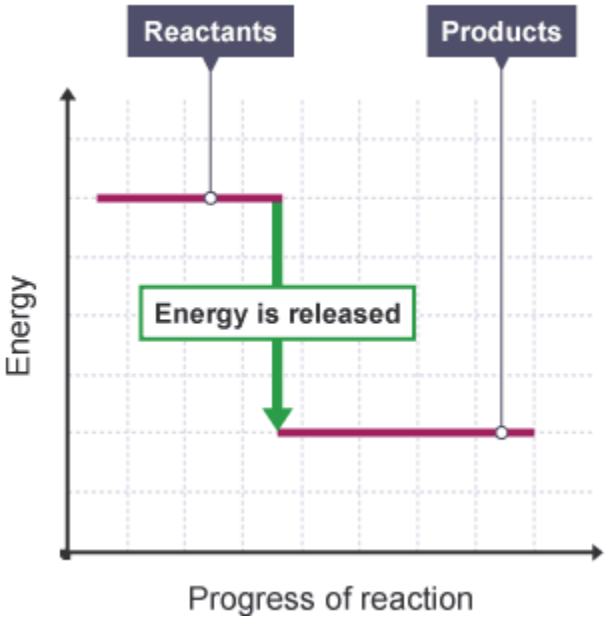
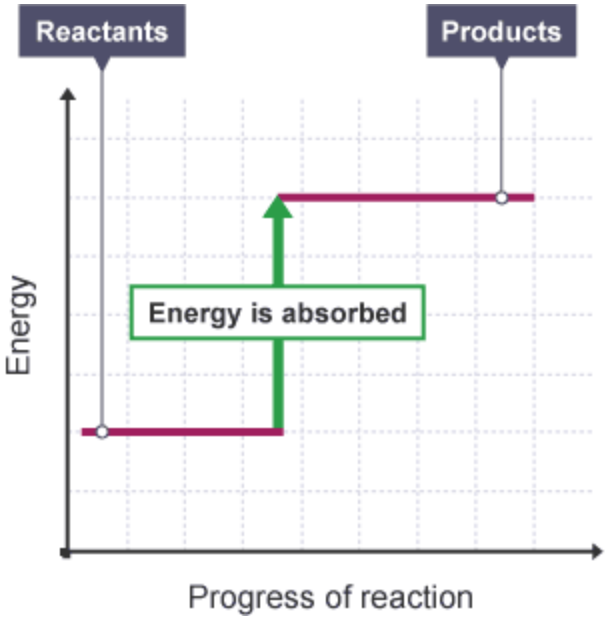


Figure caption,

A downwards arrow shows that energy is given out

Endothermic reaction

The energy level increases in an *endothermic* reaction. This is because energy is taken in from the surroundings.



An upwards arrow shows that energy is taken in

Activation energy

When a fuel burns, a spark is needed to start the reaction. A minimum amount of *energy* is needed in order for a fuel to burn.

The minimum amount of energy needed for a reaction to take place is called the *activation energy*.

High and low activation energy

In order for two substances to react, their *particles* must collide with enough energy. The particles in a substance have a range of different energies.

- A **low activation energy** means that a lot of the particles will collide with enough energy to react. The **reaction will be fast**.
- A **high activation energy** means that far fewer particles will collide with enough energy. The **reaction will be slow**.

Reaction profiles

A *reaction profile* shows how the energy in the *reactants* and *products* changes during a reaction. It includes the activation energy.

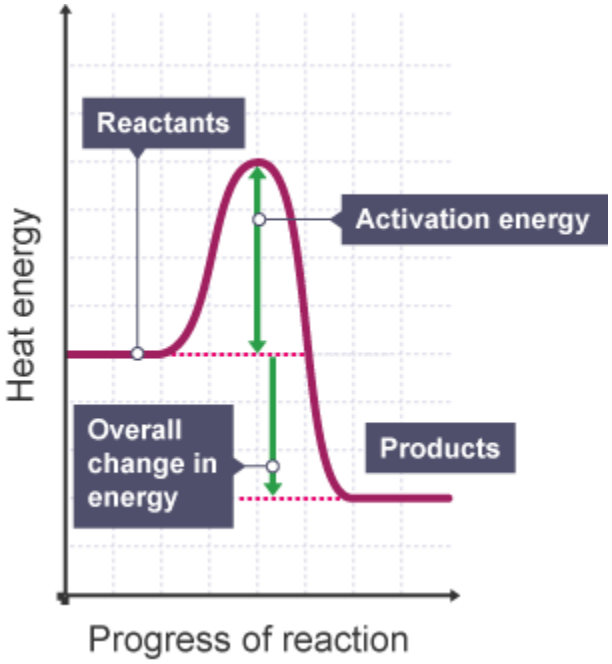
The activation energy is shown as a 'hump' in the line, which:

- starts at the energy of the reactants
- is equal to the difference in energy between the top of the 'hump' and the reactants

The overall change in energy in a reaction is the difference between the energy of the reactants and products.

Exothermic reactions

The diagram shows a reaction profile for an *exothermic* reaction.



A reaction profile for an exothermic reaction

WEEK 5

EQUILIBRIUM OF A REACTION

Key learning points

- In a reversible reaction the *products* can react to turn back into the *reactants.*
- A dynamic equilibrium is formed in a reversible reaction when the rates of the forward and reverse reaction are equal and the amounts of the reactants and the products remains constant.
- *Le Châtelier's principle* can be used to predict changes to the position of an equilibrium. This is useful when explaining the conditions used in the *Haber process* (making ammonia).

What are reversible reactions?

In many chemical reactions the reactants are converted into products and then the reaction stops.

For example:

$2Mg + O_2 \rightarrow 2MgO$

In this reaction, once the magnesium oxide has formed, it cannot go back to reform magnesium and oxygen.

This is a non-reversible reaction.

Other chemical reactions are reversible reactions.

In these reactions, the products, once formed, can change back into the original reactants.

For example:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

In this reaction, once the ammonia (NH_3) has formed, it can 'go back' to reform nitrogen and hydrogen.

Reversible reactions are shown with reversible arrows (⇌).

The reversible arrow has two 'half-arrows,' one pointing in each direction.

What is an example of a reversible reaction?

Hydrated copper(II) sulfate is blue in appearance.

When it is heated, the *water of crystallisation* is removed, leaving anhydrous copper(II) sulfate which is white in appearance.

Equation: hydrated copper(II) sulfate ⇌ anhydrous copper(II) sulfate + water

$CuSO_4.5H_2O(s) \rightleftharpoons CuSO_4(s) + 5H_2O(l)$

This reaction is reversible, as the hydrated copper(II) sulfate is reformed if water is added to anhydrous copper(II) sulfate.

***Fig. 1**. A Bunsen burner is used to heat an evaporating basin containing hydrated copper(II) sulfate.*

Fig. 2 *Water is driven off, leaving anhydrous copper(II) sulfate.*

Fig. 3 *The Bunsen burner is turned off and water is added using a pipette*

Fig. 4 *The evaporating basin now contains hydrated copper(II) sulfate again.*

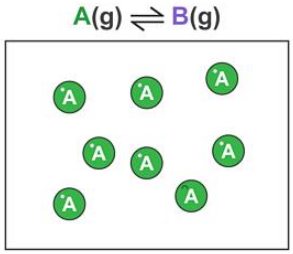
What is dynamic equilibrium?

*When a reversible reaction happens in a **closed system** (closed container) , a dynamic equilibrium can be achieved.*

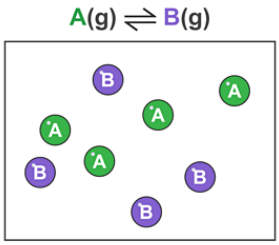
In a dynamic equilibrium:

- *the rates of the forward and reverse reactions are equal.*
- *the amounts of reactants and products remains constant.*

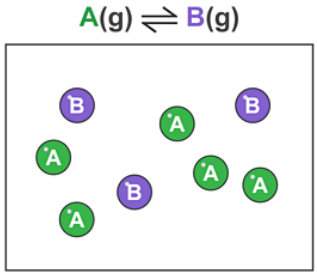
For example, consider a reversible reaction where gas A reacts to form a different gas B:



1. *At the start of the reaction, the container contains only A.*



2. As the reaction proceeds, A is converted into B.



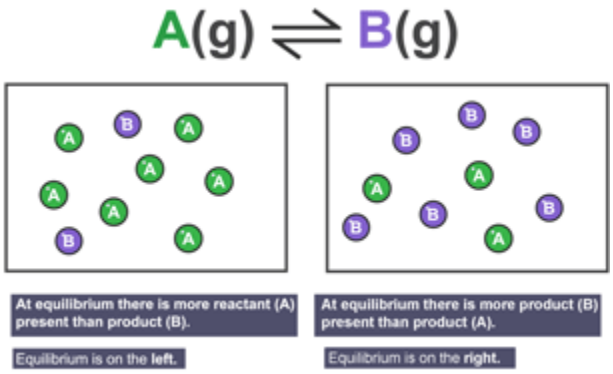
3. Once B has formed, it can reverse and turn back into A. Eventually the rate of the forward and reverse reactions are equal, and the amount of A and B remain fixed (even though both reactions are still occurring).

How to change the position of equilibrium (Higher tier only)

*The position of an equilibrium can be described as on the **'left'** or **'right'**.*

This describes how much of the reactants (on the left) and the products (on the right) are present at equilibrium.

The diagram below shows an example of an equilibrium on the left and an equilibrium on the right.



What is Le Châtelier's principle?

The equilibrium position can be changed by altering the reaction conditions, such as by:

- *changing the **pressure**.*
- *changing the **concentration**.*
- *changing the **temperature**.*

*The effect of changing a reaction condition on the position of equilibrium can be predicted using **Le Châtelier's principle**.*

This states that if a change is made to the conditions of a system at equilibrium, then the position of equilibrium moves to oppose that change in conditions.

What happens when you change the concentration?

If the concentration of a reactant or product is changed in a reaction at equilibrium, the position of the equilibrium moves to oppose the change.

For example:

$CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$

If the reaction is at equilibrium and the concentration of CH_4 is increased:

- *The equilibrium moves to decrease the concentration of CH_4.*
- *The forward reaction is favoured which reduces the concentration of CH_4.*
- *The equilibrium position moves to the right.*

Question

For the same equilibrium described above, what would happen to the position of the equilibrium if the concentration of CO was increased?

Answer

- *The equilibrium moves to decrease the concentration of CO.*
- *The reverse reaction is favoured which reduces the concentration of CO.*
- *The equilibrium position moves to the left.*

What happens when you change the pressure?

For a reaction involving gases, the greater the number of gas particles present, the greater the pressure in the system.

If the pressure exerted on a reaction at equilibrium is changed, the position of the equilibrium moves to oppose the change in pressure.

For example:

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

If the reaction is at equilibrium and the pressure is increased:

- *The equilibrium moves to decrease the pressure.*
- *The equilibrium moves in the direction of the side of the reaction with fewer gas particles to reduce the pressure.*
- *The equilibrium position moves to the right, as there are two gas particles on the right-hand side of the equation, and 4 gas particles on the left-hand side of the equation.*

Question

*For the same equilibrium described above, what would happen to the position of the equilibrium if the pressure was decreased**?***

Answer

- *The equilibrium position moves to the left, as there are four gas particles on the left-hand side of the equation, and 2 gas particles on the right-hand side of the equation.*
- *This increases the pressure, which opposes the change that has been made.*

What happens when you change the temperature?

In a reversible reaction, if the reaction is exothermic in one direction, it is endothermic in the other direction.

If the temperature of a reaction at equilibrium is changed, the position of the equilibrium moves to oppose the change in temperature.

For example:

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ (forward reaction is exothermic)

If the reaction is at equilibrium and the temperature is increased:

- *The equilibrium moves to decrease the temperature.*
- *The reverse reaction is endothermic, so this reaction is favoured to decrease the temperature.*
- *The equilibrium moves to the left.*

Question

For the same equilibrium described above, what would happen to the position of the equilibrium if the temperature was decreased?

Answer

- *The equilibrium moves to increase the temperature.*
- *The forward reaction is exothermic, so this reaction is favoured to increase the temperature.*
- *The equilibrium moves to the right.*

Adding a catalyst

- Adding a *catalyst* has no effect of the position of an equilibrium as it increases the rate of the forward and reverse reactions equally.

What is the Haber process?
Ammonia is an important industrial product used to make fertilisers, explosives and dyes.
It is manufactured using the **Haber process**.
This involves a reversible reaction between **nitrogen** and **hydrogen**:
$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
The reaction can reach a *dynamic equilibrium*.

Conditions

The following conditions are used in the Haber process:

Catalyst	*Iron*

Temperature	*450°C*
Pressure	*200 atm*

The iron catalyst is used to increase the rate of the reaction.

The temperature and pressure conditions used are selected to give a compromise between the yield of ammonia, the rate of the reaction, and the cost of the process.

Temperature

The reaction between nitrogen and hydrogen is exothermic in the forward direction.

Therefore, a low temperature would cause the equilibrium to move to the right and increase the yield of ammonia.

However, a low temperature would cause a very slow rate of reaction.

The temperature used (450°C) allows a compromise between yield and rate.

Pressure

There are fewer gas particles on the right-hand side of the reaction.

Therefore, a high pressure would cause the equilibrium to move to the right and increase the yield of ammonia.

However, a high pressure is expensive to implement and would increase the risk of explosion.

The pressure used (200 atm) allows a compromise between yield and cost.

WEEK 6

STOICHIOMETRY OF CHEMICAL REACTIONS

STOICHIOMETRY OF CHEMICAL REACTIONS

BASIC CONCEPT

The mole

Introducing the mole, the unit of measurement for the number of particles in a substance

The actual *masses* of *atoms*, *molecules* and *ions* are too small to be useful in calculations. For this reason, chemical amounts are measured in *moles*. The mole is a unit. Its symbol is mol.

The mass of one mole of a substance is equal to:

- the *relative atomic mass* (A_r) of its formula in grams if the substance is an *element*

- the *relative formula mass* (M_r) of its formula in grams if the substance is a *compound*

The table shows the masses of one mole of three substances.

Substance	Formula	Relative formula mass	Mass of 1 mol
Calcium	Ca	40	40 g
Oxygen	O_2	2 × 16 = 32	32 g
Calcium carbonate	$CaCO_3$	40 + 12 + (3 × 16) = 100	100 g

Avogadro constant

One mole of a substance contains the same number of *particles* as one mole of any other substance. The particles can be atoms, molecules or ions.

The number of atoms, molecules or ions in one mole of a substance is called the *Avogadro constant*. Its value is 6.02×10^{23} per mole, which is 602,000,000,000,000,000,000,000 per mole.

The amount in moles can apply to atoms, molecules, ions and electrons. For example, the number of atoms in 1 mol of sulfur is the same as the number of molecules in 1 mol of sulfur dioxide. This is the same as the number of sodium ions in 1 mol of sodium chloride. The number of chloride ions in 1 mol of sodium chloride is also the same.

Key fact

One mole, 1 mol, of a substance is the Avogadro constant number (6.02×10^{23}) of particles (atoms, molecules, ions or formulae) of that substance.

Calculating the number of particles

The number of particles in a substance can be calculated using:

- the Avogadro constant (6.02×10^{23})
- the amount of substance in mol

Number of particles = Avogadro constant × the amount of substance in mol

Example

Calculate the number of water molecules in 0.5 mol of water.

Number of water molecules = Avogadro constant x amount of substance in mol

$= 6.02 \times 10^{23} \times 0.5$

$= 3.01 \times 10^{23}$

It is important to state the particles involved. For example, 3.01×10^{23} water molecules contain 9.03×10^{23} atoms. This is because one water molecule, H_2O, is made up of three atoms.

Question

Calculate the number of oxygen atoms in 0.5 mol of oxygen molecules, O_2.

Answer

Number of atoms = 6.02×10^{23} per mol × 0.5 mol × 2

= 6.02×10^{23}

Moles and masses

The *mass* of a substance can be calculated from the number of *moles*, and the number of moles of a substance can be calculated from its mass. The link between the two quantities is the *relative formula mass*.

Calculating masses

Example 1

Calculate the mass of 0.25 mol of carbon dioxide molecules. (M_r of CO_2 = 44)

Mass = relative formula mass × amount

$= 44 \times 0.25$

= 11 g

Example 2

Calculate the mass of 0.10 mol of iron. (A_r of Fe = 56)

Mass = 56 × 0.10

= 5.6 g

The calculation is the same if a substance is a metal or exists as separate atoms, but its A_r is used instead of an M_r.

Question

Calculate the mass of 0.10 mol of iron. (A_r of Fe = 56)

Answer

Mass = 56 × 0.10

= 5.6 g

Calculating amounts in moles

Key fact

The amount of a given mass substance is calculated using:

$$\textit{Amount} = \frac{mass}{relative\ atomic\ or\ formular\ mass}$$

Use A_r instead of M_r for metals or separate atoms.

Example

Calculate the amount of carbon atoms in 6.0 g of carbon. (A_r of C = 12)

$$\textit{Amount} = \frac{mass}{Ar}$$

$$\textit{Amount} = \frac{6.0}{12}$$

= 0.5 mol

Question

Calculate the amount of water molecules in 36 g of water. (M_r of water = 18)

$$\textit{Amount} = \frac{mass}{relative\ formular\ mass}$$

$$\textit{Amount} = \frac{36}{18}$$

= 2.0 mol

Balanced chemical equations

A *balanced equation* models a chemical reaction using the *formulae* of the *reactants* and *products*. It shows the number of units of each substance involved.

Balancing an equation

If you just write an equation replacing names with formulae, it may not be balanced. The numbers of *atoms* of each *element* on the left must be the same as they are on the right.

To balance an unbalanced equation, you need to add numbers to the left of one or more formulae. Here is one way to work out how to do this for the reaction between nitrogen and hydrogen.

Step	**Result**
Check to see if there are equal numbers of atoms of each element on both sides. Here there aren't.	$N_2 + H_2 \rightarrow NH_3$
There are two nitrogen atoms on the left but only one on the right, so a big 2 is added to the left of the NH_3.	$N_2 + H_2 \rightarrow 2NH_3$
There are two hydrogen atoms on the left but (2 × 3) = 6 on the right, so a big 3 is placed in front of the H_2.	$N_2 + 3H_2 \rightarrow 2NH_3$
Check to see if there are equal numbers of each element of both sides. There are.	(two nitrogen atoms and six hydrogen atoms)
Add the state symbols if they are requested.	$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

Key fact

Balanced equations only show formulae, not names. A balancing number multiplies all the atoms in the substance next to it.

State symbols

Balanced equations often include *state symbols* in brackets after each formula. They show the physical state of that substance.

State symbol	**Meaning**
(s)	Solid
(l)	Liquid

State symbol	Meaning
(g)	Gas
(aq)	Aqueous solution

An *aqueous* solution forms when a substance dissolves in water.

State symbols are useful because they show what a substance is like. For example:

- $H_2O(l)$ is liquid water, but $H_2O(g)$ is steam
- HCl(g) is hydrogen chloride gas, but HCl(aq) is hydrochloric acid

- # Question
- Sodium metal reacts with water to form sodium hydroxide solution and hydrogen gas. Write a balanced equation for the reaction, including state symbols.

Answer

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ or

$Na(s) + H_2O(l) \rightarrow NaOH(aq) + ½H_2(g)$

Mole calculations

A *balanced equation* shows the amounts in *moles* of *reactants* that react and the amounts of *products* that are made. From these amounts, the *masses* of reactants and products can be calculated.

Amounts from equations

In a balanced equation, the coefficients (the numbers in front of the formulae) show the amounts of the reactants and products. If there is no coefficient, the amount is one mole.

For example, the equation below shows that one mole of zinc reacts with two moles of hydrochloric acid to make one mole of zinc chloride and one mole of hydrogen molecules.

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Example

Use the equation below to calculate the amount in moles of oxygen molecules that reacts with 2 mol of magnesium metal.

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

The equation shows that 2 mol of magnesium metal reacts with 1 mol of oxygen molecules.

Masses in equations

The masses of substances shown in a balanced equation can be calculated using the equation:

mass = relative formula mass × amount

For example, in the equation $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ the masses of the substances shown are:

Magnesium, Mg	**24 × 2 = 48 g**
Oxygen, O_2	**2 × 16 = 32 g**
Magnesium oxide, MgO	**(24 + 16) × 2 = 80 g**

Calculating masses from equations and given masses

Balanced equations and relative formula mass values can be used to calculate:

- the mass of product made from a given mass of reactant
- the mass of reactant needed to make a given mass of product

Example

In the reaction shown by the equation below, what mass of sulphur dioxide can be made from 16 g of sulphur? (M_r of SO_2 = 64)

$S(s) + O_2(g) \rightarrow SO_2(g)$

Amount of S = $\frac{mass}{relative\ atomic\ mass}$

$= \frac{16}{32}$

= 0.5 mol

The equation shows that 1 mol of sulfur reacts with 1 mol of oxygen molecules to make 1 mol of sulfur dioxide. This means that 0.5 mol of sulfur makes 0.5 mol of sulfur dioxide.

mass of SO_2 = relative formula mass × amount

= 64 × 0.5

= 32 g

Question

In the reaction shown by the equation below, what mass of nitrogen, N_2, is needed to make 120 g of nitrogen monoxide, NO? (M_r of NO = 30 and M_r of N_2 = 28)

$N_2(g) + O_2(g) \rightarrow 2NO(g)$

Amount of nitrogen dioxide = $\frac{mass}{relative\ atomic\ mass}$

$= \frac{120}{30}$

$= 4.0\ mol$

1 mol of N_2 makes 2 mol of NO.

This means that 2 mol of N_2 makes 4.0 mol of NO.

Mass of 2 mol of $N_2 = 2 \times 28$

= 56g.

Reactions and moles

Limiting reactants

A reaction finishes when one of the reactants is all used up. The other reactant has nothing left to react with, so some of it is left over:

- *the reactant that is all used up is called the limiting reactant*
- *the reactant that is left over is described as being in excess*

The mass of product formed in a reaction depends upon the mass of the limiting reactant. This is because no more product can form when the limiting reactant is all used up.

Reacting mass calculations

The maximum mass of product formed in a reaction can be calculated using:

- *the balanced equation*
- *the mass of the limiting reactant, and*
- *the A_r (relative atomic mass) or M_r (relative formula mass) values of the limiting reactant and the product*

Example

12 g of magnesium reacts completely with excess hydrochloric acid to form magnesium chloride and hydrogen:

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Calculate the maximum mass of hydrogen that can be produced. (A_r of Mg = 24, M_r of H_2 = 2)

Amount of magnesium =

Amount of magnesium =

= 0.5 mol

Looking at the equation, 1 mol of Mg forms 1 mol of H_2, so 0.5 mol of Mg forms 0.5 mol of H_2

Mass of H_2 = M_r × amount

= 2 × 0.5

= 1 g

Question

1.0 g of calcium carbonate decomposes to form calcium oxide and carbon dioxide:

$CaCO_3(g) \rightarrow CaO(s) + CO_2(g)$

Calculate the maximum mass of carbon dioxide that can be produced. (M_r of $CaCO_3$ = 100, M_r of CO_2 = 44)

Amount of calcium carbonate = $\frac{1.0}{100}$

= 0.01 mol

Looking at the equation, 1 mol of $CaCO_3$ forms 1 mol of CO_2, so 0.01 mol of $CaCO_3$ forms 0.01 mol of CO_2

Mass of CO_2 = relative formula mass × amount

= 44 × 0.01

*= **0.44 g***

Calculating balancing numbers

The balancing numbers in an equation can be worked out using masses found by experiment.

Example

6.0 g of magnesium reacts with 4.0 g oxygen to produce magnesium oxide, MgO. Deduce the balanced equation for the reaction. (A_r of Mg = 24, M_r of O_2 = 32)

Step	Action	Result	Result
1	Write the formulae of the substances	Mg	O_2
2	Calculate the amounts	$\frac{6.0}{24}$ = 0.25 mol	$\frac{4.0}{32}$ = 0.125 mol

Step	Action	Result	Result
3	Divide both by the smaller amount	$\frac{0.25}{0.125} = 2$	$\frac{0.125}{0.125} = 1$

This means that 2 mol of Mg reacts with 1 mol of O_2, so the left-hand side of the equation is:

$2Mg + O_2$

Then balancing in the normal way gives: $2Mg + O_2 \rightarrow 2MgO$

Calculating concentrations

The *concentration* of a *solution* can be shown in g/dm^3 or mol/dm^3. It is often more useful to know the concentration of a *reactant* in mol/dm^3 so that the *amount* of reactant in a given volume can be calculated.

Concentration using moles

The concentration of a solution can be calculated using:

- the amount of *dissolved solute* in *moles*, mol
- the volume of solution (or *solvent*) in cubic decimetres, dm^3

Key fact

Concentration of mol/dm^{-3} = $\frac{\textbf{\textit{amount of solute in mol}}}{\textbf{\textit{volume in}}\ \text{dm3}}$

0.5 mol of sodium hydroxide is dissolved in 2 dm^3 of water. Calculate the concentration of the sodium hydroxide solution formed.

Concentration = $\frac{\textbf{\textit{amount of solute in mol}}}{\textbf{\textit{volume in}}\ \text{dm3}}$

$= \frac{0.5\ mol}{2\ dm3}$

Concentration = **0.25 mol/dm^3**

Volume units

Volumes used in concentration calculations must be in dm^3, not in cm^3. It is useful to know that 1 dm^3 = 1000 cm^3. This means:

- divide by 1000 to convert from cm^3 to dm^3
- multiply by 1000 to convert from dm^3 to cm^3

For example, 250 cm^3 is 0.25 dm^3 (250 ÷ 1000). It is often easiest to convert from cm^3 to dm^3 before continuing with a concentration calculation.

Question

100 cm^3 of dilute hydrochloric acid contains 0.02 mol of dissolved hydrogen chloride. Calculate the concentration of the acid in mol/dm^3.

Answer

Volume of acid = 100 ÷ 1000 = 0.1 dm^3

Concentration of acid = 0.02 mol ÷ 0.1 dm^3

= 0.2 mol/dm^3

Converting between units

The *relative formula mass* of the solute is used to convert between mol/dm^3 and g/dm^3:

- to convert from mol/dm^3 to g/dm^3, multiply by the relative formula mass
- to convert from g/dm^3 to mol/dm^3, divide by the relative formula mass

Remember: the molar mass is the A_r or M_r in grams per mol.

Example

Calculate the concentration of 0.1 mol/dm^3 sodium hydroxide solution in g/dm^3. (M_r of NaOH = 40)

Concentration = 0.1 × 40

= 4 g/dm^3

Question

Calculate the concentration of 7.3 g/dm^3 hydrochloric acid in mol/dm^3. (M_r of HCl = 36.5)

Answer

Concentration = 7.3 ÷ 36.5

= 0.2 mol/dm^3

Volumes of solutions in reactions

Calculating amounts from concentration and volume

The amount in *moles* of a *solute* in a given volume of *solution* can be calculated if the *concentration* of the solution is known. The mass of solute can then be calculated.

Rearranging the equation for concentration below:

Concentration = $\frac{\textbf{amount of solute in mol}}{\textbf{volume in } dm^3}$

Gives:

Amount of solute in mol = concentration in mol/dm^3 × volume in dm^3

Example

Calculate the amount of sodium hydroxide, NaOH, in 25.0 cm3 of solution of concentration 0.1 mol/dm^3.

Converting the volume from cm^3 to dm^3, 25.0 cm^3 = 25.0 ÷ 1000 = 0.025 dm^3.

Amount of NaOH in mol = concentration in mol/dm3 × volume in dm3

= 0.1 mol/dm3 × 0.025 dm^3

= 0.0025 mol

Calculating masses from concentration and volume

If the amount in mol of a solute in a given volume of solution is known, its mass can also be calculated.

Question

For the example above, calculate the mass of NaOH in 25.0 cm^3 of solution. (*M*r = 40 g)

Answer

Mass = amount × relative formula mass

= 0.0025 × 40

= 0.1 g

More concentration calculations

If the volumes of two solutions that react completely are known, and the concentration of one of the solutions is known, then the concentration of the other solution can be calculated.

Example

25.0 cm^3 of hydrochloric acid, HCl, of concentration 1.0 mol/dm^3 reacts with 20.0 cm^3 of sodium hydroxide, NaOH, solution. Calculate the concentration of the NaOH solution. The equation for the reaction is HCl(aq) + NaOH(aq) → NaCl(aq) + H_2O(l)

Converting volumes from cm^3 to dm^3:

- volume of HCl = 25.0 cm^3 = 25.0 ÷ 1000 = 0.025 dm^3
- volume of NaOH = 20.0 cm^3 = 20.0 ÷ 1000 = 0.020 dm^3

Substituting into the equation to find the amount of HCl in the given volume:

Amount of HCl in mol = concentration in mol/dm^3 × volume in dm^3

= 1.0 mol/dm^3 × 0.025 dm^3

= 0.025 mol

The equation shows that 1 mol of HCl reacts with 1 mol of NaOH. So, the amount of NaOH in 20.0 cm^3 is 0.025 mol.

Substituting into the equation to find the concentration of NaOH:

Concentration of NaOH /dm^3 = $\frac{\textbf{amount of solute in mol}}{\textbf{volume in } dm^3}$

Concentration = $\frac{\mathbf{0.025\ mol}}{\mathbf{0.020\ dm^3}}$

= 1.25 mol/dm^3

Percentage yield

The *theoretical yield* is the maximum possible *mass* of a *product* that can be made in a chemical reaction.

It can be calculated from:

- the *balanced equation*,
- the mass and *relative formula mass* of the *limiting reactant*, and
- the relative formula mass of the product

Even though no atoms are gained or lost in a chemical reaction, it is not always possible to obtain the calculated amount of a given product. Reasons why the mass of product made is less than the maximum theoretical mass include:

- the reaction not going to completion, because it is reversible
- some of the product may be lost when it is separated from the reaction mixture by *filtering*, for example
- some of the reactants may react in ways different to the expected reaction

Calculating percentage yield

The *percentage yield* is calculated using this equation:

Key Fact

The percentage yield can vary from 100% (no product has been lost) to 0% (no product has been made).

Worked example

Copper oxide reacts with sulfuric acid to make copper sulfate and water. In an experiment, 1.6 g of dry copper sulfate crystals are made. If the theoretical yield is 2.0 g, calculate the percentage yield of copper sulfate.

Actual yield = 1.6 g

Percentage yield = $\frac{1.6}{2.0}$ × 100

Percentage yield = 80%

Question

In an experiment, the theoretical yield is 3.2 g but the actual yield is only 2.4 g. Calculate the percentage yield.

Answer

Percentage yield = $\frac{2.4}{3.2}$ × 100

Percentage yield = 75%

WEEK 7

HEAT AND CHEMICAL REACTION

HEAT OF REACTION AND CHEMICAL BONDS

During chemical reactions, chemical bonds are broken, atoms are regrouped and new bonds are formed. Bond breaking requires energy and bond forming evolves energy. The minimum amount of energy required for bond breaking is called activation energy. While bond breaking is endothermic, bond forming is exothermic. Thus, heat of reaction comes from breaking and forming of chemicals bond. Heat reaction is negative [exothermic] when bond-breaking energy is less than bond forming energy. Heat of reaction is positive [endothermic] when bond-breaking energy is more than bond forming energy.

TYPES OF HEAT CHANGES IN CHEMICAL REACTIONS

HEAT OF FORMATION

The amount of heat evolved or absorbed when one mole of a substance is formed from its elements is known as heat of formation [or enthalpy of formation].

The standard heat of formation of a substance($\Delta Hf\theta$) is the heat evolved or absorbed, when one mole of that substance is formed from its elements under standard conditions.

For the formation of 1 mole of liquid water, the equation is

$H_2(g) + 1/2O_2(g) \rightarrow H_2O(1)$ $\Delta Hf\theta = -285 kJmol^{-1}$

Thus, $\Delta Hf\theta$ of water = $-285 kJmol^{-1}$

HEAT OF NEUTRALIZATION

Neutralization is an exothermic reaction. The amount of heat evolved during a neutralization reaction in which one mole of water is formed is known as the heat of neutralization (or enthalpy of neutralization).The standard heat of neutralization $\Delta Hn\theta$ is the amount of heat evolved when 1 mole of hydrogen ions, H^+, from an acid reacts with 1 mole of hydroxide ions, OH^-, from an alkali to form 1 mole of water under standard conditions. Heat of neutralization is also known as heat of formation of one mole of water from its ionic components.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ $\Delta Hn\theta = -57.4 kJmol^{-1}$

HEAT OF COMBUSTION

Combustion reaction is always exothermic. The amount of heat evolved when one mole of a substance is burned completely in oxygen is known as the heat of combustion or enthalpy of combustion. The standard heat of combustion of a substance, $\Delta HC\theta$; is the heat evolved when one mole of the substance is burned completely in oxygen under standard conditions. A bomb calorimeter is usually used for accurate determination of heat of combustion.

Heat of combustion can be determined from the relation below:

Heat of combustion = Heat energy produced x molar mass

When the heat evolved by the burning substance is used to raise the temperature of a known mass of water, then the expression for heat of combustion can be given as:

Heat of combustion = $mC\Delta\theta \times$ molar mass

Where m = mass of water

C = Specific heat capacity of water

$\Delta\theta$ = change in temperature, that is, $\theta_2 - \theta_1$

HEAT OF SOLUTION

Heat of solution can be exothermic or endothermic. Heat of solution is the heat evolved or absorbed when one mole of a substance is dissolved in so much water that further dilution results in no detectable heat change.

Standard heat of solution, **$\Delta H_s\theta$**, is the amount of heat evolved or absorbed when 1 mole of substance is dissolved in so much water that further dilution results in no detectable heat change at standard conditions.

WEEK 8

THERMODYNAMICS

Thermodynamics is the study of relationship between heat and other forms of energy.

System in thermodynamics is any part of the universe chosen for thermodynamics consideration, i.e. the physical and chemical phenomenon or process occurring in a given environment. A system can be isolated, closed or open.

Surrounding is the environment in which a phenomenon or a process occurs.

The first law of thermodynamics states that energy can neither be created nor destroyed but may be converted from one form to another.

In thermodynamics, heat is represented by q and other forms of energy are referred to as work denoted by w. The conditions or state of a chemical system is changed when:

1. Heat is evolved or absorbed, and / or
2. Work is done on or by the system

In any case, the internal energy, U, of the system is affected and it is changed.

From first law, heat is changed into internal energy of the system it may be represented by

change in internal energy = Heat absorbed by the system + Work done by the system i.e.

$U = q + w$

Work done by the system is negative since this lead to decrease in internal energy, therefore:

$U = q - w$

For a gaseous system, w = P V

$$U = q - P V$$

$$U = H - P V$$

$$H = U - P V$$

EVALUATION

1. State the first law of thermodynamic
2. Calculate: (a) the heat adsorbed by a system when it does 72J of work and its internal energy decreases by 90J(b) U for a gas that releases 35J of heat and has 128J of work done on it.

SECOND LAW OF THERMODYNAMIC

The second law of thermodynamic states that a spontaneous process occurs only if there is an increase in the entropy of a system and its surroundings Factors which determines the spontaneously of a process are:

1. enthalpy, H: The heat content of the substances involved
2. entropy, S: The measure of degree of disorderliness or randomness of a substance
3. free energy G: The energy which is available for doing work.

ENTROPY(S) OF A SYSTEM

Entropy is the measure of degree of disorderliness or randomness of a system. The standard entropy change ($\Delta S\theta$) is a state function because it depends on the initial and final state of the system. That is:

$\Delta S\theta = S\theta products - S\theta reactants$

The S.Iunit of is JK-1mol-1

Entropy increases from solid to liquid to gaseous state because as you go from solid to liquid to gaseous state, randomness increases, that is; $\Delta S\theta$ tends to positive.

For a reversible process at constant temperature,

$$S = H/T$$

When ΔS is positive, there is increase in entropy. When ΔS is negative there is decrease in the entropy of a system.

GIBB'S FREE ENERGY

The free energy of a system is the energy which is available for doing work in the system; that is, the driving force that brings about a chemical change.

The standard free energy change ($\Delta G\theta$) is a state function because it depends on the initial and final state of the system. That is:

$\Delta G\theta = G\theta products - G\theta reactants$

Free energy takes into account the effect of the enthalpy and entropy factors as represented in the equation below:

$G = H-TS$

For a change at constant temperature,

$G = H - T S$

NOTE:

1. When G is negative, the reaction is spontaneous or feasible.
2. When G is positive, the reaction is not spontaneous, unless the resultant effect of both H and S leads to a net decrease in G
3. When G is zero, the system is in equilibrium

Example: The reaction: $C(s) + O_2(g) \longrightarrow CO2(g)$

is carried out at a temperature of 57oC. If the enthalpy change is -500J and the entropy change is +15J.Calculate the free energy change

Solution:

$G = H - T S$

$= -5000 - (57 + 273) \times (+15)$

$= -5000 - 330 \times 15$

$= -5000 - (+4950)$

$= -5000 - 4950$

$= -9950J$ or $-9.950Kj$

WEEK 9

ACID-BASE REACTION

Making salts from acids and alkalis

A *soluble salt* can be prepared by reacting an *acid* with a dilute *solution* of an *alkali* such as sodium hydroxide or ammonia. The main steps are:

1. Carry out a *titration*. This is to determine the volumes of acid and alkali that must be mixed to obtain a solution containing only salt and water.
2. Mix the acid and alkali in the correct proportions, as determined in step 1, but this time without including an *indicator*.

Pure dry *crystals* can be produced by *crystallisation*, followed by drying on a watch glass or in a warm oven.

Carrying out a titration to find out volumes of acid and alkali solutions that react

Apparatus

The apparatus needed includes:

- a *pipette* to accurately measure the volume of a reactant before transferring it to a conical flask
- a *burette* to add small, measured volumes of one *reactant* to the other reactant

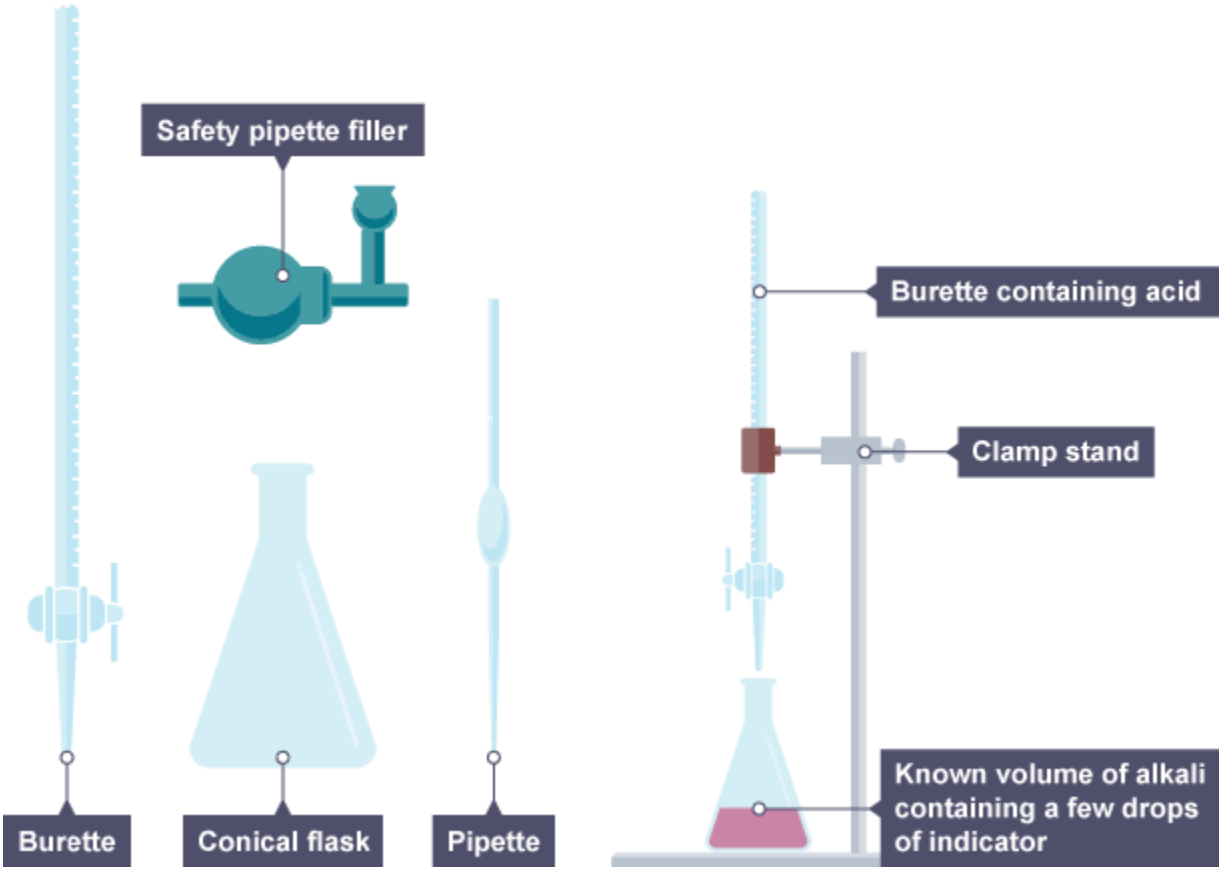


Figure caption,

A pipette filler is needed to use a pipette safely

Method

This is an outline method for carrying out a titration in which an acid is added to an alkali. The method is the same for sulfuric acid, hydrochloric acid and nitric acid.

1. Use the pipette and pipette filler to add a measured volume of sodium hydroxide solution to a clean conical flask.
2. Add a few drops of indicator and put the conical flask on a white tile.
3. Fill the burette with hydrochloric acid and note the starting volume.
4. Slowly add the acid from the burette to the alkali in the conical flask, swirling to mix.
5. Stop adding the acid when the *end-point* is reached (when the indicator first permanently changes colour). Note the final volume reading.
6. Repeat steps 1 to 5 until *concordant titres* are obtained. More accurate results are obtained if acid is added drop by drop near to the end-point.

Results

Record the results in a suitable table. The one here also shows a rough reading.

Run	**End vol**	**Start vol**	**Titre**
Rough	26.85 cm^3	1.00 cm^3	25.85 cm^3
1	24.60 cm^3	0.00 cm^3	24.60 cm^3 ✓
2	24.90 cm^3	0.60 cm^3	24.30 cm^3
3	24.00 cm^3	0.30 cm^3	24.70 cm^3 ✓

Readings should be recorded to two decimal places, ending in 0 or 5 (where the liquid level is between two graduations on the burette). The *titre* is the volume added (the difference between the end and start readings).

Analysis

At least two concordant titres should be ticked (✓) in the table above. Concordant titres often lie within 0.10 cm^3 or less of each other.

Worked example

Calculate the mean titre from the table above. Ignore the rough run, and run 2 (because they are not concordant):

Mean titre = $\frac{(24.60+24.70)}{2}$

= **24.65 cm^3**

Question

Explain why a pipette is used to measure the alkali, rather than a measuring cylinder.

Answer

The pipette allows the same volume of alkali to be added each time, helping to make the results repeatable.

Titration calculations

The results of a *titration* can be used to calculate the *concentration* of a *solution*, or the volume of solution needed.

Calculating a concentration

Worked example

In a titration, 25.0 cm^3 of 0.100 mol/dm^3 sodium hydroxide solution is exactly *neutralised* by 20.00 cm^3 of a dilute solution of hydrochloric acid. Calculate the concentration of the hydrochloric acid solution.

Step 1: Calculate the amount of sodium hydroxide in moles

Volume of sodium hydroxide solution = 25.0 ÷ 1,000 = 0.0250 dm^3

Rearrange:

Concentration in mol/dm^3 = $\frac{amount\ of\ solute\ in\ mol}{volume\ in\ dm^3}$

Amount of *solute* in mol = concentration in mol/dm^3 × volume in dm^3

Amount of sodium hydroxide = 0.100 × 0.0250

= 0.00250 mol

Step 2: Find the amount of hydrochloric acid in moles

The *balanced equation* is: NaOH(aq) + HCl(aq) → NaCl(aq) + H_2O(l)

So, the *mole ratio* NaOH:HCl is 1:1

Therefore 0.00250 mol of NaOH reacts with 0.00250 mol of HCl

Step 3: Calculate the concentration of hydrochloric acid in mol/dm^3

Volume of hydrochloric acid = 20.00 ÷ 1000 = 0.0200 dm^3

Concentration in mol/dm^3 = $\frac{amount\ of\ solute\ in\ mol}{volume\ in\ dm^3}$

Concentration in mol/dm^3 = $\frac{0.00250}{0.0200}$

= 0.125 mol/dm^3

Step 4: Calculate the concentration of hydrochloric acid in g/dm^3

Relative formula mass of HCl = 1 + 35.5 = 36.5

Mass = relative formula mass × amount

Mass of HCl = 36.5 × 0.125

= 4.56 g

So, concentration = 4.56 g/dm^3

Question

In a titration, 25.00 cm^3 of 0.200 mol/dm^3 sodium hydroxide solution is exactly neutralised by 22.70 cm^3 of a dilute solution of hydrochloric acid.

NaOH(aq) + HCl(aq) → NaCl(aq) + H_2O(l)

Calculate the concentration of the hydrochloric acid.

Answer

Volume of sodium hydroxide solution = 25.00 ÷ 1000 = 0.0250 dm^3

Amount of sodium hydroxide = 0.200 × 0.0250 = 0.005 mol

- From the equation, 0.005 mol of NaOH reacts with 0.005 mol of HCl

Volume of hydrochloric acid = 22.70 ÷ 1000 = 0.0227 dm^3

Concentration of hydrochloric acid = 0.005 mol ÷ 0.0227

= **0.220 mol/dm^3**

Calculating a volume

Worked example

25.00 cm^3 of 0.300 mol/dm^3 sodium hydroxide solution is exactly neutralised by 0.100 mol/dm^3 sulfuric acid. Calculate the volume of sulfuric acid needed.

Step 1: Calculate the amount of sodium hydroxide in moles

Volume of sodium hydroxide solution = 25.0 ÷ 1000 = 0.0250 dm^3

Amount of sodium hydroxide = concentration × volume

Amount of sodium hydroxide = 0.300 mol/dm^3 × 0.0250 dm^3

= 0.00750 mol

Step 2: Find the amount of sulfuric acid in moles

The balanced equation is:

$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

So the mole ratio NaOH:H_2SO_4 is 2:1.

Therefore 0.00750 mol of NaOH reacts with (0.00750 ÷ 2) = 0.00375 mol of H_2SO_4

Step 3: Calculate the volume of sulfuric acid

Rearrange:

Concentration in mol/dm^3 = $\frac{amount\ of\ solute\ in\ mol}{volume\ in\ dm^3}$

Volume in dm^3 = $\frac{amount\ of\ solute\ in\ mol}{concentration\ in\ \ mol/\ dm^3}$

Volume in dm^3 = $\frac{0.00375}{0.100}$

= 0.0375 dm^3 (37.5 cm^3)

Question

25.00 cm^3 of 0.100 mol/dm^3 sodium hydroxide solution is exactly neutralised by 0.125 mol/dm^3 hydrochloric acid.

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Calculate the volume of hydrochloric acid needed.

Answer

Volume of sodium hydroxide solution = 25.00 ÷ 1000 = 0.0250 dm^3

Amount of sodium hydroxide = 0.100 × 0.0250 = 0.00250 mol

From the equation, 0.00250 mol of NaOH reacts with 0.00250 mol of HCl

Volume of hydrochloric acid = 0.00250 ÷ 0.125

= 0.020 dm^3 (20 cm^3)

END