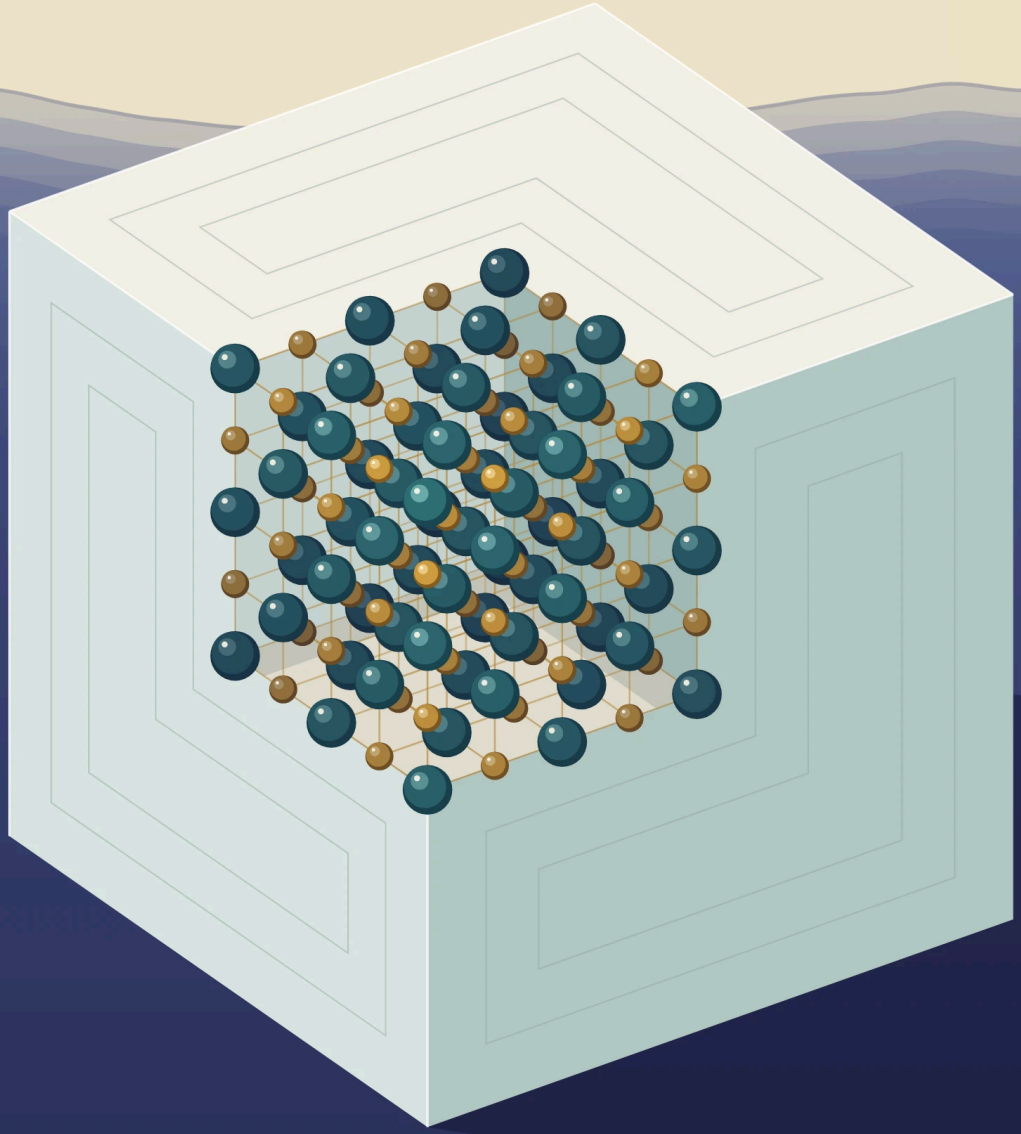
LEARNERS' ALLEY
Designed by Khow Willard

Sec 3

IP Chemistry

Book 1

For the 6092 syllabus



NAME _____

Sec 3 End-of-Year Examination

IP Chemistry, Secondary 3. There are two papers and no practical paper. The Sec 4 O Level has three papers (see the Sec 4 book).



Paper 1
Multiple Choice

DURATION	QUESTIONS	WEIGHTING	TIME PER QUESTION
45 min	30	39.6%	1.5 min

- 30 compulsory multiple choice questions.

Paper 2
Structured Questions

DURATION	MARKS	WEIGHTING	TIME PER MARK
1 h 15 min	55	60.4%	1.4 min

- **Section A** (45 marks): compulsory structured questions.
- **Section B** (10 marks): two questions, answer **one**.

No Paper 3
Practical

- There is no practical paper in the Sec 3 end-of-year examination.
- The Sec 4 O Level has a practical paper (Paper 3). It is in the Sec 4 book.

How to use these notes

Fill in the blanks

Write the missing key word on each line, e.g. electrons are added to the _____. The answers are the phrases examiners look for.

Labels

NOTE marks a special case, an exception or a point that is easy to miss. EXAM TIP marks advice on how to phrase an answer or set out working. ANSWER STRUCTURE shows how to set out the answer to a type of question you will meet again and again.

Explanations and examples

Numbered points in an explanation are separate mark points: write every one. Each WORKED EXAMPLE sets out the steps one at a time.

Sec 3 Book 1 Contents

1 — **Experimental Chemistry** — 2
Measuring quantities, collecting and drying gases, separating mixtures and testing purity.
1 Measuring Physical Quantities 2 Collecting and Drying Gases 3 Separation and Purification 4 Purity

2 — **Kinetic Particle Theory** — 8
Solids, liquids and gases, changes of state, heating curves and diffusion.
1 Solids, Liquids and Gases 2 How Particles Move

3 — **Atomic Structure** — 12
Sub-atomic particles, isotopes, electrons in shells, electronic configuration.
1 Sub-atomic Particles 2 Isotopes 3 Electrons in Shells 4 Summary

4–5 — **Chemical Bonding and Structure and Properties of Materials** — 16
Ionic, covalent and metallic bonding, then how structure decides the properties of materials.
1 Elements, Compounds and Mixtures 2 Why Do Atoms Combine? 3 Ionic Bonding 4 Covalent Bonding
5 Giant Covalent Structures 6 Metallic Bonding and Alloys 7 Structure and Properties

6–7 — **Chemical Formulae and Equations, and the Mole Concept and Stoichiometry** — 26
Writing formulae and balanced equations, then moles, concentration, reacting masses and yield.
06 · 1 Chemical Formulae 06 · 2 Constructing Formulae 06 · 3 Chemical Equations 07 · 1 Relative Masses
07 · 2 The Mole 07 · 3 Empirical and Molecular Formulae 07 · 4 Stoichiometry 07 · 5 Percentage Yield and Purity

01 Experimental Chemistry



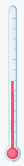
1 Measuring Physical Quantities

1.1 Quantities, units and apparatus

Quantity	SI unit	Apparatus	Other units
time	second, s	_____	minute (min), hour (h)
temperature	_____	thermometer	degree Celsius (°C)
length	metre, m	metre rule, measuring tape	cm, dm, mm
mass	kilogram, kg	_____	gram (g), tonne (t)
volume	cubic metre, m^3	burette, pipette, measuring cylinder, gas syringe	cm^3, dm^3



stopwatch


thermometer



electronic balance

- A **digital** stopwatch reads to ±0.01 s and an analogue stopwatch to ±0.1 s.
- Laboratory thermometers are alcohol, mercury or digital. A temperature probe joined to a **data logger** collects readings automatically.
- An electronic balance reads to about ±0.01 g.

1.2 Converting units

Time
1 min = 60 s 1 h = 60 min = 3600 s

Temperature
temperature in K = temperature in °C + _____

Length
1 m = 10 dm = 100 cm = 1000 mm

Mass
1 kg = 1000 g 1 g = 1000 mg

Volume
1 dm^3 = 1000 cm^3 = 1 litre

Volume
1 m^3 = _____ dm^3

NOTE 1 m = 10 dm, but 1 m^3 is not 10 dm^3. A cube of side 10 dm has a volume of 10 × 10 × 10 = 1000 dm^3.

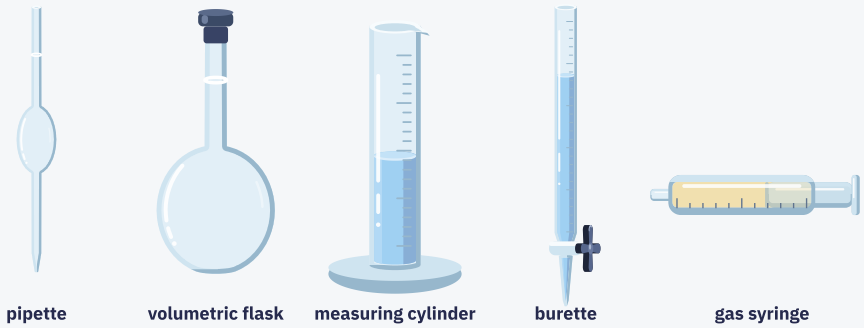
WORKED EXAMPLE Convert 250 cm^3 to dm^3.

1 dm^3 = 1000 cm^3

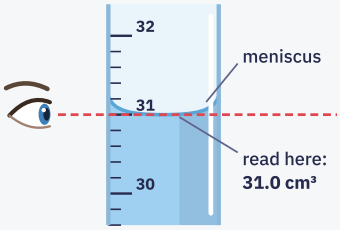
250 ÷ 1000 = **0.250 dm^3**

1 Measuring Physical Quantities (continued)

1.3 Measuring volume



Apparatus	What it measures
pipette	an accurate _____ volume, e.g. 10.0 cm^3 or 25.0 cm^3
volumetric flask	an accurate fixed larger volume, e.g. 100 cm^3 or 250 cm^3
measuring cylinder	a range of volumes, to the nearest 0.5 cm^3, e.g. 31.5 cm^3
burette	a range of volumes, to the nearest _____ cm^3, e.g. 23.00 cm^3
gas syringe	the volume of a _____, up to 100 cm^3

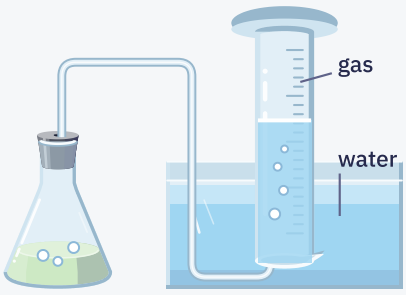


- Place your eye level with the **meniscus** to avoid parallax error.
- For water, read the _____ of the meniscus.

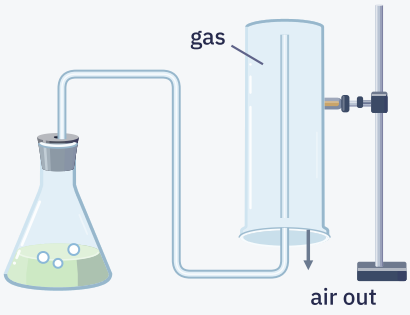
2 Collecting and Drying Gases

2.1 Methods of collecting gases

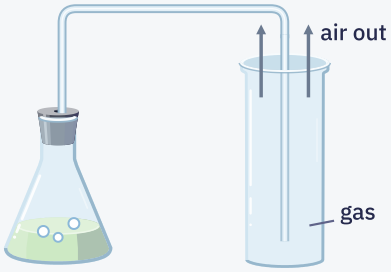
The method depends on how _____ the gas is in water and how _____ it is compared with air.



Water displacement
insoluble or slightly soluble gases



Upward delivery
gases less dense than air



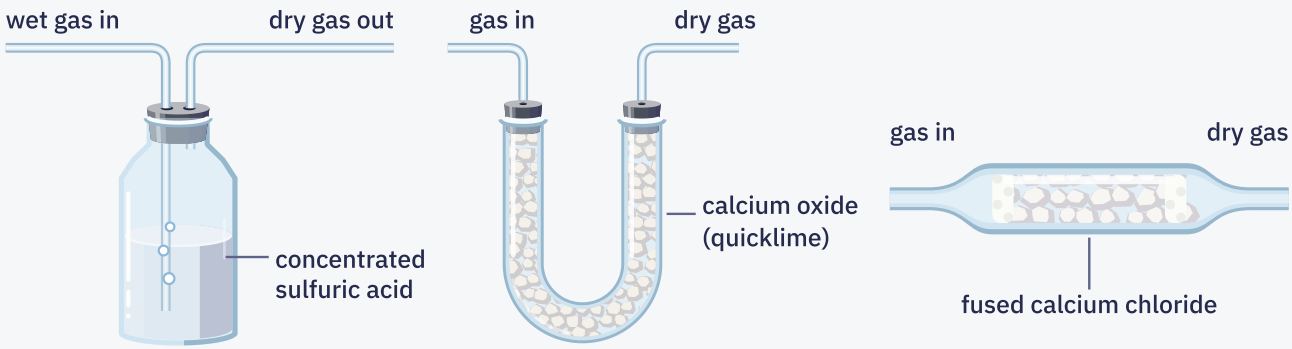
Downward delivery
gases denser than air

Method	Solubility in water	Density	Examples
water displacement	insoluble or slightly soluble	any	hydrogen, oxygen, carbon dioxide
upward delivery	soluble or insoluble	_____ than air	ammonia
downward delivery	soluble or insoluble	_____ than air	chlorine, hydrogen chloride, sulfur dioxide

NOTE To decide if a gas is denser than air, compare its M_r with the average M_r of air, about 29. A gas with M_r above 29 is **denser** than air (e.g. Cl_2, 71). A gas with M_r below 29 is **less dense** (e.g. NH_3, 17).

2 Collecting and Drying Gases (continued)

2.2 Drying agents for gases



Drying agent	Use it for	Do not use it for
concentrated sulfuric acid	most gases, e.g. chlorine and hydrogen chloride	gases that react with it, e.g. _____
calcium oxide (quicklime)	_____	gases that react with it, e.g. carbon dioxide
fused calcium chloride	most gases	ammonia

NOTE Calcium oxide and calcium chloride absorb water from the air, so they must be freshly heated before use.

3 Separation and Purification

3.1 Choosing a method

A mixture has two or more substances that are _____ chemically combined. Each is separated using a difference in a physical property: particle size, solubility, melting or boiling point, or magnetism.

What is in the mixture?

solid + solid

- one solid is magnetic → **magnet**
- different particle sizes → **sieving**
- only one solid dissolves → **dissolve, filter, evaporate**
- one solid sublimes → **sublimation**

solid + liquid

- solid does not dissolve → **filtration**
- solid dissolved, solid wanted → **evaporation to dryness**
- crystals wanted, or solid decomposes → **crystallisation**
- the liquid is wanted → **simple distillation**

liquid + liquid

- liquids do not mix (immiscible) → **separating funnel**
- miscible, boiling points differ → **fractional distillation**
- dissolved coloured substances → **paper chromatography**

NOTE **Key terms.** A **solute** dissolves in a **solvent** to make a solution. **Miscible** liquids mix completely, while **immiscible** liquids form separate layers. The **filtrate**, **residue**, **distillate** and **sublimate** are what each method collects.

3 Separation and Purification (continued)

3.2 Solid and solid mixtures

Magnetic attraction

A magnet removes _____ solids (iron, cobalt, nickel), e.g. iron filings from sulfur.

Sieving

Separates solids with different _____.

A suitable solvent

If only one solid dissolves, add the solvent and filter. For sand and salt, the salt solution is the filtrate and the sand is the residue.

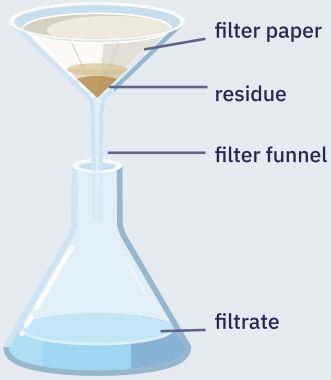
Sublimation

A solid that changes straight to a gas, e.g. _____ or naphthalene, is warmed and collected as the **sublimate** on a cold surface.



3.3 Solid and liquid mixtures

Filtration



Separates an _____ solid from a liquid. The liquid that passes through is the **filtrate**. The solid left on the paper is the **residue**.

Evaporation to dryness



Separates a _____ solid from its solvent by heating until all the solvent has boiled off.

Crystallisation

Used for a solid that _____ on strong heating, or to obtain pure crystals of a salt.

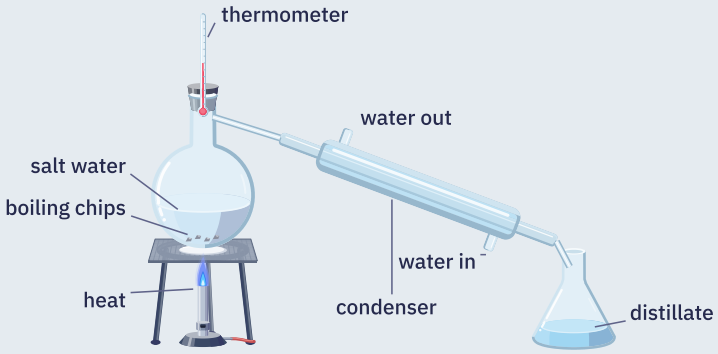
1. Heat the solution until it is _____. Stop before all the solvent evaporates.
2. Cool the solution so that crystals form.
3. Filter, wash the crystals with a little cold distilled water, and dry them between filter papers.

Why crystals form: hot water holds more solute than cold. 100 g of water holds 55 g of copper(II) sulfate at 80 °C but only 21 g at 20 °C, so on cooling the other 34 g comes out as crystals.

3 Separation and Purification (continued)

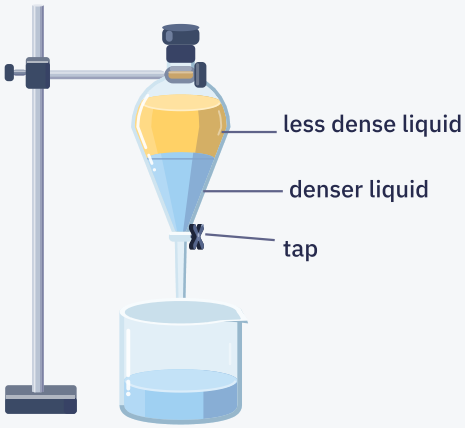
3.3 Solid and liquid mixtures (continued)

Simple distillation



- Separates a pure _____ from a solution, e.g. pure water from salt water. The solvent is collected as the **distillate**.
- The thermometer bulb is level with the **side arm**, to measure the temperature of the vapour entering the condenser.
- Cold water enters the condenser at the _____, so the condenser stays full of water and cools the vapour efficiently.
- Boiling chips give smooth boiling.

3.4 Liquid and liquid mixtures



Separating funnel

Separates _____ liquids. The less dense liquid forms the upper layer.

1. Run off the lower layer into a beaker.
2. Close the tap just as the boundary reaches it.
3. Run the upper layer into a second beaker.

WORKED EXAMPLE A mixture contains W and X, two immiscible liquids with X denser, Y, a solid that dissolves in W but not in X, and Z, a solid that is insoluble in both liquids. Outline how to separate all four.

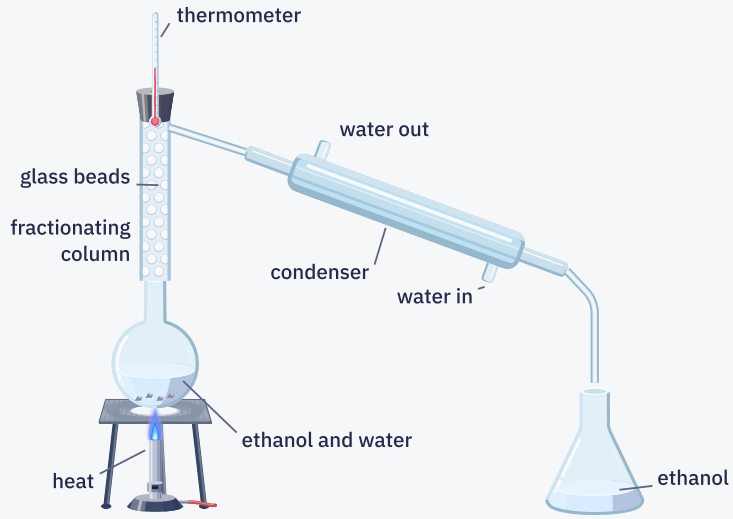
Remove the insoluble solid Z by **filtration**.

Use a **separating funnel** to run off liquid X, the lower layer.

Obtain solid Y from its solution in W by **evaporation** or **crystallisation**.

3 Separation and Purification (continued)

3.4 Liquid and liquid mixtures (continued)



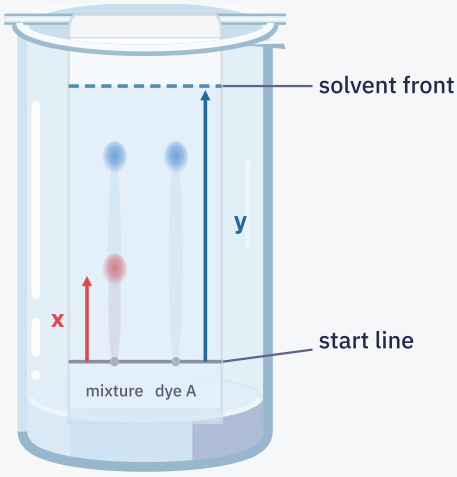
Fractional distillation

- Separates _____ liquids with different boiling points, e.g. ethanol (78 °C) and water (100 °C).
- The **glass beads** give a large surface area for the vapour to condense and boil again.
- The liquid with the _____ boiling point distils first. When the temperature rises, change the receiver.

NOTE Fractional distillation separates crude oil into fractions and liquid air into nitrogen, oxygen and argon.

3.5 Chromatography

Chromatography separates a mixture of substances that have _____ in a given solvent.



- Draw the start line in **pencil**, as graphite does not dissolve in the solvent.
- The solvent level must be _____ the start line, so the spots do not dissolve into the solvent.
- The more soluble a substance, the further it travels. A spot that matches a known pure substance is probably that substance.
- Colourless substances are shown with a **locating agent** or _____ light.

$$R_f = \frac{\text{distance moved by the } \square}{\text{distance moved by the } \square}$$ both from the **start line**

4 Purity

4.1 Pure substances and mixtures

Pure substance

Melts and boils at one _____ temperature. Gives **one spot** on a chromatogram.

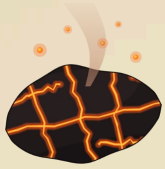
Mixture

Melts and boils over a _____ of temperatures. Impurities lower the melting point and raise the boiling point.

Purity matters in medicines and food: impurities in a medicine may cause harmful side effects.

EXAM TIP To show a sample is pure, state that it has a sharp, fixed melting point that matches the known value.

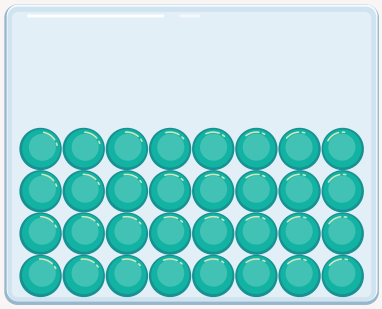
02 Kinetic Particle Theory



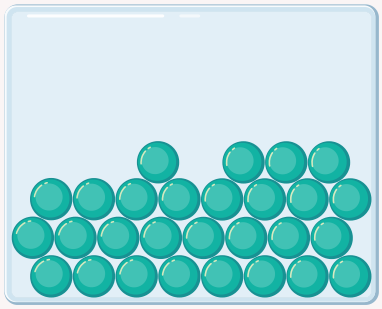
1 Solids, Liquids and Gases

1.1 The kinetic particle theory

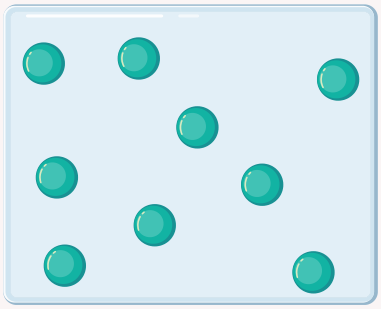
The **kinetic particle theory** states that all matter is made up of tiny _____ that are in constant, _____ motion.



solid



liquid



gas

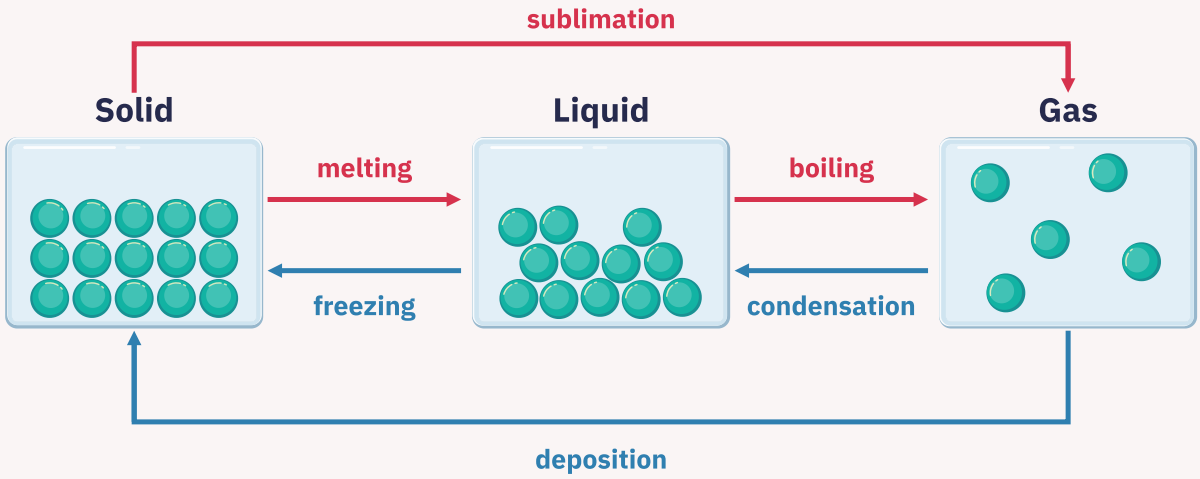
- The particles may be atoms, molecules or ions.

1.2 Comparing the three states

	Solid	Liquid	Gas
Arrangement	very closely packed, in an _____ way	closely packed, in a disorderly way	very far apart, in a _____ way
Movement	vibrate and rotate about _____ positions	slide past one another	move quickly and randomly in any direction
Kinetic energy	very low	low	high
Forces of attraction	very strong	strong	very weak
Shape and volume	definite shape and volume	no definite shape, definite volume	no definite shape or volume
Compressed?	no	no	_____

1 Solids, Liquids and Gases (continued)

1.3 Changes of state



Energy absorbed

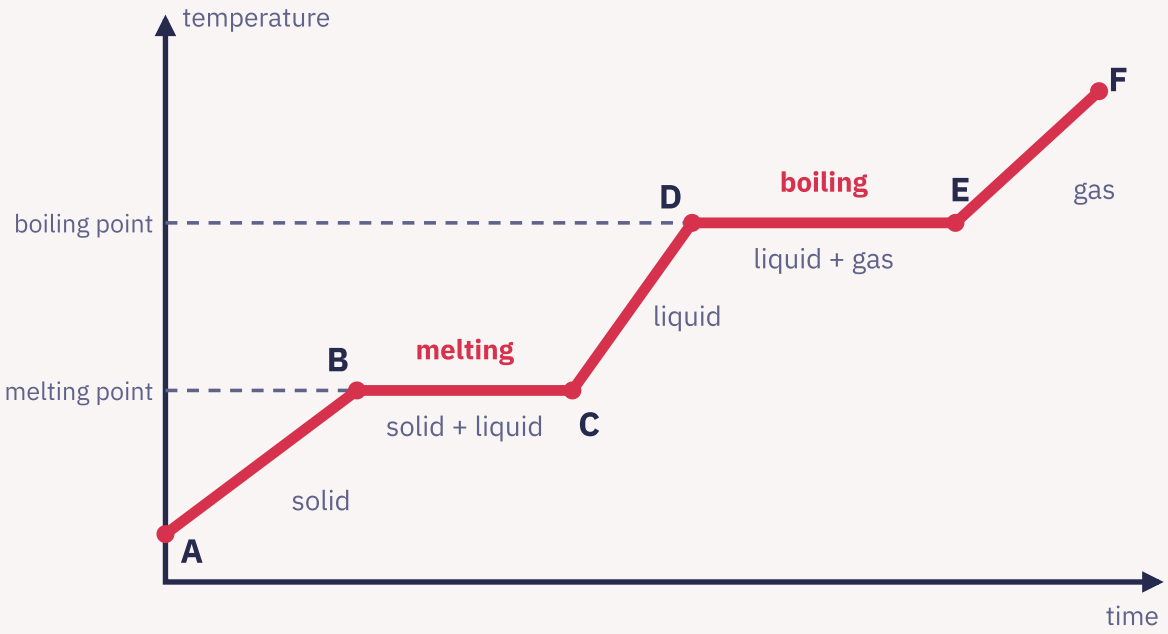
The particles gain energy: _____, boiling (and evaporation) and sublimation.

Energy released

The particles lose energy: freezing, _____ and deposition.

- **Sublimation** is a solid changing straight to a gas, e.g. dry ice (solid carbon dioxide). **Deposition** is a gas changing straight to a solid.

1.4 Heating curve

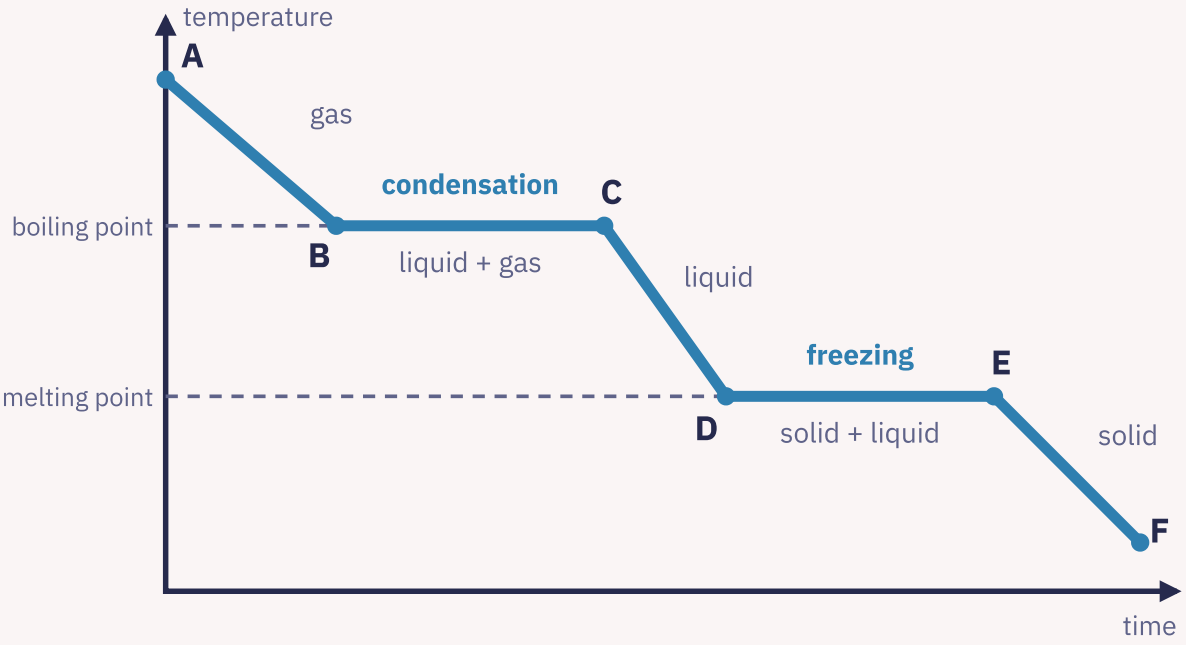


1. **A to B** (solid) and **C to D** (liquid) and **E to F** (gas): thermal energy is converted to _____, so the particles move faster and the temperature rises.
2. **B to C** (melting) and **D to E** (boiling): the energy absorbed overcomes the forces of attraction and increases the _____ of the particles as they move apart. Their kinetic energy does not change, so the temperature stays constant.

NOTE In each flat section of a heating or cooling curve, **two states** are present together.

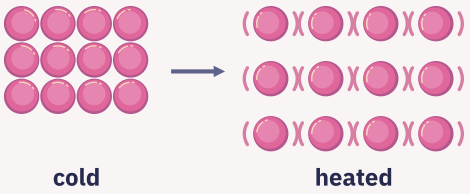
1 Solids, Liquids and Gases (continued)

1.5 Cooling curve



1. **A to B** (gas) and **C to D** (liquid) and **E to F** (solid): kinetic energy is converted to thermal energy and lost to the _____, so the particles slow down and the temperature falls.
2. **B to C** (condensation) and **D to E** (freezing): the forces of attraction pull the particles closer together, so their _____ falls and energy is released. This keeps the temperature constant.

1.6 Expansion and contraction



Expansion

Heated particles vibrate more and sit slightly further _____, so a solid expands.

Contraction

Cooled particles lose energy, vibrate less and come _____.

NOTE Water is unusual: its particles are more closely packed in liquid water than in ice, so ice is less dense than water and floats.

1.7 Evaporation and boiling

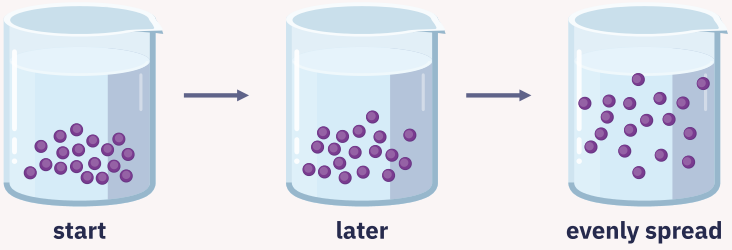
	Evaporation	Boiling
Temperature	any temperature	only at the _____
Where	at the surface only	throughout the liquid, with bubbles
Rate	slow	fast

NOTE In evaporation, only the particles at the surface with enough energy escape into the air.

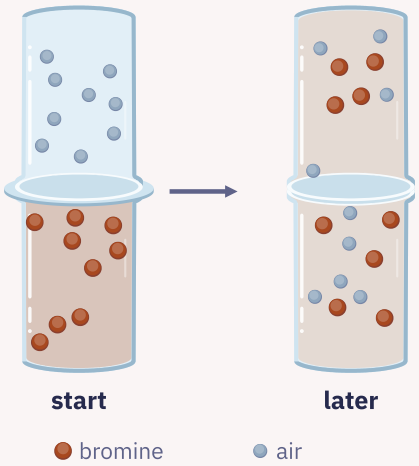
2 How Particles Move

2.1 Diffusion

Diffusion is the net movement of particles from a _____ concentration to a _____ concentration.



Ink spreading through water



Bromine spreading into air

- Diffusion happens in liquids and gases. It explains everyday effects such as a smell spreading through a room.
- The particles keep moving once evenly spread, but there is no net change in _____.

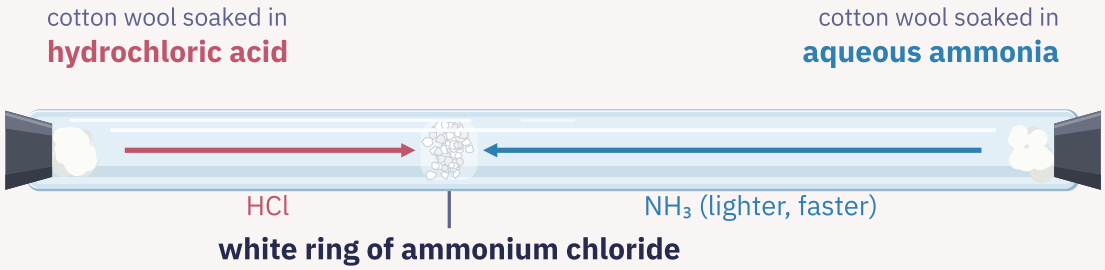
2.2 What affects the rate of diffusion

Temperature

A higher temperature gives particles more kinetic energy, so the rate of diffusion _____.

Particle mass

Heavier particles move more slowly, so the rate of diffusion _____.



- Ammonia and hydrogen chloride diffuse towards each other and react to form white ammonium chloride.
- The white ring forms nearer the _____ end, because the lighter ammonia particles diffuse faster and travel further.
- Porous pot: if the gas outside is lighter, it diffuses into the pot faster than the gas inside diffuses out, so the pressure inside **rises**.

WORKED EXAMPLE You can smell cooking food from across a room. Explain this using the kinetic particle theory.

Heating gives some particles at the surface enough kinetic energy to **escape** into the air as a vapour.
The particles are in constant random motion, so they **diffuse** through the air from the food to your nose.

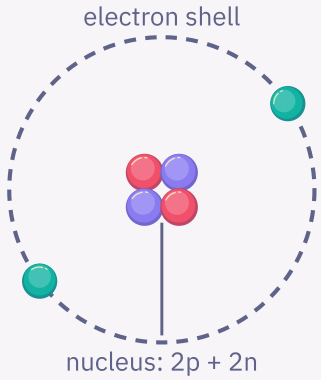
03 Atomic Structure

1 Sub-atomic Particles

1.1 What an atom is made of

An **atom** is the smallest particle of an element that still has the _____ of that element.

Particle	Relative charge	Relative mass	Location
p proton	+1	1	_____
n neutron	_____	1	nucleus
e electron	−1	_____	_____



- Protons and neutrons are together called **nucleons**. They are in the **nucleus** at the centre of the atom.
- Atoms are electrically neutral, so the number of _____ = the number of _____.

1.2 Proton number and nucleon number

$^{12}_{6}C$

12 nucleon (mass) number
6 proton (atomic) number
read as **carbon-12**

DEFINITION **Proton number**: the number of _____ in the nucleus of an atom. It tells us which _____ the atom is from.

DEFINITION **Nucleon number**: the total number of _____ in the nucleus of an atom.

number of neutrons = nucleon number − proton number

WORKED EXAMPLE Tungsten is used to make drill bits. Tungsten-184 is written $^{184}_{74}W$. Deduce the numbers of protons, neutrons and electrons in one atom of tungsten-184.

Protons = proton number = **74**

Neutrons = nucleon number − proton number = 184 − 74 = **110**

Electrons = protons = **74**

1 Sub-atomic Particles (continued)

1.3 Sub-atomic particles in ions

An **ion** is a particle formed when an atom or group of atoms _____. The numbers of protons and neutrons stay the same.

Particle	Protons	Neutrons	Electrons
calcium atom, Ca	**20**	**20**	**20**
calcium ion, Ca^{2+}	**20**	**20**	_____
oxygen atom, O	**8**	**8**	**8**
oxide ion, O^{2-}	**8**	**8**	_____

Cation (positive ion)

The atom _____ electrons, so it has fewer electrons than protons, e.g. Ca^{2+}.

Anion (negative ion)

The atom _____ electrons, so it has more electrons than protons, e.g. O^{2-}.

2 Isotopes

2.1 Isotopes of an element

DEFINITION **Isotopes** are atoms of the same element with the same _____ but different _____. They have different numbers of _____.



	hydrogen-1	hydrogen-2	hydrogen-3
Notation	$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
Protons	**1**	**1**	**1**
Neutrons	**0**	_____	_____
Electrons	**1**	**1**	**1**
Nucleon number	**1**	**2**	**3**

Chemical properties

Isotopes have the same number of _____, so they react in the _____ way.

Physical properties

Isotopes have different masses, so they may have different _____, melting points and boiling points.

NOTE Isotopes are named with their nucleon number, e.g. chlorine-35 and chlorine-37.

3 Electrons in Shells

3.1 Electron shells

- Electrons in the shell closest to the nucleus have the **lowest energy**. Electrons fill the lowest energy shells first.
- The first shell holds up to _____ electrons. The second and third shells usually hold up to _____ electrons each.
- The outermost shell is the **valence shell**, and its electrons are **valence electrons**.
- A full valence shell is stable, so atoms of helium and argon are _____. An atom with a valence shell that is not full is reactive.

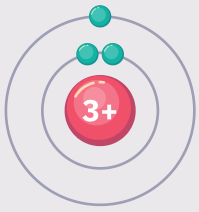
3.2 Electronic configuration

The **electronic configuration** of an atom shows the number of _____, starting from the innermost shell, e.g. sodium is 2.8.1.

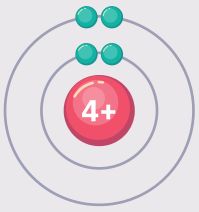
From the periodic table

- The **period** number is the number of _____
- The **group** number gives the number of valence electrons: Groups 1 and 2 have 1 and 2, and Groups 13 to 18 have 3 to 8. Helium has 2.

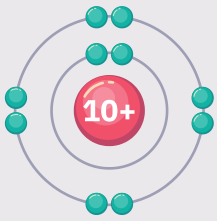
3.3 Shell diagrams



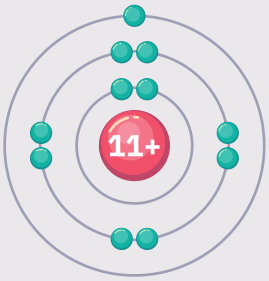
Li 2.1



Be 2.2



Ne 2.8



Na 2.8.1

3 Electrons in Shells (continued)

3.4 Configurations of the first 20 elements

Period	Group 1	2	13	14	15	16	17	18
1	**H** 1							**He** 2
2	**Li** 2.1	**Be** 2.2	**B** 2.3	**C** 2.4	**N** 2.5	**O** 2.6	**F** 2.7	**Ne** 2.8
3	**Na**	**Mg**	**Al**	**Si**	**P**	**S**	**Cl**	**Ar**
4	**K**	**Ca**						

WORKED EXAMPLE How are the electrons arranged in an atom of potassium (proton number 19)?

Potassium has 19 protons, so it has **19** electrons.

Fill the first shell with 2, then the second with 8: 2.8. That is 10 electrons.

The remaining 9 electrons: the third shell takes 8, so 1 is left for the fourth shell.

Electronic configuration = **2.8.8.1**

NOTE Electronic configurations are needed only up to calcium (proton number 20).

EXAM TIP Isotopes of an element have the **same** number of electrons, so they have the **same** electronic configuration.

4 Summary

4.1 Atomic structure at a glance

Nucleus

- Contains _____ and _____, together called nucleons.
- Protons: relative charge _____, relative mass 1. The proton number identifies the element.
- Neutrons: relative charge _____, relative mass 1.

Electron shells

- Electrons: relative charge _____, relative mass _____.
- Up to _____ electrons in the first shell and usually _____ in the second and third.
- Electronic configuration, e.g. oxygen is _____.

Isotopes

Atoms of the same element with the same _____ and different _____. They have the same chemical properties and different physical properties.

4.2 Finding the numbers

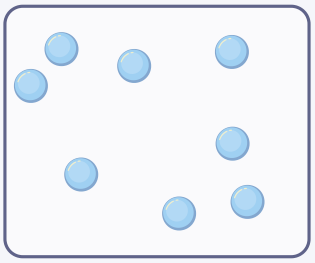
To find	Use
Protons	the _____
Neutrons	_____
Electrons in an atom	the same number as the _____
Electrons in an ion	protons – charge, e.g. Ca^{2+} has 20 – 2 = _____
Electron shells	the _____ number
Valence electrons	the _____ number (Groups 1 and 2), or 3 to 8 for Groups 13 to 18

04–05 Chemical Bonding and Materials

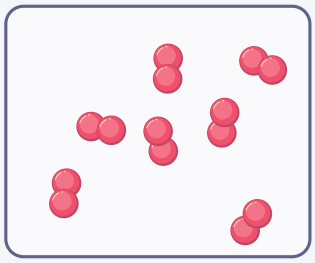
1 Elements, Compounds and Mixtures

1.1 Comparing them

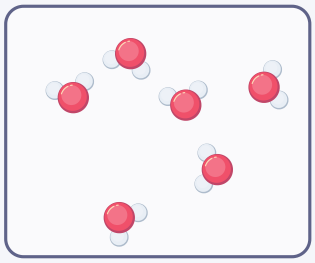
	Element	Compound	Mixture
Made of	one kind of atom	two or more elements _____	two or more substances _____ chemically combined
Ratio	not applicable	_____ ratio	no fixed ratio
Properties	its own	different from its elements	similar to its substances
Formed by	mostly occurs naturally	a **chemical reaction**	usually physical mixing
Melting and boiling	fixed points	fixed points	over a _____ of temperatures
Separated by	cannot be broken down by chemical methods	chemical methods only, e.g. **thermal decomposition** or **electrolysis**	_____ methods, e.g. filtration or distillation



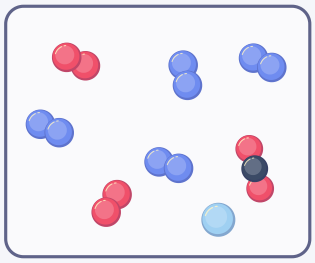
element: argon atoms



element: oxygen molecules



compound: water molecules



mixture: air

- An element can exist as single atoms (helium), molecules (O_2) or a metal lattice (copper).
- _____ all molecules are compounds: H_2 is an element and CO_2 is a compound.

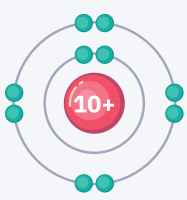
2 Why Do Atoms Combine?

2.1 The noble gases are stable

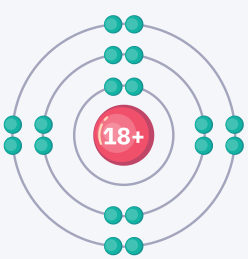
The **noble gases** (Group 18) have a _____ outermost shell, so they are stable and unreactive. This is a **noble gas electronic configuration**.



He 2



Ne 2.8



Ar 2.8.8

- Helium has a **duplet**, 2 electrons in its only shell. Neon and argon have an **octet**: 8 outer electrons.
- Noble gases are **monoatomic**. They exist as single atoms.

2 Why Do Atoms Combine? (continued)

2.2 Three ways to a full outer shell

Other atoms combine and form **bonds** to reach a noble gas electronic configuration.

Lose electrons

Metals form _____.

Gain electrons

Non-metals form _____ _____.

Share electrons

Non-metals form _____ _____.

3 Ionic Bonding

3.1 Forming ions



Metal atoms lose electrons

$Na \rightarrow Na^+ + e^-$

$Mg \rightarrow Mg^{2+} + 2e^-$

Non-metal atoms gain electrons

$Cl + e^- \rightarrow Cl^-$

$O + 2e^- \rightarrow O^{2-}$

- Na (2.8.1) loses 1 electron to become Na^+ (2.8). Cl (2.8.7) gains 1 electron to become Cl^- (2.8.8). Both ions have a _____ electronic configuration.
- Na could lose 1 electron or gain 7, and losing 1 is far more likely. Na^+ then has 11 protons and 10 electrons, so its charge is +1.
- A negative ion of a single element ends in **-ide**: chloride, oxide. Metal ions keep the name of the metal.

3 Ionic Bonding (continued)

3.2 Common ions

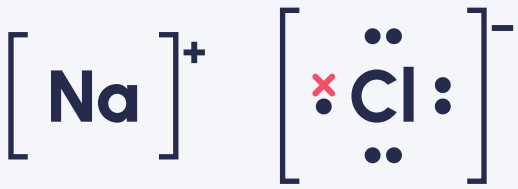
Charge	Positive ions (cations)	Negative ions (anions)
1	hydrogen H^+, sodium Na^+, potassium K^+, silver Ag^+, **ammonium NH_4^+ ★**	fluoride F^-, chloride Cl^-, bromide Br^-, iodide I^-, **hydroxide OH^- ★**, **nitrate NO_3^- ★**
2	magnesium Mg^{2+}, calcium Ca^{2+}, barium Ba^{2+}, iron(II) Fe^{2+}, copper(II) Cu^{2+}, zinc Zn^{2+}	oxide O^{2-}, **carbonate CO_3^{2-} ★**, **sulfate SO_4^{2-} ★**
3	aluminium Al^{3+}, iron(III) Fe^{3+}	phosphate PO_4^{3-}

NOTE Ions with more than one atom, such as NH_4^+ and SO_4^{2-}, are **polyatomic ions**. Learn the names, formulae and charges of the five marked **like this ★** by heart.

- The negative ions of Group 17 are **halide ions**. Ions with a charge above +3 are rare.
- Hydrogen is a non-metal that forms a positive ion. H^+ has no electrons: it is just a proton.

3.3 The ionic bond

An **ionic bond** is the mutual _____ between ions of opposite charges. It usually forms between a **metal** and a **non-metal**.



sodium chloride, NaCl

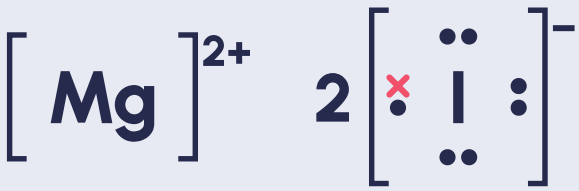
- Show only the **outer** electrons, as dots for one atom and crosses for the other.
- Put each ion in **square** brackets, with its charge outside.
- An ionic compound is neutral: total positive charge = total _____ charge.

WORKED EXAMPLE Magnesium reacts with iodine. Give the formula of the compound and draw its dot-and-cross diagram.

Magnesium forms Mg^{2+} (charge +2). Iodine forms I^- (charge −1).

To balance, one Mg^{2+} needs **two** I^-, as +2 + 2(−1) = 0.

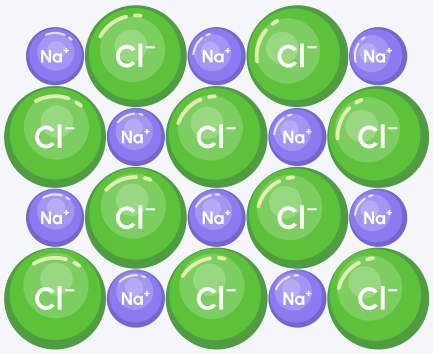
The formula is **MgI_2**. Show the number of each ion in front of its bracket.



EXAM TIP An ionic formula shows the simplest ratio of ions, with no charges.

3 Ionic Bonding (continued)

3.4 The giant ionic lattice

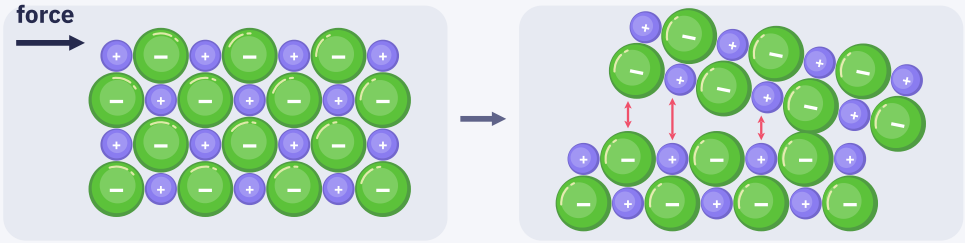


In the solid state, an ionic compound is a **giant ionic lattice**: a three-dimensional structure of _____ positive and negative ions in a regular, repeating pattern.

- Every ion attracts _____ the oppositely charged ions next to it. In sodium chloride, each Na^+ is surrounded by six Cl^-, and each Cl^- by six Na^+.

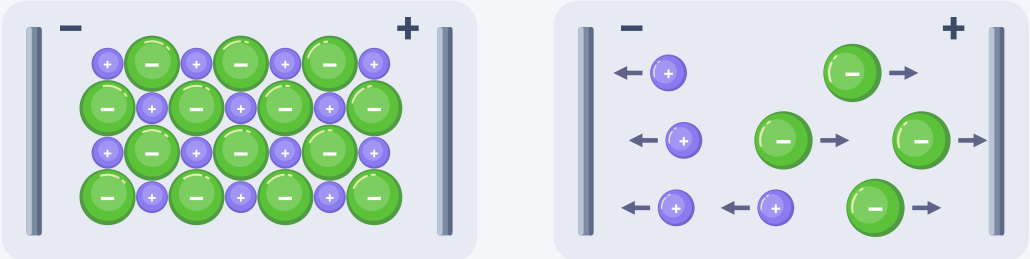
3.5 Properties of ionic compounds

Property	Explanation
High melting and boiling points, solid at room temperature	The electrostatic attraction between oppositely charged ions is strong, so a _____ is needed to overcome it.
Hard but **brittle**	A strong force moves the layers so that ions of the same charge meet. They _____ and the lattice shatters.
Usually **soluble in water**, insoluble in organic solvents	Water pulls the ions out of the lattice.
Conduct electricity when **molten or aqueous**, not as a solid	The ions are _____ only when molten or dissolved. In the solid they only vibrate about fixed positions.



before the force

like charges repel: the crystal splits



solid: ions in fixed positions

molten or aqueous: mobile ions

WORKED EXAMPLE Explain why sodium chloride has a high melting point.

Sodium chloride has a **giant ionic lattice** structure.

There are strong electrostatic forces of attraction between the oppositely charged Na^+ and Cl^- ions.

A **large amount of energy** is needed to overcome these forces, so the melting point is high.

4 Covalent Bonding

4.1 Sharing electrons

DEFINITION A **covalent bond** is formed when two atoms _____ a pair of electrons (a **bonding pair**), so each atom gains a noble gas electronic configuration. It forms between **non-metal** atoms.

Valency is the number of electrons an atom must lose, gain or share to reach a noble gas electronic configuration. For a non-metal, the bonds it forms = 8 − number of valence electrons (hydrogen: 2 − 1 = 1).

Element	Group	Valence electrons	Valency
hydrogen	1	1	1
carbon, silicon	14	4	4
nitrogen, phosphorus	15	5	3
oxygen, sulfur	16	6	2
fluorine, chlorine, bromine, iodine	17	7	1

4.2 Dot-and-cross diagrams of molecules



Cl_2: single bond, Cl–Cl



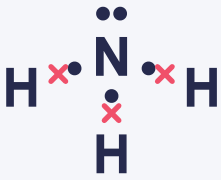
O_2: double bond, O=O



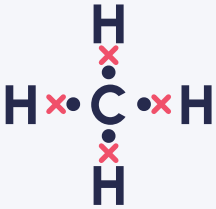
N_2: triple bond, N≡N



water, H_2O



ammonia, NH_3



methane, CH_4



carbon dioxide, CO_2

- Draw each shared pair **between** the two symbols, and the other outer electrons in pairs.
- Cl_2, O_2 and N_2 are **diatomic**: each molecule has two atoms. A **structural formula** shows each shared pair as a line, as in Cl–Cl, O=O and N≡N.

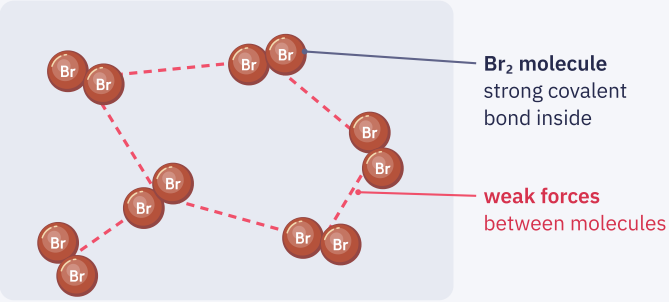
WORKED EXAMPLE Draw a dot-and-cross diagram for hydrogen sulfide, H_2S.

Hydrogen has valency 1. Sulfur has 6 outer electrons, so its valency is 2.
Sulfur forms **two** single bonds, one to each hydrogen atom.
Sulfur keeps its other 4 outer electrons as 2 lone pairs.



4 Covalent Bonding (continued)

4.3 Properties of simple covalent substances



- **Strong** covalent bonds _____ each molecule.
- **Weak** intermolecular forces of attraction _____ the molecules.
- Larger molecules have stronger forces between them, so higher melting and boiling points. At room temperature methane is a gas, bromine a liquid and iodine a solid.
- At low temperatures, some simple molecules form a lattice, as water does in ice.

Property	Explanation
Low melting and boiling points, often **volatile**	A _____ amount of energy overcomes the weak intermolecular forces.
Usually insoluble in water but soluble in organic solvents	They are neutral molecules.
Do not conduct electricity	No mobile ions or electrons. Hydrogen chloride conducts only after it dissolves in water and forms H^+ and Cl^-.

EXAM TIP When a simple covalent substance melts or boils, the weak intermolecular forces **between** molecules are overcome. The covalent bonds inside the molecules do not break.

5 Giant Covalent Structures

5.1 Simple and giant molecules

Simple molecules

A _____ number of atoms in a fixed ratio, e.g. H_2O, CO_2 and Br_2.

Giant molecules

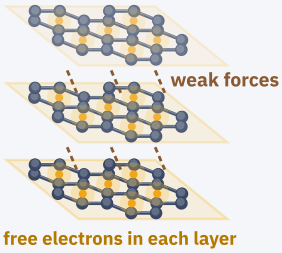
An _____ number of atoms all joined by covalent bonds, e.g. diamond and silicon dioxide.

5.2 Diamond, graphite and silicon dioxide

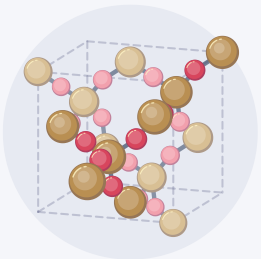
Allotropes are different forms of the same element with different arrangements of atoms. Diamond and graphite are allotropes of carbon.



diamond



graphite



silicon dioxide: Si larger, O smaller

	Diamond	Graphite	Silicon dioxide
Bonding	each C bonded to **4** C, tetrahedral	each C bonded to **3** C, in layers of hexagons	each Si bonded to 4 O, SiO_2, as in quartz and sand
Hardness	_____	soft and slippery	hard
Conducts?	no	_____	no

- Diamond coats drill bits and cutting tools. Graphite is used as a solid lubricant and in pencil lead.

5 Giant Covalent Structures (continued)

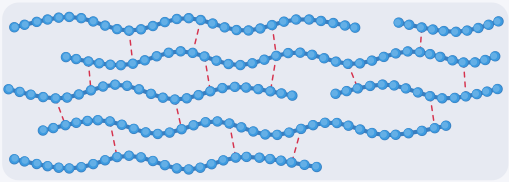
5.3 Properties of giant covalent substances

Property	Explanation
Very high melting and boiling points, solids	Many strong covalent bonds link all the atoms, so a _____ is needed to break them.
Graphite is **soft and slippery**	Only weak forces hold the layers together, so the layers _____ over each other.
Graphite **conducts** electricity	Each carbon uses 3 of its 4 outer electrons in bonds. The fourth is _____ and moves along the layers.
Diamond and silicon dioxide do not conduct	All outer electrons are used in bonds, so there are no mobile electrons.

- Giant covalent substances are insoluble in water and in organic solvents.

NOTE Graphene (a single layer of graphite), carbon nanotubes and fullerenes are other allotropes of carbon.

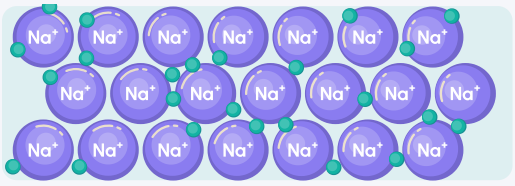
5.4 Macromolecules



- A **polymer** is many small molecules joined into long chains, e.g. poly(ethene), nylon, starch and silk.
- The chains have different lengths, so polymers have _____ _____ and soften gradually on heating.
- Most do not conduct electricity. Most are insoluble in water but soluble in organic solvents.

6 Metallic Bonding and Alloys

6.1 A sea of delocalised electrons



A **metallic bond** is the mutual electrostatic attraction between the _____ and the _____ _____.

- Metal atoms lose their outer electrons and form a **giant metallic lattice** of positive ions.
- The electrons no longer belong to any one atom. They are **delocalised** and move freely between the ions.
- Metal atoms form only positive ions, so ionic bonding between them is unlikely.

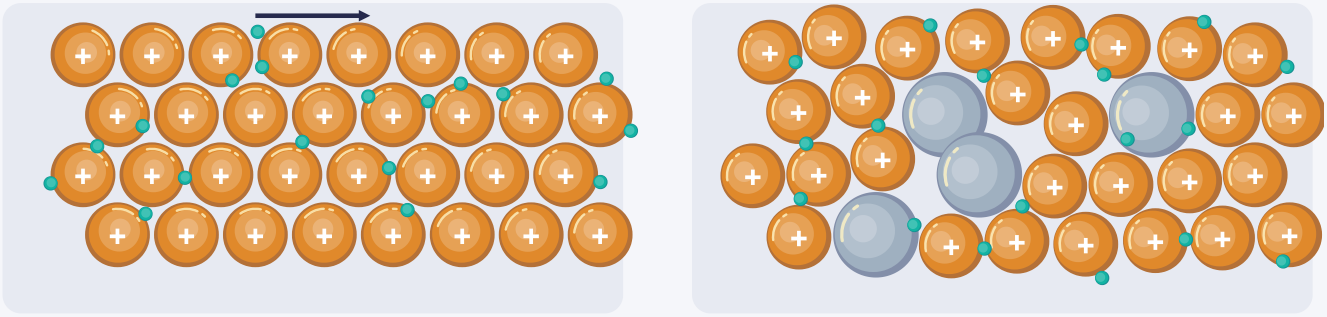
6 Metallic Bonding and Alloys (continued)

6.2 Properties of metals

Property	Explanation
High melting and boiling points	The strong attraction between the ions and the delocalised electrons needs a _____ to overcome.
Conduct **electricity** as a solid and when molten	The delocalised electrons are _____ and carry the current.
Conduct **heat**	The delocalised electrons pass on thermal energy through the lattice.
Malleable and ductile	The layers of identical ions _____ over one another, and the sea of electrons keeps holding them together.

6.3 Alloys

An **alloy** is a mixture of a metal with one or more other elements.



pure metal: layers slide **alloy: different-sized atoms stop sliding**

- Different-sized atoms _____ the orderly layers, so the layers cannot slide easily. Alloys are **harder and stronger** than the pure metal.
- An alloy is a mixture, so it melts over a _____ of temperatures. Its make-up is given as approximate percentages by mass, not as a formula.

WORKED EXAMPLE Brass is an alloy of copper and zinc. Explain why brass is harder than pure copper.

In pure copper, the layers of identical copper ions can **slide** over one another.

In brass, the zinc atoms are a **different size** and disrupt the orderly layers.

The layers cannot slide over one another easily, so brass is **harder**.

6 Metallic Bonding and Alloys (continued)

6.4 Uses of alloys

Alloy	Made from	Why it is used
bronze	copper and tin	harder and stronger, e.g. statues and bells
brass	copper and zinc	harder than copper
solder	tin and lead	lower melting point
steel	iron and carbon	harder than iron
stainless steel	iron, carbon, chromium, nickel	resists corrosion

7 Structure and Properties

7.1 The three types of bond

	Ionic	Covalent	Metallic
Between	metal and non-metal	non-metals	metal atoms
Electrons	_____	_____	_____
Held by attraction between	oppositely charged ions	a shared pair of electrons and both nuclei	positive ions and delocalised electrons

EXAM TIP To state the type of bonding, look at the elements: a metal and a non-metal are ionic, non-metals only are covalent, and a metal element is metallic.

7 Structure and Properties (continued)

7.2 Which type of structure is it?

	Ionic	Simple covalent	Giant covalent	Metallic
Particles	ions	molecules	atoms	ions and electrons
Melting point	high	_____	very high	high
Conducts	molten, aqueous	no	no (graphite yes)	solid and molten
Soluble in water	usually	usually not	no	no

1. Is the melting point low? Yes: a _____ substance. No: go on.

2. Does it conduct as a solid? Yes: a _____ or graphite. No: go on.

3. Does it conduct when molten or aqueous? Yes: an _____. No: a _____ substance.

NOTE A substance that melts over a range of temperatures is a mixture, such as an alloy, or a polymer.

WORKED EXAMPLE Solid A is an insulator but conducts when dissolved in water. Solid B is insoluble in water but conducts as a solid. What bonding do they have?

A conducts only when dissolved, so its ions are mobile in solution. A has **ionic** bonding.

B conducts as a solid, so it has mobile electrons. B has **metallic** bonding.

WORKED EXAMPLE Solids X, Y and Z are graphite, sodium chloride and silicon dioxide. X is soft and slippery and conducts electricity. Y and Z have very high melting points. Y conducts when dissolved in water but not as a solid. Z does not conduct. Identify X, Y and Z.

X conducts as a solid and is slippery, so X is **graphite**.

Y conducts only when dissolved, so it is ionic: Y is **sodium chloride**.

Z has a very high melting point but never conducts, so it is giant covalent: Z is **silicon dioxide**.

06 Chemical Formulae and Equations

1 Chemical Formulae

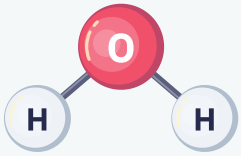
1.1 What a formula tells you

A **chemical formula** shows the elements in a substance and the _____ of their atoms.



carbon dioxide, CO_2

1 carbon atom, 2 oxygen atoms



water, H_2O

2 hydrogen atoms, 1 oxygen atom

- The _____ show which elements are present. The _____ show how many atoms of each.
- CO_2: one carbon atom and two oxygen atoms. H_2O: two hydrogen atoms and one oxygen atom.
- The prefix **di-** means _____. Carbon **di**oxide has two oxygen atoms.

1.2 Formulae of elements

Monoatomic

Single, _____ atoms. The Group 18 elements, such as neon (Ne) and argon (Ar), have _____ outer shells.

Diatomic

_____ atoms joined: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2.

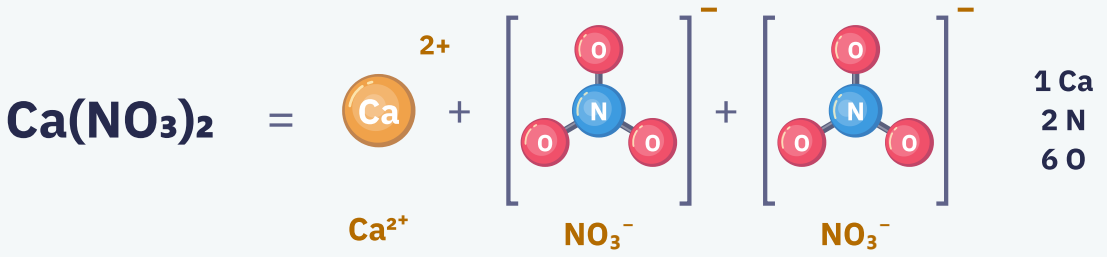
Polyatomic

Three or more atoms joined, such as phosphorus and ozone (O_3).

1 Chemical Formulae (continued)

1.3 Formulae of compounds

- A subscript outside **brackets** multiplies _____ inside them. $Ca(NO_3)_2$ has 1 calcium, _____ nitrogen and _____ oxygen atoms.



2 Constructing Formulae

2.1 Valency

Valency is the number of electrons an atom must lose, gain or share to get a noble gas configuration. The valency of an ion is its charge _____.

Group	1	2	13	14	15	16	17
Examples	Li, Na, K	Be, Mg, Ca	B, Al	C, Si	N, P	O, S	F, Cl, Br, I
Valency	1	2	_____	_____	_____	_____	1

NOTE Transition metals can form more than one ion. A _____ shows the valency: iron(II) is Fe^{2+}, iron(III) is Fe^{3+}, copper(II) is Cu^{2+}, copper(I) is Cu^{+}. Zinc (Zn^{2+}) and silver (Ag^{+}) have only one common ion.

2.2 Polyatomic ions

Name	Formula	Valency	Name	Formula	Valency
ammonium ★	NH_4^+	1	**sulfate** ★	SO_4^{2-}	2
hydroxide ★	OH^-	1	**carbonate** ★	CO_3^{2-}	2
nitrate ★	NO_3^-	1	phosphate	PO_4^{3-}	3

NOTE A polyatomic ion stays together as one unit. Put it in brackets when more than one is needed, e.g. $Mg(NO_3)_2$.

2 Constructing Formulae (continued)

2.3 The crossing method



1 write the valencies, then cross them over: C_2O_4

2 simplify 2 : 4 to 1 : 2

1. Write the symbols and their valencies below.
2. _____ the valencies: each valency becomes the _____ of the other element.
3. _____ the ratio if possible. C_2O_4 becomes CO_2.

2.4 Worked examples

WORKED EXAMPLE Write the formula of sodium phosphate.

Sodium has valency 1 (Na^+). Phosphate is a polyatomic ion with valency 3 (PO_4^{3-}).
Three Na^+ are needed to cancel one PO_4^{3-}.
The formula is **Na_3PO_4**.

WORKED EXAMPLE Write the formula of magnesium nitrate.

Magnesium has valency 2 (Mg^{2+}). Nitrate has valency 1 (NO_3^-).
Two NO_3^- are needed. Put the polyatomic ion in _____ before the subscript 2.
The formula is **$Mg(NO_3)_2$**.

2.5 Finding a charge from a formula

WORKED EXAMPLE Find the charge on the Al ion in $KAl(SO_4)_2$.

The compound is neutral, so the total charge is 0.
Let the charge on Al be x: 0 = (+1) + x + 2(−2).
So x = **+3**.

2 Constructing Formulae (continued)

2.6 Practice with formulae

Compound	Ions and valencies	Formula
calcium chloride	Ca^{2+} (2) and Cl^- (1)	**$CaCl_2$**
silver phosphate	Ag^+ (1) and PO_4^{3-} (3)	
iron(III) sulfate	Fe^{3+} (3) and SO_4^{2-} (2)	**$Fe_2(SO_4)_3$**
aluminium oxide	Al^{3+} (3) and O^{2-} (2)	

3 Chemical Equations

3.1 Word and chemical equations

sodium + chlorine → sodium chloride
$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

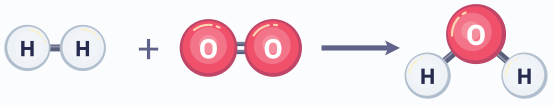


2Na Cl_2 2NaCl

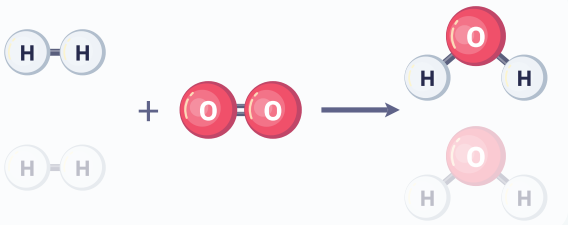
- _____ go on the left of the arrow and _____ on the right.
- The number in front of a formula is the number of units reacting or formed.
- **State symbols** show the physical state: (s) solid, (l) liquid, (g) gas, (aq) aqueous, which means _____.

NOTE The same substance can have a different state in different reactions. Sodium chloride is (s) when it forms from sodium and chlorine, but (aq) when it forms in a solution.

3.2 Balancing an equation



$H_2 + O_2 \rightarrow H_2O$: 2 O on the left, 1 O on the right



$2H_2 + O_2 \rightarrow 2H_2O$: 4 H and 2 O on each side

1. From the word equation (hydrogen + oxygen → water), write the correct formulae for reactants and products.
2. Count the atoms of each element on both sides.
3. Add numbers _____ formulae to balance. Never change a _____.
4. Add the state symbols: $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$.

3 Chemical Equations (continued)

3.3 Balancing with a solid and a gas

WORKED EXAMPLE Iron(III) oxide reacts with carbon monoxide to form iron and carbon dioxide. Write the balanced equation with state symbols.

Formulae: $Fe_2O_3 + CO \rightarrow Fe + CO_2$.

Two iron atoms on the left need **2Fe**. Three oxygen atoms in Fe_2O_3 are removed as CO_2, so use **3CO** and **3CO_2**.

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

3.4 A harder equation

WORKED EXAMPLE Aqueous iron(III) sulfate reacts with aqueous barium nitrate to form aqueous iron(III) nitrate and solid barium sulfate. Write the equation with state symbols.

Formulae: $Fe_2(SO_4)_3 + Ba(NO_3)_2 \rightarrow Fe(NO_3)_3 + BaSO_4$.

Balance the **ions**. Two Fe^{3+} need $2Fe(NO_3)_3$. Three SO_4^{2-} need $3BaSO_4$, so $3Ba(NO_3)_2$.

$Fe_2(SO_4)_3(aq) + 3Ba(NO_3)_2(aq) \rightarrow 2Fe(NO_3)_3(aq) + 3BaSO_4(s)$

3.5 Ionic equations

1. Write the balanced equation with state symbols.
   $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
2. _____ the aqueous substances into ions.
   $Na^+ + OH^- + H^+ + Cl^- \rightarrow Na^+ + Cl^- + H_2O(l)$
3. **Cancel** the _____, which appear _____ on both sides.
4. What remains is the ionic equation: **$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$**



spectator ions Na^+ and Cl^- are unchanged, so they cancel

- _____, gases and liquids such as water are **not** split into ions.
- Write the quantity in front of an ion: $H_2SO_4(aq)$ splits into $2H^+(aq) + SO_4^{2-}(aq)$.
- Aqueous covalent substances such as aqueous chlorine are not split: $Cl_2(g) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$.

3 Chemical Equations (continued)

3.6 An ionic equation with a solid and a gas

WORKED EXAMPLE Write the ionic equation for $ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + CO_2(g) + H_2O(l)$.

Split only the aqueous substances: $2H^+ + 2NO_3^-$ and $Zn^{2+} + 2NO_3^-$. Keep $ZnCO_3(s)$, $CO_2(g)$ and $H_2O(l)$ whole.

Cancel the spectator ions, $2NO_3^-(aq)$.

$ZnCO_3(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + CO_2(g) + H_2O(l)$

3.7 A precipitation reaction

WORKED EXAMPLE Write the ionic equation for $AgNO_3(aq) + HCl(aq) \rightarrow AgCl(s) + HNO_3(aq)$.

Split the aqueous substances: $Ag^+ + NO_3^- + H^+ + Cl^- \rightarrow AgCl(s) + H^+ + NO_3^-$.

Cancel the spectator ions, H^+ and NO_3^-. The solid AgCl stays whole.

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

3.8 Common mistakes to avoid

- Changing a subscript to balance an equation. Subscripts belong to the formula of the substance, so only change the number in **front**.
- Forgetting the brackets in a formula such as $Mg(NO_3)_2$.
- Writing hydrogen as H instead of H_2. The diatomic elements are H_2, N_2, O_2, F_2, Cl_2, Br_2 and I_2.
- Splitting a solid, a gas or water into ions in an ionic equation.
- Leaving the state symbols out when the question asks for them.

EXAM TIP After you balance, count the atoms of every element one last time. Check the charges too in an ionic equation.

07 Mole Concept and Stoichiometry



1 Relative Masses

1.1 Relative atomic mass, A_r

DEFINITION The **relative atomic mass (A_r)** of an element is the average mass of one atom of the element compared with _____ of the mass of a carbon-12 atom.

- The A_r values are in the Periodic Table. They have **no units**.

WORKED EXAMPLE Chlorine is 75 % chlorine-35 and 25 % chlorine-37. Find its A_r.

$A_r = (75/100 \times 35) + (25/100 \times 37) = \mathbf{35.5}$

1.2 Relative molecular and formula mass, M_r

DEFINITION The **relative molecular mass (M_r)** of a molecule is the _____ _____ of all the atoms in its formula. For an ionic compound it is called the **relative formula mass** and has the same symbol, M_r.

Substance	Formula	Working	M_r
water	H_2O	2(1) + 16	18
carbon dioxide	CO_2	12 + 2(16)	44
calcium carbonate	$CaCO_3$	40 + 12 + 3(16)	100
iron(II) nitrate	$Fe(NO_3)_2$	56 + 2(14) + 6(16)	180

NOTE Multiply the A_r by the number of atoms of that element, including the number outside any brackets.

1 Relative Masses (continued)

1.3 Percentage mass of an element

$$\text{percentage mass} = \frac{\text{number of atoms} \times A_r \text{ of the element}}{M_r \text{ of compound}} \times 100\ \%$$

WORKED EXAMPLE Find the percentage mass of nitrogen in $(NH_4)_2SO_4$.

$M_r = 2(14 + 4) + 32 + 4(16) = 132$.

Percentage of N $= \frac{2 \times 14}{132} \times 100\ \% =$ **21.2 %**.

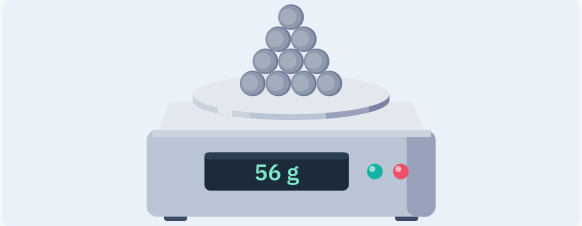
2 The Mole

2.1 Counting particles



a bank counts coins by weighing them, not one by one

same idea →



56 g of iron (1 mol)
= 6.02 × 10²³ Fe atoms

DEFINITION One **mole** of any substance contains _____ particles. This number is the **Avogadro constant**. The symbol for the mole is **mol**.

- The particles can be atoms, molecules, ions or even electrons.

$$\text{number of moles} = \frac{\text{number of particles}}{6.02 \times 10^{23}}$$

2 The Mole (continued)

2.2 Particles in a mole

Amount	Particles	Number
1 mol of Ne	Ne atoms	6.02×10^{23}
1 mol of H_2	H_2 molecules	6.02×10^{23}
1 mol of H_2	H atoms (2 mol)	1.204×10^{24}
2 mol of $Mg(NO_3)_2$	NO_3^- ions (4 mol)	2.408×10^{24}

EXAM TIP Do not mix up the number of **moles** of a substance with the number of **particles**. Look at what the question asks you to count.

2.3 Moles and mass

DEFINITION The mass of one mole is the **molar mass**. It equals the A_r or M_r in _____. The unit is g/mol.

$$\text{number of moles} = \frac{\text{mass in g}}{\text{molar mass in g/mol}}$$

WORKED EXAMPLE Find the mass of 4.5 mol of lead(II) sulfate, $PbSO_4$.

Molar mass = 207 + 32 + 4(16) = 303 g/mol.

Mass = 4.5 × 303 = **1360 g** (3 s.f.).

2.4 Moles and gas volume

DEFINITION One mole of **any gas** occupies _____ at r.t.p. (25 °C and 1 atmosphere).

$$\text{number of moles of gas} = \frac{\text{volume in dm}^3}{24\ \text{dm}^3}$$

WORKED EXAMPLE Find the volume of 2 mol of oxygen at r.t.p.

Volume = 2 × 24 = **48 dm^3**.

2 The Mole (continued)

2.5 Concentration of a solution

DEFINITION The **concentration** of a solution is the amount of solute dissolved in a unit volume of solution. Two units are used: **g/dm³** and **mol/dm³**.

$$\text{concentration in g/dm}^3 = \frac{\text{mass of solute in g}}{\text{volume of solution in dm}^3}$$

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

$$\text{concentration in mol/dm}^3 = \frac{\text{concentration in g/dm}^3}{\text{molar mass in g/mol}}$$

NOTE 1 dm^3 = 1000 cm^3. Change cm^3 to dm^3 by dividing by 1000.

WORKED EXAMPLE 500 cm^3 of solution contains 2 mol of NaCl. Find the concentration in g/dm^3.

Mass of NaCl = 2 × 58.5 = 117 g.

$$\text{Concentration} = \frac{117\text{ g}}{0.500\text{ dm}^3} = \mathbf{234\ g/dm^3}.$$

WORKED EXAMPLE 1.74 g of K_2SO_4 is dissolved to make 200 cm^3 of solution. Find the concentration in mol/dm^3.

Molar mass = 2(39) + 32 + 4(16) = 174 g/mol. Moles = $\frac{1.74}{174}$ = 0.01 mol.

$$\text{Concentration} = \frac{0.01\text{ mol}}{0.200\text{ dm}^3} = \mathbf{0.05\ mol/dm^3}.$$

2.6 The formula triangles



NOTE Cover the quantity you want. A horizontal line means **divide** and a vertical line means **multiply**.

3 Empirical and Molecular Formulae

3.1 Two kinds of formula

Empirical formula

The _____ of the atoms of each element in a compound.

Molecular formula

The _____ of atoms of each element in one molecule, a multiple of the empirical one.

Compound	Molecular	Empirical
water	H_2O	H_2O
phosphorus(V) oxide	P_4O_{10}	P_2O_5
ethene	C_2H_4	CH_2

3.2 Empirical formula from percentage mass

A compound is 40.0 % carbon, 6.70 % hydrogen and 53.3 % oxygen by mass.

Step	C	H	O
1 Assume 100 g, so mass in g	40.0	6.70	53.3
2 Divide by A_r to get moles	40.0 ÷ 12 = 3.33	6.70 ÷ 1 = 6.70	53.3 ÷ 16 = 3.33
3 Divide by the smallest	1	2	1

The empirical formula is **CH_2O**.

3.3 Molecular formula

$$n = \frac{M_r \text{ of compound}}{M_r \text{ of empirical formula}}$$

WORKED EXAMPLE A compound has $M_r = 60$ and empirical formula CH_2O. Find its molecular formula.

M_r of $CH_2O = 12 + 2(1) + 16 = 30$.

$n = 60 \div 30 = 2$.

Molecular formula = $(CH_2O)_2$ = **$C_2H_4O_2$**.

3.4 When the ratio is not a whole number

WORKED EXAMPLE Plexiglass is 59.9 % carbon, 8.1 % hydrogen and 32.0 % oxygen by mass. Find its empirical formula.

Moles: C = 59.9 ÷ 12 = 4.99, H = 8.1 ÷ 1 = 8.1, O = 32.0 ÷ 16 = 2.00.

Divide by the smallest (2.00): C = 2.5, H = 4.05, O = 1.

2.5 is not a whole number, so multiply all by 2: C = 5, H = 8, O = 2. The formula is **$C_5H_8O_2$**.

4 Stoichiometry

4.1 The mole ratio

Stoichiometry is the ratio between the quantities of substances, in _____, in a chemical reaction.

$MgCO_3$	+	$2HCl$	→	$MgCl_2$	+	CO_2	+	H_2O
1 mol		2 mol		1 mol		1 mol		1 mol

EXAM TIP The ratio of **moles** equals the ratio of the numbers in front of the formulae, not the ratio of masses.

4.2 Reacting masses and gas volumes

1. Find the **moles** of the substance you know (from mass, gas volume or concentration and volume).
2. Use the **mole ratio** from the balanced equation to find the moles of the substance you want.
3. Change the moles to the **mass**, **volume** or **concentration** you want.

WORKED EXAMPLE Find the mass of magnesium carbonate needed to produce 100 dm^3 of carbon dioxide at r.t.p. $MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + CO_2(g) + H_2O(l)$

Moles of CO_2 = 100 ÷ 24 = 4.17 mol.

1 mol of $MgCO_3$ gives 1 mol of CO_2, so moles of $MgCO_3$ = 4.17 mol.

Mass = 4.17 × (24 + 12 + 48) = **350 g**.

WORKED EXAMPLE Find the mass of iron made from 2000 g of iron(III) oxide. $Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

M_r of Fe_2O_3 = 160, so moles = 2000 ÷ 160 = 12.5 mol.

1 mol of Fe_2O_3 gives 2 mol of Fe, so moles of Fe = 25.0 mol.

Mass = 25.0 × 56 = **1400 g**.

4.3 Gas volumes in a reaction

NOTE **Avogadro's law:** at the same temperature and pressure, equal volumes of gases contain the **same number of particles**. So the mole ratio of gases is also their **volume ratio**.

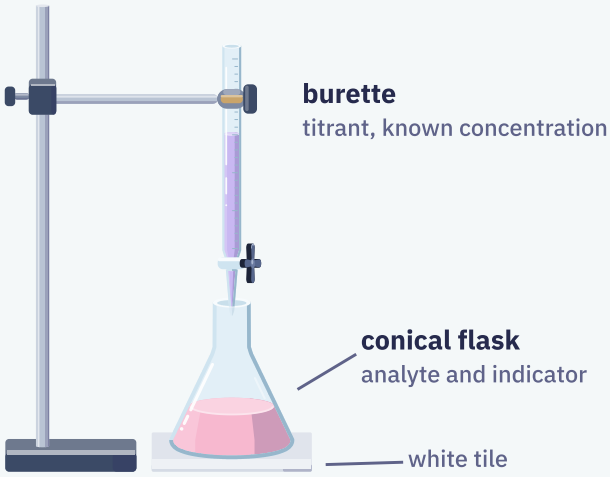
WORKED EXAMPLE $N_2(g) + O_2(g) \rightarrow 2NO(g)$. Find the volume of nitrogen needed to make 800 cm^3 of nitric oxide.

Volume ratio N_2 : NO = 1 : 2.

Volume of N_2 = 800 ÷ 2 = **400 cm^3**.

4 Stoichiometry (continued)

4.4 Titration



- The **titrant** has a known concentration. The **analyte** has an unknown concentration.
- Add a fixed volume of analyte (usually 25.0 cm^3, by pipette) and a few drops of indicator to the flask. Add titrant from the burette, swirling, until the colour changes permanently. This is the **end-point**.
- Repeat until two titres agree to within 0.10 cm^3, and use the average of those titres.

WORKED EXAMPLE 20.0 cm^3 of H_2SO_4 reacts with 30.0 cm^3 of 0.100 mol/dm^3 KOH. $2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$. Find the concentration of the acid.

Moles of KOH = 0.100 × 0.0300 = 0.003 mol.

2 mol KOH react with 1 mol H_2SO_4, so moles of acid = 0.0015 mol.

Concentration = 0.0015 ÷ 0.0200 = **0.075 mol/dm^3**.

4.5 A gas reacting with a solution

WORKED EXAMPLE 250 cm^3 of calcium hydroxide solution reacts completely with 6 dm^3 of carbon dioxide at r.t.p. $CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$. Find the concentration of the calcium hydroxide and the mass of calcium carbonate formed.

Moles of CO_2 = 6 ÷ 24 = 0.25 mol, so moles of $Ca(OH)_2$ = 0.25 mol (ratio 1 : 1).

Concentration = 0.25 ÷ 0.250 = **1.0 mol/dm^3**.

Moles of $CaCO_3$ = 0.25 mol, so mass = 0.25 × 100 = **25 g**.

4 Stoichiometry (continued)

4.6 Limiting reactants

recipe	**2 bread**	+	**1 cheese**	→ **1 sandwich**
we have	10 bread		4 cheese	4 sandwiches
	2 left over: in **excess**		all used: **limiting**	set by the cheese

- The reactant that is **completely used up** is the _____. The other is in **excess**.
- The amount of product depends on the limiting reactant, so calculate from it.
- Example: 5 mol of NaOH and 3 mol of HCl react 1 : 1. All 3 mol of HCl are used up and 2 mol of NaOH remain. HCl is limiting and 3 mol of NaCl form.

4.7 Finding the limiting reactant

WORKED EXAMPLE 2.00 g of calcium carbonate powder is added to 50.0 cm^3 of 2.00 mol/dm^3 hydrochloric acid. Find the volume of carbon dioxide at r.t.p. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Moles of HCl = 2.00 × 0.0500 = 0.100 mol.

Moles of $CaCO_3$ = 2.00 ÷ 100 = 0.0200 mol.

0.100 mol of HCl needs 0.0500 mol of $CaCO_3$, but only 0.0200 mol is present. **$CaCO_3$ is limiting.**

1 mol of $CaCO_3$ gives 1 mol of CO_2, so volume = 0.0200 × 24 = **0.480 dm^3**.

4.8 Common mistakes in mole calculations

- Forgetting to change cm^3 to dm^3: divide by 1000.
- Using the ratio of **masses** instead of the ratio of **moles** from the equation.
- Using the M_r of the wrong substance, or leaving out the number in front of a formula.
- Using 24 dm^3 for a solid, a liquid or a solution. It is only for gases at r.t.p.

5 Percentage Yield and Purity

5.1 Percentage yield

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical maximum yield}} \times 100\ \%$$

- The **theoretical maximum** is worked out from the mole ratio in the equation.

WORKED EXAMPLE 100 g of $CaCO_3$ reacts with excess HNO_3: $CaCO_3 + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2O + CO_2$. 129.6 g of $Ca(NO_3)_2$ is collected. Find the percentage yield.

Moles of $CaCO_3$ = 100 ÷ 100 = 1 mol, so 1 mol of $Ca(NO_3)_2$ is the theoretical maximum.

M_r of $Ca(NO_3)_2$ = 164, so the theoretical mass is 164 g.

Percentage yield = 129.6 ÷ 164 × 100 % = **79.0 %**.

5.2 Percentage purity

$$\text{percentage purity} = \frac{\text{mass of pure substance in the sample}}{\text{total mass of the sample}} \times 100\ \%$$

WORKED EXAMPLE 50.0 g of impure aspirin, $C_9H_8O_4$, contains 0.20 mol of pure aspirin. Find the percentage purity.

Molar mass = 9(12) + 8(1) + 4(16) = 180 g/mol.

Mass of pure aspirin = 0.20 × 180 = 36.0 g.

Percentage purity = 36.0 ÷ 50.0 × 100 % = **72.0 %**.

5.3 Purity from a reaction

WORKED EXAMPLE 200 g of impure zinc is added to excess copper(II) chloride solution and 171 g of copper is obtained. $Zn + CuCl_2 \rightarrow ZnCl_2 + Cu$. Find the percentage purity of the zinc.

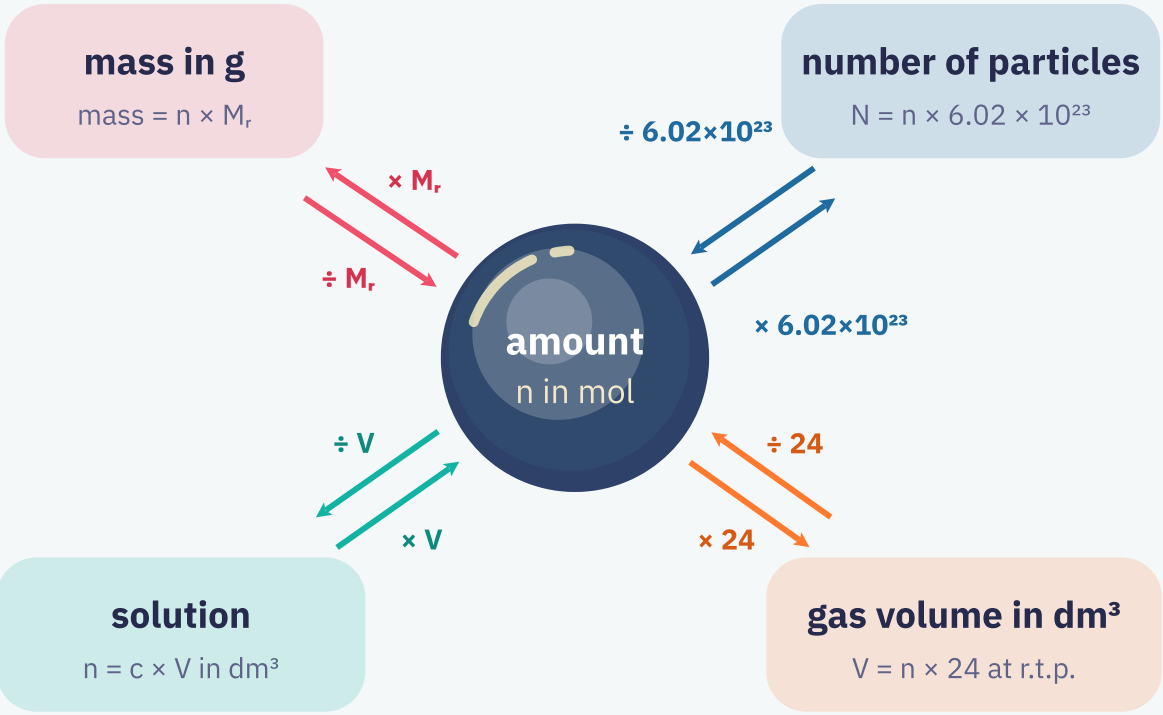
Moles of Cu = 171 ÷ 64 = 2.67 mol, so moles of Zn = 2.67 mol (ratio 1 : 1).

Mass of pure zinc = 2.67 × 65 = 174 g.

Percentage purity = 174 ÷ 200 × 100 % = **86.8 %**.

6 Summary

6.1 Moles at the centre



- To go between any two quantities, change to **moles** first, then change to the quantity you want.

6.2 Formulae to know

To find	Use
M_r	add the A_r of every atom in the formula
moles from mass	mass in g ÷ molar mass in g/mol
moles of a gas	volume in dm^3 ÷ 24
moles from particles	number of particles ÷ 6.02 × 10^{23}
moles in solution	concentration in mol/dm^3 × volume in dm^3
concentration in g/dm^3	concentration in mol/dm^3 × M_r
molecular formula	$(\text{empirical formula})_n$, n = M_r ÷ M_r of empirical formula
percentage yield	actual ÷ theoretical maximum × 100 %
percentage purity	pure mass ÷ sample mass × 100 %

EXAM TIP Write the equation, then moles, then the mole ratio, then the answer. Show every step and give the units.