

Year 9 Science · Unit 1

Knowledge Book

Atomic Structure & Chemical Change

*The **why** behind your workbook — read this to understand and to revise.*

v2.1

How to use this book

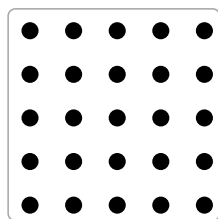
Your **workbook** is where you **write**. This **Knowledge Book** is where you **read** — to understand the ideas and to revise for the test. Each chapter has the same shape: **Core knowledge** (the facts) → **Why it works** (the reasons) → worked reasoning → **Key words** → **Watch out** (common mistakes) → **Going deeper** (beyond the test).

1.1 Modelling matter & classifying substances

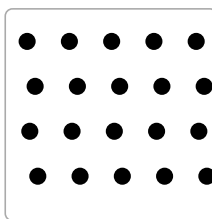
What matter is, and how we model it



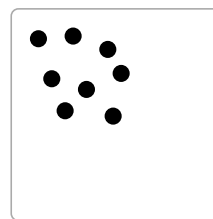
Matter is the 'stuff' that everything is made of. We define it as anything that has **mass** and **volume**. All matter is ultimately built from **atoms**. To explain how it behaves, scientists use the **particle model**: matter is made of tiny particles in constant motion, with forces of attraction between them. How much **energy** those particles have decides the **state** — solid, liquid or gas.



Solid — fixed, packed



Liquid — close, can move



Gas — far apart, fast

We also sort matter by **what its particles are made of**. Three categories cover everything:

Elemental substances — made of **one type of atom only**. They may still be molecules of two or more atoms, as long as the atoms are the same, e.g. O_2 and H_2 .

Compound substances — **two or more types of atom chemically combined in fixed ratios**, e.g. H_2O , $NaCl$.

Mixtures — different substances **mixed but not chemically combined**, in **no fixed ratio**, e.g. air, salt water.

A **homogeneous** mixture looks the same throughout (salt solution, air); a **heterogeneous** mixture has visibly separate parts (oil and water, a rock).

Why these categories actually differ



The dividing line between a **compound** and a **mixture** is the words **chemically combined**. In a compound, the atoms are joined by chemical bonds, so the only way to separate them is **another chemical reaction**. In a mixture nothing is bonded, so the parts keep their own properties and can be separated **physically** — by filtering, evaporating or letting them settle.

That single idea explains a lot:

Fixed ratio?

Compound **yes** (always H_2O).
Mixture **no** (salt water can be weak or strong).

New properties?

A compound behaves like a **new substance**, unlike the elements it formed from. A mixture keeps each part's properties.

Separated how?

Compound — by **reaction** only.
Mixture — by **physical** means.

The same reasoning sits underneath the states of matter: heating gives particles more energy, so they break free of their fixed positions (solid → liquid) and then of each other (liquid → gas). The **particles do not change** — only their arrangement and movement do.

WORKED EXAMPLE

Is salt water a compound or a mixture? It contains water molecules **and** sodium chloride, but they are simply mixed — the salt is dissolved, not bonded to the water, and the ratio can vary. You can get the salt back by **evaporating** the water (a physical change, no reaction). So it is a **mixture** — and a **homogeneous** one, because it looks uniform throughout.

Key words



matter	anything that has mass and volume; all matter is made of atoms
particle model	the idea that matter is made of tiny moving particles with forces between them
elemental substance (element)	a substance made of only one type of atom
compound substance (compound)	two or more types of atom chemically combined in fixed ratios
mixture	different substances mixed but not chemically combined, in no fixed ratio
homogeneous / heterogeneous	uniform throughout / made of visibly separate parts
chemically combined	joined by chemical bonds; separable only by another chemical reaction

Watch out



Trap Truth A mixture and a compound are the same kind of thing. No — a compound's atoms are **bonded** in a **fixed ratio**; a mixture's substances are just **mixed**, in any ratio, and keep their own properties.

Trap Air is a compound because it is “one thing”.

Truth Air is a **mixture** of gases (mostly nitrogen and oxygen). They are not bonded and the proportions can vary.

Trap An element can only be a single atom.

Truth An elemental substance can be a **molecule** (O_2 , H_2) – what matters is that **all the atoms are the same type**.

Trap Melting or boiling makes a new substance.

Truth These are **physical** changes – the particles gain energy and move apart, but ice, water and steam are all still H_2O .

Going deeper



For 2000 years people believed everything was made from four “elements” – earth, air, fire and water. We now recognise about **118** real elements, and the difference between a mixture and a compound was only settled once chemists could weigh reactants and products carefully. The habit to carry forward is the one this whole unit is built on: **a model is kept only as long as the evidence supports it**. Next year, the bonding that holds a compound together (ionic and covalent) is exactly what you will study to explain **why** those fixed ratios appear.

1.2 Symbols, formulae & representing substances

The shorthand of chemistry

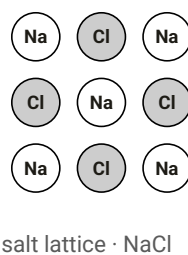
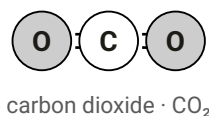
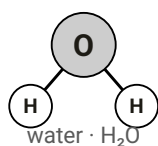


Chemists write substances in a **shorthand** that is the same in every language. Each element has a **chemical symbol** of one or two letters — the **first always a capital**, the second (if there is one) **always lowercase**: C, He, Na, Fe. Compounds are written as a **chemical formula**: symbols with **subscripts** that show **how many of each atom** are present.

Reading a formula

- A **subscript** applies only to the symbol **immediately before** it: H_2O = 2 hydrogen, 1 oxygen.
- **Brackets** group atoms; a subscript **outside** the bracket multiplies **everything inside**: $\text{Ca}(\text{OH})_2$.
- A **big number in front** (a coefficient) multiplies the **whole formula**: 2CO_2 = 2 carbon, 4 oxygen.

The formula is just a count of the atoms in the picture



Why the rules are written so strictly



The capital-then-lowercase rule is not fussiness — it removes **ambiguity**. **Co** is the element cobalt, but **CO** is the **compound** carbon monoxide (one carbon, one oxygen). One letter case changes the meaning entirely, so the convention has to be exact.

Subscripts and coefficients do **different jobs**, and mixing them up is the commonest slip:

Subscript (small, after a symbol)

changes **what the substance is** — H_2O and H_2O_2 are different substances.

Coefficient (big, in front)

changes **how many copies** you have — $2 \text{H}_2\text{O}$ is two of the **same** water molecule.

WORKED EXAMPLE

How many atoms of each element are in $\text{Ca}(\text{OH})_2$?

The bracket (OH) is multiplied by 2 → 1 calcium, 2 oxygen, 2 hydrogen — **5 atoms in total**.

WORKED EXAMPLE

How many oxygen atoms are in **2 H₂SO₄**?

One H₂SO₄ has 4 oxygen. The coefficient 2 multiplies the whole formula → **8** oxygen atoms (and 4 hydrogen, 2 sulfur).

Key words



chemical symbol	one or two letters for an element; first letter capital, second lowercase
chemical formula	symbols and subscripts showing the atoms in a substance
subscript	the small number after a symbol; counts the atoms of that element only
coefficient	the large number in front of a formula; multiplies the whole formula
diatomic molecule	a molecule of two atoms, e.g. O ₂ , H ₂ , Cl ₂

Watch out



Trap CO and Co mean the same thing.

Truth No – **CO** is carbon monoxide (two atoms); **Co** is the element cobalt. Letter case is part of the meaning.

Trap A subscript can multiply a whole formula.

Truth A subscript only counts the symbol **before** it. To multiply the **whole** formula, use a **coefficient** in front (2 H₂O).

Trap 2 H₂O has 2 oxygen atoms in each molecule.

Truth Each water molecule still has **1** oxygen. The 2 means **two molecules**, so 2 oxygen atoms **in total**.

Trap You can change a subscript to balance an equation.

Truth Never – changing a subscript changes the **substance**. You only adjust **coefficients** (see 1.7).

Going deeper



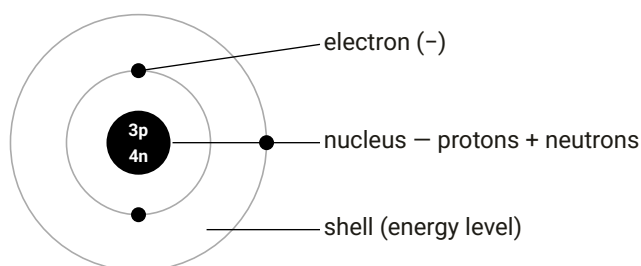
Some symbols look nothing like the English name because they come from **Latin**: **Na** (sodium) from **natr**ium, **Fe** (iron) from **ferr**um, **Pb** (lead) from **plumb**um — the root of the word “plumber”. This shared shorthand means a chemist anywhere in the world reads H_2SO_4 the same way, which is exactly why science can be checked and built on across borders. Next you will use these formulae inside **equations**, where the counting you practised here becomes the basis of **balancing** and the **conservation of mass**.

1.3 Inside the atom – the Bohr model

What an atom is made of



Everything around you is made of **atoms**. An atom has a tiny central **nucleus** containing **protons** and **neutrons**, surrounded by **electrons** that occupy **shells** (energy levels). The nucleus is unimaginably small but holds almost all of the atom's mass; the rest of the atom is **mostly empty space**.



Particle	Where it is	Relative charge	Relative mass
Proton	in the nucleus	+1	1
Neutron	in the nucleus	0	1
Electron	in shells around it	- 1	almost 0 (about 1/1836)

Why the atom looks the way it does



Why is an atom neutral? It contains equal numbers of positive protons and negative electrons, so the two charges cancel out exactly. (This is the rule we lean on again and again: in a **neutral atom, electrons = protons**.)

Why do we say it is “mostly empty space”? In Ernest Rutherford's famous gold-foil experiment, nearly all the tiny particles fired at a thin sheet of gold passed **straight through**; just a few bounced sharply back. The only way to explain that: almost all the mass and all the positive charge sit in a **very small, dense nucleus**, with the electrons far away and empty space in between.

Why is the Bohr model just a “model”? The **Bohr model** — electrons drawn in neat circles — is a **simplification**. Real electrons are not on tidy rings; they occupy fuzzy regions called **orbitals**, where we can only say where an electron is **likely** to be. We keep the Bohr model because it is simple and it still **predicts** how atoms behave. Our picture has improved each time new evidence arrived:

Dalton (solid sphere) → **Thomson** (plum pudding) → **Rutherford** (nucleus) → **Bohr** (shells) → **quantum** (orbitals)

WORKED EXAMPLE

Reading a helium atom. Helium has 2 protons, 2 neutrons and 2 electrons. The 2 protons and 2 neutrons sit in the nucleus; the 2 electrons fill the first shell. Check the charge: $2 \times (+1)$ from protons and $2 \times (-1)$ from electrons gives **0** – neutral, exactly as a normal atom should be.

Key words



atom the smallest particle of an element that still behaves like that element

nucleus the tiny, dense centre of an atom; holds the protons and neutrons

proton positive particle in the nucleus – relative charge +1, relative mass 1

neutron neutral particle in the nucleus – relative charge 0, relative mass 1

electron negative particle in the shells – relative charge -1 , almost no mass

shell (energy level) a region around the nucleus where electrons are found

Bohr model a simple, useful model with electrons in fixed circular shells

orbital the more accurate idea: a region where an electron is **likely** to be

Watch out



Trap Electrons sit inside the nucleus.
Truth Only protons and neutrons are in the nucleus. Electrons are in **shells around** it.

Trap An atom is a solid little ball.
Truth An atom is **mostly empty space** – the nucleus is tiny compared with the whole atom.

Trap The Bohr model is exactly how atoms are.
Truth It is a **model** – useful but simplified. Real electrons occupy **orbitals** (probability clouds), not neat rings.



If the whole atom were the size of a sports stadium, the nucleus would be about the size of a **pea** in the centre — that is how much of an atom is empty space. Later (IB Diploma) you will replace shells with **orbitals** and energy **sub-levels**, and you will see direct evidence for them. The key habit to take forward: **a scientific model is judged by how well it explains and predicts — and it is always open to improvement.**

1.4 Atomic notation, ions & isotopes

Counting the particles



Every atom is described by two numbers:

Atomic number, Z

the number of **protons**. It **defines the element** — change it and you have a different element.

Mass number, A

the number of **protons + neutrons** — the count of the heavy particles in the nucleus.

From these two numbers you can find every particle:

$$\text{neutrons} = A - Z \quad \text{electrons} = \text{protons (in a neutral atom)}$$

We write them around the symbol like this: ${}^7_3\text{Li}$ — mass number on top, atomic number below.

The one part everyone mixes up



This is the idea most students get tangled, so go slowly.

Electrons = protons — always (for a neutral atom). An atom has no overall charge, so the negative electrons must exactly balance the positive protons. If there are 3 protons, there are 3 electrons. This one is **guaranteed**.

Neutrons are the independent one. Neutrons add mass and help hold the nucleus together, but their number is **not** fixed by the protons — which is exactly why **isotopes** exist (same element, different neutrons). So you must **work neutrons out** with $A - Z$. Never assume neutrons = protons; sometimes they happen to be equal (carbon-12 has 6 and 6), but often they are not (lithium-7 has **3** protons and **4** neutrons).

So what does each particle control?

change the **protons**
→ a **different element**

change the **neutrons**
→ a different **isotope** (same element)

change the **electrons**
→ an **ion** (same element, now charged)

WORKED EXAMPLE

Lithium-7 ${}^7_3\text{Li}$. $Z = 3$, $A = 7$.

Protons = $Z = 3$. Electrons = protons = **3**. Neutrons = $A - Z = 7 - 3 = 4$.

Notice: electrons match protons (both 3), but neutrons (4) do **not** — you had to calculate them.

WORKED EXAMPLE

A sodium ion ${}^{23}_{11}\text{Na}^{+}$. An ion has **lost or gained electrons**, so it is charged.

Protons = 11 (unchanged). Neutrons = $23 - 11 = 12$. The “+” means it **lost one electron**, so electrons = 10.

Check the charge: protons – electrons = $11 - 10 = +1$. ✓

WORKED EXAMPLE

Two isotopes of chlorine ${}^{35}_{17}\text{Cl}$ and ${}^{37}_{17}\text{Cl}$.

Both have $Z = 17$, so both have **17 protons and 17 electrons** – they are the same element. But neutrons = $35 - 17 = 18$ versus $37 - 17 = 20$. Same element, different neutrons = **isotopes**.

Key words



atomic number (Z) the number of protons; also called the **proton number** – defines which element it is

mass number (A) the number of protons + neutrons; also called the **nucleon number** (protons and neutrons together are called nucleons)

ion an atom that has lost or gained electrons, so it carries a charge

isotope atoms of the **same** element (same Z) with **different** numbers of neutrons

relative atomic mass the average mass of an element's atoms, allowing for its isotopes

electron arrangement how the electrons are spread across the shells, e.g. 2, 8, 1 for sodium

Watch out



Trap Truth Neutrons always equal protons.
Not true. Only sometimes (carbon-12: 6 and 6). Lithium-7 has 3 protons but 4 neutrons. **Always use $A - Z$.**

Trap Electrons equal neutrons.

Truth No – **electrons equal protons** in a neutral atom. Neutrons are separate and can differ.

Trap Isotopes are different elements.

Truth No – same element, same protons. Only the **neutron** number (and so the mass number) differs.

Trap The mass number is the atom's mass in grams.

Truth No – it is a **count** of protons + neutrons (a relative figure), not a mass in grams.

Trap A “+” ion has gained something.

Truth A **positive** ion has **lost** electron(s); a **negative** ion has **gained** electron(s). Charge = protons – electrons.

Going deeper



Because a real sample of chlorine is a **mixture** of $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$ (about 3 to 1), its **relative atomic mass** is a weighted average – **35.5**, not a whole number. That is why the periodic table shows decimals. Later you will measure these proportions directly with a **mass spectrometer** (IB Diploma). And the idea of ions leads straight into **ionic bonding** next year: metals lose electrons and non-metals gain them, and the charges you worked out here decide the formulae.

1.5 The Periodic Table — groups & periods

How the table is organised



The Periodic Table lists every element in order of **proton number** (atomic number). That order is not just tidy — it groups elements so that ones with **similar properties line up**.

Groups are the **vertical columns** (1–18).
Elements in the same group have the **same number of outer-shell electrons**, so they have **similar chemical properties**.

Periods are the **horizontal rows** (1–7). Across a period the proton number rises by one each step, and a new outer shell starts with each new period.

Metals sit to the **left and centre**; **non-metals** to the **right**. The **noble gases** (He, Ne, Ar, Kr, Xe, Rn) fill **Group 18** — they have **complete outer shells**, so they barely react. (Older books number this group as Group 0.)

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	4 He
2	7 Li	9 Be											11 B	12 C	14 N	16 O	19 F	20 Ne
3	23 Na	24 Mg											27 Al	28 Si	31 P	32 S	35.5 Cl	40 Ar
4	39 K	40 Ca	45 Sc	48 Ti	51 V	52 Cr	55 Mn	56 Fe	59 Co	58 Ni	64 Cu	65 Zn	70 Ga	73 Ge	75 As	79 Se	80 Br	84 Kr
5	85 Rb	88 Sr	89 Y	91 Zr	93 Nb	96 Mo	[98] Tc	101 Ru	103 Rh	106 Pd	108 Ag	112 Cd	115 In	119 Sn	122 Sb	128 Te	127 I	131 Xe
6	133 Cs	137 Ba	175 Lu	178 Hf	181 Ta	184 W	186 Re	190 Os	192 Ir	195 Pt	197 Au	201 Hg	204 Tl	207 Pb	209 Bi	[209] Po	[210] At	[222] Rn
7	[223] Fr	[226] Ra	[262] Lr	[267] Rf	[270] Db	[271] Sg	[270] Bh	[277] Hs	[278] Mt	[281] Ds	[282] Rg	[285] Cn	[286] Nh	[289] Fl	[290] Mc	[293] Lv	[294] Ts	[294] Og
	87	88	103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118
*	139 La	140 Ce	141 Pr	144 Nd	[145] Pm	150 Sm	152 Eu	157 Gd	159 Tb	163 Dy	165 Ho	167 Er	169 Tm	173 Yb				
**	[227] Ac	232 Th	231 Pa	238 U	[237] Np	[244] Pu	[243] Am	[247] Cm	[247] Bk	[251] Cf	[252] Es	[257] Fm	[258] Md	[259] No				
	89	90	91	92	93	94	95	96	97	98	99	100	101	102				

☐ metals ☐ metalloids ☐ non-metals ☐ noble gases

The bold staircase line separates metals (left) from non-metals (right).

Metals vs non-metals — typical physical properties

Typical metal (solid)	Typical non-metal (solid)
Good conductor of heat and electricity	Poor conductor of heat and electricity
Shiny surface	Dull surface

Typical metal (solid)	Typical non-metal (solid)
High density	Low density
Malleable and ductile	Brittle

Why position predicts behaviour



The whole power of the table is this: **an element's place tells you how it behaves**, because **position reflects electron arrangement**.

Same group → **similar reactions**. Chemical reactions are about **outer-shell electrons**. Every Group 1 atom has **one** outer electron, every Group 2 atom **two**, every Group 17 atom **seven**. Because reactions depend on those outer electrons, elements in a group react in the **same way** – which is why we can learn a group's behaviour once and apply it to all its members (1.6).

Group number → **outer electrons**. For the main groups, the group number gives the outer-electron count (Group 1 → 1; Group 17 → 7; Group 18 → a full 8, except helium's 2). **Period number** → **number of shells**. So sodium (period 3, group 1) has **three** shells with **one** electron in the outermost: 2, 8, 1.

Why the noble gases are inert. A full outer shell is already stable, so a noble-gas atom has no tendency to lose or gain electrons – and reacting **means** losing or gaining electrons. A complete shell is the “goal” that drives every other element's reactions.

WORKED EXAMPLE

Place chlorine (Cl, Z = 17). Its electron arrangement is 2, 8, 7 – **three shells**, so it is in **period 3**; **seven** outer electrons, so it is in **group 17**. Seven outer electrons (one short of a full shell) tells you it is a **reactive non-metal** that tends to **gain** one electron.

Key words



group	a vertical column; same number of outer-shell electrons → similar properties
period	a horizontal row; the period number equals the number of occupied shells
metal / non-metal	metals (left and centre) conduct and are malleable; non-metals (right) do not and are brittle
noble gas	a Group 18 element with a complete outer shell, so very unreactive

proton number
(atomic number)

the number of protons; sets each element's place in the table

Watch out



Trap Elements in a group have the same number of shells.

Truth No — that is a **period**. A **group** shares the same number of **outer-shell electrons**, but each element down a group has **one more shell**.

Trap The table is ordered by atomic mass.

Truth It is ordered by **proton number**. Mass mostly rises too, but a few pairs (e.g. Ar and K) only sit correctly when ordered by protons.

Trap Noble gases are unreactive because they have no electrons.

Truth They are unreactive because their outer shell is **full** — the most stable arrangement, with no drive to gain or lose electrons.

Trap There is a sharp line between metals and non-metals.

Truth The change is gradual; a “staircase” of **metalloids** (e.g. silicon) shows in-between behaviour.

Going deeper



When Dmitri Mendeleev built his table in 1869 he **left gaps** for elements not yet discovered — and **predicted their properties** from the pattern. Gallium and germanium were later found almost exactly as he described: striking evidence that **a good model predicts, not just describes**. Today the table's true order — proton number — only made sense once the **nucleus** (1.3) was understood, so this chapter quietly depends on the one before it. Next year, the link between group position and electron arrangement becomes the basis of **bonding**.

1.6 Reactivity trends – groups 1, 2 & 17

Three groups, three patterns



Because elements in a group share the same number of outer electrons (1.5), each group reacts in a **characteristic way**, and reactivity changes **smoothly down the group**.

Group 1 – alkali metals

One outer electron. React with water, **fizzing** and giving off **hydrogen**, forming soluble **hydroxides**. Reactivity **increases** down the group (Li < Na < K).

Group 2 – alkaline earth metals

Two outer electrons. React with water and acid to give **hydrogen**; form oxides and hydroxides. Reactivity also **increases** down the group.

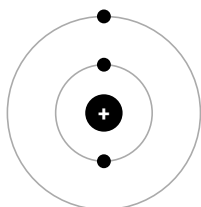
Group 17 – halogens

Seven outer electrons; coloured and **diatomic** (Cl₂, Br₂, I₂). React by **gaining** one electron. Reactivity **decreases** down the group (F > Cl > Br).

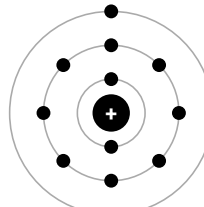
Why reactivity changes down a group



Every trend here comes from one idea: **how easily an atom loses or gains its outer electron**, which depends on **how far the outer shell is from the nucleus**.



Lithium 2,1 – outer electron **close** to the nucleus



Sodium 2,8,1 – outer electron **further out**

Metals (Groups 1 and 2) react by losing electrons. The further the outer electron is from the nucleus – i.e. the **lower down** the group – the **weaker** the pull holding it, so it is lost **more easily** and the metal is **more reactive**. That is why potassium beats sodium beats lithium.

Halogens (Group 17) react by gaining an electron. The same distance argument runs the **opposite** way: lower down the group, the incoming electron would sit **further from the nucleus**, where the pull is **weaker**, so it is captured **less easily**. Reactivity therefore **decreases** down the group, with **fluorine** the most reactive at the top.

WORKED EXAMPLE

Why is potassium more reactive than lithium? Both are Group 1, so both react by losing their **one** outer electron. In potassium that electron is in a shell **further from the nucleus** than in lithium, so it is held **less tightly** and lost **more easily**. Easier electron loss = **more reactive**.

WORKED EXAMPLE

Fluorine or iodine – which is more reactive? Both are Group 17 and react by **gaining** one electron. Fluorine's outer shell is **closest** to the nucleus, so it pulls in an electron **most strongly**. Fluorine is therefore **more reactive** than iodine.

Key words



reactivity	how readily a substance takes part in a chemical reaction
alkali metal (Group 1)	a soft, very reactive metal with one outer electron; reacts with water to give hydrogen
alkaline earth metal (Group 2)	a reactive metal with two outer electrons
halogen (Group 17)	a reactive non-metal with seven outer electrons; reacts by gaining one electron
diatomic	existing as molecules of two atoms, e.g. Cl_2

Watch out



Trap All groups get more reactive down the group.
Truth **Metals** (1 and 2) get **more** reactive down; **halogens** (17) get **less** reactive down. The direction depends on whether the atom **loses** or **gains** electrons.

Trap Group 1 metals gain electrons.
Truth They **lose** their single outer electron. It is the **non-metals** (like halogens) that **gain** electrons.

Trap Reactivity is about how heavy the atom is.
Truth It is about how **easily the outer electron is lost or gained** – i.e. its **distance from the nucleus**, not the atom's mass on its own.

Trap Fluorine is least reactive because it is at the top.
Truth The opposite – for halogens the **top** is **most** reactive, because the outer shell is closest to the nucleus.



The trend you predict on paper plays out dramatically in the lab: drop **caesium** (low in Group 1) into water and it reacts **explosively**, far more violently than lithium at the top. The deeper reason — **shielding** by inner shells and the **increasing nuclear charge** — is refined in the IB Diploma, but the Year-9 picture (distance from the nucleus) already predicts every trend correctly. These electron gains and losses are also the direct lead-in to **ionic bonding**: a Group 1 metal hands its electron to a Group 17 non-metal, and the opposite charges hold the compound together.

1.7 Physical & chemical change

Two kinds of change



In a **physical change**, the **arrangement and movement** of the particles change, but the **particles themselves stay the same**. Changes of state — melting, freezing, boiling — are the classic examples: **no new substance** is made and the change is usually **reversible**.

In a **chemical change**, the particles **rearrange** and **one or more new substances** form. It is usually **hard to reverse**. Tell-tale signs: a **colour change**, a **gas** given off, a **temperature change**, or a **solid (precipitate)** forming.

We write reactions as **reactants** → **products** and add **state symbols** to show each substance's physical state:

(s) = solid

(l) = liquid

(g) = gas

(aq) = dissolved in water

The law of conservation of mass: the **total mass of the products always equals the total mass of the reactants**. Around 1789 the French chemist **Antoine Lavoisier** established it by careful weighing; decades later **John Dalton** explained **why** — atoms are only **rearranged**, never created or destroyed.

Why mass is conserved — and why we balance



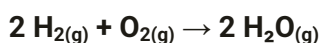
A chemical reaction does **not** make or destroy atoms. It takes the atoms in the reactants, **separates** them, and **recombines** them into the products. So the **number and type of each atom** is exactly the same before and after — and since each atom keeps its mass, the **total mass cannot change**.

That is the whole reason we **balance equations**. A balanced equation is simply a statement that **the atoms tally on both sides**. We make them tally by changing **coefficients** (the big numbers in front) — **never subscripts**, because changing a subscript would change the **substance** itself (1.2).

WORKED EXAMPLE

Balance $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$.

The right side has 1 oxygen but the left has 2, so we need **2** water molecules — which then needs **2** hydrogen molecules:



Tally: 4 H and 2 O on each side. Same atoms before and after — **mass is conserved**.

WORKED EXAMPLE

Why does a flask of fizzing acid seem to lose mass — but a sealed one does not? When hydrochloric acid reacts with sodium hydrogen carbonate, **carbon dioxide gas** escapes from an open flask, so the balance

reads lighter. Seal the flask (e.g. with a balloon) and the gas is trapped: the mass stays **the same**, exactly as Lavoisier's law predicts. The mass was never lost — it just **floated away**.

Key words



physical change particles rearrange but stay the same substance; usually reversible (e.g. melting)

chemical change particles rearrange into one or more new substances; usually hard to reverse

reactant / product the substances that react / the new substances formed

state symbol (s), (l), (g) or (aq), showing a substance's physical state in an equation

law of conservation of mass total mass of products = total mass of reactants (Lavoisier; explained by Dalton)

coefficient the number in front of a formula, adjusted to balance an equation

Watch out



Trap Mass is lost when a gas is given off.
Truth No — the gas still has mass; it just **leaves** an open container. In a **sealed** container the mass stays the same.

Trap You balance an equation by changing the little subscripts.
Truth Never — that changes the **substance** (H_2O vs H_2O_2). Only change the **coefficients** in front.

Trap Dissolving sugar is a chemical change.
Truth It is **physical** — the sugar is still sugar and can be recovered by evaporating the water. No new substance forms.

Trap A colour change always means a chemical reaction.
Truth It is a **strong clue**, but check for a **new substance**. Always weigh the evidence (gas, precipitate, temperature change) together.



A rusting iron nail gets **heavier**, not lighter — because iron atoms combine with **oxygen from the air**, and that oxygen's mass is now part of the rust. It looks like mass appears from nowhere until you remember the air. This is the law of conservation of mass in everyday life: mass is never lost or gained, only **rearranged**. The same accounting — atoms in = atoms out — scales all the way up to industrial chemistry and even (with energy added in) to the reactions inside stars.

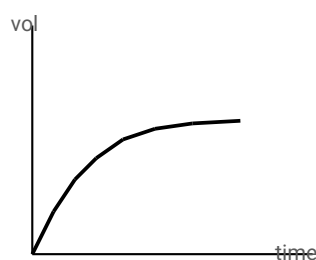
1.8 Rates of reaction

How fast — and what changes it



The **rate of a reaction** is how fast it happens — defined as the **change in concentration of a reactant or product over time**. A faster reaction uses up reactants and makes products **more quickly**.

Collision theory explains rate. For a reaction to occur, particles must **collide**, and the collision must be **successful** — the particles need **enough energy** to react. Anything that makes **successful collisions happen more often** speeds the reaction up.



volume of gas vs time

Reading a rate graph: a **steeper** line means a **faster** rate. The slope is **steepest at the start**, when concentration is highest, and the line **levels off** once a reactant runs out and the reaction stops.

The main factors are **concentration**, **temperature**, **surface area** and **catalysts**.

Why each factor works — all through collisions



Every factor changes the rate for the **same underlying reason**: it changes **how often successful collisions happen**.

Higher concentration

more particles in the same space → **more frequent** collisions.

Higher temperature

particles move **faster** and carry **more energy** → more collisions, and more of them **successful**.

Larger surface area (e.g. a powder)

more particles **exposed** at the surface → more collisions per second.

A catalyst

provides an **easier path** needing less energy → more collisions succeed. It is **not used up**.

This also explains the **shape of the graph**. At the start, reactant concentration is **highest**, so collisions are most frequent and the line is **steepest**. As reactants are consumed, collisions become **less frequent**, the line **flattens**, and when a reactant runs out the reaction **stops**.

WORKED EXAMPLE

Why is the reaction fastest at the very start? At the start the reactant concentration is at its **maximum**, so particles collide **most often**. The most successful collisions per second means the **fastest rate** – which is why the graph is **steepest** there and then gradually levels off.

Key words



rate of reaction	the change in concentration of a reactant or product over time
collision theory	particles must collide with enough energy (a successful collision) to react
successful collision	a collision with enough energy for a reaction to occur
catalyst	a substance that speeds up a reaction without being used up
surface area	the amount of a solid's surface exposed to react; greater for a powder than a lump

Watch out



Trap A catalyst is used up in the reaction.

Truth No – a catalyst speeds the reaction up but is **recovered unchanged** at the end, so a little goes a long way.

Trap Any collision causes a reaction.

Truth Only **successful** collisions do – the particles must have **enough energy**. Most collisions simply bounce off.

Trap A flat line on the graph means the reaction is fast.

Truth A **flat** line means the reaction has **slowed or stopped**. A **steep** line means it is fast.

Trap More concentration means bigger particles.

Truth Concentration is about **how many** particles are packed into a space, not their size – more particles means more frequent collisions.

Going deeper



Surface area can be dramatic: a solid lump of flour barely burns, but flour **dust** exposes so much surface that flour mills have **exploded**. Catalysts are everywhere too — a car's **catalytic converter** uses them to clean up exhaust gases. In the IB Diploma you will quantify all of this with the **Maxwell–Boltzmann** distribution and **activation energy**, but the Year-9 picture — **more frequent, more energetic, more successful collisions** — already explains every rate change you will meet.