

Topic 1 (C)

Study then cover and check

1. **Choose:** select one lesson and briefly study its notes and worked example.
2. **Cover:** turn to the quiz. Cover the whole right-hand answer column with opaque paper and keep the notes out of sight.
3. **Attempt:** write numbered answers on separate paper or in your exercise book. Complete a small group of questions before checking.
4. **Check:** uncover the matching answers. Compare the scientific meaning, vocabulary and working. Equivalent correct wording is acceptable.
5. **Mark:** use C for correct, P for partly correct, and R for reattempt. A partly correct answer still needs improvement.
6. **Correct:** add missing reasoning or fix the first incorrect step. Cover the answer again and reattempt the question.
7. **Revisit:** return after a delay and include some earlier questions. Seeing and recognising an answer is not independent recall.

An example 18 minute session

3 minutes: understand the notes. **8 minutes:** answer a chosen group of questions with answers covered. **4 minutes:** check and correct. **3 minutes:** reattempt missed questions and record your next revisit.

Improve a partly correct answer

Question: A sodium ion has $Z = 11$, $A = 23$ and charge $+1$. Give its protons, neutrons and electrons.

First attempt: “11 protons, 12 neutrons, 11 electrons.” The proton and neutron counts are correct, so mark P.

Correction: “11 protons; $23 - 11 = 12$ neutrons; $11 - 1 = 10$ electrons. The $+1$ ion has lost one electron; its nucleus is unchanged.” Now hide this and repeat the question.

Plan a flexible revisit

For example, quiz again **the next day**, **later that week** and **the following week**. These are suggested intervals, not a fixed rule. Bring difficult questions back sooner; keep checking secure material periodically.

Record on your answer paper: **date | lesson | questions to retry | next revisit**. Use later results to adjust the plan. Work in a quiet place and keep distractions out of reach.



Elements, compounds and mixtures

Classify the substance

Atom: the smallest part of an element that can exist. An **element** contains one type of atom. A **molecule** contains atoms joined by chemical bonds; O_2 is a molecule of an element.

A **compound** contains different elements chemically bonded in fixed proportions, with properties different from its elements. A **mixture** contains substances not chemically bonded to each other; proportions can vary and components retain their properties.

Read formulae and balance equations

Symbols use a capital first letter and a lower-case second letter: Na, Cl, Mg. $MgCl_2$ is magnesium chloride, with $Mg:Cl = 1:2$. CO_2 is carbon dioxide. In $2H_2O$ there are four H atoms and two O atoms.

Worked example: magnesium + oxygen $\rightarrow$ magnesium oxide. Start with $Mg + O_2 \rightarrow MgO$. Put 2 before MgO to balance oxygen, then 2 before Mg: $2Mg + O_2 \rightarrow 2MgO$. Each side now has $Mg = 2$ and $O = 2$. Change coefficients, never formula subscripts.

Choose a physical separation

Method	Use and reason
Filtration	Insoluble solid from a liquid. Solid is trapped as residue; liquid passes through as filtrate.
Crystallisation	Dissolved solid from solution. Evaporate some solvent, cool, then filter and dry the crystals.
Simple distillation	Recover a solvent. Evaporate it, then cool and condense its vapour; dissolved solid remains.
Fractional distillation	Separate mixed liquids by different boiling points, using repeated evaporation and condensation.
Chromatography	Separate dissolved substances by different solubilities in the solvent and attractions to the stationary phase.

Apply the method

Sand, salt and water: filter, rinse and dry the sand residue. Crystallise salt from the filtrate. If water must be collected, distil the filtrate. Dissolved salt passes through filter paper. Physical separation makes no new substance; separating a compound into its elements requires a chemical reaction.



Elements, compounds and mixtures

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. Define an element and a compound.	1. An element has one type of atom. A compound has different elements chemically bonded in fixed proportions.
2. Why is pure O ₂ an element although its atoms are bonded?	2. Both atoms are oxygen. A compound must contain different elements.
3. Key: ● and ○ are different atoms; a line is a bond. Classify this sample: ●-○ ●-○ ○-○ Explain your answer.	3. Mixture: compound particles (●-○) and molecules of an element (○-○) are present.
4. Give two differences between a compound and a mixture.	4. Compound: bonded elements, fixed proportions. Mixture: components not bonded to each other, variable proportions. Compounds require chemical separation; mixtures can be separated physically.
5. Name MgCl ₂ . State Mg:Cl and the H and O atom totals in 2H ₂ O.	5. Magnesium chloride; Mg:Cl = 1:2. 2H ₂ O contains H = 4 and O = 2.
6. Write a word equation for hydrogen reacting with oxygen to form water. Balance H ₂ + O ₂ → H ₂ O.	6. hydrogen + oxygen → water 2H ₂ + O ₂ → 2H ₂ O Both sides: H = 4, O = 2.
7. Why is changing H ₂ O to H ₂ O ₂ an incorrect way to balance the water equation?	7. It changes water into hydrogen peroxide, a different substance. Use coefficients to change amounts.
8. How would you obtain dry sand and salt from sand, salt and water? Explain each separation.	8. Filter: insoluble sand is trapped; rinse and dry it. Evaporate some water from the filtrate, cool to crystallise salt, then filter and dry the crystals.
9. How would you collect pure water from salt solution? Why does filtration fail?	9. Simple distillation: boil water and condense the vapour. Salt stays behind. Dissolved salt passes through a filter with water.
10. Name the method for separating mixed liquids with different boiling points, then the method for separating soluble ink dyes.	10. Fractional distillation: different boiling points. Chromatography: different attractions to the stationary phase and solubilities in the solvent.



Developing the atomic model

Models explain evidence

A **scientific model** is a simplified representation used to explain observations and make predictions. It is not a photograph. An **observation** is what is detected; an **inference** is a conclusion drawn from evidence.

New evidence can support, refine or challenge a model. A replacement must explain the important evidence the earlier model could not explain.

How the model developed

Stage	Idea and evidence
Solid spheres	Atoms were thought to be tiny, indivisible spheres.
Thomson	Electron discovery showed atoms contain smaller negative particles. His plum-pudding model put electrons in a sphere of spread-out positive charge.
Rutherford	Alpha scattering supported a tiny, dense, positive nucleus surrounded by mostly empty space, with electrons outside it.
Bohr	Electrons occupy specific energy levels at particular distances from the nucleus. His calculations agreed with observations.
Protons and neutrons	Later work identified units of positive nuclear charge as protons. Chadwick provided evidence for neutral neutrons in the nucleus.

Read the scattering evidence

Observation at thin gold foil	Inference about atoms
Most alpha particles passed straight through.	Most of the atom is empty space.
Some positive alpha particles were deflected.	Concentrated positive charge repels them: like charges repel.
Very few were deflected backwards.	Nearly all mass and positive charge are in a tiny, dense nucleus.

Worked explanation of a model change

In plum pudding, positive charge and mass were spread out. This could not explain the rare, very large alpha-particle deflections. Concentrating charge and mass in a tiny nucleus explained those observations, so the nuclear model replaced plum pudding.

Avoid the error: Rutherford's team inferred a nucleus from detected particle paths. They did not directly see it, and most alpha particles did not bounce back.



Developing the atomic model

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. What is a scientific model?	1. A simplified representation that explains observations and supports predictions.
2. Put these in order: nuclear model; solid spheres; energy levels; plum pudding.	2. Solid spheres → plum pudding → nuclear model → electrons in specific energy levels.
3. How did electron discovery challenge the solid-sphere model?	3. Electrons are smaller particles inside atoms, so atoms cannot be indivisible spheres.
4. Describe charge in the plum-pudding model.	4. Negative electrons are embedded in spread-out positive charge. The charges balance in a neutral atom.
5. Most alpha particles passed through thin gold foil. What does this imply?	5. Most of an atom is empty space, so most particles do not encounter the concentrated central region.
6. Why did deflection of positive alpha particles suggest a positive nucleus?	6. Like charges repel. A concentrated positive region repelled the positive alpha particles.
7. Very few alpha particles were deflected backwards. What did this imply, and why did it challenge plum pudding?	7. Mass and positive charge are concentrated in a tiny, dense nucleus. Spread-out charge could not explain such large deflections.
8. Compare the plum-pudding and nuclear models. Give a similarity and a difference.	8. Both contain electrons and positive charge. Plum pudding spreads mass and positive charge out; the nuclear model concentrates them in a tiny nucleus.
9. What did Bohr add? What did Chadwick establish?	9. Bohr: electrons occupy specific energy levels. Chadwick: evidence for neutrons in the nucleus.
10. Correct: "Rutherford saw a nucleus and changed the model because he preferred it."	10. The nucleus was inferred from scattering paths. The model changed because new evidence could not be explained by plum pudding.
11. Revisit Lesson 1: Why can filtration separate sand from water but not water into its elements?	11. Sand is an insoluble part of a mixture. Water is a compound; separating it into its elements requires a chemical reaction.



Subatomic particles and atomic number

The nuclear atom

A tiny nucleus contains protons and neutrons. Electrons occupy energy levels around it. Most of the atom is empty space; almost all its mass is in the nucleus.

Particle properties

Particle	Location	Relative charge	Relative mass
Proton	Nucleus	+1	1
Neutron	Nucleus	0	1
Electron	Energy levels	-1	About 1/2000

Use atomic and mass numbers

Atomic number, Z = protons. It identifies the element. **Mass number, A = protons + neutrons.** Therefore **neutrons = $A - Z$.** Electron mass is negligible in mass-number calculations.

In nuclear notation, ${}^{23}_{11}\text{Na}$ has mass number 23 upper left and atomic number 11 lower left. A neutral atom has electrons = protons: equal opposite charges cancel.

An ion is charged because electrons have been gained or lost. Positive ions lose electrons; negative ions gain them. Proton and neutron counts stay unchanged when an ion forms.

Worked particle counts

Species	Protons	Neutrons	Electrons
Na: $Z = 11$, $A = 23$	11	$23 - 11 = 12$	11
Na^+ : $Z = 11$, $A = 23$	11	$23 - 11 = 12$	$11 - 1 = 10$
Cl^- : $Z = 17$, $A = 35$	17	$35 - 17 = 18$	$17 + 1 = 18$

Check the charge and scale

Charge check: protons - electrons = ionic charge. Na^+ has $11 - 10 = +1$. Changing electron number does not change the element.

Atom radius $\approx 0.1 \text{ nm} = 1 \times 10^{-10} \text{ m}$. Nucleus radius is of order 10^{-14} m , typically less than one ten-thousandth of the atom's radius. Atom diagrams are not to scale.

Worked ratio using approximate radii: $10^{-10} \div 10^{-14} = 10^4 = 10,000$. The radii differ by about four orders of magnitude. A negative exponent means a small positive distance, not a negative distance.



Lesson 3 | Self quiz

Subatomic particles and atomic number

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. State the relative charge and mass of each subatomic particle.	1. Proton: +1, mass 1. Neutron: 0, mass 1. Electron: -1, mass about 1/2000, treated as negligible here.
2. Where are protons, neutrons and electrons found? Why is most mass in the nucleus?	2. Protons and neutrons are in the nucleus; electrons are in surrounding energy levels. Protons and neutrons contribute almost all the mass.
3. Define atomic number and mass number.	3. Atomic number = number of protons. Mass number = protons + neutrons.
4. A neutral aluminium atom has $Z = 13$ and $A = 27$. Find protons, neutrons and electrons.	4. Protons = 13. Neutrons = $27 - 13 = 14$. Electrons = 13 because the atom is neutral.
5. A calcium ion has $Z = 20$, $A = 40$ and charge $2+$. Find all three particle counts.	5. Protons = 20. Neutrons = $40 - 20 = 20$. Electrons = $20 - 2 = 18$: two were lost.
6. An oxygen ion has $Z = 8$, $A = 18$ and charge $2-$. Find all three particle counts.	6. Protons = 8. Neutrons = $18 - 8 = 10$. Electrons = $8 + 2 = 10$: two were gained.
7. A sulfur species has 16 protons, 16 neutrons and 18 electrons. Give its mass number and ion formula.	7. Mass number = $16 + 16 = 32$. Charge = $16 - 18 = -2$. Ion: S^{2-} .
8. Why does an atom remain the same element when it forms an ion?	8. Its proton number, which identifies the element, is unchanged. Only electron number changes.
9. Convert 0.1 nm to metres. Use $1 \text{ nm} = 10^{-9} \text{ m}$.	9. $0.1 \times 10^{-9} \text{ m} = 1 \times 10^{-10} \text{ m}$.
10. Atom radius = 10^{-10} m ; nucleus radius $\approx 10^{-14} \text{ m}$. Calculate their radius ratio.	10. Atom $\div$ nucleus = $10^{-10} \div 10^{-14} = 10^4 = 10,000$. About four orders of magnitude.
11. Revisit Lesson 2: Which scattering observation suggested mostly empty space?	11. Most alpha particles passed straight through the thin gold foil.



Isotopes and relative atomic mass

Same element with different masses

Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons. Different neutron numbers give different mass numbers.

Neutral isotopes have the same electron arrangement, so their chemical properties are very similar. Proton number determines identity; changing neutrons changes mass.

Compare an isotope pair

Neutral atom	Protons	Neutrons	Electrons
Chlorine-35, Z = 17	17	$35 - 17 = 18$	17
Chlorine-37, Z = 17	17	$37 - 17 = 20$	17

Relative atomic mass

Relative atomic mass, A_r , is the weighted mean mass of an element's atoms compared with $1/12$ of a carbon-12 atom's mass. It accounts for isotope abundance and has no units.

Abundance means the proportion of each isotope in a sample. A more abundant isotope contributes more to the mean. A_r is not normally the mass number of a single atom.

Method: multiply each isotope mass by its abundance, add the products, then divide by total abundance. Divide by 100 for percentages; for ratios, divide by the total number of parts.

Worked examples

Percentage sample: 75% mass 35 and 25% mass 37.

$$A_r = [(35 \times 75) + (37 \times 25)] \div 100$$

$$= (2625 + 925) \div 100 = \mathbf{35.5}.$$

Ratio sample: masses 10 and 11 occur in a 4:1 ratio.

$$A_r = [(10 \times 4) + (11 \times 1)] \div (4 + 1)$$

$$= 51 \div 5 = \mathbf{10.2}.$$

Three-isotope sample: 92% mass 28, 5% mass 29, 3% mass 30.

$$A_r = [(28 \times 92) + (29 \times 5) + (30 \times 3)] \div 100$$

$$= (2576 + 145 + 90) \div 100 = \mathbf{28.11}.$$

Check before accepting an answer

The mean lies between the isotope masses. For two isotopes, it is nearer the more abundant one. Do not divide a percentage-weighted total by the number of isotopes. An ordinary mean works only when abundances are equal.



Isotopes and relative atomic mass

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. Define isotopes.	1. Atoms of the same element with the same proton number but different neutron numbers.
2. Carbon-12 and carbon-14 both have $Z = 6$. Give their proton and neutron counts and write their nuclear symbols.	2. Both: 6 protons. Carbon-12: $12 - 6 = 6$ neutrons; $^{12}_6\text{C}$. Carbon-14: $14 - 6 = 8$ neutrons; $^{14}_6\text{C}$.
3. Why do neutral isotopes have similar chemical properties but different masses?	3. They have the same electron arrangement, which controls reactions. Different neutron numbers change their masses.
4. Define A_r and state whether it has units.	4. The isotope-weighted mean atomic mass relative to 1/12 of a carbon-12 atom's mass. It has no units.
5. A sample is 60% isotope mass 24 and 40% mass 26. Calculate A_r and check it.	5. $[(24 \times 60) + (26 \times 40)] \div 100 = (1440 + 1040) \div 100 = 24.8$. It is between 24 and 26, nearer 24.
6. Masses 63 and 65 occur in the ratio 7:3. Calculate A_r .	6. $[(63 \times 7) + (65 \times 3)] \div 10 = (441 + 195) \div 10 = 63.6$. Divide by 7 + 3 = 10 parts.
7. A sample has masses 54, 56, 57 at 6%, 92%, 2%. Calculate A_r .	7. $[(54 \times 6) + (56 \times 92) + (57 \times 2)] \div 100 = (324 + 5152 + 114) \div 100 = 55.90$.
8. For 80% mass 24 and 20% mass 26, a pupil uses $(24 + 26) \div 2 = 25$. Correct the method.	8. Unequal abundances need weighting. $[(24 \times 80) + (26 \times 20)] \div 100 = 2440 \div 100 = 24.4$.
9. Only masses 35 and 37 occur in a sample. Why is $A_r = 38.2$ impossible?	9. A weighted mean must lie between 35 and 37. 38.2 exceeds both masses.
10. Revisit Lesson 3: A chloride ion has $Z = 17$, $A = 37$ and charge -1 . Find its neutrons and electrons.	10. Neutrons = $37 - 17 = 20$. Electrons = $17 + 1 = 18$. Its 17 protons stay unchanged.



The periodic table and Mendeleev

Read the modern periodic table

Elements are arranged in increasing **atomic number**, the number of protons. A **group** is a vertical column; a **period** is a horizontal row. Elements in a group have similar chemical properties.

Periodic means similar properties recur at regular intervals. For example, lithium, sodium and potassium occupy corresponding positions in successive periods and have similar reactions.

First 20 elements

Period	1	2	3	4	5	6	7	0
1	1 H							2 He
2	3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca						

Columns are AQA GCSE Groups 1–7 and 0. Each entry gives atomic number and symbol. Blank cells are outside this first-20 extract.

Why early tables had problems

Before subatomic particles were known, scientists arranged elements by **atomic weight**. Many elements were undiscovered, so the tables were incomplete. Strict weight order sometimes put elements with different properties in the same group.

Mendeleev used patterns and predictions

Mendeleev arranged mainly by atomic weight but sometimes changed the order to keep similar properties together. He left deliberate gaps for undiscovered elements and predicted their properties.

Worked explanation: leaving a gap alone did not show the table was correct. When later discoveries matched properties predicted beforehand, this provided independent evidence supporting the table. A prediction could have failed, so a successful test increased confidence.

Why mass order is not always correct

Isotope masses and abundances affect relative atomic mass without changing proton number. This later explained why atomic-weight order does not always match atomic-number order. The modern table uses proton number.

Avoid the errors: Mendeleev did not use proton number or discover every missing element. His gaps were predictions, not accidental omissions.



The periodic table and Mendeleev

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. What determines the order of elements in the modern periodic table?	1. Increasing atomic number, which is the number of protons.
2. Define a group, a period and “periodic”.	2. Group: vertical column. Period: horizontal row. Periodic: similar properties recur at regular intervals.
3. Li, Na and K are in Group 1. Predict whether their chemical properties are similar.	3. They have similar chemical properties because they belong to the same group.
4. An element is in the third horizontal row and the column labelled Group 7. State its period and group.	4. Period 3 and Group 7. Rows give periods; columns give groups.
5. Give two problems with early tables arranged strictly by atomic weight.	5. Undiscovered elements made them incomplete. Strict weight order sometimes grouped elements with different properties or separated similar elements.
6. How did Mendeleev address each of those problems?	6. He left gaps for undiscovered elements and changed the order where needed to keep similar properties together.
7. Hypothetical A and C form chlorides ACl and CCl and react vigorously with water. B forms BCl_2 and reacts slowly. Which pair belongs together? Give evidence.	7. A and C: both have one chlorine per atom in their chloride and both react vigorously with water. B differs in both properties.
8. Why did later discoveries support Mendeleev’s table?	8. Discovered elements had properties close to predictions made before their discovery. This independently tested and supported his pattern.
9. Why can relative atomic mass order differ from proton-number order?	9. Isotope masses and abundances affect relative atomic mass. It does not always increase as proton number increases.
10. Revisit Lesson 4: Atoms X and Y each have 17 protons but have 18 and 20 neutrons. Are they isotopes? Explain.	10. Yes. They are the same element because proton numbers match, but different neutron numbers make them isotopes.



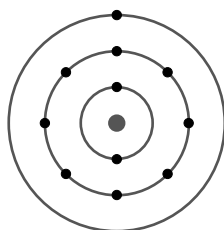
Metals, non-metals and electron arrangement

Fill shells from the inside out

In a neutral atom, electrons = atomic number. Electrons occupy the lowest available energy levels first. For the **first 20 elements**, use shell populations up to 2, then 8, then 8, before starting the fourth shell. This is not a universal rule for all atoms.

Worked example: sodium, $Z = 11$. Put 2 electrons in the first shell, 8 in the second and the remaining 1 in the third: **2,8,1**. Check: $2 + 8 + 1 = 11$.

Sodium electron shells



Dots represent electrons; the centre is the nucleus. Not to scale.

Link electrons to position

Period = occupied shells. In Groups 1–7, **group number = outer electrons**. Group 0 has a full outer shell: 2 for helium, 8 for the others. Do not apply these group rules to transition metals.

Calcium, $Z = 20$: **2,8,8,2**. Four occupied shells give Period 4; two outer electrons give Group 2. Similar outer-electron arrangements explain similar reactions within a group.

Compare metals and non-metals

Metals	Non-metals
Mainly left and lower in the table; usually few outer electrons. Form positive ions by losing electrons.	Mainly right and upper in the table. Do not form positive ions; many gain or share electrons.
Generally conduct heat and electricity and are malleable.	Generally poor conductors and brittle when solid. There are exceptions to physical trends.

Group 0

Noble gases are very unreactive because their outer shells are full and stable. They have little tendency to gain, lose or share electrons and exist as single atoms: **monatomic**.

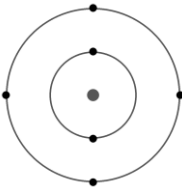
Boiling points increase down Group 0 as relative atomic mass increases: neon $\approx -246^\circ\text{C}$; argon $\approx -186^\circ\text{C}$; krypton $\approx -153^\circ\text{C}$. A less negative temperature is higher.

Avoid the error: helium has two electrons, not none. Two fill its first shell, making it stable.



Metals, non-metals and electron arrangement

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. Draw the electron shells of carbon, $Z = 6$. Show all electrons and give the number arrangement.	1. 2,4: 2 electrons in the inner shell, 4 in the outer shell.  Dots = electrons; centre = nucleus. Not to scale.
2. Write arrangements for chlorine, $Z = 17$, and calcium, $Z = 20$.	2. Cl: 2,8,7. Ca: 2,8,8,2. Totals are 17 and 20 electrons.
3. An atom has arrangement 2,8,3. State its atomic number, group and period.	3. $Z = 2 + 8 + 3 = 13$. Group 3: three outer electrons. Period 3: three occupied shells.
4. Why do lithium (2,1) and potassium (2,8,8,1) have similar reactions?	4. Both have one outer electron and form $1+$ ions by losing it.
5. Mg has arrangement 2,8,2. Give its usual ion and the ion's arrangement.	5. Mg^{2+} , arrangement 2,8. It loses two electrons; forming positive ions is metal behaviour.
6. Compare metals and non-metals: position, ion formation and two physical properties.	6. Metals: mainly left/lower, form positive ions, usually conducting and malleable. Non-metals: mainly right/upper, do not form positive ions, usually poor conductors and brittle when solid.
7. Why is helium, arrangement 2, unreactive? Is it monatomic or molecular?	7. Its first shell is full and stable with two electrons. It is monatomic: separate atoms.
8. Neon boils at -246°C ; argon at -186°C . Krypton is below argon. Predict whether its boiling point is higher or lower than -186°C .	8. Higher: boiling point increases down Group 0 as relative atomic mass increases.
9. Revisit Lesson 5: Why did Mendeleev sometimes change atomic-weight order?	9. To keep elements with similar chemical properties together.



Lesson 7 | Key knowledge

Group 1

Alkali metals and their ions

Lithium, sodium and potassium each have one outer electron: **2,1; 2,8,1; 2,8,8,1**. They react by losing it and form 1+ ions, explaining their similar reactions. Hydrogen is a non-metal, not an alkali metal.

Worked ion example: $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$. Sodium changes from 2,8,1 to 2,8. Its 11 protons remain; 10 electrons give charge $11 - 10 = +1$.

Alkali metals are soft and have relatively low melting points. Li, Na and K have low densities and float on water. They must be protected from air and moisture.

Reactions with water

Metal	Typical observations
Lithium	Floats and fizzes steadily; moves slowly.
Sodium	Fizzes rapidly; often melts into a moving ball.
Potassium	Reacts very vigorously; often burns with a lilac flame.

Products and equations

metal + water $\rightarrow$ metal hydroxide + hydrogen

The hydroxide solution is alkaline. Fizzing shows gas production but does not alone identify the gas.

Worked balance: $\text{Na} + \text{H}_2\text{O} \rightarrow \text{NaOH} + \text{H}_2$ becomes

$2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2$. Both sides: Na = 2, O = 2, H = 4. Lithium and potassium follow the same water-reaction pattern.

Oxygen and chlorine

Li, Na and K react with oxygen and tarnish in air. Lithium forms Li_2O : **$4\text{Li} + \text{O}_2 \rightarrow 2\text{Li}_2\text{O}$** . Oxygen products depend on the metal and conditions; do not assume every product is M_2O .

All three react with chlorine to form white chlorides, LiCl, NaCl and KCl. Example: **$2\text{K} + \text{Cl}_2 \rightarrow 2\text{KCl}$** . A 1+ metal ion balances a 1- chloride ion.

Explain and predict reactivity

Down Group 1, more occupied shells increase the outer electron's distance from the nucleus and shielding by inner electrons. Nuclear attraction is weaker, so the electron is lost more easily. **Reactivity increases.**

Rubidium is below potassium: predict more vigorous reactions and Rb^+ ions. Its chloride is RbCl. These are predictions to explain on paper; alkali-metal demonstrations are teacher-only.



Lesson 7 | Self quiz

Group 1

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. Why do Li, Na and K have similar reactions and form 1+ ions?	1. Each has one outer electron. Losing it leaves one more proton than electrons.
2. Write the electron-loss equation for potassium and give the ion's arrangement. K atoms are 2,8,8,1.	2. $K \rightarrow K^+ + e^-$. K^+ has arrangement 2,8,8. The nucleus is unchanged.
3. Describe typical observations when Li, Na and K react with water.	3. Li floats and fizzes steadily. Na fizzes rapidly, often melting into a moving ball. K is very vigorous, often with a lilac flame.
4. Name the products of sodium reacting with water. What does a purple universal indicator show afterwards?	4. Sodium hydroxide and hydrogen. Purple indicator shows the solution is alkaline.
5. Balance $Na + H_2O \rightarrow NaOH + H_2$. Give the atom totals.	5. $2Na + 2H_2O \rightarrow 2NaOH + H_2$. Each side: Na = 2, O = 2, H = 4.
6. Lithium forms Li_2O with oxygen. Balance $Li + O_2 \rightarrow Li_2O$ and describe fresh alkali-metal surfaces in air.	6. $4Li + O_2 \rightarrow 2Li_2O$. Both sides: Li = 4, O = 2. Fresh surfaces tarnish as they react.
7. Balance $K + Cl_2 \rightarrow KCl$. Name the product and describe its colour.	7. $2K + Cl_2 \rightarrow 2KCl$. Each side: K = 2, Cl = 2. Potassium chloride is a white solid.
8. Explain why potassium is more reactive than sodium.	8. More occupied shells increase distance and shielding. Attraction to its outer electron is weaker, so potassium loses it more easily.
9. Rubidium lies below potassium. Predict its water reaction, ion and chloride formulae.	9. More vigorous, producing rubidium hydroxide and hydrogen. Ion Rb^+ ; chloride $RbCl$. Its outer electron is lost more easily.
10. Revisit Lesson 6: Compare Group 0 reactivity with Group 1 and explain the difference.	10. Group 0 is very unreactive: full stable outer shells. Group 1 reacts readily by losing its single outer electron.



Group 7

Halogens and halide ions

Halogens are non-metals with seven outer electrons. Familiar halogens form diatomic molecules: F_2 , Cl_2 , Br_2 , I_2 . A halogen atom gains one electron to form a 1- **halide ion**, for example $Cl + e^- \rightarrow Cl^-$.

Cl_2 is a neutral molecule containing two chlorine atoms. Cl^- is a single chloride ion. All halogens have similar reactions because of their outer-electron arrangements.

Physical trends

At room temperature	Down Group 7
Chlorine: pale green gas Bromine: red-brown liquid Iodine: grey-black solid	Relative molecular mass, melting point and boiling point increase. Iodine forms purple vapour when heated.

Reactivity and compounds

Reactivity decreases down Group 7. More shells increase distance and shielding, weakening attraction for an incoming electron. Gaining an electron becomes less easy.

With metals, halogens form ionic compounds containing halide ions: $2Na + Cl_2 \rightarrow 2NaCl$. With non-metals they can form covalent molecules: $H_2 + Cl_2 \rightarrow 2HCl$, hydrogen chloride.

Predict displacement

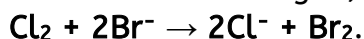
A more reactive halogen displaces a less reactive one from an aqueous halide. Order: **chlorine > bromine > iodine**.

Worked example: chlorine + potassium bromide $\rightarrow$ potassium chloride + bromine.
 $Cl_2 + 2KBr \rightarrow 2KCl + Br_2$. Chlorine gains electrons from bromide ions. Bromine forms; its aqueous solution is orange.

Iodine cannot displace bromine from potassium bromide: iodine is less reactive. The same halogen as the halide also gives no displacement.

Higher Tier electron transfer

K^+ ions are unchanged, so omit them from the ionic equation:



Reduction is gain: $Cl_2 + 2e^- \rightarrow 2Cl^-$.

Oxidation is loss: $2Br^- \rightarrow Br_2 + 2e^-$.

Add the half equations and cancel the two electrons. Each balances atoms and total charge.

Avoid the error: a molecule X_2 needs two electrons to form two X^- ions. Group 7 reactivity decreases because electron gain becomes harder, unlike electron loss in Group 1.



Lesson 8 | Self quiz

Group 7

Cover the answers. Keep your notes out of sight. Write your numbered answers on separate paper, then check and correct them.

Questions	Answers
1. Distinguish Br ₂ and Br ⁻ by particle type and charge.	1. Br ₂ : neutral molecule of two bromine atoms. Br ⁻ : bromide ion, charge -1, formed by gaining one electron.
2. State the room-temperature states of chlorine, bromine and iodine. Give the mass, melting-point and boiling-point trends down the group.	2. Chlorine gas; bromine liquid; iodine solid. Relative molecular mass, melting point and boiling point increase.
3. Explain why iodine is less reactive than chlorine.	3. More shells increase distance and shielding, weakening attraction for an incoming electron. Iodine gains an electron less readily.
4. Compare the compounds formed by chlorine with sodium and with hydrogen.	4. Sodium chloride contains Na ⁺ and Cl ⁻ ions. Hydrogen chloride, HCl, consists of covalent molecules.
5. Chlorine solution is added to aqueous KBr. Predict the products and write the balanced symbol equation.	5. Potassium chloride and bromine form: chlorine is more reactive. $\text{Cl}_2 + 2\text{KBr} \rightarrow 2\text{KCl} + \text{Br}_2$
6. Will iodine solution displace chlorine from aqueous KCl? Explain.	6. No. Iodine is less reactive than chlorine and cannot displace it from chloride ions.
7. For $\text{Cl}_2 + 2\text{KBr} \rightarrow 2\text{KCl} + \text{Br}_2$, write the ionic equation. Which ions are unchanged?	7. $\text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2$. K ⁺ ions are unchanged and omitted.
8. Write both half equations for $\text{Cl}_2 + 2\text{Br}^- \rightarrow 2\text{Cl}^- + \text{Br}_2$. Identify oxidation and reduction.	8. Reduction: $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$. Oxidation: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$. Each side of each half equation has total charge -2.
9. X is chlorine, bromine or iodine. It displaces iodine from KI but not bromine from KBr. Identify X.	9. Bromine. Chlorine would displace both; iodine would displace neither.
10. Revisit Lesson 7: Why does weaker attraction increase Group 1 reactivity but decrease Group 7 reactivity?	10. Group 1 loses an electron: weaker attraction makes loss easier. Group 7 gains an electron: weaker attraction makes gain harder.



