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NCERT Solutions for Class 10 Science Chapter 1 — Chemical Reactions and Equations

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These **NCERT Solutions for Class 10 Science Chapter 1** cover every intext and exercise question with concept-first explanations, worked steps, and common-error warnings for the 2026-27 session. Each answer explains the principle before the working, so you can apply the same idea to a fresh question in the exam.

Use the jump links below to reach any section quickly. Every equation is balanced and uses proper state symbols where the textbook requires them. You can also cross-check the activities and equations directly on the official NCERT textbook page for Chapter 1.

For deeper revision, pair this page with our Class 10 Science notes, the Class 10 notes hub, or the CBSE notes collection on this site.

Intext Solutions (Page 6): Chemical Equations Basics

Question 1: Why should a magnesium ribbon be cleaned before burning in air?

Magnesium is reactive enough to combine with oxygen in ordinary air, so a ribbon left on a shelf is already coated with a thin layer of **magnesium oxide**. That oxide layer is stable and does not burn — it sits between the flame and the fresh metal underneath.

Cleaning the ribbon with sandpaper removes this coating and exposes fresh magnesium so that the reaction you want to observe can actually start (Activity 1.1, p. 1).

Step 1: Magnesium reacts with atmospheric oxygen over time: $2Mg + O_2 \rightarrow 2MgO$.

Step 2: The MgO coating is dull and prevents uniform ignition.

Step 3: Rubbing with sandpaper strips this layer, exposing shiny magnesium metal for clean, uniform burning.

Common error: Students write that the ribbon is cleaned "to remove dust." The dust is irrelevant — the real reason is the magnesium oxide layer.

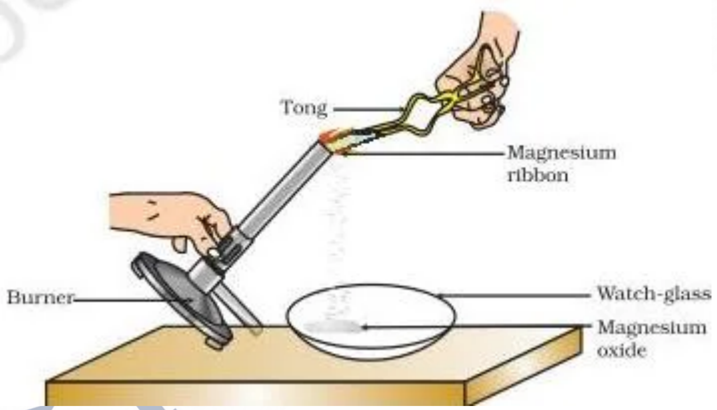

Figure 1.1 Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass. Source: NCERT

Question 2: Write balanced equations for (i) hydrogen + chlorine, (ii) barium chloride + aluminium sulphate, (iii) sodium + water

Each reaction combines or exchanges elements in fixed atomic ratios. Balancing places whole-number coefficients so that atoms of every element are equal on both sides — you never alter a formula's subscripts to force equality.

Part (i): Hydrogen and chlorine are both diatomic, so each starts as H_2 and Cl_2. They combine in a 1:1 ratio to give two molecules of HCl:

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

Part (ii): Barium chloride is $BaCl_2$; aluminium sulphate is $Al_2(SO_4)_3$. Exchange the ions: Ba^{2+} pairs with SO_4^{2-}, and Al^{3+} pairs with Cl^-. Three $BaCl_2$ units supply the six chloride ions that two $AlCl_3$ units need:

$$3BaCl_2(aq) + Al_2(SO_4)_3(aq) \rightarrow 3BaSO_4(s) + 2AlCl_3(aq)$$

Part (iii): Sodium is Na, water is H_2O. Sodium displaces hydrogen from water to form sodium hydroxide (NaOH) and hydrogen gas. Two Na atoms and two water molecules give two NaOH and one H_2, which balances hydrogen (4 on each side):

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

Common error: Writing Na_2O instead of NaOH — sodium reacts with water to form a hydroxide and hydrogen gas, not a simple oxide.

Question 3: Write balanced equations with state symbols for (i) barium chloride + sodium sulphate, (ii) sodium hydroxide + hydrochloric acid

State symbols make an equation informative: (s) for solid, (l) for liquid, (g) for gas, (aq) for aqueous solution. A precipitate — an insoluble solid formed when two solutions mix — must carry (s), not (aq), because it drops out of solution.

Part (i): Both reactants are in aqueous solution. Ba^{2+} meets SO_4^{2-} and forms insoluble barium sulphate, which precipitates as a white solid. Sodium and chloride remain dissolved:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$$

Part (ii): An acid-base neutralisation: H^+ and OH^- join to form water. Sodium chloride stays in solution:

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

Common error: Writing $BaSO_4(aq)$ — barium sulphate is the precipitate, so it must be (s).

Intext Solutions (Page 10): Whitewashing and Electrolysis

Question 1: A solution of substance X is used for whitewashing. Name X, write its formula, and its reaction with water.

The substance used for whitewashing is **calcium oxide** (CaO), also called *quick lime* (p. 6). When you add water to quick lime, it reacts vigorously to form *slaked lime* — calcium hydroxide, $Ca(OH)_2$ — and releases a large amount of heat.

This is the solution that is applied to walls; over two to three days it reacts with atmospheric CO_2 to form a shiny layer of calcium carbonate ($CaCO_3$), the same compound as marble (p. 7).

Step 1: X = calcium oxide (quick lime), formula CaO.

Step 2: Its vigorous reaction with water:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + Heat$$

Common error: Naming X as calcium hydroxide instead of calcium oxide. Slaked lime is the *product* of the water reaction, not the starting substance used for whitewashing.

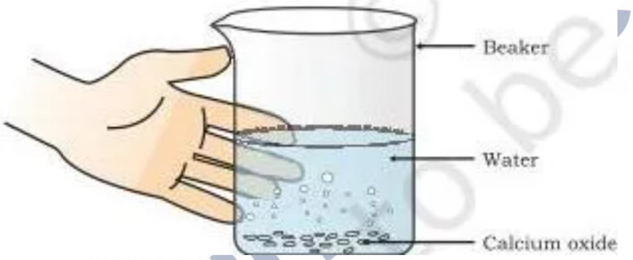

Figure 1.3 Formation of slaked lime by the reaction of calcium oxide with water. Source: NCERT

Question 2: Why is the amount of gas in one test tube in Activity 1.7 double the other? Name the gas.

Water has the formula H_2O — two hydrogen atoms for every one oxygen atom. Electrolytic decomposition splits water into its elemental gases in that same atomic ratio. So the test tube collecting hydrogen holds twice the volume of the one collecting oxygen (Activity 1.7, p. 9).

Step 1: Water molecule: H_2O — 2 H atoms and 1 O atom.

Step 2: Splitting two water molecules gives two H_2 and one O_2:

$$2H_2O(l) \xrightarrow{Electricity} 2H_2(g) + O_2(g)$$

Step 3: Hydrogen collects at the cathode (negative electrode), oxygen at the anode (positive electrode).

Common error: Swapping the electrodes — hydrogen appears at the cathode, oxygen at the anode, not the other way around.

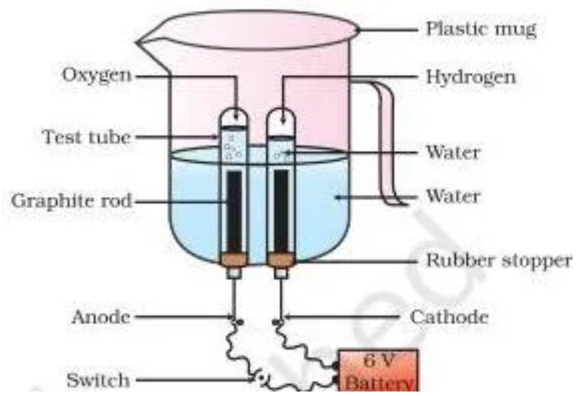

Figure 1.6 Electrolysis of water. Source: NCERT

Intext Solutions (Page 13): Displacement, Double Displacement, Redox

Question 1: Why does the colour of copper sulphate solution change when an iron nail is dipped in it?

A more reactive metal displaces a less reactive one from its salt solution. Iron is above copper in the reactivity series, so iron pulls copper out of copper sulphate. The blue colour of the Cu^{2+} solution fades as copper ions leave the solution, and the iron nail acquires a brown coating of deposited copper metal (Activity 1.9, p. 10).

$$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$$

Common error: Saying the colour changes because iron dissolves. The colour fades specifically because blue Cu^{2+} ions are removed from solution and replaced by pale green Fe^{2+} ions.

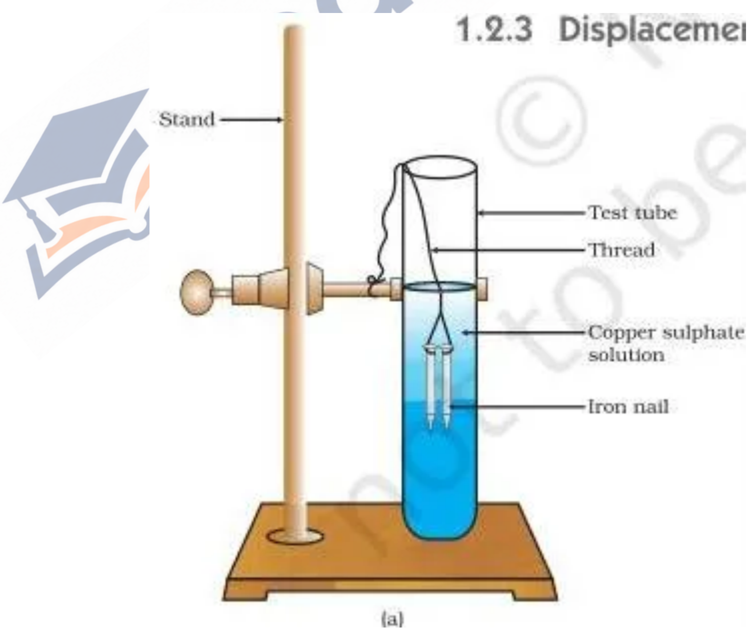

Figure 1.8(a) Iron nails dipped in copper sulphate solution. Source: NCERT

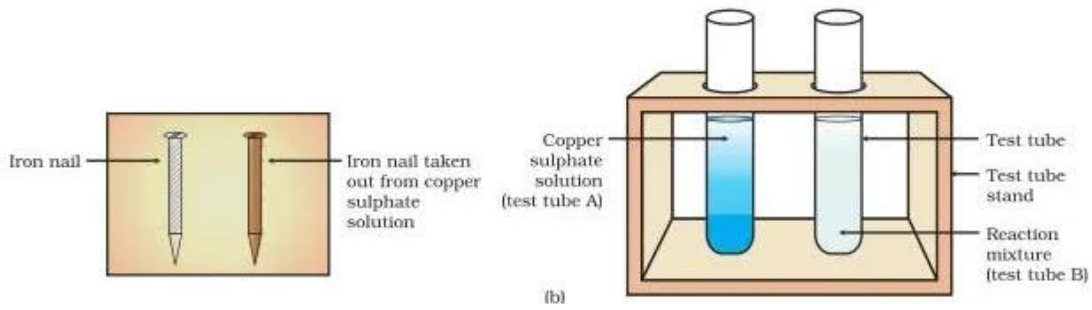

Figure 1.8(b) Iron nails and copper sulphate solutions compared before and after the experiment. Source: NCERT

Question 2: Give an example of a double displacement reaction other than the one in Activity 1.10.

In a double displacement reaction, two compounds in solution exchange ions — the positive ion of one compound pairs with the negative ion of the other, and vice versa. If one of the new combinations is insoluble, it drops out as a precipitate. The textbook's Activity 1.2 (p.

2) uses lead nitrate and potassium iodide; mixing these gives a yellow precipitate of lead iodide.

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

Common error: Writing this as a displacement reaction because metals are involved. It is double displacement because two whole ionic compounds swap partners — no free element is displaced.

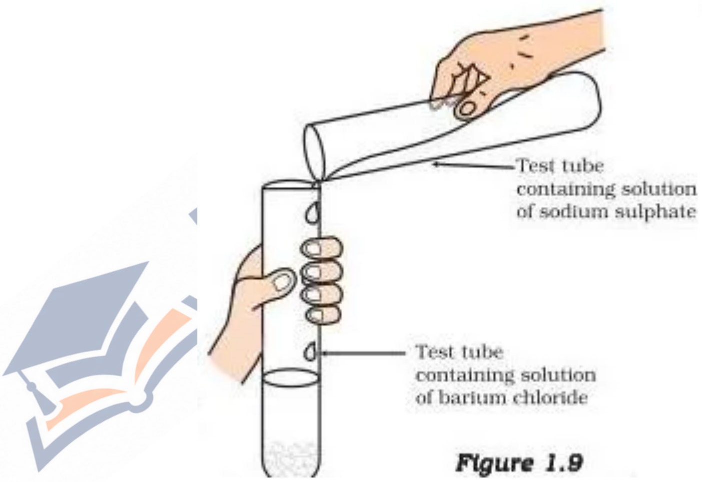

Figure 1.9 Formation of barium sulphate and sodium chloride. Source: NCERT

Question 3: Identify oxidised and reduced substances in (i) $4Na + O_2 \rightarrow 2Na_2O$ and (ii) $CuO + H_2 \rightarrow Cu + H_2O$

Oxidation and reduction are coupled: one substance gains oxygen (is oxidised) while another loses oxygen (is reduced). They always occur together, which is why such reactions are called **redox reactions** (p. 12).

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

Sodium gains oxygen to become Na_2O, so **sodium is oxidised**. Oxygen is the oxidising agent — it causes the oxidation and is itself reduced.

Part (ii): $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

Copper oxide loses its oxygen to become copper, so **CuO is reduced**. Hydrogen gains oxygen to become water, so **hydrogen is oxidised**.

Common error: Calling the oxidising agent "oxidised." The oxidising agent *causes* oxidation in the other substance but is itself *reduced* — keep these roles separate.

Exercise Q1–Q3: Redox Identification and Reaction Types

Exercise Q1: Which statements about $2PbO + C \rightarrow 2Pb + CO_2$ are incorrect?

Track oxygen to each substance separately. Lead oxide loses oxygen to become lead — so lead is reduced and lead oxide is reduced. Carbon gains oxygen to become carbon dioxide — so carbon is oxidised. CO_2 is a *product* of carbon's oxidation, not itself undergoing further oxidation.

- **Statement (a)** "Lead is getting reduced" — *correct* (Pb gains electrons, loses O).
- **Statement (b)** "Carbon dioxide is getting oxidised" — *incorrect*; CO_2 is the product, not a reactant being further oxidised.
- **Statement (c)** "Carbon is getting oxidised" — *correct*.
- **Statement (d)** "Lead oxide is getting reduced" — *correct*.

> **Final answer:** The incorrect statements are (a) and (b), so the answer option is **(i)**.

Common error: Marking statement (a) as incorrect because the wording feels tricky. Statement (a) is factually correct — lead *is* getting reduced. The only incorrect statement among the first two is (b).

Exercise Q2: $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$ is an example of which reaction type?

Aluminium is more reactive than iron, so it displaces iron from iron oxide. Because a single element (aluminium) takes the place of another element (iron) in a compound, the textbook classifies this as a **displacement reaction** (p. 10). It is also a redox reaction, but among the given choices, displacement is the classification the board expects.

> **Final answer:** (d) displacement reaction.

Common error: Calling this a combination reaction because aluminium oxide forms. Aluminium oxide forms by combination, but the *overall* reaction is displacement — aluminium replaces iron.

Exercise Q3: What happens when dilute HCl is added to iron filings?

Iron is an active metal. It displaces hydrogen from dilute hydrochloric acid, producing iron(II) chloride and hydrogen gas. Because HCl is dilute and iron forms Fe^{2+} here, the correct salt is $FeCl_2$, not $FeCl_3$.

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$

> **Final answer:** (a) Hydrogen gas and iron chloride are produced.

Common error: Writing $FeCl_3$ — iron forms Fe^{2+} with dilute HCl, so the product is $FeCl_2$. Students also sometimes think chlorine gas (Cl_2) is produced; it is not.

Exercise Q4–Q5: Balancing and Translating Statements

Exercise Q4: What is a balanced chemical equation? Why should equations be balanced?

A **balanced chemical equation** has equal numbers of atoms of each element on both sides of the arrow (p. 3). Equations must be balanced because of the **Law of Conservation of Mass**: mass can neither be created nor destroyed in a chemical reaction. If the atom counts were unequal, the reaction would imply that atoms appeared or vanished — which is impossible.

Common error: Defining a balanced equation as "same number of molecules on both sides." Molecules need not be equal — atoms of each element must be equal.

Exercise Q5: Translate four statements into balanced chemical equations

Part (a) Hydrogen + nitrogen → ammonia. Nitrogen and hydrogen are both diatomic. Ammonia is NH_3, so three H_2 molecules supply the six hydrogen atoms needed by two NH_3:

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

Part (b) Hydrogen sulphide burns in air to give water and sulphur dioxide. H_2S provides H and S; O_2 supplies oxygen for both water and SO_2. Balance S first, then H, then O, giving:

$$2H_2S(g) + 3O_2(g) \rightarrow 2SO_2(g) + 2H_2O(l)$$

Part (c) Barium chloride + aluminium sulphate → aluminium chloride + barium sulphate precipitate. Three $BaCl_2$ units supply six chlorides for two $AlCl_3$; the sulphate groups transfer to barium:

$$3BaCl_2(aq) + Al_2(SO_4)_3(aq) \rightarrow 2AlCl_3(aq) + 3BaSO_4(s)$$

Part (d) Potassium + water → potassium hydroxide + hydrogen. Potassium behaves like sodium: it forms a hydroxide and releases hydrogen gas:

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

Common error in (b): Forgetting oxygen is diatomic — you need a coefficient of 3 on O_2 to supply six oxygen atoms. **Common error in (d):** Writing K_2O instead of KOH — potassium in water gives a hydroxide and hydrogen, not a simple oxide.

Exercise Q6–Q7: Balancing and Writing Equations

Exercise Q6: Balance four equations

Part (a): Calcium has two nitrate groups in $Ca(NO_3)_2$, so we need two HNO_3 on the left, which gives two water molecules:

$$2HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2H_2O$$

Part (b): Sodium sulphate (Na_2SO_4) needs two sodium atoms, so place 2 before NaOH, which gives two waters:

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

Part (c): Each ion already matches one-to-one, so the equation is already balanced:

$$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$$

Part (d): $BaSO_4$ takes one sulphate group. The two chlorines in $BaCl_2$ pair with two hydrogens in 2HCl:

$$BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$$

Common error in (c): Adding unnecessary coefficients (e.g. a 2 on every term) to an equation that is already balanced.

Exercise Q7: Write balanced equations for four reactions

Part (a): Calcium hydroxide + carbon dioxide → calcium carbonate + water. This is the whitewashing follow-up reaction (p. 7). One $Ca(OH)_2$ has two OH groups; one CO_2 supplies the carbonate; the leftover H and O form one water:

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

Part (b): Zinc + silver nitrate → zinc nitrate + silver. Zinc forms Zn^{2+}, so it needs two $AgNO_3$ to balance charge and nitrate count:

$$Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$$

Part (c): Aluminium + copper chloride → aluminium chloride + copper. Aluminium forms Al^{3+}; two $AlCl_3$ need six chlorides, so three $CuCl_2$ supply them, giving three copper atoms:

$$2Al(s) + 3CuCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Cu(s)$$

Part (d): Barium chloride + potassium sulphate → barium sulphate + potassium chloride. One Ba^{2+} pairs with one SO_4^{2-}; two K^+ need two Cl^-, which one $BaCl_2$ provides:

$$BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2KCl(aq)$$

Common error in (b) and (d): Forgetting to multiply $AgNO_3$ by 2 or KCl by 2, which leaves the anion or cation count unbalanced.

Exercise Q8: Balanced Equations with Reaction-Type Identification

Exercise Q8: Write balanced equations and identify the reaction type for four reactions

Part (a): Potassium bromide + barium iodide → potassium iodide + barium bromide. Two ionic compounds exchange partners — K^+ swaps with Ba^{2+}, and Br^- swaps with I^-. This is a **double displacement** reaction.

$$2KBr(aq) + BaI_2(aq) \rightarrow 2KI(aq) + BaBr_2(s)$$

Part (b): Zinc carbonate → zinc oxide + carbon dioxide. A single reactant breaks into two simpler products on heating — a **thermal decomposition** reaction (p. 8).

$$ZnCO_3(s) \xrightarrow{Heat} ZnO(s) + CO_2(g)$$

Figure 1.4 Correct way of heating the boiling tube containing crystals of ferrous sulphate and of smelling the odour. Source: NCERT

Part (c): Hydrogen + chlorine → hydrogen chloride. Two elements combine to form a single compound — a **combination** reaction (p. 6).

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

Part (d): Magnesium + hydrochloric acid → magnesium chloride + hydrogen. Magnesium, a metal, displaces hydrogen from the acid — a **displacement** reaction (p. 10).

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

Common error: Misclassifying (a) as a displacement reaction because metal symbols appear. It is double displacement — two ionic compounds swap partners, and no free element is displaced from the start.

Exercise Q9–Q10: Exothermic, Endothermic, Respiration

Exercise Q9: What are exothermic and endothermic reactions? Give examples.

Reactions either release energy to the surroundings or absorb energy from them. The difference is the direction of energy flow when bonds break and form.

- **Exothermic reaction:** Heat is released along with the products — the mixture warms up. Examples: burning of natural gas $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ (p. 7); quick lime + water $CaO + H_2O \rightarrow Ca(OH)_2 + \text{Heat}$; respiration.
- **Endothermic reaction:** Heat or energy is absorbed — the mixture cools or needs a steady energy supply. Examples: thermal decomposition of limestone $CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$ (p. 8); silver chloride in sunlight $2AgCl \xrightarrow{\text{Sunlight}} 2Ag + Cl_2$ (p. 9); electrolysis of water.

Common error: Calling the decomposition of vegetable matter into compost endothermic — it releases heat, so it is *exothermic* (p. 7).

Exercise Q10: Why is respiration considered an exothermic reaction?

During digestion, food breaks down into simpler substances, including glucose. Inside our cells, glucose combines with oxygen and releases energy that the body uses to stay alive. Because energy flows *out* of the reaction into the body, respiration is exothermic (p. 7, eq. 1.18).

$$C_6H_{12}O_6(aq) + 6O_2(aq) \rightarrow 6CO_2(aq) + 6H_2O(l) + \text{energy}$$

Common error: Forgetting the energy term on the right side. Writing respiration without "+ energy" hides the very point the question asks for.

Exercise Q11–Q12: Decomposition vs Combination

Exercise Q11: Why are decomposition reactions called the opposite of combination reactions? Write equations.

A combination reaction builds a single product from two or more reactants. A decomposition reaction does the reverse — one reactant splits into two or more products. One builds up; the other breaks apart. That is why they are opposites (p. 8).

- **Combination:** $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$ — two reactants form one product.
- **Decomposition:** $CaCO_3(s) \xrightarrow{\text{Heat}} CaO(s) + CO_2(g)$ — one reactant forms two products.

Exercise Q12: Write one decomposition equation each driven by heat, light, and electricity.

Decomposition always needs an energy input because chemical bonds in the reactant must be broken. That energy can arrive as heat, light, or electricity, and each form drives a different reaction.

Heat (thermal decomposition) — lead nitrate breaks into lead oxide, nitrogen dioxide, and oxygen (Activity 1.6, p. 8):

$$2Pb(NO_3)_2(s) \xrightarrow{\text{Heat}} 2PbO(s) + 4NO_2(g) + O_2(g)$$

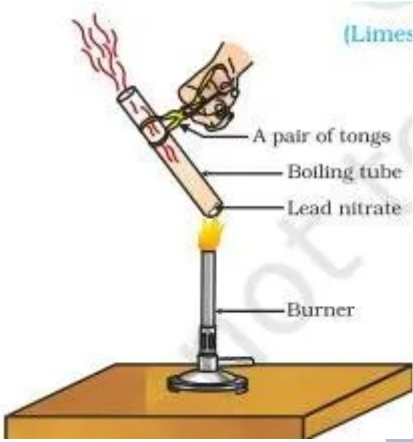

Figure 1.5 Heating of lead nitrate and emission of nitrogen dioxide. Source: NCERT

Light (photolytic decomposition) — silver chloride decomposes in sunlight (Activity 1.8, p. 9):

$$2AgCl(s) \xrightarrow{\text{Sunlight}} 2Ag(s) + Cl_2(g)$$

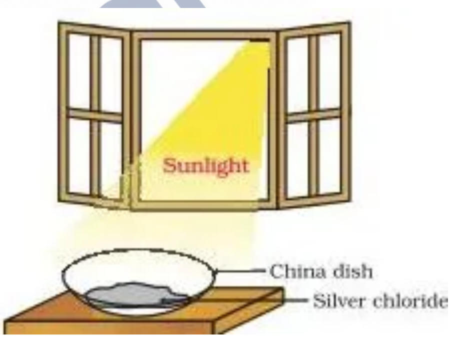

Figure 1.7 Silver chloride turns grey in sunlight to form silver metal. Source: NCERT

Electricity (electrolytic decomposition) — water splits into hydrogen and oxygen (Activity 1.7, p. 9):

$$2H_2O(l) \xrightarrow{\text{Electricity}} 2H_2(g) + O_2(g)$$

Common error: Writing a thermal decomposition equation without the “Heat” label above the arrow. The energy source is part of the answer — omitting it loses the point of the question.

Exercise Q13–Q15: Displacement vs Double Displacement, Silver, Precipitation

Exercise Q13: Difference between displacement and double displacement reactions, with equations.

The difference is whether a single free element swaps into a compound, or whether two whole ionic compounds exchange partners. One involves an element entering a compound; the other involves ions trading places between two compounds.

Feature	Displacement	Double Displacement
Direction of exchange	One-way: a free element pushes another element out	Two-way: ions swap partners between two compounds
Reactant make-up	Element + compound	Compound + compound
Typical products	New element + new compound	Two new compounds (often one precipitate)
Example	$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$	$Na_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2NaCl$

Displacement equation: $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

Double displacement equation: $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Exercise Q14: Write the reaction for recovery of silver from silver nitrate using copper.

Copper is more reactive than silver, so it displaces silver from silver nitrate solution. Copper dissolves as copper nitrate while silver metal precipitates out — this is how silver is recovered in refining (p. 11).

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

Common error: Writing only one $AgNO_3$. Copper forms Cu^{2+}, so two nitrate ions must pair with it — two $AgNO_3$ are required on the left.

Exercise Q15: What is a precipitation reaction? Explain with examples.

A **precipitation reaction** is one in which an insoluble solid — the *precipitate* — forms when two soluble solutions mix (p. 11). The precipitate drops out of solution and can be filtered off.

Example 1: Barium chloride + sodium sulphate → white barium sulphate precipitate:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s)\downarrow + 2NaCl(aq)$$

Example 2: Lead nitrate + potassium iodide → yellow lead iodide precipitate:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s)\downarrow + 2KNO_3(aq)$$

Common error: Calling *any* double displacement a precipitation reaction. Only those that form an insoluble product qualify — if both products stay dissolved, there is no precipitate.

Exercise Q16–Q17: Oxidation, Reduction, Identifying Element X

Exercise Q16: Explain oxidation and reduction in terms of gain or loss of oxygen, with two examples each.

Oxidation and reduction always occur together. If one substance gains oxygen, another must lose it — they are coupled, which is the defining feature of redox reactions (p. 12).

Oxidation = gain of oxygen.

- $2Cu + O_2 \xrightarrow{Heat} 2CuO$ — copper gains oxygen; copper is oxidised.
- $4Na + O_2 \rightarrow 2Na_2O$ — sodium gains oxygen; sodium is oxidised.

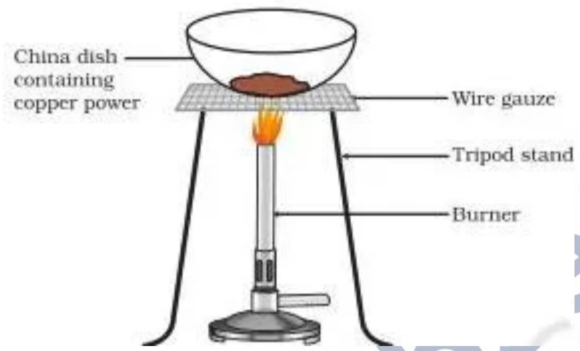

Figure 1.10 Oxidation of copper to copper oxide. Source: NCERT

Reduction = loss of oxygen.

- $CuO + H_2 \xrightarrow{Heat} Cu + H_2O$ — copper oxide loses oxygen; CuO is reduced.
- $ZnO + C \rightarrow Zn + CO$ — zinc oxide loses oxygen; ZnO is reduced.

Exercise Q17: A shiny brown element X on heating in air becomes black. Name X and the black compound.

Copper is a *shiny brown* metal. On heating in air, it combines with oxygen to form **copper(II) oxide** (CuO), which is black (Activity 1.11, p. 12). Though copper is less reactive than many metals, it still oxidises when heated directly.

$$2Cu(s) + O_2(g) \xrightarrow{Heat} 2CuO(s)$$

> **Final answer:** X = copper; black compound = copper oxide (CuO).

Common error: Guessing X as iron. Iron is grey, not shiny brown, and it forms reddish-brown rust only in the presence of moisture — not a black coating on simple heating.

Exercise Q18–Q20: Corrosion and Rancidity

Exercise Q18: Why do we apply paint on iron articles?

Corrosion is the gradual attack on a metal by substances around it — moisture, acids, or gases — which forms a surface coating (p. 13). For iron, that means rusting, a reddish-brown powder. Paint creates a physical barrier between iron and the moisture and oxygen that cause rust, so the metal underneath is protected.

Common error: Saying paint makes iron "look new." Appearance is secondary; the functional reason is preventing contact with moisture and oxygen.

Exercise Q19: Why are oil and fat containing food items flushed with nitrogen?

Nitrogen is an unreactive gas. Flushing it into a food bag displaces oxygen, so fats and oils cannot oxidise and turn **rancid** (p. 13). The food keeps its smell and taste for longer.

Common error: Saying nitrogen prevents bacterial or microbial spoilage. The reason specified in this chapter is *oxidation* of fats — keep your answer to that point.

Exercise Q20: Explain corrosion and rancidity with one example each.

Corrosion — a metal is gradually attacked by moisture, acids, or gases, forming a surface coating. Example: reddish-brown rust on iron; black coating on silver; green coating on copper (p. 13).

Rancidity — fats and oils are oxidised by air, changing their smell and taste. Example: old chips or stale cooking oil smelling unpleasant (p. 13).

Aspect	Corrosion	Rancidity
What is attacked	Metal surface	Fats and oils in food
Cause	Moisture, acids, gases	Oxidation by air (oxygen)
Visible effect	Coating on metal (rust, tarnish)	Change in smell and taste
Example	Rust on iron	Stale chips smelling bad

Activity Insights: Key Observations and Common Errors

Each textbook activity demonstrates a specific reaction type and a key observation. The table below summarises Activities 1.1 through 1.11 and flags the mistake students most often make when describing them in an exam.

Activity	What it shows	Equation (reference)	Common student error
1.1	Magnesium burns with dazzling white flame to form white MgO ash	$2Mg + O_2 \rightarrow 2MgO$	Saying the ash is "magnesium" — it is magnesium oxide
1.2	Lead nitrate + KI produces a yellow precipitate; zinc + acid produces gas	$Pb(NO_3)_2 + 2KI \rightarrow PbI_2\downarrow + 2KNO_3$	Mixing up the two results – remember Fig 1.2 shows the gas from zinc + acid
1.3	Zinc + dilute acid releases hydrogen gas; flask warms up	$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\uparrow$	Forgetting the flask warms — this reaction is exothermic
1.4	Quick lime + water → slaked lime; beaker becomes hot	$CaO + H_2O \rightarrow Ca(OH)_2 + \text{Heat}$	Calling the product quick lime — it is slaked lime
1.5	Ferrous sulphate changes from green to reddish-brown; smell of burning sulphur	$2FeSO_4 \xrightarrow{\text{Heat}} Fe_2O_3 + SO_2 + SO_3$	Writing the colour change as "black" — it is reddish-brown (ferric oxide)
1.6	Brown fumes of nitrogen dioxide rise from heated lead nitrate	$2Pb(NO_3)_2 \xrightarrow{\text{Heat}} 2PbO + 4NO_2 + O_2$	Calling the brown fumes SO_2 — SO_2 is colourless; the brown fumes are NO_2
1.7	Electrolysis of water: gas volume in one tube is double the other	$2H_2O \rightarrow 2H_2 + O_2$	Swapping which electrode gives hydrogen — hydrogen is at the cathode
1.8	White silver chloride turns grey in sunlight	$2AgCl \xrightarrow{\text{Sunlight}} 2Ag + Cl_2$	Writing the energy source as heat instead of sunlight
1.9	Iron nail in copper sulphate: blue fades, brown coating on nail	$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$	Saying iron "dissolves" — copper is what deposits on the nail
1.10	Sodium sulphate + barium chloride → white barium sulphate precipitate	$Na_2SO_4 + BaCl_2 \rightarrow BaSO_4\downarrow + 2NaCl$	Calling it a simple displacement — it is double displacement
1.11	Copper powder turns black on heating; black turns brown in hydrogen	$2Cu + O_2 \rightarrow 2CuO$; $CuO + H_2 \rightarrow Cu + H_2O$	Forgetting that the reverse reaction (reduction) also happens when hydrogen is passed over heated CuO

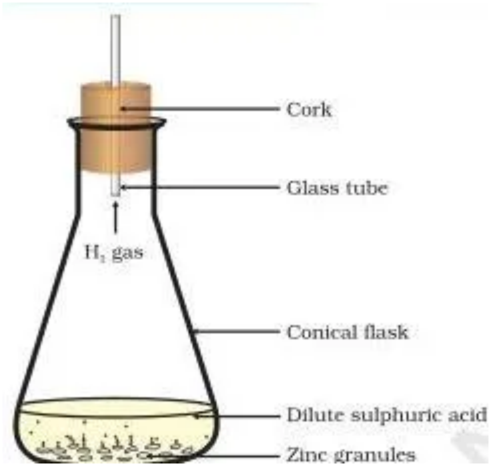

Figure 1.2 Formation of hydrogen gas by the action of dilute sulphuric acid on zinc. Source: NCERT

Method Recap: How to Balance Any Equation

The textbook's hit-and-trial method (pp. 4–5) works in seven steps. Follow them in order on any unfamiliar equation:

1. **Box each formula.** Draw boxes around every reactant and product. You can change only the numbers *outside* the boxes (coefficients) — never the subscripts inside.
2. **List atom counts.** For each element, count atoms on the left and on the right.
3. **Start with the biggest.** Pick the compound that has the most atoms, and within it the element with the most atoms. Balance that element first.
4. **Balance with coefficients.** Place a whole-number coefficient in front of the formula that is short. Never change a subscript to force balance.
5. **Move to the next element.** Balance each remaining element one at a time using new coefficients.
6. **Verify all counts.** Recount every element on both sides; they must match exactly.
7. **Add state symbols.** Write (s), (l), (g), or (aq) for each substance when the question asks for them.

The single rule that catches most students: **formulae are fixed; only coefficients change**. To balance oxygen you write $4H_2O$, never H_2O_4 or $(H_2O)_4$ — those are different chemicals entirely.

Worked example 1: Methane combustion

Balance $CH_4 + O_2 \rightarrow CO_2 + H_2O$. Carbon appears once on each side, so balance C first (already 1:1). Then balance H: four H on the left (in CH_4) need two H_2O on the right. Now count oxygen: two in CO_2 plus two in $2H_2O$ = four O on the right, so we need $2O_2$ on the left.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

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Worked example 2: Calcium phosphate and silica

Balance $Ca_3(PO_4)_2 + SiO_2 \rightarrow CaSiO_3 + P_2O_5$. Start with the largest compound, $Ca_3(PO_4)_2$: three Ca on the left mean three $CaSiO_3$ on the right. That gives three Si on the right, so three SiO_2 on the left.

Now count oxygen: left has eight (in phosphate) plus six (in three SiO_2) = fourteen; right has nine (in three $CaSiO_3$) plus five (in P_2O_5) = fourteen. P already balances at 2:2.

$$Ca_3(PO_4)_2 + 3SiO_2 \rightarrow 3CaSiO_3 + P_2O_5$$

Common error: Changing subscripts inside a formula (e.g. writing $Ca_2(PO_4)_2$) instead of adding a coefficient. The formula of a compound is fixed; only coefficients are adjustable when balancing.

FAQs: Chemical Reactions and Equations

What is the difference between a skeletal equation and a balanced chemical equation?

A skeletal equation shows the correct reactants and products with correct formulae, but the atom counts on the two sides are not equal. A balanced equation has equal numbers of every element on both sides, satisfying the Law of Conservation of Mass (p. 3).

How can I tell if a reaction is exothermic or endothermic without being told in the question?

Look for clues: if the equation has "+ Heat" or "+ energy" on the product side, the reaction is exothermic. If heat, sunlight, or electricity is written above the arrow as a required input, the reaction is endothermic. A warming flask indicates exothermic; a cooling flask or an external energy supply indicates endothermic (pp. 7, 9).

Do all double displacement reactions produce a precipitate?

No. A double displacement reaction exchanges ions between two compounds. If one of the new combinations is insoluble in water, a precipitate forms and the reaction is also a precipitation reaction. If both products stay dissolved, there is no precipitate — it is still double displacement, just not a precipitation reaction (p. 11).

What is the difference between rancidity and corrosion in one line each?

Corrosion is the slow attack on a metal surface by moisture or gases, forming a coating (e.g. rust on iron). Rancidity is the oxidation of fats and oils in food, which changes their smell and taste (p. 13).

When should I write physical state symbols like (s), (l), (g), and (aq) in a chemical equation?

Write state symbols whenever the question asks for them, whenever a precipitate forms (to show it is solid), or whenever knowing the state clarifies the reaction (e.g. steam is $H_2O(g)$, not liquid). The textbook says states are not usually written unless necessary (p. 5).

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How do I identify the oxidising agent and reducing agent in a redox reaction?

Find which substance gains oxygen (or loses hydrogen) — that substance is oxidised and is the reducing agent. Find which substance loses oxygen (or gains hydrogen) — that substance is reduced and is the oxidising agent. The agent is named for what it *causes*, not what happens to it (p. 12).

Mistake	Correct rule	How to check
Writing $FeCl_3$ with dilute HCl	Iron forms Fe^{2+} with dilute HCl, so use $FeCl_2$	Check the charge: dilute acid gives Fe^{2+}, not Fe^{3+}
Saying hydrogen is at the anode in electrolysis	Hydrogen collects at the cathode, oxygen at the anode	Count H atoms in H_2O: there are two, so hydrogen has double volume
Calling a double displacement a displacement reaction	Two ionic compounds swapping partners is double displacement	Ask: is a free element pushing another out? If not, it is double displacement
Guessing element X as iron	Shiny brown element that turns black on heating is copper	Iron is grey; copper is shiny brown and forms black CuO on heating
Changing subscripts while balancing	Use coefficients only; formulae are fixed	Box each formula first, then only change numbers outside the boxes

Reference: NCERT Class 10 Science textbook, chapter Chemical Reactions and Equations.

About the Author

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Greya Lakshmi is the academic editor behind LearnCBSE.net. She works with subject teachers to turn the official CBSE syllabus, NCERT textbooks and previous-year question papers into clear, exam-ready study material: step-by-step NCERT solutions, worked examples, marking-scheme tips and common-mistake guides for CBSE Classes 1 to 12. Every page is reviewed for syllabus alignment and accuracy before it is published.

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