

Ages 15–16 · Science

LESSON SLIDES

Chemical Reactions and Equations

Chemistry · Writing, balancing and classifying reactions

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Chemical Reactions and Equations

SPEAKER NOTES

Open with a quick question: what changes when milk turns sour, or when an iron gate rusts? Collect answers and point out that in both cases a new substance forms. Tell the class that today they will learn to describe such changes in the short, exact language of chemistry.

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By the end of this lesson you can

- 🎯 Spot the signs of a chemical reaction
- 🎯 Write and balance a chemical equation
- 🎯 Name the four main types of reactions
- 🎯 Tell exothermic from endothermic reactions
- 🎯 Explain oxidation, reduction, corrosion and rancidity

By the end of this lesson you can

SPEAKER NOTES

Read the objectives aloud and ask students which one sounds hardest. Promise that balancing, usually the scariest, will be done step by step with a familiar example.

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What is a chemical reaction?

- Starting substances (reactants) change into new substances (products)
- Products have different properties from the reactants
- Atoms are rearranged; no atom is created or destroyed
- Examples: milk souring, food cooking, iron rusting

What is a chemical reaction?

SPEAKER NOTES

Contrast with a physical change such as ice melting, where no new substance forms. Ask: is dissolving sugar in water a chemical reaction? It is not, because the sugar can be recovered by evaporating the water.

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Signs that a reaction has happened

- A change of colour
- A gas is given off (bubbles or fizzing)
- The mixture gets hotter or colder
- A change of state, or a solid (precipitate) appears

Signs that a reaction has happened

SPEAKER NOTES

Burn a small magnesium ribbon only as a teacher demonstration with tongs and eye protection: it gives a dazzling white flame and leaves white magnesium oxide. Warn students never to look straight at the flame. A sign such as a colour change hints at a reaction but does not prove one on its own.

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From words to a chemical equation

- Word equation: magnesium + oxygen → magnesium oxide
- Replace names with formulas: $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$
- Reactants go on the left, products on the right
- The arrow means 'forms' or 'changes into'

From words to a chemical equation

SPEAKER NOTES

Show that the formula version is shorter and tells us exactly which atoms are involved. Ask a student to count the oxygen atoms on each side of $\text{Mg} + \text{O}_2 \rightarrow \text{MgO}$; they will find two on the left and one on the right, which sets up the next slide.

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Why equations must be balanced

- Law of conservation of mass: mass is neither created nor destroyed
- Each element needs the same number of atoms on both sides
- Balance with coefficients (numbers in front of formulas)
- Never change the subscripts inside a formula



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Why equations must be balanced

SPEAKER NOTES

Balanced form: $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$. Explain that changing O_2 to O or MgO to MgO_2 would describe different substances, which is why only coefficients may change. A common mistake is to write the coefficient as a subscript.

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 Example

Balancing the burning of methane (natural gas)

- 1 Skeleton equation: $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- 2 Carbon: one on each side, already balanced
- 3 Hydrogen: 4 on the left, so write $2\text{H}_2\text{O}$
- 4 Oxygen: now 4 on the right, so write 2O_2
- 5 Balanced: $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$

Balancing the burning of methane (natural gas)

SPEAKER NOTES

Natural gas, used for cooking and heating in millions of homes, is mostly methane, so this reaction happens in every gas flame. Work through each step on the board and recount every element at the end. Ask students to balance $\text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O}$ on their own before moving on.

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Making equations more informative

- State symbols: (s) solid, (l) liquid, (g) gas
- (aq) means dissolved in water, an aqueous solution
- Conditions go above or below the arrow: heat, light, catalyst
- Example: $\text{Zn(s)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{ZnSO}_4\text{(aq)} + \text{H}_2\text{(g)}$



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Making equations more informative

SPEAKER NOTES

State symbols are often optional unless a question asks for them, but they make an equation much clearer. Ask: in the zinc example, which product would you see as bubbles? The answer is hydrogen gas.

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Combination and decomposition

Combination

- Two or more reactants form one product
- $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$ (slaked lime)
- Gives out a lot of heat
- Slaked lime is used to whitewash walls

Decomposition

- One reactant breaks into simpler products
- $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ on heating
- Needs energy: heat, light or electricity
- Used to make quicklime from limestone

Combination and decomposition

SPEAKER NOTES

Point out that one reaction is roughly the reverse pattern of the other: many things combine into one, or one thing splits into many. Adding quicklime to water must be done carefully because the mixture gets very hot. Ask students to classify $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$; it is a combination reaction.

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Three ways to break a compound down

- Heat (thermal): $2\text{FeSO}_4 \rightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2 + \text{SO}_3$
- Electricity (electrolysis): $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$
- Light (photolysis): $2\text{AgCl} \rightarrow 2\text{Ag} + \text{Cl}_2$
- White silver chloride turns grey in sunlight



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Three ways to break a compound down

SPEAKER NOTES

In the electrolysis of water, the volume of hydrogen collected is twice that of oxygen, which matches the 2 : 1 ratio in the equation. The action of light on silver salts was the basis of black-and-white photography. Heating ferrous sulphate gives off gases with a sharp smell, so it should be done only under supervision.

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Displacement and double displacement

Displacement

- A more reactive metal pushes out a less reactive one
- $\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$
- Blue solution fades; iron gets a brown copper coat

Double displacement

- Two compounds swap their ions
- $\text{Na}_2\text{SO}_4 + \text{BaCl}_2 \rightarrow \text{BaSO}_4 + 2\text{NaCl}$
- White BaSO_4 precipitate forms

Displacement and double displacement

SPEAKER NOTES

Iron is more reactive than copper, so it displaces copper from copper sulphate solution. A reaction that produces an insoluble solid is also called a precipitation reaction. Ask: would copper displace iron from iron sulphate? No, because copper is less reactive.

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Exothermic and endothermic reactions

Exothermic

- Release heat to the surroundings
- Burning natural gas or LPG
- Respiration in our cells
- Quicklime reacting with water

Endothermic

- Absorb energy from the surroundings
- Heating limestone to make quicklime
- Electrolysis of water
- Silver chloride splitting in sunlight

Exothermic and endothermic reactions

SPEAKER NOTES

Respiration is exothermic: glucose reacts with oxygen and releases the energy our body uses. Most decomposition reactions are endothermic because energy is needed to break the reactant apart. Ask students which type a burning candle is.

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Oxidation and reduction

- Oxidation: a substance gains oxygen or loses hydrogen
- Reduction: a substance loses oxygen or gains hydrogen
- $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$
- Here CuO is reduced and H₂ is oxidised
- Both happen together: a redox reaction



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Oxidation and reduction

SPEAKER NOTES

Heating copper powder in air forms black copper oxide, $2\text{Cu} + \text{O}_2 \rightarrow 2\text{CuO}$, which is oxidation. Passing hydrogen over hot copper oxide turns it back to brown copper. A common misconception is that oxidation always needs oxygen gas; losing hydrogen also counts.

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 Example

Oxidation at home: corrosion and rancidity

- 1 An iron gate rusts faster in damp, humid weather
- 2 Silver jewellery and cutlery slowly turn black
- 3 Copper vessels develop a green coating
- 4 Old crisps and cooking oil smell and taste stale (rancid)
- 5 Crisp (chip) packets are filled with nitrogen to slow rancidity

Oxidation at home: corrosion and rancidity

SPEAKER NOTES

Corrosion is the slow damage of a metal by air, moisture or chemicals around it; rusting of iron is the most costly example. Rancidity is the oxidation of fats and oils in food. Ask students how their families prevent these: painting and oiling gates, storing food in airtight containers, and adding antioxidants.

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 Quick check

In $\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}$, which substance is oxidised?

A CuO**B** H₂**C** Cu**D** H₂O

Quick check

SPEAKER NOTES

Answer: B, hydrogen, because it gains oxygen to become water. Copper oxide loses oxygen, so it is reduced. Let students vote before revealing the answer, then ask them to explain their choice.

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Key takeaways

- ✓ Reactions rearrange atoms into new substances
- ✓ Balanced equations obey conservation of mass
- ✓ Four types: combination, decomposition, displacement, double displacement
- ✓ Reactions either release heat or absorb energy
- ✓ Oxidation and reduction always happen together

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Key takeaways

SPEAKER NOTES

Ask each student to write one balanced equation and label its reaction type as an exit ticket. Then move to the five-question quiz. Homework: find two examples of corrosion or rancidity at home and suggest how to prevent each.