CBSE CLASS X

Science (086)

ANSWER KEY

AI-generated question paper

Code: KTEOTB **Questions: 31** **Maximum Marks: 83** **Generated: 2026-06-25 17:42**

SELECTIONS USED

Subject	Science
Lessons	3 Metals and Non-metals
Level of understanding	Thorough understanding
Question selection	Curated chapter coverage (~5 questions per section + 8 synthesis)
Model	claude-sonnet-4-6

Composition — Difficulty: 1 straightforward · 18 medium · 12 deep | Types: 24 Short · 5 MCQ · 2 Long

Q1. deep | thorough-understanding | § Chapter Introduction **[3]**

Graphite, a form of carbon, conducts electricity like a metal, whereas sulfur — also a non-metal — is a brittle solid that does not conduct electricity at all. Yet both are classified as non-metals. (i) What does this suggest about using physical properties alone to classify elements as metals or non-metals? (ii) Which category of properties provides a more reliable basis for this classification? Justify your answer with one example.

◆ Metals and Non-metals

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Model Answer

(i) Physical properties alone are **not reliable** for classifying elements as metals or non-metals, because there are many exceptions. Graphite (a non-metal) conducts electricity like a metal, and iodine (a non-metal) is lustrous. Such exceptions show that no single physical property is universal to all metals or non-metals.

(ii) Chemical properties provide a more reliable basis for classification.

Justification: Metals form **basic oxides** when burnt in air, while non-metals form **acidic oxides**. For example, magnesium burns to form MgO, which is basic, whereas sulphur burns to form SO_2, which dissolves in water to give an acidic solution. This chemical behaviour is consistent and does not show the exceptions seen in physical properties.

Source: Chapter 3, Section 3.1.2 and Activity 3.8

Explanation

- Part (i) expects you to state that physical properties have **too many exceptions** — the textbook itself lists graphite conducting electricity, iodine being lustrous, and alkali metals being very soft as key examples.
- Part (ii) is directly from the textbook conclusion after Activity 3.8: *"Elements can be more clearly classified as metals and non-metals on the basis of their chemical properties."* Always support with the metal oxide / non-metal oxide example (basic vs. acidic oxide) as the examiner expects it.
- Don't just say "chemical properties are better" — you must **justify with an example** as the question demands.

Q2. deep | thorough-understanding | § 3.1 PHYSICAL PROPERTIES **[3]**

Iodine is a non-metal, yet it has a shiny surface, and graphite is a non-metal that conducts electricity. What do these two exceptions tell us about the reliability of physical properties alone for classifying elements as metals or non-metals?

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Model Answer

The two exceptions — iodine being lustrous and graphite conducting electricity — show that physical properties alone are **not reliable** for classifying elements as metals or non-metals.

Lustre (shiny surface) is generally considered a property of metals, yet iodine, a non-metal, is lustrous. Similarly, non-metals are generally bad conductors of electricity, yet graphite (an allotrope of carbon) conducts electricity. These exceptions prove that physical properties have limitations and can be misleading.

Therefore, **chemical properties** provide a more accurate and reliable basis for classifying elements as metals or non-metals.

Source: Chapter 3, Section 3.1.2 (Physical Properties — Non-metals)

Explanation

- Examiners expect you to **name both exceptions clearly**, explain what rule each one breaks, and then give the **conclusion** that chemical properties are more reliable.
- The textbook explicitly states: *"we cannot group elements according to their physical properties alone, as there are many exceptions."* Use this idea directly.
- Don't just list facts — connect them to what they *tell us* (the reliability question), since that is exactly what the question asks.

Q3. medium | thorough-understanding | § 3.1 PHYSICAL PROPERTIES **[2]**

A metal worker shapes copper into various forms for industrial use — flattening it into thin sheets and stretching it into long, thin wires. Name the two physical properties of copper being utilised in each case and explain how each property makes the respective process possible.

◆ **Metals and Non-metals**

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Model Answer

Flattening into thin sheets — Malleability: Copper is malleable, meaning it can be beaten or rolled into thin sheets without breaking. This property allows the metal to change shape under compressive force without cracking.

Stretching into wires — Ductility: Copper is ductile, meaning it can be drawn into long, thin wires. This property allows the metal to be stretched under tensile force without breaking.

Source: Chapter 3, Section 3.1.1

Explanation

- The examiner expects **both property names** (malleability, ductility) to be correctly matched to the right process — each named property with a brief explanation earns 1 mark each.
- A common error is confusing the two: **malleable = beaten into sheets; ductile = drawn into wires**. Remember: **D**uctile → **D**raw wires.
- You do not need to give examples of other metals; just define and link each property to the process described.

Q4. medium | thorough-understanding | § 3.1 PHYSICAL PROPERTIES **[3]**

Electric wires used in homes are made of copper or aluminium but are always covered with PVC or a rubber-like coating. Why is the metal core chosen for the inner part, and why must the outer covering be made of a non-metal material?

◆ Metals and Non-metals

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Model Answer

Metal core (copper/aluminium): Metals are good conductors of electricity, allowing electric current to flow easily through the wire with low resistance. Copper and aluminium are preferred because they are highly conductive, ductile, and relatively affordable.

Outer covering (PVC/rubber): The outer coating must be a non-metal because non-metals are bad conductors (insulators) of electricity. This insulating layer prevents electric shock to the user, avoids short-circuiting if two wires touch, and protects against current leakage.

Source: Chapter 12, Section 12.4 (Domestic Electric Circuits); Chapter 3 (Metals and Non-metals — What You Have Learnt)

Explanation

- The question tests **two distinct concepts** — conductivity of metals and insulating property of non-metals — so address both clearly for full marks (likely 1½ + 1½ or 1 + 2 split).
- Key phrase from Chapter 3: *"Metals are good conductors of heat and electricity"* and *"Non-metals are bad conductors of heat and electricity."*
- Examiners expect you to link the property (conductivity / non-conductivity) to the function (carrying current / preventing shock/short circuit). Don't just name the property — state **why it matters** for the wire.

Q5. medium thorough-understanding § 3.1 PHYSICAL PROPERTIES **[1]**

Which of the following statements about the physical properties of metals is INCORRECT?
(A) Mercury is the only metal that exists as a liquid at room temperature.
(B) Alkali metals such as sodium and potassium are hard solids with high melting points.
(C) Silver and copper are the best conductors of heat and electricity among all metals.
(D) Gold is highly ductile — one gram of it can be drawn into a very long, thin wire.

A Mercury is a metal that exists as a liquid at room temperature.
B Alkali metals such as sodium and potassium have high densities and high melting points.
C Silver and copper are the best conductors of heat among metals.
D Gold is the most ductile metal, and a wire of about 2 km can be drawn from one gram of it.

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Model Answer

(B) is incorrect. Alkali metals such as sodium and potassium are **soft** (can be cut with a knife) and have **low densities and low melting points** — not high densities and high melting points.

Source: Chapter 3, Section 3.1 Physical Properties

Explanation

The textbook explicitly states: *"Alkali metals (lithium, sodium, potassium) are so soft that they can be cut with a knife. They have low densities and low melting points."* All other options are factually correct as per the chapter — mercury is the only liquid metal at room temperature; silver and copper are the best conductors of heat; and one gram of gold can be drawn into ~2 km of wire. In MCQs, find the one statement that directly contradicts the textbook.

Q6. medium | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[3]**

Sodium and potassium react violently with cold water, while calcium's reaction is comparatively less vigorous. State two observations that distinguish sodium's reaction with cold water from calcium's reaction, and give reasons for these differences.

◆ Metals and Non-metals

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Model Answer

Two observations distinguishing sodium's reaction from calcium's:

1. **Vigour/Speed:** Sodium reacts very vigorously (violently) with cold water, often catching fire, whereas calcium's reaction is comparatively slower and less vigorous.

1. **Movement on water:** Sodium moves rapidly/erratically on the water surface (may melt into a ball) due to the large amount of heat produced, while calcium sinks and reacts steadily, producing bubbles.

Reasons: Sodium is higher in reactivity than calcium in the activity series — it loses its single valence electron more readily. Calcium has two valence electrons to lose, making it less reactive. Both form hydroxides and hydrogen gas:

$$2\text{Na} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2 \uparrow$$

$$\text{Ca} + 2\text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 + \text{H}_2 \uparrow$$

Source: Chapter 3, Activity series / reactivity of metals with water

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Explanation

- Examiners expect **two distinct observations** (vigour, fizzing/fire, movement) and a **reason linked to reactivity/electronic configuration** — losing 1 electron (Na) vs 2 electrons (Ca).
- Writing the balanced equations earns marks and shows understanding.
- "Sodium catches fire" is a valid observation for full credit.
- Do not confuse CaO + H_2O (combination reaction) with Ca metal + H_2O — the question is about the metal reacting with water.

Q7. medium | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[3]**

When iron reacts with steam, the product is Fe_3O_4, not Fe_2O_3. Write the balanced chemical equation for this reaction and explain why iron cannot be used instead of magnesium in a reaction with cold water to produce hydrogen gas in the laboratory.

◆ Metals and Non-metals

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Model Answer

Balanced chemical equation for iron reacting with steam:

$$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$$

Why iron cannot replace magnesium in a reaction with cold water:

Iron is much less reactive than magnesium. Magnesium reacts readily with hot water to produce hydrogen gas, whereas iron does not react with cold water or even hot water — it reacts only with steam at high temperatures. Therefore, iron cannot be used in the laboratory to produce hydrogen gas by reacting with cold water.

Source: Chapter 3, Section 3.2.2

Explanation

- The equation for iron + steam is a standard board exam question. Ensure the formula $\mathbf{Fe_3O_4}$ (not Fe_2O_3) is written correctly and the equation is balanced: 3 Fe, 4 H_2O, 1 Fe_3O_4, 4 H_2.
- The second part tests your understanding of the **reactivity series**: Mg > Fe. Iron reacts only with steam (needs high temperature), not cold or hot water, so it cannot produce H_2 conveniently in the lab.
- Examiners award 1 mark for the balanced equation and 2 marks for the explanation comparing reactivity of iron and magnesium with water.

Q8. medium | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[1]**

Which of the following correctly explains why hydrogen gas is NOT produced when copper reacts with dilute hydrochloric acid?
(A) Copper reacts with HCl to form an insoluble salt that blocks further reaction.
(B) Copper lies below hydrogen in the reactivity series and therefore cannot displace hydrogen from an acid.
(C) Copper reacts with HCl but the hydrogen produced immediately reacts back with copper.
(D) Copper requires concentrated acid, not dilute acid, to produce hydrogen.

A Copper reacts with HCl to form an insoluble salt that blocks further reaction.
B Copper lies below hydrogen in the reactivity series and therefore cannot displace hydrogen from an acid.
C Copper reacts with HCl but the hydrogen produced immediately reacts back with copper.
D Copper requires concentrated acid, not dilute acid, to produce hydrogen.

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Model Answer

(B) Copper lies below hydrogen in the reactivity series and therefore cannot displace hydrogen from an acid.

Explanation

The reactivity series determines which metals can displace hydrogen from dilute acids. Only metals placed **above hydrogen** in the reactivity series (like Zn, Fe, Mg) can displace it. Copper is below hydrogen, so it cannot reduce H^+ ions to H_2 gas. This is the standard CBSE reason; options A, C, and D are factually incorrect.

Q9. medium | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[3]**

A student adds a piece of zinc to copper sulphate solution and observes that the blue colour of the solution gradually fades and a reddish-brown deposit forms on the zinc. (i) Write the balanced chemical equation for this reaction. (ii) Identify the type of reaction. (iii) What does this result tell you about the relative reactivities of zinc and copper?

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Model Answer

(i) Balanced chemical equation:

$$\mathrm{Zn}(s) + \mathrm{CuSO_4}(aq) \rightarrow \mathrm{ZnSO_4}(aq) + \mathrm{Cu}(s)$$

(ii) Type of reaction:
This is a **displacement reaction**, as zinc displaces copper from copper sulphate solution.

(iii) Relative reactivities:
Zinc is more reactive than copper. A more reactive metal can displace a less reactive metal from its salt solution. Since zinc displaces copper from copper sulphate, zinc is higher than copper in the reactivity series.

Source: Chapter 1, Section 1.2.3 Displacement Reaction; Chapter 3, Section 3.2.4

Explanation

- The equation (1.25) is directly given in the textbook — write it with state symbols for full marks.
- "Displacement reaction" is the exact NCERT term; do not write "single displacement" or "substitution."
- For part (iii), the key idea is: a more reactive metal displaces a less reactive one. The reactivity series (Table 3.2) places Zn above Cu, confirming this.
- The fading blue colour indicates $CuSO_4$ is being consumed; the reddish-brown deposit is metallic copper being formed — good details to mention if asked to describe observations.

Q10. medium | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[3]**

Aluminium oxide reacts with both hydrochloric acid and sodium hydroxide solution. Write the chemical equations for both reactions and explain what property of aluminium oxide these reactions demonstrate.

◆ Metals and Non-metals

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Model Answer

Reaction with hydrochloric acid (HCl):

$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

Reaction with sodium hydroxide (NaOH):

$$Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$$

(Sodium aluminate)

Property demonstrated:
Since aluminium oxide reacts with both acids and bases to produce salt and water, it behaves as both a basic oxide (with HCl) and an acidic oxide (with NaOH). This property is called **amphoteric nature**. Aluminium oxide is therefore an **amphoteric oxide**.

Source: Chapter 2, Section 2.1.5 & 2.1.6

Explanation

- Examiners expect **both balanced equations** (1 mark each) and a clear explanation of the amphoteric property (1 mark).
- The key term to use is **amphoteric** — an oxide that reacts with both acids and bases. This is the exact terminology CBSE expects.
- Note that with NaOH, the product is **sodium aluminate ($NaAlO_2$)**, not aluminium hydroxide — make sure to write the correct product.
- Don't forget: metallic oxides are basic oxides, but Al_2O_3 is an exception because it also behaves as an acidic oxide, hence amphoteric.

Q11. deep | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[3]**

Hydrogen gas is not evolved when most metals react with nitric acid, yet it IS evolved when magnesium reacts with very dilute nitric acid. What property of nitric acid accounts for the usual absence of hydrogen gas, and what does the exception with very dilute HNO_3 and magnesium suggest about the conditions needed for this property to operate?

◆ **Metals and Non-metals**

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Model Answer

Property of HNO_3: Nitric acid is a strong oxidising agent. It oxidises the hydrogen gas produced to water and itself gets reduced to nitrogen oxides (N_2O, NO, NO_2). Therefore, H_2 is not evolved when metals react with HNO_3.

Exception with Mg and very dilute HNO_3: When the acid is extremely dilute, its oxidising ability is insufficient to oxidise all the hydrogen produced. Magnesium is also highly reactive (above H in activity series), reacting fast enough that H_2 escapes before it can be oxidised. This shows that the oxidising property of HNO_3 operates effectively only at sufficient concentration.

Source: Chapter 3, Section 3.2.3

Explanation

- The key term examiners look for is **"strong oxidising agent"** — this is the direct reason H_2 is not evolved.
- Mention that H_2 is oxidised to **water** and HNO_3 is reduced to **nitrogen oxides**.
- For the exception part, link it to **very dilute** conditions reducing the oxidising power, and **high reactivity of Mg**. Both parts together earn full marks.
- Don't write lengthy chemistry; one equation is not required here — the explanation in words suffices for 3 marks.

Q12. deep | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[2]**

Metal X displaces metal Y from Y's sulphate solution, but metal Y cannot displace metal Z from Z's sulphate solution, whereas metal X can displace metal Z. Arrange X, Y and Z in decreasing order of reactivity, justifying each step of your reasoning using the concept of displacement reactions.

◆ **Metals and Non-metals**

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Model Answer

Since X displaces Y from Y's sulphate solution, X is more reactive than Y.
Since Y cannot displace Z from Z's sulphate solution, Y is less reactive than Z.
Since X can displace Z, X is more reactive than Z.

Decreasing order of reactivity: X > Z > Y

Source: Metals and Non-metals, Activity Series concept; Chapter 1, Displacement Reactions

Explanation

- The key rule: **a more reactive metal displaces a less reactive metal from its salt solution** (from the activity series concept, Chapter 3).
- Work through each clue logically: X > Y (clue 1), Y < Z (clue 2), X > Z (clue 3).
- This fixes Z between X and Y: X > Z > Y.
- Examiners expect you to state the reasoning for each step, not just the final order — that's where the 2 marks lie.

Q13. deep | thorough-understanding | § 3.2 CHEMICAL PROPERTIES OF METALS **[2]**

During the burning of metals in air, iron filings burn vigorously when sprinkled into a flame, yet a solid piece of iron does not burn even on strong heating. What does this tell you about the factor(s) that determine the rate at which a metal reacts with oxygen, beyond just its position in the reactivity series?

◆ **Metals and Non-metals**

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Model Answer

The rate of reaction of a metal with oxygen depends not only on its reactivity but also on the **surface area** exposed. Iron filings have a very large surface area, allowing vigorous burning, whereas a solid piece of iron has a much smaller surface area, so it does not burn even on strong heating.

Source: Chapter 3, Section 3.2.1

Explanation

The textbook explicitly states: "Iron does not burn on heating but iron filings burn vigorously when sprinkled in the flame." This points to **surface area** as the key additional factor. Examiners want you to name this factor clearly and link it to the observation given. Do not just repeat reactivity — the question asks what is *beyond* position in the reactivity series.

Q14. medium | thorough-understanding | § 3.3 HOW DO METALS AND NON-METALS REACT? **[3]**

Using electronic configurations, explain why magnesium chloride has the formula $MgCl_2$ while sodium chloride has the formula NaCl. (Atomic numbers: Na = 11, Mg = 12, Cl = 17)

◆ Metals and Non-metals

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Model Answer

Electronic configurations:

- Na (11): 2, 8, 1 — has **1** electron in outermost shell
- Mg (12): 2, 8, 2 — has **2** electrons in outermost shell
- Cl (17): 2, 8, 7 — needs **1** electron to complete octet

NaCl: Sodium loses 1 electron to form Na^+. One Cl atom gains that 1 electron to form Cl^-. So one Na combines with one Cl → formula **NaCl**.

$MgCl_2$: Magnesium loses 2 electrons to form Mg^{2+}. Each Cl atom can accept only 1 electron. So two Cl atoms are needed to accept both electrons, forming $2Cl^-$. Thus one Mg combines with two Cl → formula **$MgCl_2$**.

$$\mathrm{Mg} \rightarrow \mathrm{Mg}^{2+} + 2e^- \quad ; \quad 2\mathrm{Cl} + 2e^- \rightarrow 2\mathrm{Cl}^-$$

Source: Chapter 3, Section 3.3

Explanation

Examiners look for:

1. Correct electronic configurations of all three elements written explicitly.
2. Clear statement of how many electrons Na and Mg **lose** (valency).
3. Linking Mg's 2 electrons lost to needing **2** Cl atoms — this is the key reasoning for $MgCl_2$ vs NaCl.
4. Equations showing ion formation are a bonus that fetch full marks.

Remember: the formula of an ionic compound is determined by balancing the charges of the ions formed, which directly follows from the number of electrons lost/gained based on electronic configuration.

Q15. medium | thorough-understanding | § 3.3 HOW DO METALS AND NON-METALS REACT? **[3]**

An ionic compound dissolves readily in water but does not conduct electricity in the solid state. However, when melted, the same compound conducts electricity. Why does changing the physical state from solid to liquid enable electrical conduction?

◆ **Metals and Non-metals**

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Model Answer

In the **solid state**, ions in an ionic compound are held in a fixed, rigid lattice structure. Although ions are present, they cannot move freely, so no electrical conduction occurs.

When the compound is **melted (molten state)**, the lattice structure breaks down. The ions (cations and anions) become free to move. These freely moving ions act as charge carriers, allowing electricity to be conducted.

Thus, electrical conduction requires *mobile charged particles* — solid ionic compounds have ions but they are immobile, while molten ionic compounds have freely moving ions that carry current.

Source: Chapter 3 (ionic compounds), referenced in Chapter 4, Section 4.1

Explanation

- The key concept examiners look for: **ions must be free to move** to conduct electricity.
- In solid state → ions are in fixed positions (lattice) → no conduction.
- In molten state → lattice breaks → ions are mobile → conduction occurs.
- This question links Chapter 3 (ionic compounds) content; the passage in Chapter 4 confirms ionic compounds "conduct electricity in solution or in the molten state."
- Avoid confusing this with covalent compounds, which don't conduct even in molten state because they form no ions.

Q16. straightforward | thorough-understanding | § 3.3 HOW DO METALS AND NON-METALS REACT? **[1]**

Which of the following statements correctly explains why ionic compounds have high melting points?
(A) The molecules of ionic compounds are very large and tightly packed.
(B) A large amount of energy is needed to overcome the strong electrostatic forces between oppositely charged ions.
(C) Ionic compounds consist of neutral atoms held together by weak intermolecular forces.
(D) The electrons shared between ions create a strong covalent bond that is hard to break.

A The molecules of ionic compounds are very large and tightly packed.
B A large amount of energy is needed to overcome the strong electrostatic forces between oppositely charged ions.
C Ionic compounds are insoluble in water, so they require more heat to melt.
D The electrons shared between ions create a strong covalent bond that is hard to break.

◆ **Metals and Non-metals**

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Model Answer

(B) A large amount of energy is needed to overcome the strong electrostatic forces between oppositely charged ions.

Explanation

The textbook (Chapter 3, Section 3.3.1) states: "Ionic compounds have high melting and boiling points because a considerable amount of energy is required to break the strong inter-ionic attraction." Options A, C, and D are incorrect — ionic compounds do not consist of molecules, solubility is unrelated to melting point, and ionic bonds are not covalent bonds.

Q17. deep | thorough-understanding | § 3.3 HOW DO METALS AND NON-METALS REACT? **[5]**

A student claims that since aluminium (atomic number 13) is a reactive metal, it should form an ionic compound with fluorine (atomic number 9) in which one aluminium ion bonds with one fluoride ion. Identify the error in this claim and show, using electronic configurations, the correct formula of the ionic compound formed.

◆ **Metals and Non-metals**

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Model Answer

Error in the student's claim:

Aluminium has atomic number 13, so its electronic configuration is 2, 8, 3. It has **3** valence electrons and loses all three to form Al^{3+} (not Al^{+}). The student incorrectly assumed aluminium forms a 1+ ion.

Electronic configurations:

- Al (2, 8, 3) → Al^{3+} (2, 8) + $3e^{-}$
- F (2, 7) + $1e^{-}$ → F^{-} (2, 8)

Correct formula:

Each Al^{3+} ion loses 3 electrons. Each F^{-} ion gains only 1 electron. Therefore, **3 fluoride ions** are needed to accept the 3 electrons lost by one aluminium ion.

$$Al^{3+} + 3F^{-} \rightarrow AlF_3$$

The correct formula of the ionic compound formed is **AlF_3** (aluminium fluoride).

Source: Chapter 3, Section 3.3 — How do Metals and Non-metals React?

Explanation

- Examiners award marks for: (1) identifying the specific error (Al forms Al^{3+}, not Al^{+}), (2) showing both electronic configurations with electron transfer, and (3) deriving the correct formula using charge balance.
- Always write electronic configurations in shell notation (2, 8, 3) and show the ion formed after electron loss/gain.
- The formula follows from charge neutrality: one Al^{3+} needs exactly three F^{-} ions → AlF_3. This mirrors the $MgCl_2$ example in the textbook.

Q18. medium | thorough-understanding | § 3.4 OCCURRENCE OF METALS **[2]**

Gold and silver are found in nature as free elements, while metals like sodium and calcium are never found in free form. What fundamental property of a metal determines whether it occurs in nature as a free element or in a combined state?

◆ Metals and Non-metals

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Model Answer

The fundamental property is **reactivity** (position in the activity series).

Metals at the bottom of the activity series (like gold and silver) are **least reactive**, so they do not combine with other elements and are found in the free state. Metals at the top (like sodium and calcium) are **highly reactive** and readily combine with elements like oxygen, sulphur, etc., so they are never found in the free state in nature.

Source: Chapter 3, Section 3.4.1

Explanation

- The examiner wants you to name the property (reactivity / position in activity series) and then link it to both examples given in the question.
- Key phrase: "least reactive → free state"; "highly reactive → combined state."
- Avoid vague answers like "chemical property" — be specific: **reactivity** or **position in the activity series**.

Q19. medium | thorough-understanding | § 3.4 OCCURRENCE OF METALS **[3]**

[short_answer] Zinc sulphide (ZnS) and zinc carbonate ($ZnCO_3$) are two common ores of zinc. (i) Name the process used to convert each of these ores into zinc oxide (ZnO). (ii) Why must these ores be converted to the oxide form before the metal can be extracted by reduction with carbon, rather than reducing them directly?

◆ **Metals and Non-metals**

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Model Answer

(i)

- **ZnS (zinc sulphide):** The process used is **roasting** — heating strongly in excess air.

$$2\text{ZnS}(s) + 3\text{O}_2(g) \xrightarrow{\text{Heat}} 2\text{ZnO}(s) + 2\text{SO}_2(g)$$

- **$ZnCO_3$ (zinc carbonate):** The process used is **calcination** — heating strongly in limited air.

$$\text{ZnCO}_3(s) \xrightarrow{\text{Heat}} \text{ZnO}(s) + \text{CO}_2(g)$$

(ii) It is easier to obtain a metal from its **oxide** than from its sulphide or carbonate, because carbon can readily reduce metal oxides. If sulphides or carbonates were reduced directly, the reaction would be less efficient and carbon cannot displace the metal as easily from these compounds.

Source: Chapter 3, Section 3.4.4

Explanation

- Examiners award 1 mark each for naming roasting and calcination correctly (with equations if space allows), and 1 mark for the reason.
- Key distinction: **roasting = excess air; calcination = limited/no air**. Don't mix these up.
- The reason must state that metal oxides are easier to reduce (with carbon) compared to sulphides/carbonates — this is the textbook's exact reasoning. Vague answers like "to purify the ore" will not score.

Q20. deep | thorough-understanding | § 3.4 OCCURRENCE OF METALS **[3]**

Carbon (coke) is an effective reducing agent for extracting iron and zinc from their oxides, but it cannot be used to extract sodium or aluminium from their oxides. Why not? How are sodium and aluminium extracted instead?

◆ Metals and Non-metals

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Model Answer

Carbon cannot reduce the oxides of sodium and aluminium because these metals have **more affinity for oxygen than carbon**. Since they are higher than carbon in the reactivity series, carbon cannot displace them from their oxides.

Extraction of Sodium: Sodium is extracted by **electrolytic reduction** of molten sodium chloride (NaCl).

- At cathode: $Na^+ + e^- \rightarrow Na$
- At anode: $2Cl^- \rightarrow Cl_2 + 2e^-$

Extraction of Aluminium: Aluminium is extracted by **electrolytic reduction** of molten aluminium oxide (Al_2O_3).

Source: Chapter 3, Section 3.4.5

Explanation

- The key reason is **relative affinity for oxygen**: Na and Al bind oxygen more strongly than carbon does, so carbon cannot act as a reducing agent for them.
- The examiner expects you to name **electrolysis/electrolytic reduction** as the method and mention the molten compound used (NaCl for Na; Al_2O_3 for Al).
- Writing the electrode reactions for sodium earns extra credit and shows depth — include them if marks allow.
- Do not confuse *electrolytic refining* (purification) with *electrolytic reduction* (extraction). These are different processes.

Q21. medium | thorough-understanding | § 3.4 OCCURRENCE OF METALS **[1]**

[mcq] In the thermit reaction, iron(III) oxide is reduced to produce molten iron. Which of the following correctly identifies the reducing agent used and the reason carbon cannot be used instead?
((A)) Carbon; it has a higher affinity for oxygen than aluminium at high temperatures
((B)) Aluminium; it is more reactive than iron and has a higher affinity for oxygen than carbon at the reaction temperature
((C)) Aluminium; it is less reactive than iron, so it does not interfere with the product
((D)) Carbon; it is cheaper and more readily available than aluminium

A Aluminium is cheaper and more readily available than carbon.
B Aluminium is more reactive than iron and has a greater affinity for oxygen than carbon does at the temperatures involved.
C Carbon would produce a poisonous gas, making it unsafe to use.
D Aluminium has a lower melting point than carbon, making it easier to ignite.

◆ Metals and Non-metals

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Model Answer

Answer: B — Aluminium is more reactive than iron and has a greater affinity for oxygen than carbon does at the temperatures involved.

Explanation

The thermit reaction is: $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$. Aluminium is the reducing agent because it is more reactive than iron and displaces it. At the high temperatures of this reaction, aluminium has a greater affinity for oxygen than carbon, so carbon cannot reduce Fe_2O_3 effectively here. The textbook (Chapter 1, Exercise Q2) classifies this as a displacement reaction. Examiners expect you to correctly name aluminium as the reducing agent and justify it by reactivity/oxygen affinity — not by cost or melting point.

Q22. deep | thorough-understanding | § 3.4 OCCURRENCE OF METALS **[3]**

In electrolytic refining of copper, explain what happens at each electrode during the process. Why do the insoluble impurities not get deposited on the cathode along with the pure copper?

◆ Metals and Non-metals

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Model Answer

At the Anode (impure copper): The impure copper dissolves into the electrolyte (acidified copper sulphate solution). Copper atoms lose electrons and go into solution as Cu^{2+} ions.

At the Cathode (pure copper strip): Cu^{2+} ions from the electrolyte gain electrons and get deposited as pure copper.

Why insoluble impurities don't deposit on cathode: The impurities present in impure copper (anode) that are insoluble do not dissolve into the electrolyte. They simply fall to the bottom of the anode and collect as **anode mud**. Since they never enter the solution as ions, they cannot be deposited at the cathode. Only copper ions migrate and deposit on the cathode.

Source: Chapter 3, Section 3.4.6 – Refining of Metals

Explanation

- Examiners expect you to clearly name both electrodes and state what happens at each — dissolution at anode, deposition at cathode.
- The key point for the second part is the word **"insoluble"** — impurities that don't dissolve can't form ions, so they can't migrate to the cathode. They settle as **anode mud** (this term is expected).
- Avoid vague statements like "they are heavy" — the correct reason is they don't dissolve into the electrolyte.

Q23. medium | thorough-understanding | § 3.5 CORROSION **[3]**

Iron nails are placed in three separate test tubes: one containing ordinary water open to air, one containing boiled distilled water sealed with a layer of oil, and one containing only dry air with a desiccant. After several days, predict in which test tube rusting will be observed and in which it will not. Give reasons for your answer and state the conditions necessary for iron to rust.

◆ Metals and Non-metals

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Model Answer

Observation:

Rusting will be observed only in **Test Tube A** (ordinary water open to air). No rusting will occur in Test Tubes B and C.

Reasons:

- **Test Tube A:** Both air (oxygen) and water are present → rusting occurs.
- **Test Tube B:** Boiled distilled water removes dissolved air; oil layer prevents air from entering → only water present, no rusting.
- **Test Tube C:** Desiccant (anhydrous $CaCl_2$) absorbs all moisture → only dry air present, no rusting.

Conditions necessary for rusting:

The presence of **both oxygen (from air) and water** simultaneously is essential for iron to rust.

Source: Chapter 3, Section 3.5

Explanation

- The examiner expects you to name the correct test tube (A) and give a reason for each of the three tubes — that's the core of the answer.
- Key phrase: "both air and water are necessary" — this must appear explicitly.
- Mentioning the role of boiled water (removes dissolved air) and oil layer shows deeper understanding and secures full marks.
- Anhydrous $CaCl_2$ as desiccant removes moisture — state this clearly for Test Tube C.

Q24. deep thorough-understanding § 3.5 CORROSION **[3]**

Zinc is used to coat iron articles to protect them from rusting (galvanisation). Even if the zinc coating gets scratched or broken, the iron underneath still does not rust immediately. Using your knowledge of the reactivity series, explain why this is so.

◆ **Metals and Non-metals**

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Model Answer

Zinc is more reactive than iron in the reactivity series. When the zinc coating is scratched, both zinc and iron are exposed to moisture and air. Since zinc is more reactive, it preferentially reacts with oxygen and moisture instead of iron. Zinc acts as a **sacrificial metal**, corroding itself while protecting the iron beneath. Thus, iron does not rust until all the zinc in the surrounding area is consumed.

Source: Chapter 3, Section 3.5.1 – Prevention of Corrosion

Explanation

- The key concept examiners want is **reactivity series**: zinc is higher (more reactive) than iron.
- The term **sacrificial protection** (or sacrificial metal) is expected — it often fetches a dedicated mark.
- Avoid just saying "zinc coats iron"; the question specifically asks about what happens *after* the coating breaks, so you must explain the electrochemical/reactivity reasoning.
- Three marks map to roughly three points: (1) zinc is more reactive than iron, (2) zinc reacts preferentially with oxygen/moisture, (3) iron is thus protected until zinc is fully consumed.

Q25. medium | thorough-understanding | § 3.5 CORROSION **[1]**

Pure iron is too soft and corrodes easily, so it is alloyed to make stainless steel. Which of the following correctly identifies the elements added and the primary properties they impart?

(A) Nickel and copper — to increase malleability and electrical conductivity
(B) Carbon and chromium — to increase hardness and resistance to corrosion
(C) Carbon and manganese — to lower melting point and increase lustre
(D) Chromium and copper — to increase ductility and make it lighter

- A To increase its melting point and make it lustrous
- B To make it hard and resistant to corrosion
- C To increase its electrical conductivity and ductility
- D To make it lighter and more malleable

◆ Metals and Non-metals

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Model Answer

(B) Carbon and chromium — to increase hardness and resistance to corrosion

When iron is mixed with nickel and chromium, we get stainless steel, which is hard and does not rust.

Explanation

The textbook (Section 3.5) explicitly states: "if it is mixed with a small amount of carbon (~0.05%), it becomes hard and strong. When iron is mixed with nickel and chromium, we get stainless steel, which is hard and does not rust." Option (B) is the closest correct match — carbon increases hardness and chromium provides corrosion resistance. Examiners expect you to link the alloying elements directly to the properties they impart.

Q26. medium | thorough-understanding | § (whole-chapter synthesis) **[3]**

A student claims: 'Aluminium is highly reactive, so it should corrode rapidly in air and be useless for making cooking utensils.' Identify the flaw in this reasoning and explain, using aluminium's chemical behaviour, why this claim is incorrect.

◆ Metals and Non-metals

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Model Answer

The flaw in the student's reasoning is ignoring the **protective oxide layer** that forms on aluminium's surface.

Although aluminium is highly reactive, when exposed to air it reacts with oxygen to form a thin, tough layer of aluminium oxide (Al_2O_3):

$$4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$$

This oxide layer is stable and strongly adherent. It **prevents further reaction** of the underlying aluminium with air or moisture, thereby protecting it from corrosion. Because of this self-protective behaviour, aluminium does not corrode rapidly and is safe and durable for making cooking utensils.

Source: Chapter 3, Section 3.2.1

Explanation

- The key concept examiners look for is the **protective oxide layer** — name it, explain it forms immediately on exposure to air, and state it prevents further corrosion.
- Writing the balanced equation for aluminium oxide formation earns a mark and shows chemical understanding.
- Avoid just saying "aluminium doesn't corrode" — you must explain *why* (the oxide coat acts as a barrier). That reasoning is the heart of the answer.
- The word "amphoteric" is not required here; stay focused on the corrosion-resistance aspect.

Q27. medium thorough-understanding § (whole-chapter synthesis) **[3]**

Zinc can be extracted from zinc oxide by heating with carbon, but sodium cannot be extracted from sodium oxide the same way. Explain why these two metals require fundamentally different extraction methods, linking your answer to a property that distinguishes them.

◆ **Metals and Non-metals**

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Model Answer

Zinc is a metal of **medium reactivity** (middle of the activity series). Carbon can reduce zinc oxide because carbon has a greater affinity for oxygen than zinc at high temperatures:

$$\mathrm{ZnO(s)} + \mathrm{C(s)} \rightarrow \mathrm{Zn(s)} + \mathrm{CO(g)}$$

Sodium, however, is a **highly reactive metal** (top of the activity series). It has a much greater affinity for oxygen than carbon does, so carbon cannot reduce sodium oxide. Therefore, carbon reduction fails for sodium.

Instead, sodium is extracted by **electrolytic reduction** — electrolysis of molten sodium chloride, where sodium is deposited at the cathode:

$$\mathrm{Na}^{+} + e^{-} \rightarrow \mathrm{Na}$$

The distinguishing property is **reactivity (position in the activity series)**.

Source: Chapter 3, Sections 3.4.4 and 3.4.5

Explanation

Examiners look for three things: (1) identifying that zinc is mid-series and can be reduced by carbon, with the equation; (2) explaining *why* carbon fails for sodium — sodium's greater affinity for oxygen; (3) naming electrolysis as sodium's extraction method. Linking everything to "position in the activity series / reactivity" earns the reasoning mark. Don't just say "sodium is more reactive" — explain what that means for the reduction process.

Q28. deep | thorough-understanding | § (whole-chapter synthesis) **[3]**

An alloy of copper and zinc does not conduct electricity as well as pure copper, even though both copper and zinc are good conductors individually.
(i) What general rule about alloys does this observation illustrate?
(ii) Despite being a good conductor of heat, copper is alloyed with zinc to make brass for many household items instead of being used as pure copper. State one advantage of using brass over pure copper for such applications.

◆ **Metals and Non-metals**

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Model Answer

(i) This observation illustrates that **alloys generally have lower electrical (and thermal) conductivity than their pure parent metals**. The irregular arrangement of different atoms in an alloy disrupts the free flow of electrons, thereby reducing conductivity.

(ii) One advantage of using brass (copper-zinc alloy) over pure copper for household items is that **brass is harder and stronger than pure copper**, making it more durable and resistant to wear and deformation in everyday use.

Source: Metals and Non-metals, Chapter 3 (alloy definition and properties)

Explanation

- Part (i) is worth ~1 mark: state the rule clearly — alloys have lower conductivity than pure constituent metals. Mention *why* briefly (disruption of electron flow) if you want to secure the mark fully.
- Part (ii) is worth ~2 marks: the key advantage examiners expect is **greater hardness/strength/durability**. Avoid repeating conductivity (that's the disadvantage given in the question). The textbook notes alloys are made to improve mechanical properties like hardness.
- Keep answers concise; one clear sentence per part is sufficient for the marks allocated.

Q29. medium | thorough-understanding | § (whole-chapter synthesis) **[5]**

Trace the complete journey of zinc from its naturally occurring sulphide ore to a pure, usable metal. Your answer must include: the reason its ore form must be converted before reduction, the two processes used for this conversion and what they involve, the reduction step, and how the final impure metal is purified.

◆ **Metals and Non-metals**

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Model Answer

Zinc is a moderately reactive metal found mainly as zinc sulphide (ZnS) or zinc carbonate ($ZnCO_3$) in nature.

Why conversion is needed: It is easier to reduce a metal oxide than a sulphide or carbonate. So the ore must first be converted to zinc oxide (ZnO).

Two conversion processes:

- **Roasting** – ZnS is heated strongly in excess air, converting it to ZnO:

$$2ZnS + 3O_2 \xrightarrow{\Delta} 2ZnO + 2SO_2$$

- **Calcination** – $ZnCO_3$ is heated strongly in limited air, converting it to ZnO:

$$ZnCO_3 \xrightarrow{\Delta} ZnO + CO_2$$

Reduction: ZnO is heated with carbon (coke), which acts as the reducing agent:
$$ZnO + C \rightarrow Zn + CO$$

Purification: The impure zinc is refined by **electrolytic refining**. Impure zinc is made the anode, pure zinc the cathode, and zinc salt solution the electrolyte. On passing current, pure zinc deposits on the cathode; insoluble impurities collect as anode mud.

Source: Metals and Non-metals, Section 3.4.4 and 3.4.6

Explanation

- The question has four clearly labelled parts — examiners will award marks for each, so don't skip any.
- The key distinction between **roasting** (excess air) and **calcination** (limited air) is a favourite examiner point — get those details right.
- Writing the chemical equations earns marks and shows precision; keep them balanced.
- For electrolytic refining, remember the three essentials: impure metal = anode, pure metal = cathode, metal salt solution = electrolyte. Mentioning "anode mud" shows complete knowledge.
- Don't confuse calcination (for carbonates) with roasting (for sulphides) — a common error.

Q30. deep | thorough-understanding | § (whole-chapter synthesis) **[3]**

Both the formation of ionic compounds (e.g., NaCl from sodium and chlorine) and the extraction of highly reactive metals by electrolysis of their molten ores share the same fundamental chemical tendency of metals. Identify this tendency and explain how it drives both processes — one leading to compound formation and the other to isolation of the pure metal.

◆ **Metals and Non-metals**

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Model Answer

The fundamental tendency is that **metals lose electrons easily to form positive ions** (cations). This is driven by their need to attain a stable electronic configuration.

In ionic compound formation (e.g., NaCl): Sodium loses one electron to chlorine, forming Na^+ and Cl^- ions. The oppositely charged ions attract each other, forming the ionic compound NaCl.

In electrolytic extraction of reactive metals: Highly reactive metals like sodium or magnesium have such strong affinity for oxygen that carbon cannot reduce their compounds. Electrolysis of molten chlorides is used; the metal cations (Na^+) gain electrons at the cathode and are **deposited as pure metal**: $Na^+ + e^- \rightarrow Na$.

Thus, the same tendency — losing electrons to form cations — leads to compound formation in nature and enables isolation through forced electron gain during electrolysis.

Source: Chapter 3, Sections 3.3 and 3.4.5

Explanation

- The examiner wants you to **name the tendency** (metals lose electrons / form positive ions) and then **link it explicitly to both processes**.
- For NaCl: electron transfer → ionic bond. For electrolysis: metal ions already exist in molten ore; cathode supplies electrons to reduce them back to metal.
- The contrast is key: in NaCl formation the metal *gives away* electrons; in electrolysis, those same cations are *given back* electrons artificially — both driven by the metal's tendency to exist as cations.
- Always write the cathode half-reaction when discussing electrolytic extraction; it fetches a dedicated mark.

Q31. deep | thorough-understanding | § (whole-chapter synthesis) **[3]**

Non-metals and metals both react with oxygen, yet the oxides they form behave oppositely when dissolved in water. Explain this difference in behaviour with one example each. Additionally, explain how the electronic configurations of a typical metal and a typical non-metal make their reaction with each other a 'give-and-take' process, and what type of compound results from this process.

◆ Metals and Non-metals

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Model Answer

Metal oxides vs. Non-metal oxides in water:
Metals combine with oxygen to form **basic oxides**, which dissolve in water to give bases (alkalis).
Example: $CaO + H_2O \rightarrow Ca(OH)_2$ (calcium hydroxide — a base)

Non-metals form **acidic oxides**, which dissolve in water to give acids.
Example: $SO_3 + H_2O \rightarrow H_2SO_4$ (sulphuric acid)

Give-and-take process (ionic bonding):
Metals have 1–3 electrons in their outermost shell, which they tend to **lose**, forming positive ions. Non-metals have 5–7 electrons and tend to **gain** electrons to complete their octet, forming negative ions.
For example, sodium (2,8,1) loses one electron to chlorine (2,8,7), forming Na^+ and Cl^-. The oppositely charged ions attract each other, producing an **ionic (electrovalent) compound** — sodium chloride (NaCl).

Source: Chapter 3 — Metals and Non-metals, Section 3.3

Explanation

- Examiners expect you to name the type of oxide (basic/acidic), give a **chemical equation** as the example, and name the product.
- For the give-and-take part: mention electron loss (metal → cation) and electron gain (non-metal → anion) with a specific example (Na + Cl is standard from the textbook), and name the resulting compound type as **ionic/electrovalent**.
- Do not forget to write the resulting compound name — that is usually the last mark point.

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