YouLearn

Occurrence of Metals in Nature

Metals are found in the Earth's crust either in their **free state** (native or uncombined form) or **combined state** (as compounds), depending on their reactivity.

- **Low reactivity metals** (at the bottom of the **reactivity series**) exist in free state because they are chemically inert and do not react easily—like silver (Ag), copper (Cu), gold (Au), and platinum (Pt).
- **Highly reactive metals**, such as sodium (Na) and magnesium (Mg), are found only in combined state as compounds because they must react to form stable compounds.
- **Copper and silver** are exceptions, occurring in both free and combined states.

> 💡 **Key Insight:** Understanding this distinction matters because it determines how easily a metal can be extracted—low reactivity metals are already "free," while others require breaking compounds.

This foundation sets up why extraction processes vary by reactivity.

Minerals and Ores

Minerals are naturally occurring elements or compounds found in the Earth's crust.

Ores are specific minerals with a high percentage of metal where extraction is economically profitable (easy and cost-effective).

- Not all minerals are ores; for example, among rocks containing metal X, only the one with high metal content becomes an ore.
- **Key relationship:** All ores are minerals, but not all minerals are ores.

> ℹ️ **Note:** Analogy—All KitKats are chocolates, but not all chocolates (Dairy Milk, Silk, etc.) are KitKats. Similarly, ores are a subset of minerals selected for high metal yield.

Ores come in types like **oxide ores**, **sulphide ores**, **carbonate ores**, **sulphate ores**, and **halide ores**, each named by composition and metal source.

This distinction is crucial because only ores are mined and processed for metal extraction.

Metallurgy: Overview of Extraction Steps

Metallurgy refers to the series of processes used to extract metals from their ores.

There are **three major steps** in metallurgy, building progressively from raw ore to pure metal:

1. **Concentration of ore** (or enrichment): Remove impurities.
2. **Conversion of concentrated ore to metal oxide**.
3. **Refining the impure metal** (detailed later, but involves purification).

These steps ensure efficient, profitable extraction by first purifying, then simplifying the chemical form, and finally isolating the metal.

Step 1: Concentration of Ore (Enrichment)

Raw ore contains the metal compound plus **gangue** (unwanted impurities like sand, rocky material, earth particles, or animal remains).

Concentration removes gangue to increase metal percentage, using methods based on physical or chemical property differences between ore and gangue (e.g., magnetic, density).

- Result: **Concentrated ore** with much higher metal content.
- Specific methods (beyond Class 10 scope here) clean the ore effectively.

Without this step, extraction would be inefficient due to low metal yield amid impurities.

Step 2: Conversion to Metal Oxide

Regardless of ore type (sulphide, carbonate, etc.), concentrated ore is first converted to **metal oxide** because obtaining metal from oxide is easier via simple reduction (removal of oxygen).

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Two key processes:

- **Roasting**: Heat concentrated **sulphide ore** strongly in excess air (presence of O_2). Sulphide undergoes oxidation to metal oxide + SO_2 gas.

$$2\backslash\text{ce}ZnS + 3\backslash\text{ce}O2 \xrightarrow{\text{heat}} 2\backslash\text{ce}ZnO + 2\backslash\text{ce}SO2$$

…

- **Calcination**: Heat concentrated **carbonate ore** in absence or limited supply of air. Carbonate decomposes to metal oxide + CO_2 gas.

$$\backslash\text{ce}ZnCO3 \xrightarrow{\text{heat}} \backslash\text{ce}ZnO + \backslash\text{ce}CO2$$

…

> ⚠️ **Warning:** Roasting requires excess air (for oxidation), while calcination uses limited/no air (to avoid oxidation, focus on decomposition). s s

This step unifies ore forms into oxides, simplifying the next reduction. s

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Extraction by Reactivity: Low Reactivity Metals

Metals low in the **reactivity series** (e.g., Hg, Cu) are least reactive; their oxides reduce to metal by **heating alone** (thermal decomposition or self-reduction). …

Example: Mercury from cinnabar (HgS, sulphide ore):

1. Roasting: $\backslash\mathrm{ce}2HgS + 3O2 - > 2HgO + 2SO2$
2. Heat HgO: $\backslash\mathrm{ce}2HgO - > 2Hg + O2$ → Mercury vapor condenses to metal.

Example: Copper from Cu_2S:

- Roasting gives mixture of Cu_2S and Cu_2O.
- Self-reduction on heating: $\backslash\mathrm{ce}2Cu2O + Cu2S - > 6Cu + SO2$

These metals need no external reducing agent because their oxides are unstable at high heat—building on low reactivity from occurrence section.

Extraction by Reactivity: Moderately Reactive Metals

Moderately reactive metals (middle of reactivity series, e.g., iron, zinc) occur as sulphides/carbonates; first convert to oxide via roasting/calcination, then reduce using **reducing agents**.

Primary method: Carbon (coke):

- Heat metal oxide with carbon: $\backslash\mathrm{ce}ZnO + C - > Zn + CO$ → Zinc metal.

Alternative: Displacement using highly reactive metals (more reactive metal displaces less reactive one from oxide):

- Example: Al reduces MnO_2 (Al more reactive than Mn): Produces molten Mn due to exothermic heat.

> 💡 **Key Insight:** Displacement reactions release so much heat that metals form in **molten state**.

Application: Thermite reaction

- Fe_2O_3 + Al (Al more reactive than Fe): $\backslash\mathrm{ce}Fe2O3 + 2Al - > 2Fe + Al2O3$ → Molten iron.
- Used to join railway tracks or repair cracked machine parts (molten Fe fills cracks).

This connects to reactivity series: More reactive metals act as reducers for less reactive oxides. [s] [s]

Reactivity Series Grouping for Extraction

The reactivity series is divided into three groups, dictating reduction methods:

Group	Examples	Reduction Method	Why It Works
Low reactivity (bottom)	Cu, Hg, Ag, Au	Heating alone (self-reduction)	Oxides unstable, decompose easily [s]
Moderate reactivity (middle)	Fe, Zn	Carbon or highly reactive metals (displacement) [s]	Need external reducers [s]
High reactivity (top)	Na, Ca, Mg, Al	(Advanced electrolytic methods, not detailed here) [s]	Too reactive for simple reduction [s]

Different techniques per group ensure efficient extraction based on reactivity. [s] [s]

Chemical Definitions of Metals and Non-Metals

Metals are elements that lose electrons to form positive ions (**cations**), also known as **electropositive elements**. [s] [s] This tendency to readily donate electrons drives their reactivity, as atoms seek stability by achieving a complete octet (or duet for K shell) in their valence shell, mimicking noble gases. [s]

For example, **sodium** (atomic number 11, configuration $2, 8, 1$) loses one electron to form Na^{+} (stable $2, 8$ configuration), while **magnesium** (atomic number 12, configuration $2, 8, 2$) loses two electrons to form Mg^{2+}. [s] [s] Think of it like two kids drinking bitter kadha: Chintu (sodium) finishes one glass quickly with enthusiasm, while Pintu (magnesium) takes longer for two glasses—sodium reacts faster because it loses electrons more readily. [s] [s]

Non-metals are elements that gain electrons to form negative ions (**anions**), known as **electronegative elements**. ⓢ ⓢ Unlike metals, they accept electrons for stability.

Oxygen (atomic number 8, configuration $2, 6$) gains two electrons to form O^{2-} (stable $2, 8$). ⓢ This opposite behavior explains why metals and non-metals react differently with substances. ⓢ

Reactivity Series: Arrangement and Significance

The **reactivity series** arranges metals in decreasing order of reactivity: Potassium (K), Sodium (Na), Calcium (Ca), Magnesium (Mg), Aluminium (Al), Zinc (Zn), Iron (Fe), Lead (Pb), Hydrogen (H), Copper (Cu), Mercury (Hg), Silver (Ag), Gold (Au)—mnemonic: "Kaku Nana Chaat Maange Aalu Jara Fake PaRe Hain Chaap". ⓢ ⓢ

Reactivity depends on how easily a metal loses electrons to form cations—the easier the loss, the higher in the series. ⓢ ⓢ Metals above hydrogen lose electrons more readily than hydrogen, making them more reactive; those below lose electrons less readily, so less reactive. ⓢ ⓢ

> 💡 **Key Insight:** Hydrogen is included despite not being a metal because, like metals, it loses an electron to form H^+. ⓢ ⓢ This allows comparison: metals above H displace H from acids/water; those below do not. ⓢ ⓢ

This series predicts reactions: a more reactive metal displaces a less reactive one from compounds. ⓢ For instance, any metal picked (e.g., Zn) displaces those below it but not above. ⓢ ⓢ

Reactions of Metals with Oxygen, Water, and Acids

Metals react with oxygen to form **metal oxides**, which are basic in nature. ⓢ ⓢ

- $2Mg + O_2 \rightarrow 2MgO$ (magnesium oxide, reacts with water to form base $Mg(OH)_2$). ⓢ
- Reactivity varies: Highly reactive like Na, K react at room temperature; Mg, Zn need heating; low reactivity like Cu need strong heating. ⓢ ⓢ **Amphoteric oxides** (e.g., Al_2O_3, ZnO) show both acidic and basic behavior. ⓢ ⓢ

Metals react with water/steam to evolve hydrogen, forming hydroxide (cold water) or oxide (steam). [s] [s] Intensity depends on position in reactivity series—K, Na react violently with cold water (H_2 catches fire); Mg with hot water/steam; Cu, Ag, Au do not react. [s] [s]

- Reason low reactivity metals don't react: They are below H, cannot displace H from water. [s] [s]

Metals react with dilute acids to form salt + H_2, displacing H (only if above H in series). [s] [s] Cu, Ag below H do not react. [s] Nitric acid (strong oxidant) oxidizes H_2 to water/NO_x. [s]

Reactions of Non-Metals

Non-metals react with oxygen to form **non-metallic oxides**, acidic or neutral. [s]

- $C + O_2 \rightarrow CO_2$ (acidic); under different conditions, $2C + O_2 \rightarrow 2CO$ (neutral). [s]

Non-metals do not react with water/steam to evolve H_2 or with dilute acids to displace H, because they gain (not lose) electrons—cannot reduce H^+. [s] [s] Analogy: Acid's H^+ demands electrons; metals donate, non-metals need them themselves, so no reaction. [s] [s]

A more reactive non-metal displaces a less reactive one from its salt solution (e.g., Cl_2 displaces Br^- from NaBr). [s]

How Metals and Non-Metals React: Ionic Bond Formation

Metals lose electrons (form cations); non-metals gain (form anions), leading to **electrostatic attraction** forming **ionic bonds** when electron transfer occurs. [s] [s] This achieves noble gas configuration for stability. [s]

Example: Na $(2,8,1)$ loses 1e^- → Na^+ $(2,8)$; Cl $(2,8,7)$ gains 1e^- → Cl^- $(2,8)$; Na^+ + Cl^- → NaCl (ionic compound). [s] [s]

$MgCl_2$: Mg loses 2e^- to two Cl atoms. [s] Ionic bonds form between cations (metals) and anions (non-metals). [s] Covalent bonds form via sharing (non-metal + non-metal). [s]

Properties of Ionic Compounds

Ionic compounds are solids (strong electrostatic forces), hard, with high melting/boiling points (energy needed to break forces). [s] [s]

Soluble in polar solvents (e.g., water) like dissolves like; insoluble in non-polar (e.g., CCl_4). ⓢ

Conduct electricity in molten/aqueous state (ions free to move) but not solid (ions fixed). ⓢ ⓢ

Extraction of Metals: Metallurgy Overview

Metals occur as **minerals** (natural compounds/elements in Earth's crust); economically viable high-metal-content minerals are **ores**. ⓢ ⓢ All ores are minerals, but not vice versa (analogy: All KitKats are chocolates, but not all chocolates are KitKats). ⓢ ⓢ

Steps in metallurgy (extraction from ore): ⓢ

1. **Concentration/enrichment of ore**: Remove gangue (impurities like sand) using physical/chemical differences. ⓢ ⓢ
2. **Conversion to metal oxide**: Roasting (sulphide ores, heat in excess air → oxide + SO_2, e.g., $2ZnS + 3O_2 \rightarrow 2ZnO + 2SO_2$); Calcination (carbonate ores, heat limited air → oxide + CO_2, e.g., $CaCO_3 \rightarrow CaO + CO_2$). ⓢ ⓢ Easier to reduce oxides than sulphides/carbonates. ⓢ
3. **Reduction to metal** (depends on reactivity):

Reactivity Category	Examples	Reduction Method	Details ⓢ ⓢ
Low (bottom of series)	Cu, Hg, Ag, Au	Simple heating	Cinnabar (HgS) roasted to HgO, heated → Hg; Cu_2S roasted + self-reaction → Cu. ⓢ ⓢ
Moderate (middle)	Fe, Zn	Reducing agents (carbon/coke) or displacement	$ZnO + C \rightarrow Zn + CO$; Thermit: $3MnO_2 + 4Al \rightarrow 2Al_2O_3 + 3Mn + \text{heat}$ (molten metal for welding). ⓢ ⓢ
High (top)	Na, K, Ca, Mg, Al	Electrolytic reduction of molten chloride/oxide	Molten NaCl: Cathode ($^-$) $Na^+ + e^- \rightarrow Na$ (reduction); Anode ($^+$) $2Cl^- \rightarrow Cl_2 + 2e^-$ (oxidation). Al from molten $\mathrm{Al_2O}$\$_3\$. ⓢ ⓢ Metals have higher oxygen affinity than carbon. ⓢ

4. **Refining**: Purify impure metal via **electrolytic refining** (e.g., Cu). Impure Cu anode dissolves (oxidation → Cu^{2+}), pure Cu deposits on cathode (reduction); soluble impurities in solution, insoluble as anode mud. [s] [s]

Aqua regia (3:1 conc. $HCl:HNO_3$) dissolves noble metals like Au, Pt (fuming, corrosive). [s] [s]

Flowchart summary: Ore → Concentration → Oxide (roast/calcine) → Reduction (heating/reducing agent/electrolysis) → Refining. [s]

Corrosion and Prevention

Corrosion is slow oxidation by air/moisture forming oxide/carbonate/sulphide layers that destroy metal. [s]

- Ag → black Ag_2S. [s]
- Cu → green basic $CuCO_3 \cdot Cu(OH)_2$. [s]
- Fe → reddish-brown hydrated $Fe_2O_3 \cdot xH_2O$ (rust). [s]

Prevention: Block air/moisture contact via painting, oiling, greasing, galvanising (Zn coating), chrome plating, anodising, alloying. [s] [s]

Alloys: Homogeneous Mixtures for Enhanced Properties

Alloys are homogeneous mixtures of two+ metals or metal + non-metal, improving hardness, tensile strength, corrosion resistance; less ductile/malleable; melting point higher/lower than constituents. [s] [s] Prepared by melting primary metal, dissolving others, cooling. [s]

Alloy	Composition	Uses ⓢ
Brass	Cu + Zn	Utensils, cartridges
Bronze	Cu + Sn	Statues, medals, coins, machines
Solder	Pb + Sn	Soldering wires
Duralumin	Al + Cu + Mn + Mg	Aircraft bodies, kitchenware
Steel	Fe + C (small %)	Tools, buildings

Alloys harder, corrosion-resistant than pure metals (e.g., duralumin stronger than Al for aircraft). ⓢ If Hg present, called **amalgam**. ⓢ Color often differs (e.g., pink brass from white Cu/Zn). ⓢ

Corrosion: The Chemical Attack on Metals

Corrosion is the process where metals deteriorate due to an attack by air and moisture, leading to the formation of unwanted layers like oxides, carbonates, or sulphides that destroy the metal's surface. ⓢ This happens because the metal reacts chemically with oxygen, water, or other atmospheric components, weakening its structure over time—this is why metals like those in bridges or vehicles need protection to last longer. ⓢ

Think of corrosion like rust eating away at a nail left outside: the "attack" creates a flaky layer that flakes off, exposing more metal to further damage. ⓢ

Corrosion of Silver: Formation of Black Sulphide Layer

Silver articles turn black after some time when exposed to air because silver reacts with sulphur present in the air to form a coating of **silver sulphide** (Ag_2S), which appears black. ⓢ

This black layer deposits on the surface, tarnishing the shine—unlike pure oxidation, here sulphur from pollutants like hydrogen sulphide in air plays a key role. ⓢ As a result, silver jewelry or utensils loses its luster unless cleaned or protected. ⓢ

Corrosion of Copper: Green Patina Development

Copper reacts with moist **carbon dioxide** in the air, slowly losing its shiny brown surface and gaining a green coat known as **basic copper carbonate** ($CuCO_3 \cdot Cu(OH)_2$). ⓢ

Building on the general corrosion mechanism, this reaction involves both moisture and CO_2 forming a protective yet unsightly patina—for example, the green layer on the Statue of Liberty is this basic copper carbonate. ⓢ This matters because it shows how corrosion can alter appearance and, over time, structural integrity, though the green layer somewhat slows further corrosion. ⓢ

Corrosion of Iron: Rust Formation

Iron undergoes rusting when attacked by air and moisture, forming a reddish-brown flaky substance called **rust** with the formula $Fe_2O_3 \cdot xH_2O$. ⓢ

This reddish-brown coating flakes off easily, unlike copper's stable patina, exposing fresh iron to more attack and accelerating damage—this is why iron objects like tools or railings rust quickly in humid environments. ⓢ Rusting connects directly to the core corrosion process, as oxygen and water oxidize iron into hydrated iron(III) oxide. ⓢ

Preventing Rusting of Iron: Blocking Air and Moisture

To prevent rusting, coat iron articles with a layer that blocks contact between the iron and air/moisture, stopping the corrosion reaction at its source. ⓢ ⓢ

Common methods include:

- **Painting**: Applies a paint layer as a barrier. ⓢ
- **Oiling or greasing**: Uses oil/grease to repel water. ⓢ
- **Galvanization**: Deposits a zinc layer (detailed below). ⓢ
- **Chrome plating**: Coats with chromium for shine and protection. ⓢ
- **Anodizing**: Creates an oxide layer on metals like aluminum. ⓢ
- **Alloying**: Mixes iron with other metals to enhance resistance (detailed below). ⓢ

These techniques work *because* they physically prevent the reactants (O_2 and H_2O) from reaching the iron, directly addressing the corrosion mechanism explained earlier. [s]

Galvanization: Zinc Coating for Iron Protection

Galvanization (or galvanised iron) protects iron from rusting by coating it with a layer of **zinc**. [s]

Zinc acts as a sacrificial layer—even if scratched, zinc corrodes preferentially (due to its higher reactivity), shielding the iron underneath; this builds on prevention by providing both barrier and electrochemical protection. [s] For instance, galvanized buckets or wires resist rust longer than plain iron. [s]

Alloys: Improving Metal Properties to Resist Corrosion

Alloying is an effective method to enhance or improve the properties of metals, including resistance to corrosion. [s]

By mixing a metal with other elements, alloys gain better strength, hardness, or durability—for example, stainless steel (iron + chromium) resists rust better than pure iron, connecting to prevention strategies by altering the metal's inherent reactivity. [s] This matters because alloys allow metals to be used in harsh environments without additional coatings. [s]

Physical Properties of Metals

Metals exhibit distinct physical characteristics that make them useful in everyday applications, such as in wires, utensils, and jewelry. [s]

These properties include:

- **Hardness**: Metals are generally hard, though the degree varies from metal to metal (e.g., not all are equally hard). [s]
- **Malleability**: Metals can be beaten into thin sheets without breaking (e.g., gold and silver are among the best malleable metals). [s] [s]

- **Ductility**: Metals can be drawn into thin wires (e.g., copper wires used in electric circuits). [s] [s]
- **High melting and boiling points**: Metals require high temperatures to melt or boil (e.g., copper melts at $1084^\circ\mathrm{C}$). [s]
- **State at room temperature**: Metals are solid at room temperature. [s]
- **Sonorous**: Metals produce a ringing sound when struck (e.g., school bells or temple bells make a "ting" sound). [s]
- **Good conductors of heat**: Metals transfer heat efficiently (e.g., silver is the best conductor, so cooking utensils are made of copper or aluminum). [...]
- **Good conductors of electricity**: Metals allow electric current to flow (e.g., copper and silver; silver is the best, copper next). [...]
- **Lustrous**: Metals have a shiny appearance due to **metallic lustre**. [s] [s]

These properties arise because of the metallic bonding in metals, where electrons are delocalized, allowing flexibility, conductivity, and shine—this is why metals are ideal for construction and electrical uses. [s]

Physical Properties of Non-Metals

Non-metals have properties opposite to metals, making them brittle and poor conductors, which explains their use in insulators like plastics. [s] Building on metals' properties, simply negate them for non-metals to remember easily. [s]

Key properties include:

- **Brittle**: Non-metals are not malleable or ductile; they crumble or break easily when hammered. [s]
- **Poor conductors of heat and electricity**: Non-metals do not conduct well (unlike metals). [s]
- **Dull appearance**: Non-metals lack lustre (no shiny surface). [s]

- **Generally soft**: Non-metals are softer than metals. [s]
- **Non-sonorous**: Non-metals do not produce a ringing sound when hit. [s]
- **Variable states at room temperature**: Non-metals exist as solids, liquids, or gases (unlike metals, which are solids). [s]
- **Low melting and boiling points**: Non-metals melt and boil at lower temperatures (compared to metals). [s]

This contrast highlights why non-metals are used for insulation (e.g., plastic covering on wires prevents shocks). [s]

Property	Metals	Non-Metals
Malleability/Ductility	Yes (sheets/wires)	No (brittle)
Conductivity (heat/electricity)	Good	Poor
Lustre	Shiny	Dull
Hardness	Generally hard	Generally soft
Sonorous	Yes (ringing sound)	No
State at Room Temp	Solid	Solid, liquid, or gas
Melting/Boiling Pt	High	Low
...		

Key Exceptions in Physical Properties

While general rules hold, chemistry has exceptions—crucial for exam questions, as they test nuanced understanding. [s] These build on the properties above, showing not all metals/non-metals fit perfectly. [s]

For Metals:

- **Softness**: Alkali metals like lithium (Li), sodium (Na), potassium (K) are soft enough to cut with a knife. [s] [s]
- **Liquid at room temperature**: Mercury (Hg) is the only metal liquid at room temp. [s] [s]
- **Low melting points**: Sodium, potassium, gallium, caesium have low melting points (they melt slightly above room temp, unlike general high points). [s]
- Metals lose shine over time due to formation of oxide/carbonate layer from air gases (**corrosion**), acquiring a dull appearance. [s] [s]

For Non-Metals:

- **Conducts electricity**: Graphite (form of carbon) conducts due to delocalized electrons. [s]
- **Lustre**: Iodine and graphite have a shiny appearance. [s] [s]
- **Hardness and high melting point**: Diamond (form of carbon) is the hardest natural substance with very high melting/boiling points. [...]

> ℹ️ **Note:** Only two elements are liquid at room temperature: **Mercury** (metal) and **bromine** (non-metal). Common exam questions: (1) Which elements are liquid at room temp? (Ans: Hg, Br); (2) Which metal? (Hg); (3) Which non-metal? (Br). [...]

These exceptions matter because they appear in "which of the following" questions—memorize them as starred points for quick recall. [s]

Chemical Definition of Metals and Non-Metals

Chemically, metals are defined by their behavior in forming ions, building on physical properties to explain reactivity. [s]

Metals are elements that **lose electrons** to form **positive ions** (cations), also called **electropositive elements.** [s] [s]

- Example: Sodium (atomic no. 11, config. $2, 8, 1$) loses 1 electron → Na^+ (octet complete, stable). [s]

- Magnesium (atomic no. 12, config. $2, 8, 2$) loses 2 electrons → Mg^{2+}. [s]

This electron loss tendency makes metals reactive, contrasting non-metals (which gain electrons to form negative ions, **electronegative**). Why it matters: Explains all chemical reactions of metals. [s]

Reactions of Metals with Oxygen

Metals react with oxygen to form **metal oxides**, which are generally **basic in nature** (react with acids to form salt + water). [s] [s] Reactivity varies per **reactivity series** (more reactive metals react easily). [...]

General reaction: $\text{Metal} + \text{O}_2 \rightarrow \text{Metal Oxide}$ [s]

- Examples:
  - $4Na + \text{O}_2 \rightarrow 2Na_2\text{O}$ (sodium oxide, basic). [s]
- $2Mg + \text{O}_2 \rightarrow 2Mg\text{O}$ (magnesium oxide). [s]
- $2Zn + \text{O}_2 \rightarrow 2Zn\text{O}$. [s]

Conditions depend on reactivity (top of series react faster):

- Highly reactive (e.g., Na, K): React even at room temp, often catch fire; stored in kerosene to prevent. [s]
- Moderately reactive (Mg, Zn): React on heating. [s]
- Less reactive (Cu): Form CuO on strong heating (amphoteric—shows both acidic/basic behavior). [s]

Metal oxides are basic: E.g., $Mg\text{O} + \text{H}_2\text{O} \rightarrow Mg(\text{OH})_2$ (base). [s] Mnemonic: "Mogo" (Metal Oxide = Basic Oxide). [s]

Amphoteric oxides (show acidic + basic nature): $Al_2\text{O}_3$, $Zn\text{O}$.

- With acid: $Al_2\text{O}_3 + 6\text{HCl} \rightarrow 2AlCl_3 + 3\text{H}_2\text{O}$ (basic behavior). [s]

- With base: $Al_2O_3 + 2NaOH \rightarrow 2NaAlO_2 + H_2O$ (sodium aluminate, acidic behavior). [s] Similar for ZnO (forms sodium zincate with NaOH). [s]

Exam Q: "What are amphoteric oxides? Give 2 examples." Ans: Oxides showing both behaviors; Al_2O_3, ZnO. [s]

Reactions of Metals with Water

Metals displace hydrogen from water to form metal hydroxide/oxide + H2 gas; product depends on water form and metal reactivity. [s] [s] Only metals above H in reactivity series react. [s]

General: $\mathrm{Metal} + H_2O \rightarrow \mathrm{Metal\ Hydroxide/Oxide} + H_2$

- **Liquid water (cold/hot)** → Metal hydroxide.
    - K: Violent with cold water → $KOH + H_2$ (H2 catches fire). [s]
- Na: Vigorous → $NaOH + H_2$ (fire). [s]
- Ca: Less violent → $Ca(OH)_2 + H_2$. [s]
- Mg: With hot water → $Mg(OH)_2 + H_2$. [s]

Steam (gaseous water) → Metal oxide (needs more energy).

- Fe, Zn, Al: E.g., $3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$. [s]

Metals below H (Cu, Ag, Au) do not react, even with steam—they're less reactive than H. [s] [s] This connects to displacement: More reactive metal pushes out less reactive H. [s]

Reactions of Metals with Acids and Special Cases

Metals (above H in series) react with **dilute acids** to form salt + H2 by displacing H+. [s]

General: $\mathrm{Metal} + \mathrm{Acid} \rightarrow \mathrm{Salt} + H_2$

- Examples: $Mg + 2HCl \rightarrow MgCl_2 + H_2$; $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$. [s]

- Highly reactive (Na, Mg, Ca): Violent. [s]
- Less reactive (Fe): Slow. [s]
- Cu, Ag, Au: No reaction (below H). [s] [s]

Nitric acid (HNO3) exception: Strong oxidant; oxidizes H2 to H2O, itself reduces to nitrogen oxides (NOx like N2O)—no H2 gas. [s]

Aqua regia ("royal water"): Freshly prepared 1:3 conc. HNO3:HCl mixture; highly corrosive, dissolves noble metals like Au, Pt (alone, acids can't). [...]

Reactions with Salt Solutions and Reactivity Series

More reactive metal displaces less reactive metal from its salt solution (**displacement reaction**). [s] [s]

- E.g., Fe + $CuSO_4$ → $FeSO_4$ + Cu (Fe > Cu in series). [s]
- Zn + $CuSO_4$ → $ZnSO_4$ + Cu (Zn > Cu). [s]
- Reverse (Cu + $ZnSO_4$): No reaction (Cu < Zn). [s]

Reactivity series: Metals arranged in **decreasing order of reactivity** (top to bottom: reactivity decreases). [...]

Mnemonic: **K**a**K**u **N**a**N**a **C**haat **M**an**g**e **A**alu **J**ara **F**ik**e** **P**a**D**e **H**ai **C**ha**P** (K, Na, Ca, Mg, Al, Zn, Fe, Pb/H, Cu, Hg, Ag, Au). [s]

- Any metal displaces all below it from compounds/solutions; cannot displace above it. [s]
- H included: Metals above H displace H from acids/water; below don't. [s] [s]
- Purpose: Predicts reaction feasibility (e.g., why Na reacts violently, Cu inert). [s] [s]

> 💡 **Key Insight:** Reactivity decreases top-to-bottom because higher metals lose electrons more easily. [s]

This series unifies all reactions—master it to solve 80% of chapter problems. [s]

Defining Metals and Non-Metals

Metals are elements that lose electrons to form positive ions, known as **electropositive elements** because they donate electrons easily. ⓢ This tendency allows metals to achieve stability by emptying their valence shell, mimicking noble gas configurations. ⓢ

In contrast, **non-metals** are elements that gain electrons to form negative ions, called **electronegative elements**. ⓢ ⓢ For example, oxygen (atomic number 8, configuration 2,6) gains two electrons to form O^{2-}, completing its octet and becoming stable. ⓢ ⓢ This electron gain creates a negative charge, distinguishing non-metals from metals. ⓢ

> 💡 **Key Insight:** Reactivity in elements stems from their atoms' tendency to achieve a stable, noble gas-like octet in their valence shell. ⓢ

Physical Properties of Metals

Metals exhibit unique physical traits that make them useful in daily applications, building on their electropositive nature which allows free-moving electrons.

Malleability is the ability of metals to be beaten into thin sheets without breaking. ⓢ ⓢ Gold and silver are among the best examples, used historically for ornaments. ⓢ

Ductility allows metals to be drawn into thin wires. ⓢ This property enables copper wires in electrical circuits, combining with conductivity. ⓢ

Metals are **good conductors of heat** and **electricity** due to delocalized electrons. ⓢ ⓢ Silver is the best heat conductor, used in cooking utensils; copper excels in electricity conduction for wires, insulated with PVC to prevent shocks. ⓢ ⓢ

Lustrous (shiny) surfaces define metals' metallic luster. ⓢ ⓢ However, exposure to air causes corrosion, forming oxide/carbonate layers that dull the shine, like silver turning black from Ag_2S. ⓢ ⓢ

Chemical Properties of Non-Metals

Non-metals' electronegative nature leads to distinct reactions, opposite to metals which form basic oxides.

Non-metals react with oxygen to form **non-metallic oxides**, which are acidic or neutral in nature. [s] [s] For instance, carbon + oxygen yields CO_2 (acidic) or CO (neutral under different conditions). [s] [s] This contrasts with metal oxides, which are basic. [s]

Unlike metals, non-metals **do not react with water or dilute acids** to produce hydrogen gas. [s] [s] The reason: acids release H^+ ions needing electrons to form H_2, but non-metals cannot donate electrons—they gain them instead. [⋯] Metals succeed because they lose electrons readily. [s] [s]

> ⚠️ **Warning:** Non-metals fail to displace hydrogen from acids as they are electron acceptors, not donors—hence no reaction occurs. [s] [s]

A more reactive non-metal displaces a less reactive one from its salt solution. [s] [s] Chlorine (more reactive) displaces bromine: $Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$. [s] This mirrors metal reactivity series but for non-metals. [s]

How Metals React with Non-Metals: Ionic Bond Formation

Metals and non-metals react via **electron transfer**, forming **ionic bonds** between cations and anions, driven by the drive for octet stability like noble gases. [⋯]

The process builds on definitions: metals lose electrons (form cations), non-metals gain (form anions). [s] [s]

1. Sodium (Na, config 2,8,1) loses 1 electron → Na^+ (2,8). [s]
2. Chlorine (Cl, config 2,8,7) gains 1 electron → Cl^- (2,8,8). [s]
3. Opposite charges attract, forming **NaCl** with ionic bond. [s] [s]

For **MgCl_2**: Mg (2,8,2) loses 2 electrons → Mg^{2+}; two Cl atoms each gain 1 → two Cl^-. [⋯]

> ℹ️ **Note:** Ionic bonds form only via complete electron transfer (loss and gain), between metals and non-metals, unlike covalent bonds from sharing. [s] [s]

Definition: An ionic bond forms when one atom donates electrons (metal) and another accepts (non-metal) to achieve noble gas configuration. [s] [s]

Properties of Ionic Compounds

Ionic compounds, formed from metal-non-metal reactions, have properties arising from strong electrostatic attractions between oppositely charged ions.

They are **solid and hard** at room temperature due to these forces. [s] [s]

High melting and boiling points result because breaking these forces requires significant energy. [s] [s]

Solubility: Polar (like dissolves like), soluble in polar solvents (e.g., water) but insoluble in non-polar (e.g., benzene). […] Ions' charges make them polar. [s]

Conductivity: Do not conduct in solid state (ions fixed by attractions). [s] [s] Conduct in molten or aqueous states as ions are free to move. […]

Property	Reason	Example
Solid/Hard	Strong inter-ionic forces	NaCl crystals [s]
High MP/BP	Energy needed to overcome forces	NaCl melts at 801°C [s]
Soluble in water	Polar-polar interaction	NaCl dissolves [s]
Conducts when molten	Free ion movement	Molten NaCl [s]

> 💡 **Key Insight:** Conductivity depends on charged particle (ion) movement—absent in solid ionic compounds. [s]

Extraction of Metals: From Ore to Pure Metal

Extraction involves converting ores to metals via steps tailored to reactivity, ensuring comprehensive coverage from concentration to refining.

1. **Ore Concentration:** Clean ore to get concentrated ore. [s]
2. **Conversion to Oxide:** Roasting (sulfides → oxide + SO2), calcination (carbonates → oxide + CO2). [...]
3. **Reduction to Metal:** Depends on reactivity series position. [...]

Reactivity Level	Examples	Reduction Method	Why It Works
Low (e.g., Hg from cinnabar)	Hg, Cu	Simple heating	Weak metal-oxygen bond [s] [s]
Moderate	Fe, Zn	Carbon (coke) as reducing agent	Carbon displaces oxygen [s] [s]
High	Na, K, Ca, Mg, Al	Electrolytic reduction of molten chlorides/oxides	Stable oxides; metals have higher oxygen affinity than carbon [...]

Electrolytic Reduction for high reactivity: Pass current through molten salt. [s]

- **Molten NaCl:** At cathode (negative): $Na^{+} + e^{-} \rightarrow Na$ (reduction, gain e-). [...] At anode (positive): $2Cl^{-} \rightarrow Cl_2 + 2e^{-}$ (oxidation, lose e-). [s] [s]
- **Molten Al2O3:** Cathode: $Al^{3+} + 3e^{-} \rightarrow Al$; Anode: $2O^{2-} \rightarrow O_2 + 4e^{-}$. [...]

This separates pure metal at cathode. [s] [s]

Refining purifies metal via electrolytic refining. [...] Impure metal (anode) dissolves (oxidation: loses e- → ions), pure metal deposits at cathode (reduction: gains e-). [...] Soluble impurities dissolve; insoluble form **anode mud**. [s] Example: Cu with CuSO4 electrolyte—impure Cu → pure Cu. [...]

Step	Low Reactivity	Moderate	High
Ore to Oxide	Roasting/Calcination [s]	Roasting/Calcination [s]	Roasting/Calcination [s]
Reduction	Heating [s]	Carbon [s]	Electrolysis [s]
Refining	Electrolytic [s]	Electrolytic [s]	Electrolytic [s]

Alloys: Improving Metal Properties

Alloys are homogeneous mixtures of two or more metals or a metal and non-metal, formed by melting the primary metal and dissolving others, then cooling. [s] [s] This enhances properties like strength and corrosion resistance. [s]

Key properties:

- Harder, less ductile/malleable than constituents. [s]
- Resistant to corrosion. [s]
- Melting point higher or lower than constituents. [s]
- Pronounced properties vs. pure metals (e.g., duralumin stronger than Al for aircraft). [s] [s]

If mercury is involved, called **amalgam**. [s] [s] Color often differs (e.g., brass pinkish). [s]

Alloy	Composition	Uses
Brass	Cu + Zn	Utensils, cartridges [s]
Bronze	Cu + Sn	Statues, medals, coins [s]
Solder	Pb + Sn	Soldering wires [s]
Duralumin	Al + Cu + Mn + Mg	Aircraft bodies, kitchenware [s] [s]

Alloys prevent corrosion and improve utility, connecting to metals' luster loss issue. [s]

> 🔗 **Connection:** Alloys build on extraction by refining metals for better performance. s