METALS

NANYANG
GIRLS' HIGH
勤慎端樸

METALS

What do I need to know and understand about *metals*?

METALS

Properties of Metals – O’ Level Syllabus

- Describe the general physical properties of metals as solids having high melting and boiling points, malleable, good conductors of heat and electricity in terms of their structure.
- Describe alloys as a mixture of a metal with another element, *e.g.* brass and stainless steel.
- Identify representations of metals and alloys from diagrams of structures.
- Explain why alloys have different physical properties to their constituent elements.

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Reactivity Series – O' Level Syllabus

- Place in order of reactivity calcium, copper, (hydrogen), iron, lead, magnesium, potassium, silver, sodium and zinc by reference to:
  - → The reactions, if any, of the metals with water, steam and dilute hydrochloric acid.
  - → The reduction, if any, of their oxides by carbon and/or by hydrogen.
- Describe the reactivity series as related to the tendency of a metal to form its positive ion, illustrated by its reaction with:
  - → The aqueous ions of the other listed metals.
  - → The oxides of the other listed metals.
- Deduce the order of reactivity from a given set of experimental results.
- Describe the action of heat on the carbonates of the listed metals and relate thermal stability to the reactivity series.

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Extraction of Metals – O' Level Syllabus

- Describe the ease of obtaining metals from their ores by relating the elements to their positions in the reactivity series.

Recycling of Metals – O' Level Syllabus

- Describe metal ores as a finite resource and hence the need to recycle metals, *e.g.* recycling of iron.
- Discuss the social, economic and environmental issues of recycling metals.

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Iron – O' Level Syllabus

- Describe and explain the essential reactions in the extraction of iron using haematite, limestone and coke in the blast furnace.
- Describe steels as alloys which are a mixture of iron with carbon or other metals and how controlled use of these additives changes the properties of the iron, *e.g.* high carbon steels are strong but brittle whereas low carbon steels are softer and more easily shaped.
- State the uses of:
  - → Mild steel, *e.g.* car bodies, machinery.
  - → Stainless steel, *e.g.* chemical plants, cutlery and surgical instruments.

METALS

Iron – O' Level Syllabus

- Describe the essential conditions for the corrosion (rusting) of iron as the presence of oxygen and water. Prevention of rusting can be achieved by placing a barrier around the metal, *e.g.* painting, greasing, plastic coating and galvanising.
- Describe the sacrificial protection of iron by a more reactive metal in terms of the reactivity series where the more reactive metal corrodes preferentially, *e.g.* underwater pipes have a piece of magnesium attached to them.

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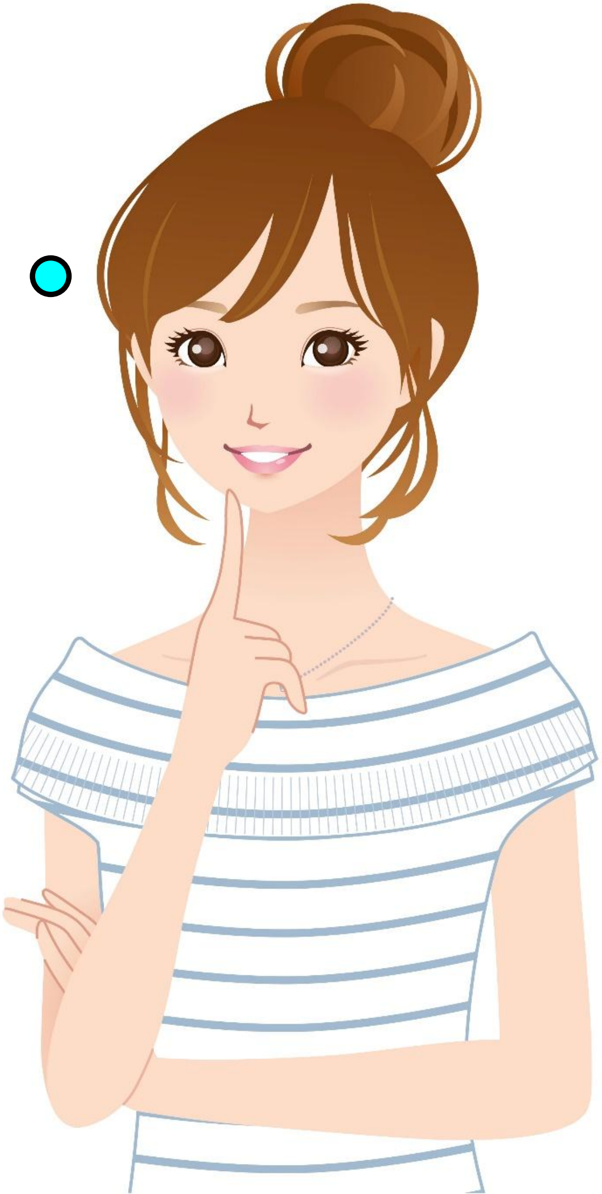
What are the general *properties* of metals? How are these properties related to their use?



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- Metals have *high melting points* and *high boiling points* because the electrostatic force of attraction between the metal cations and “sea” of delocalised electrons is strong.

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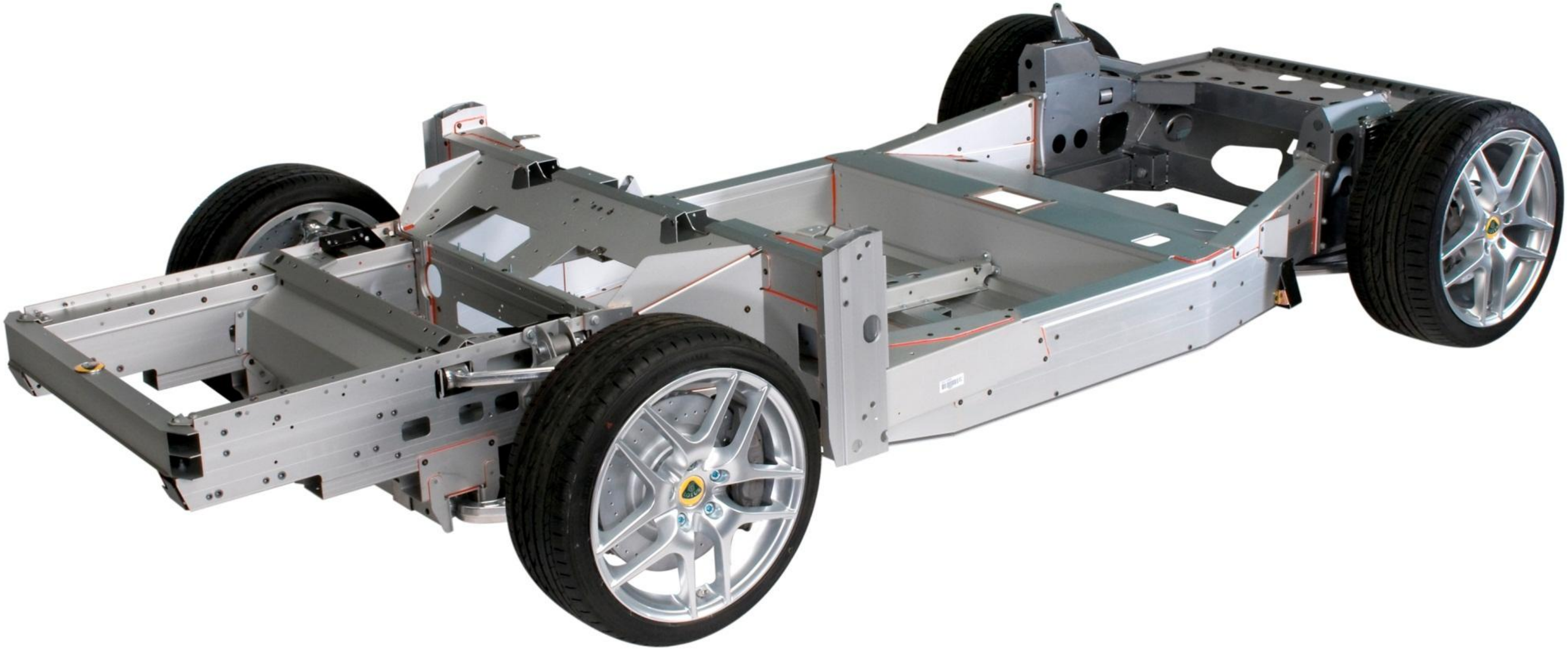

- Metals are *hard* and *strong* because the electrostatic force of attraction between the metal cations and “sea” of delocalised electrons is strong.

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- Metals are *malleable*. This means that they can be hammered or rolled into thin sheets. This is because the metal cations are able to slide over each other without the metallic bonds breaking.

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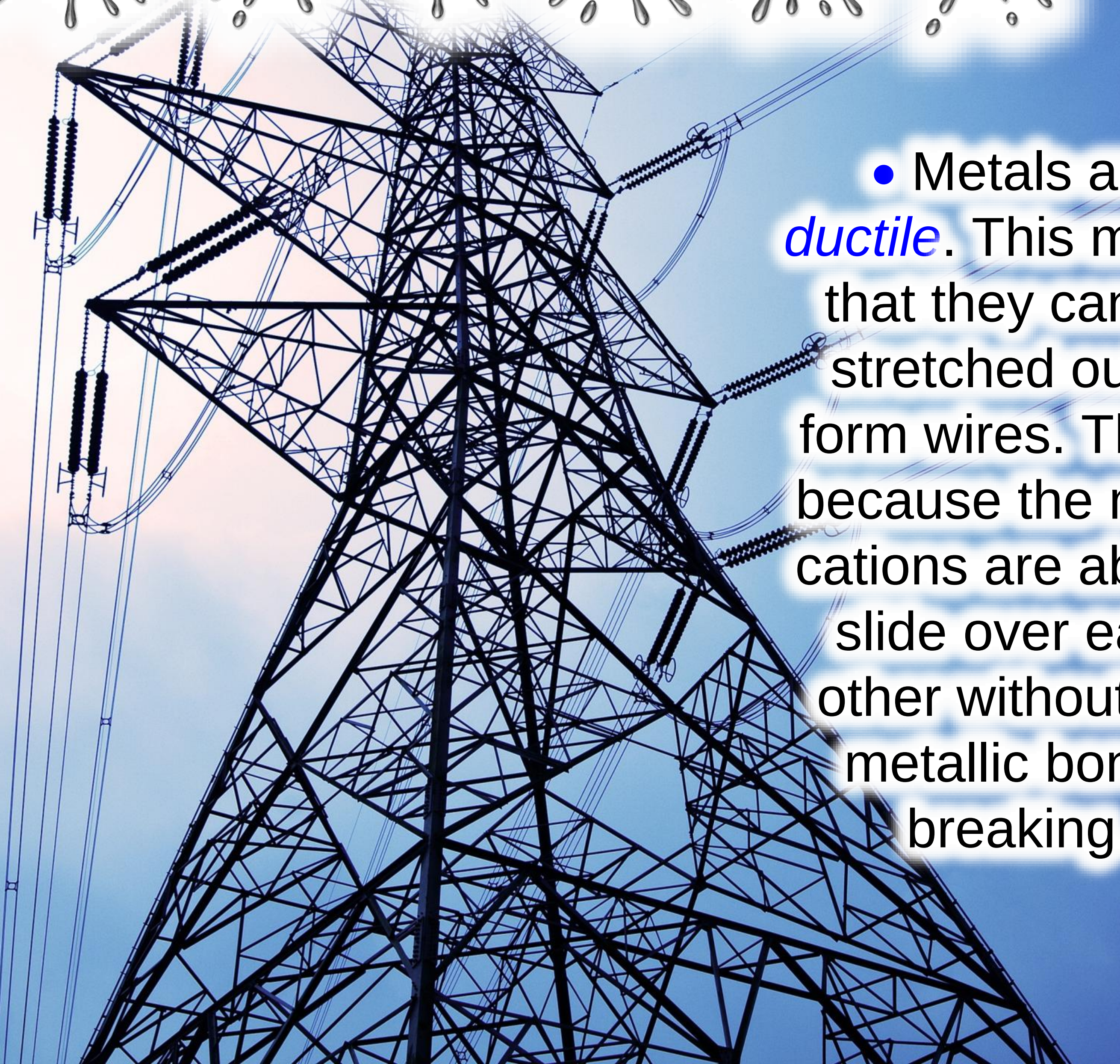

- Metals are *ductile*. This means that they can be stretched out to form wires. This is because the metal cations are able to slide over each other without the metallic bonds breaking.

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- Metals are *good conductors* of *heat*. This is because both the close packed cations in the crystal lattice, as well as the “sea” of delocalised electrons, are able to transfer kinetic energy through the metal.

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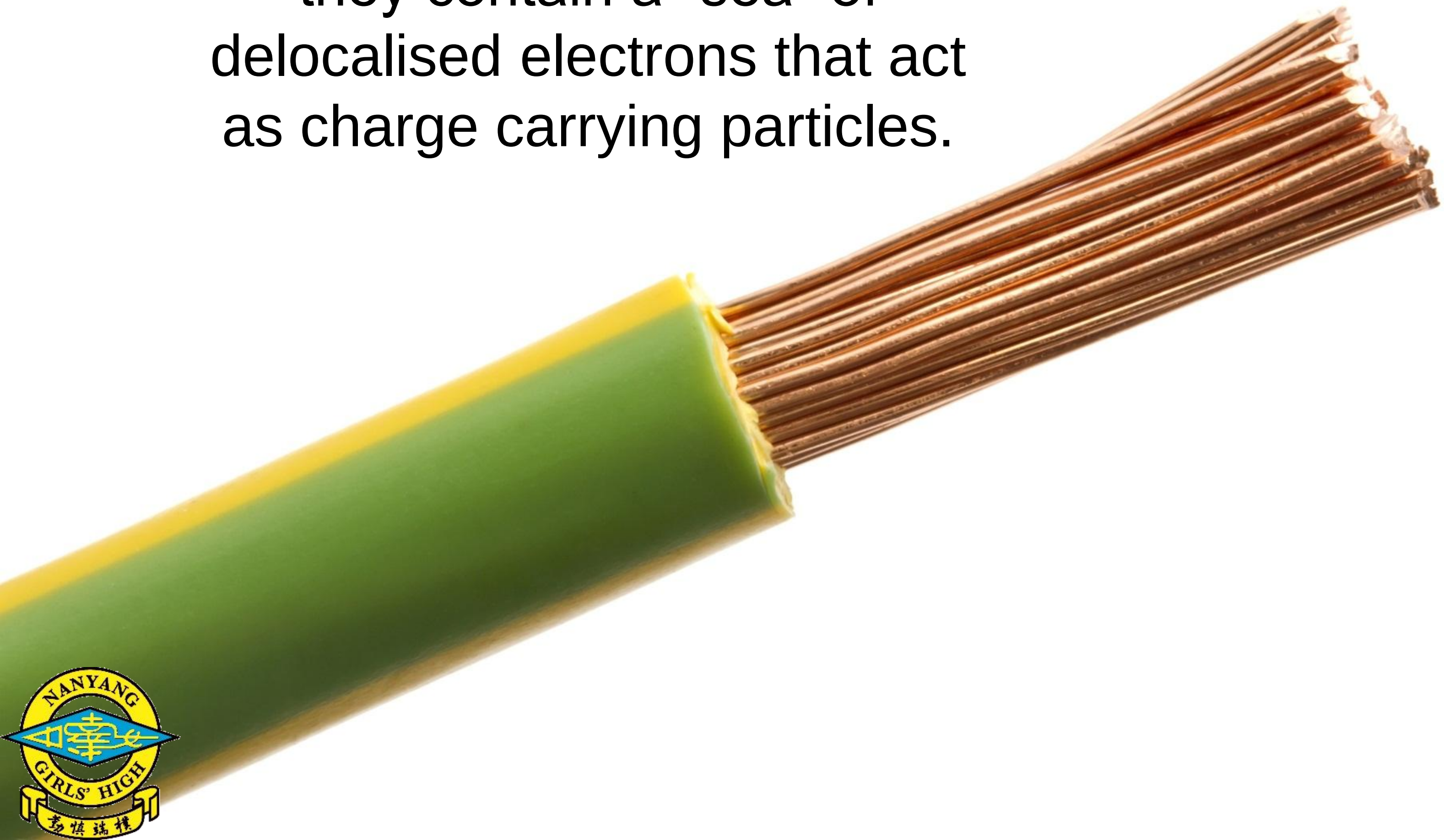
- Metals are *good conductors* of *electricity*. This is because they contain a “sea” of delocalised electrons that act as charge carrying particles.



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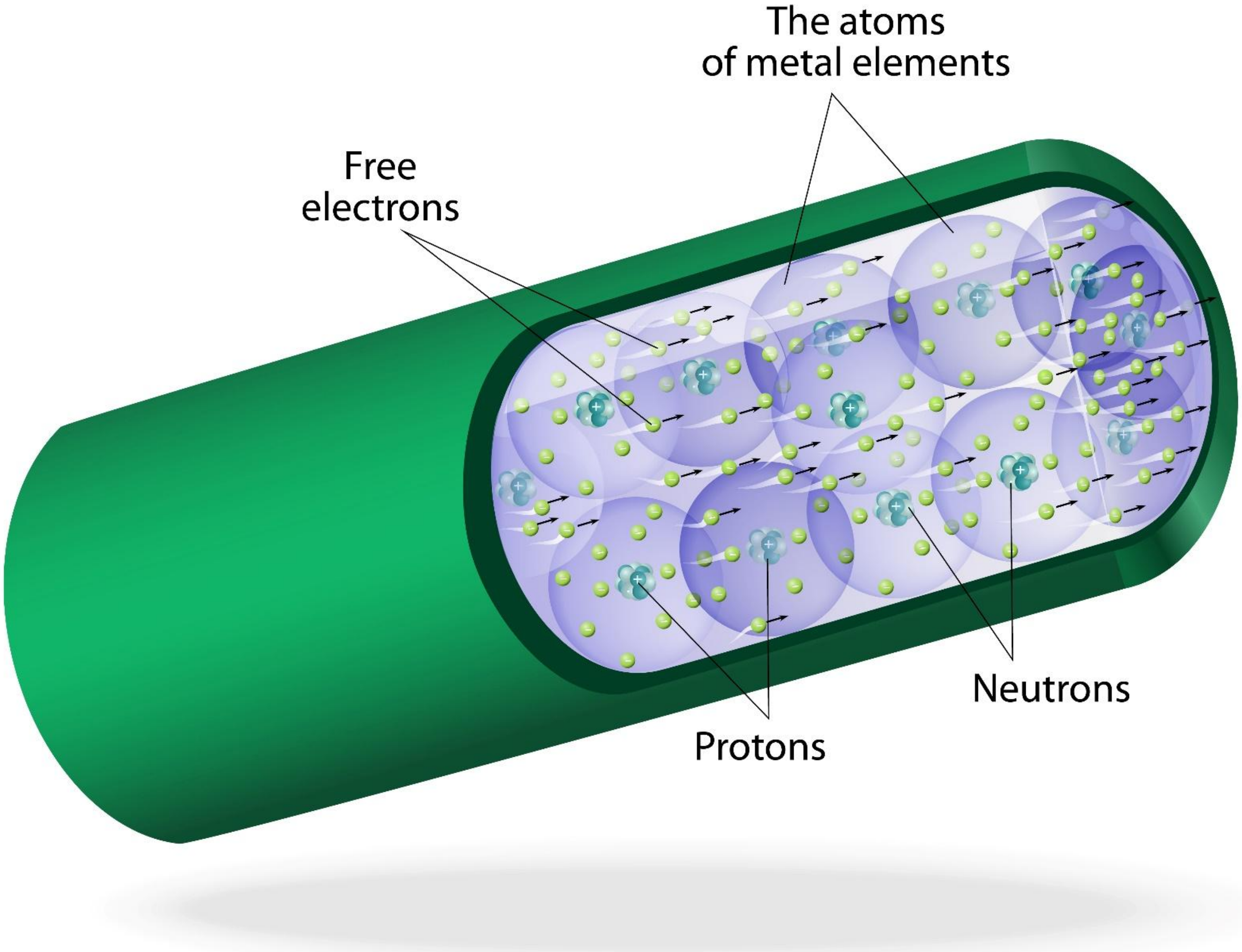
The atoms of metal elements
Free electrons
Neutrons
Protons

METALS

- Metals are *sonorous*. This means that metals produce a ringing sound when struck. This is due to the stiff, yet slightly elastic nature of the metal lattice. This means that the metal lattice can deform slightly, but will then return to its original shape, *i.e.* vibrate.

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- Metals have a *metallic lustre*. This is due to the way in which their delocalised electrons interact with light.

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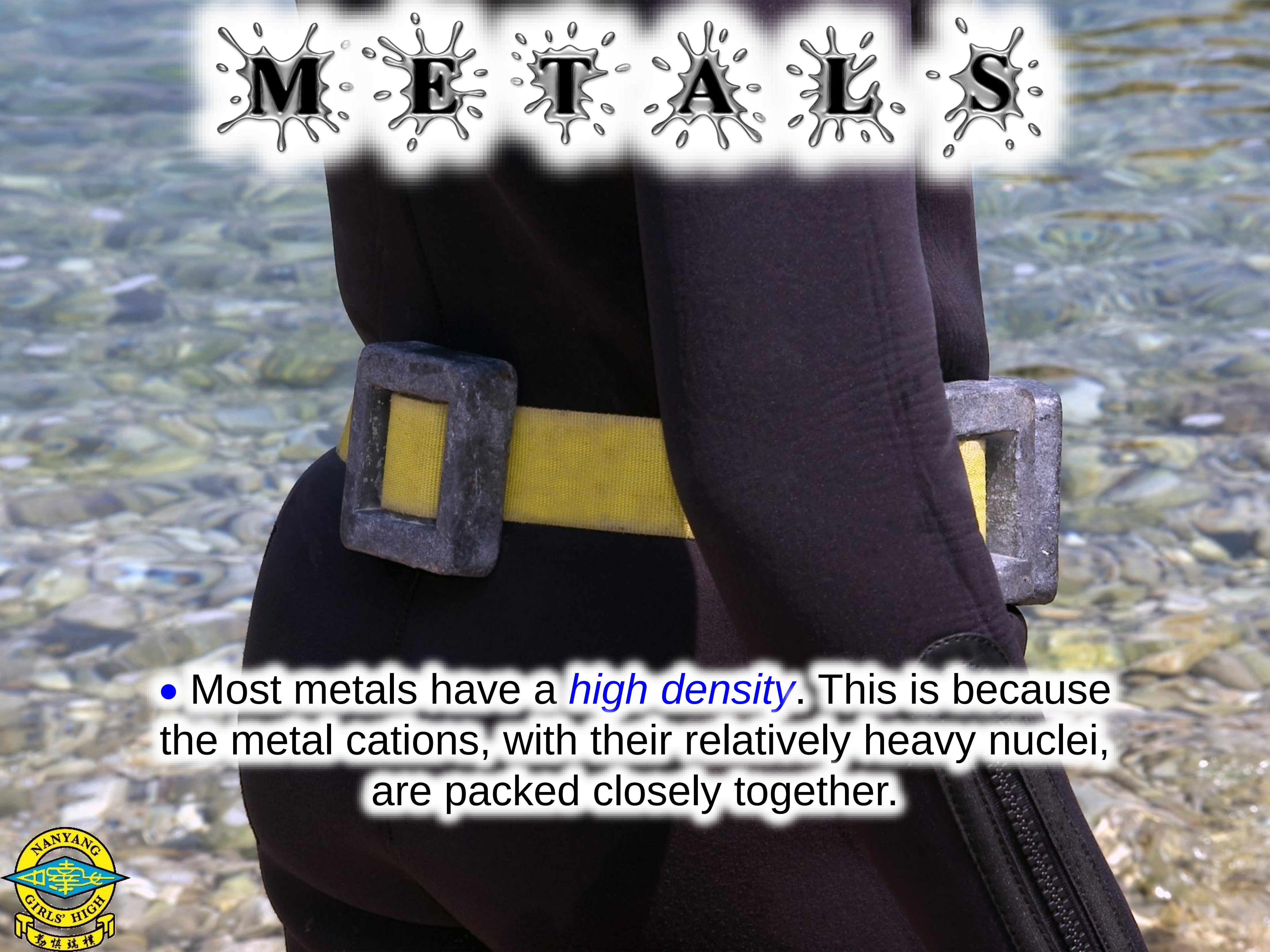

- Most metals have a *high density*. This is because the metal cations, with their relatively heavy nuclei, are packed closely together.

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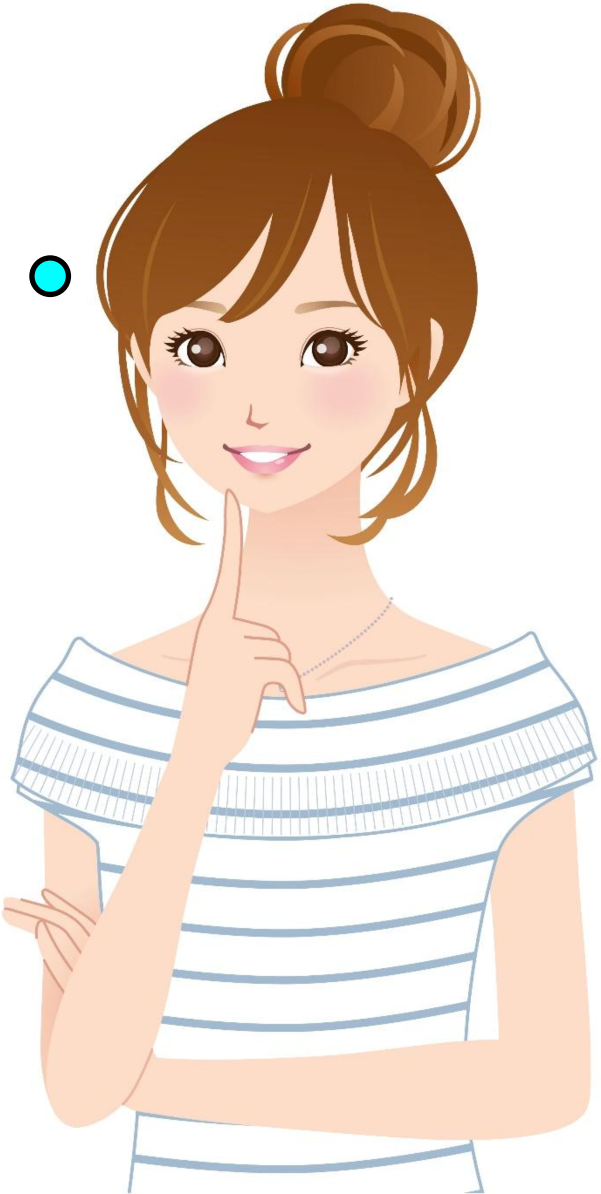

Those are all *physical* properties? What are the typical *chemical* properties of metals?

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Metal oxides are *basic**.

acid + base → salt + water

*Note: some metal oxides (Al_2O_3, PbO and ZnO) are *amphoteric*.

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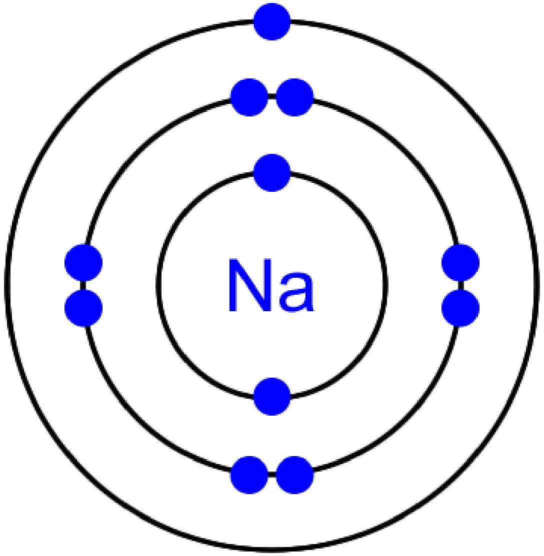

- Atoms of the metallic elements react by *losing* their valence electrons to form *cations* with the electronic configuration of a noble gas.

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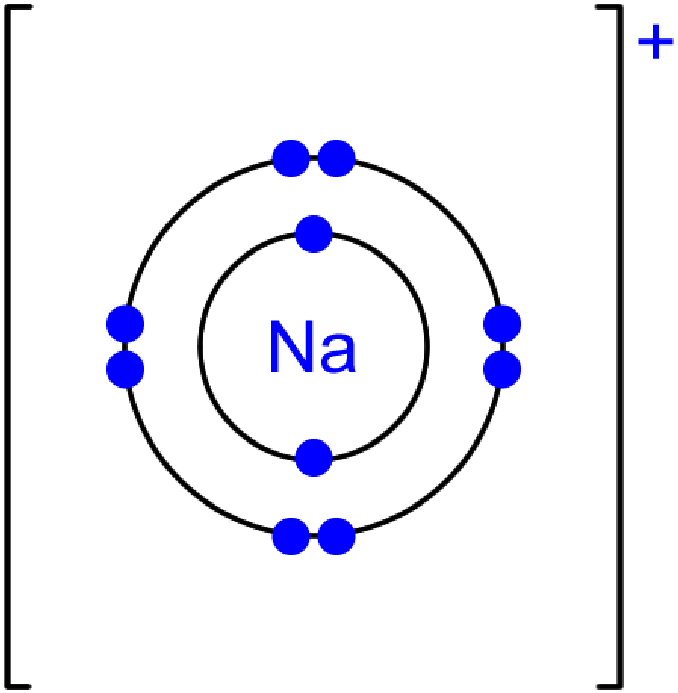
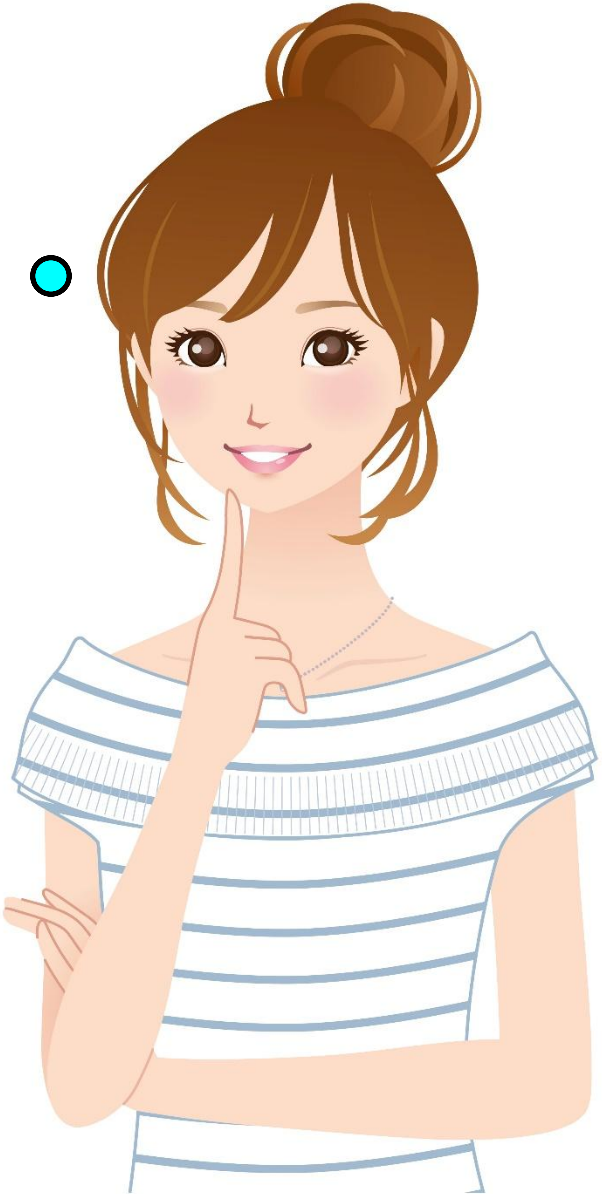


- Atoms of the metallic elements react by *losing* their valence electrons to form *cations* with the electronic configuration of a noble gas.

METALS

Could I please have a *summary* of the properties of metals?



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Property	Application
High melting points and boiling points	Cooking utensils, combustion engines
Hard and strong	Manufacture of aircraft, bridges, cars
Malleable	Motorcar bodies, household water pipes
Ductile	Electrical wires
Good conductors of heat	Cooking utensils
Good conductors of electricity	Electrical wires
Sonorous	Bells and tuning forks
Shiny (metallic lustre)	Jewellery and mirrors
High density	Diving weights and sports equipment
Form basic oxides	Regulate pH of acidic soils
React by losing electrons to form cations	Reducing agents

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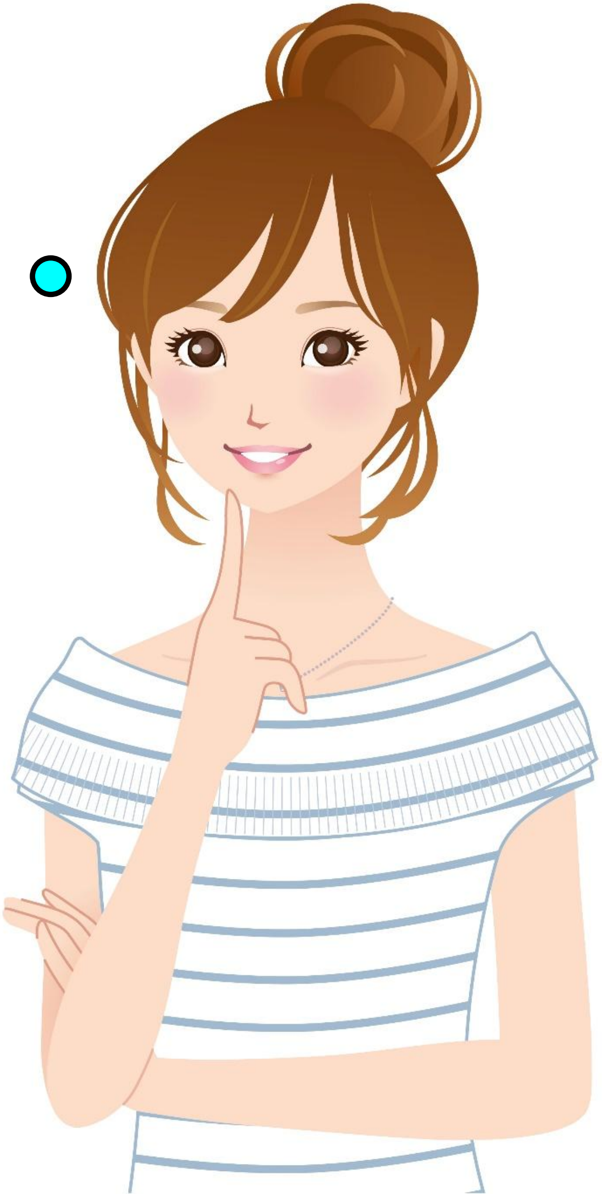
What is an *alloy*?
How are the properties of alloys different from those of a *pure* metal?



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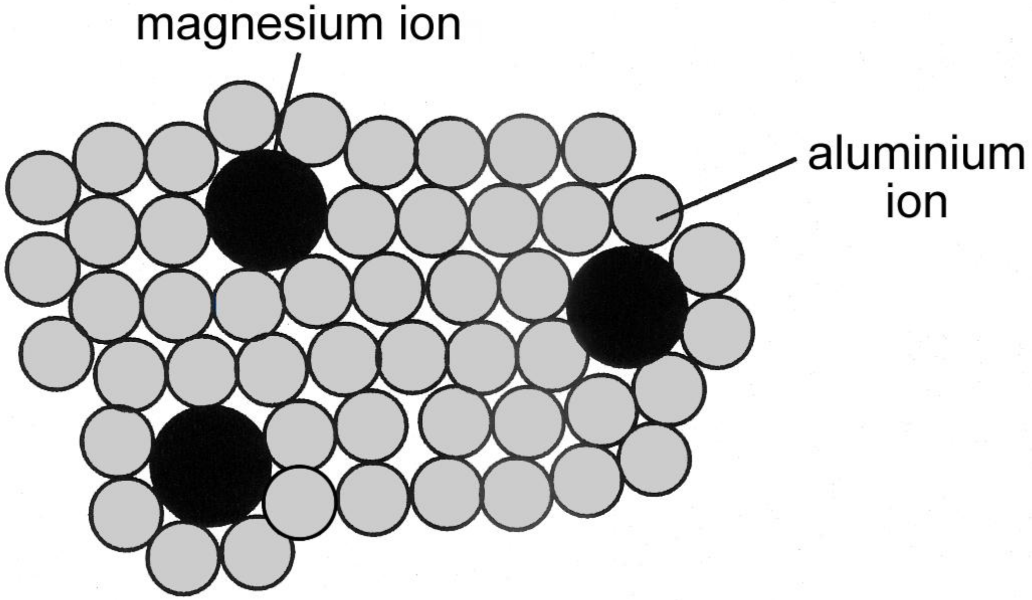

- An *alloy* is a mixture of a metal with another element. Alloying a metal is done by combining a metal with another metallic or non-metallic element in order to improve the properties of the original metal.

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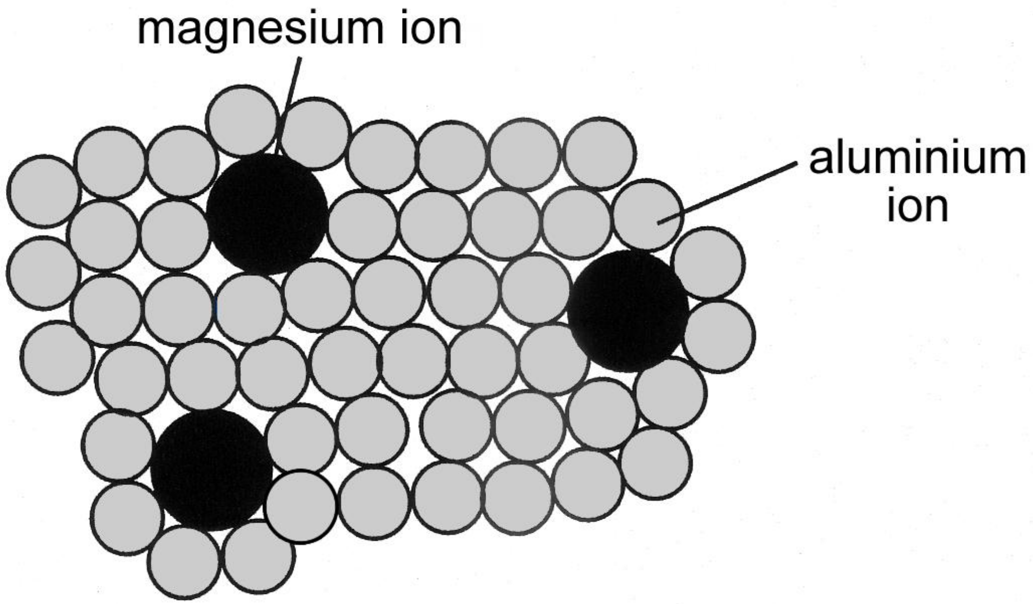

- Brass is an alloy of copper and zinc.
- Bronze is an alloy of copper and tin.

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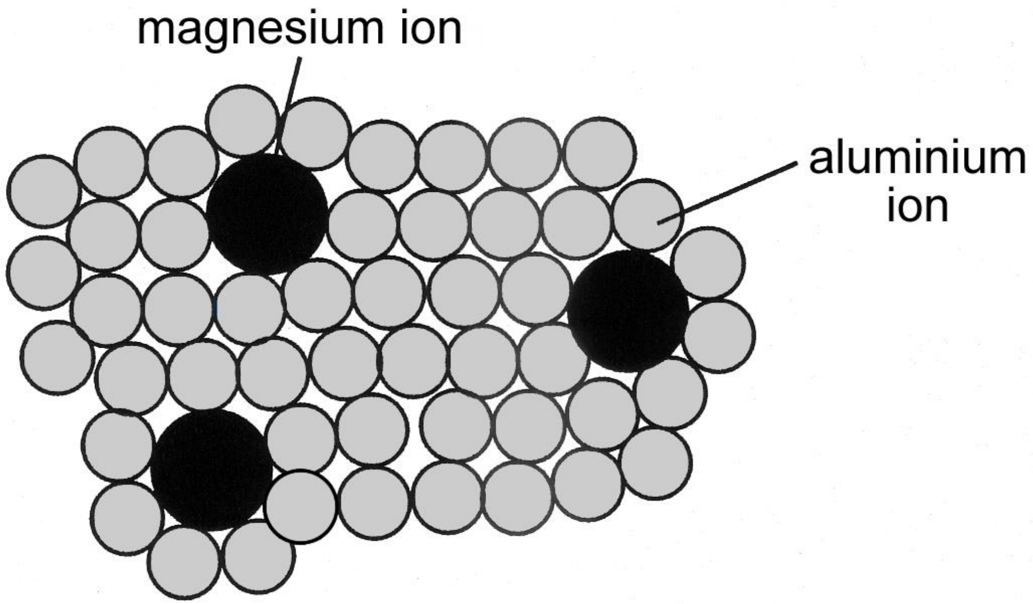

- Steel is an alloy of iron and (most commonly) carbon.
- Stainless steel is an alloy of iron and chromium.

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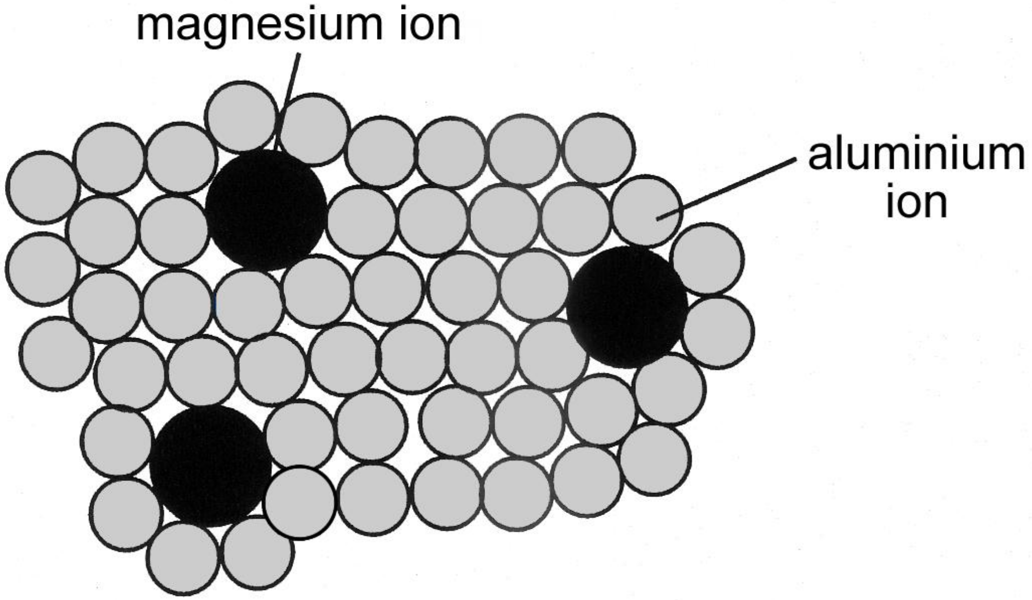

- The diagram above represents the structure of an alloy known as *duralumin*. The major component of this alloy is aluminium, with the minor components being copper, magnesium and manganese.

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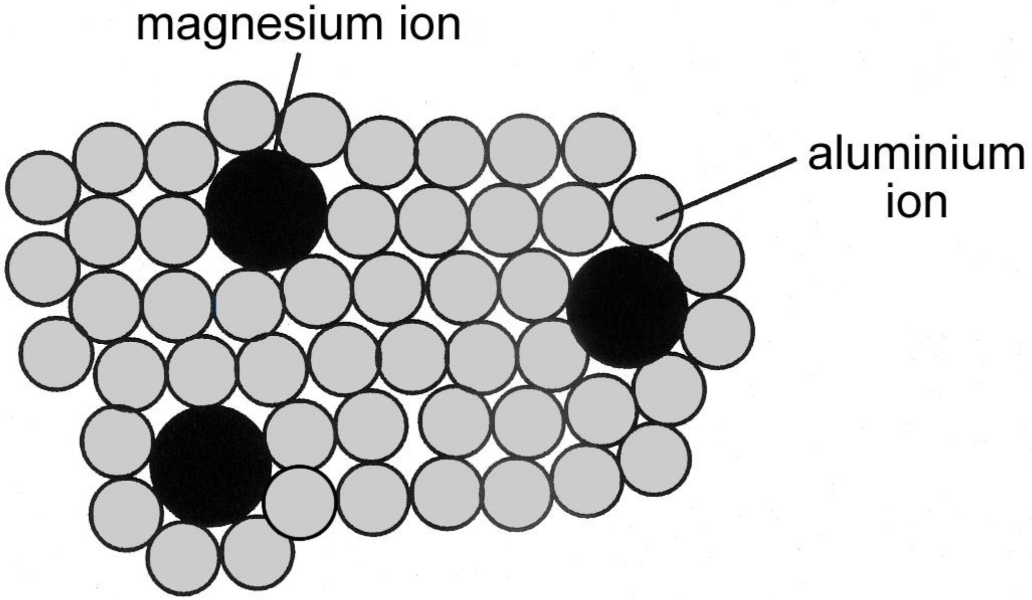

- The duralumin will be *harder* and *stronger* (less malleable and ductile) than the original aluminium because the large magnesium ions disrupt the regular, ordered, crystalline structure of the aluminium, making it difficult for the layers of aluminium ions to side over each other.

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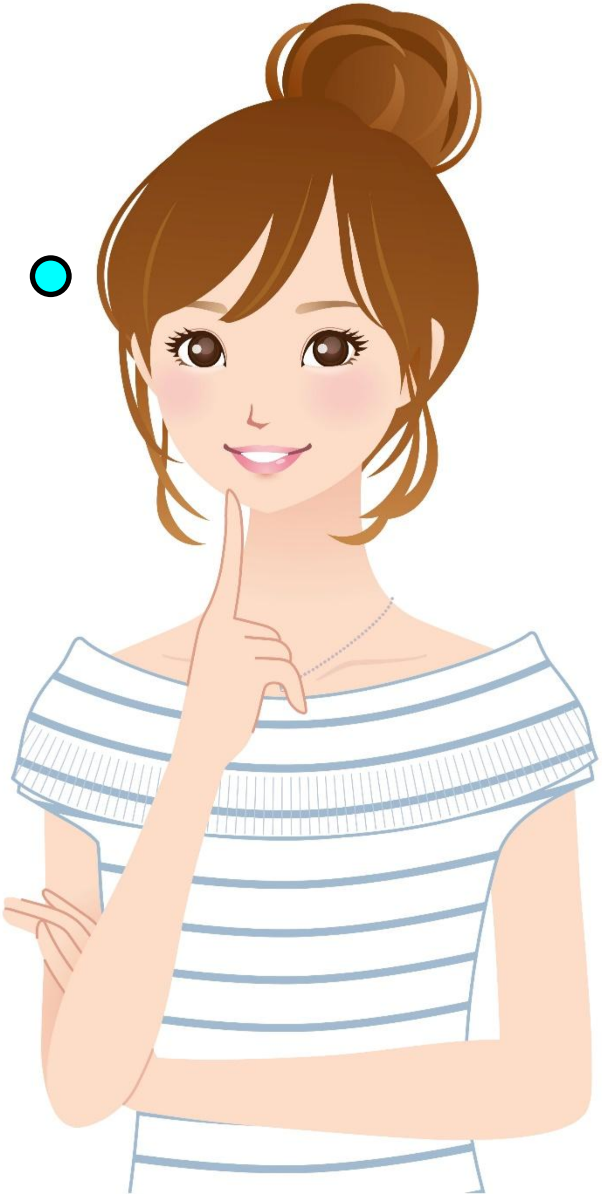
How can the reactivity of a metal be inferred from its *atomic structure*?



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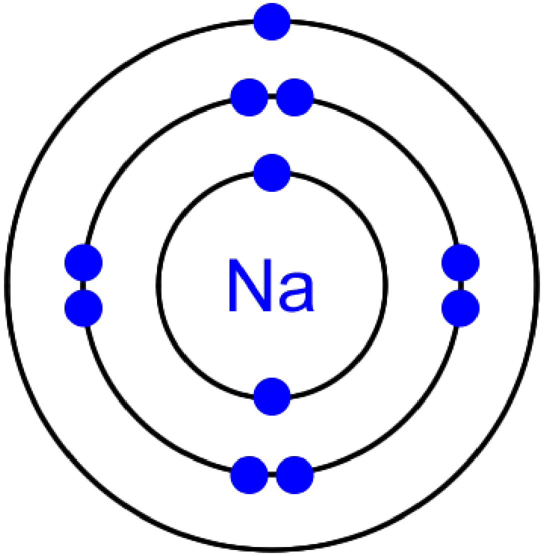

- Atoms of the metallic elements react by *losing* their valence electrons to form *cations* with the electronic configuration of a noble gas.

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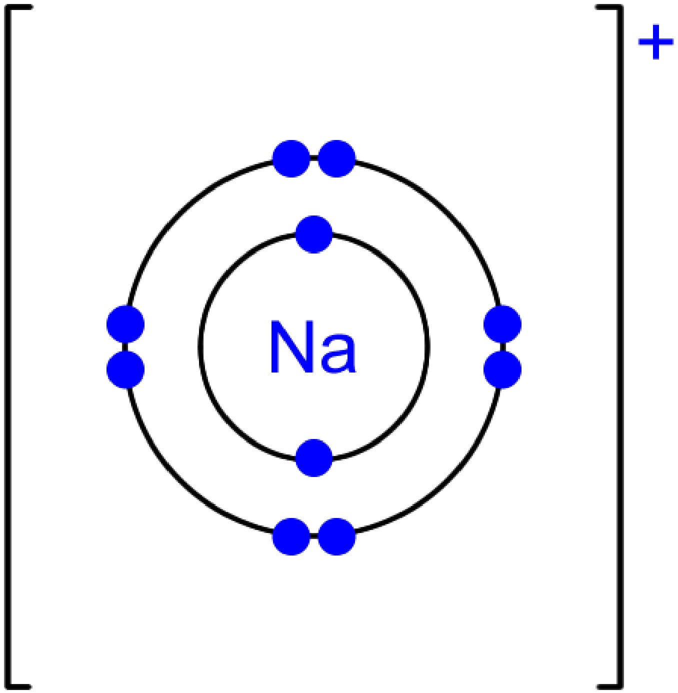

- The *more readily* the atoms of a metallic element lose their valence electrons (*i.e.* the *lower the energy* that is required to remove the valence electrons) then the *more reactive* the metal will be.

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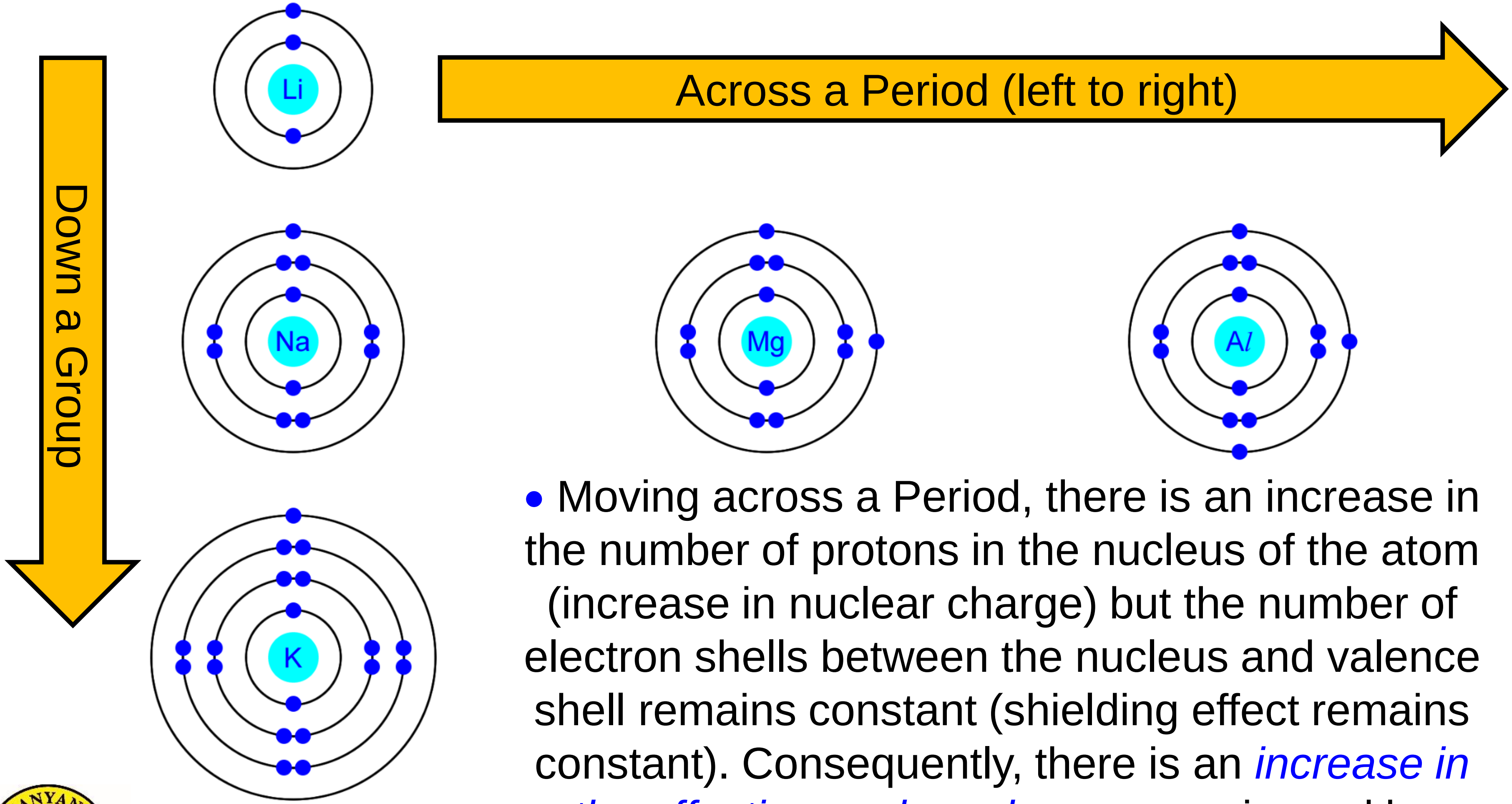

- Moving across a Period, there is an increase in the number of protons in the nucleus of the atom (increase in nuclear charge) but the number of electron shells between the nucleus and valence shell remains constant (shielding effect remains constant). Consequently, there is an *increase in the effective nuclear charge* experienced by electron(s) in the valence shell of the atom.

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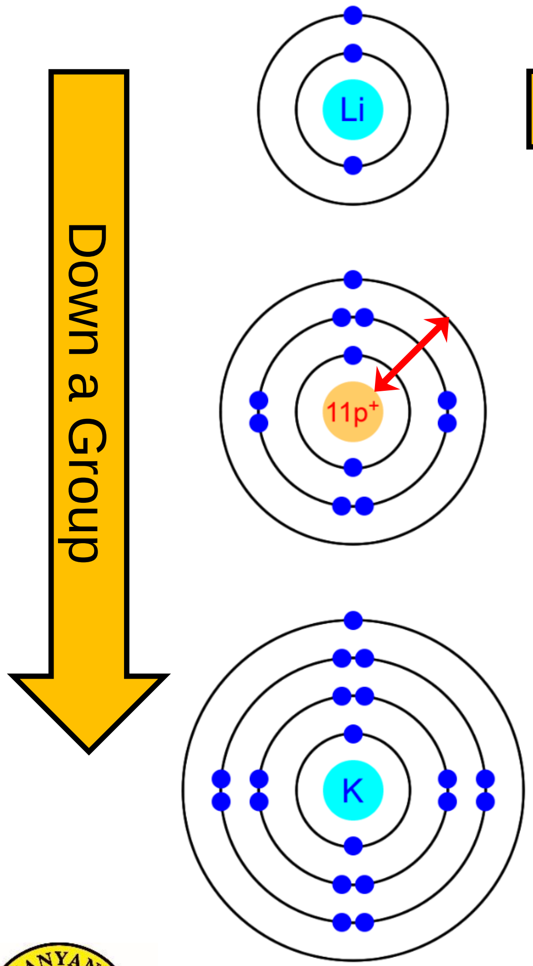

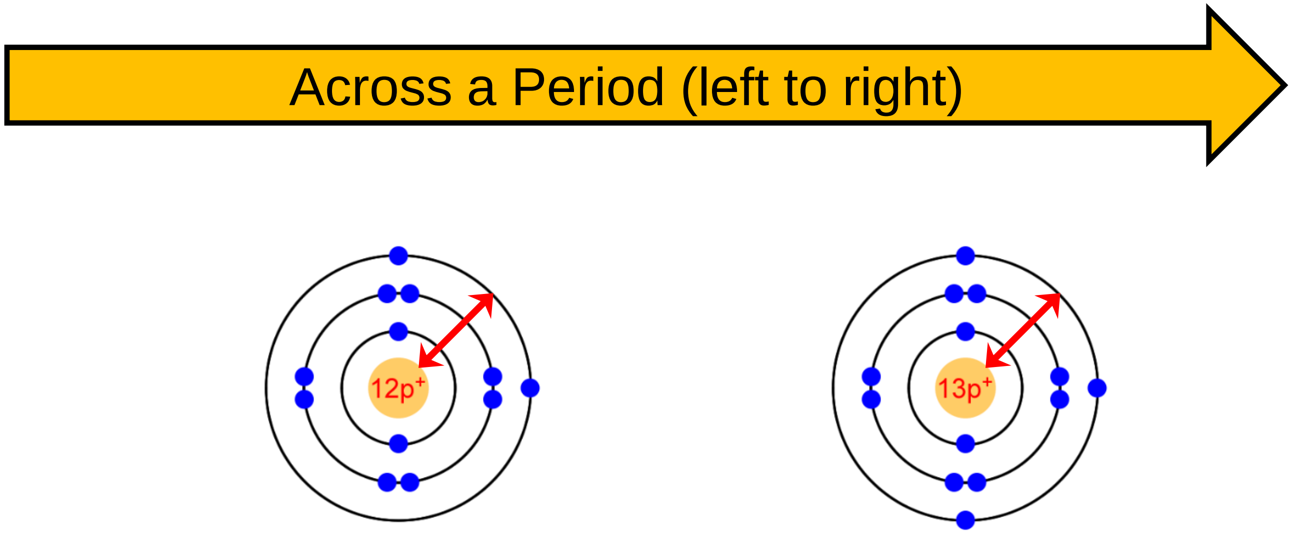

- Due to the increasing effective nuclear charge, the *electrostatic force of attraction* between the positively charged nucleus and negatively charged electron(s) in the valence shell *increases across a Period* from left to right. Consequently, *more energy* is required to remove an electron from the valence shell of the atom and the metals become *less reactive* across a Period from left to right.

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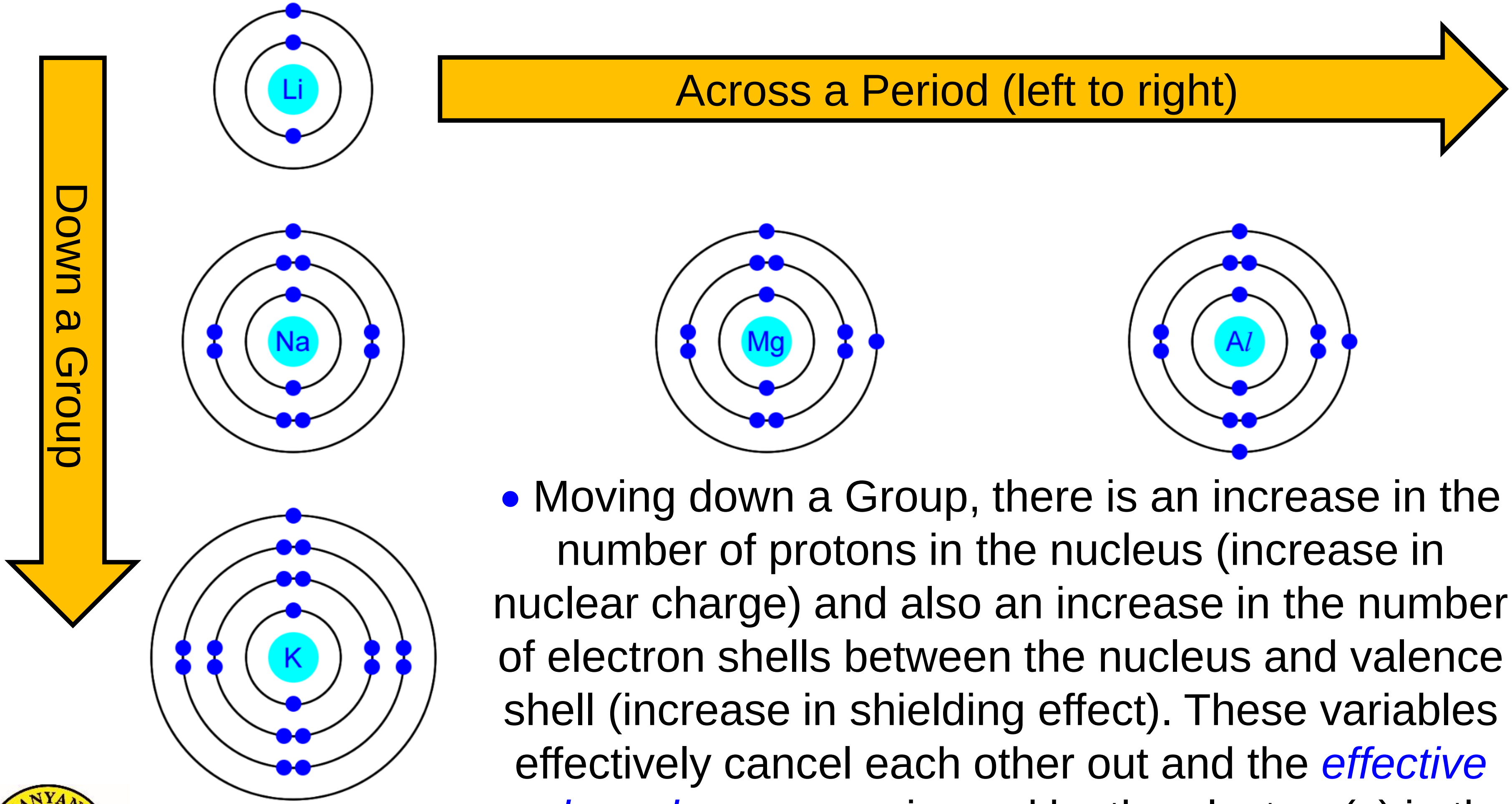

Across a Period (left to right)

Down a Group

- Moving down a Group, there is an increase in the number of protons in the nucleus (increase in nuclear charge) and also an increase in the number of electron shells between the nucleus and valence shell (increase in shielding effect). These variables effectively cancel each other out and the *effective nuclear charge* experienced by the electron(s) in the valence shell *remains constant down a Group*.

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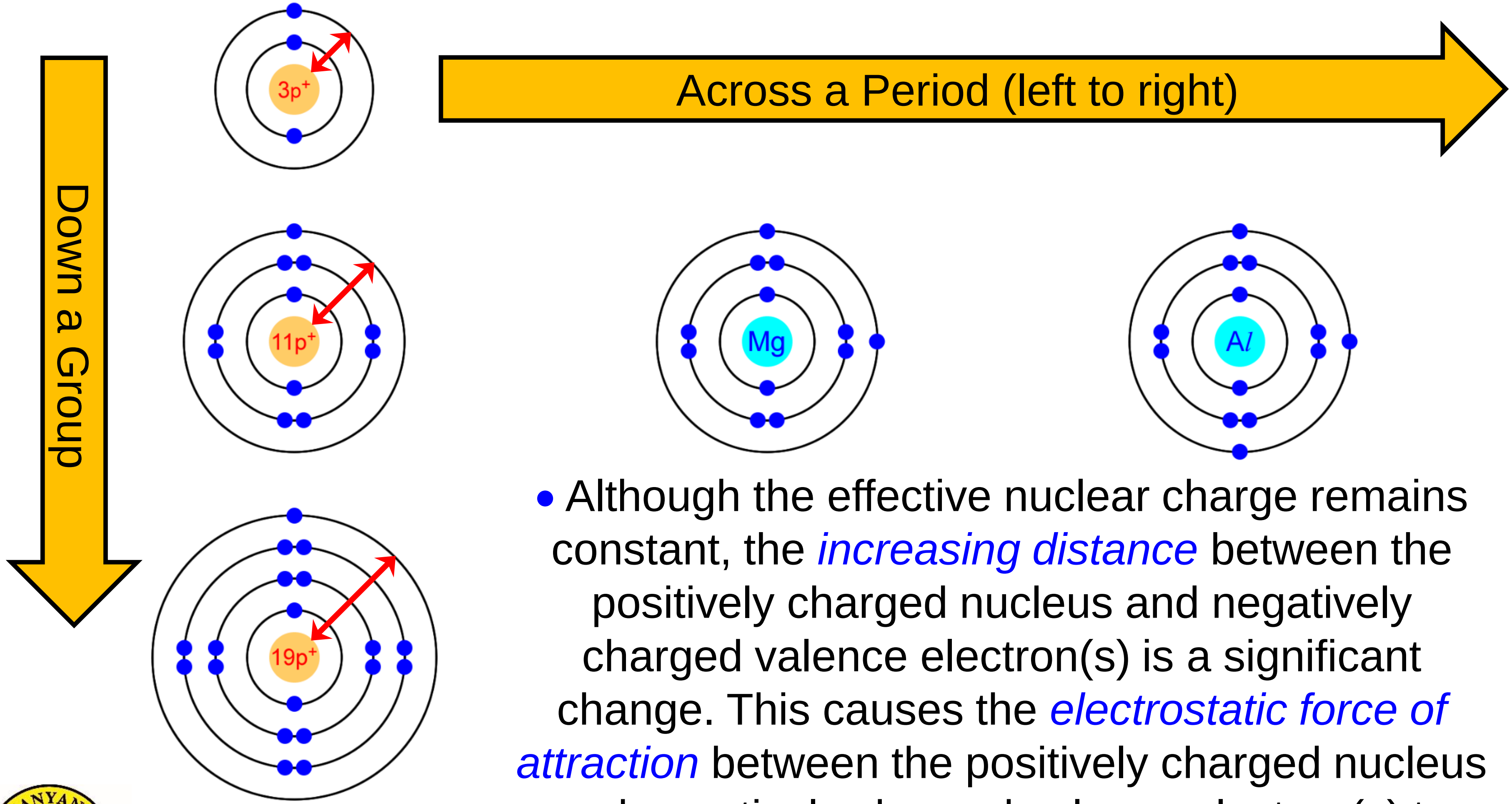

Across a Period (left to right)

Down a Group

- Although the effective nuclear charge remains constant, the *increasing distance* between the positively charged nucleus and negatively charged valence electron(s) is a significant change. This causes the *electrostatic force of attraction* between the positively charged nucleus and negatively charged valence electron(s) to *decrease down a Group*.

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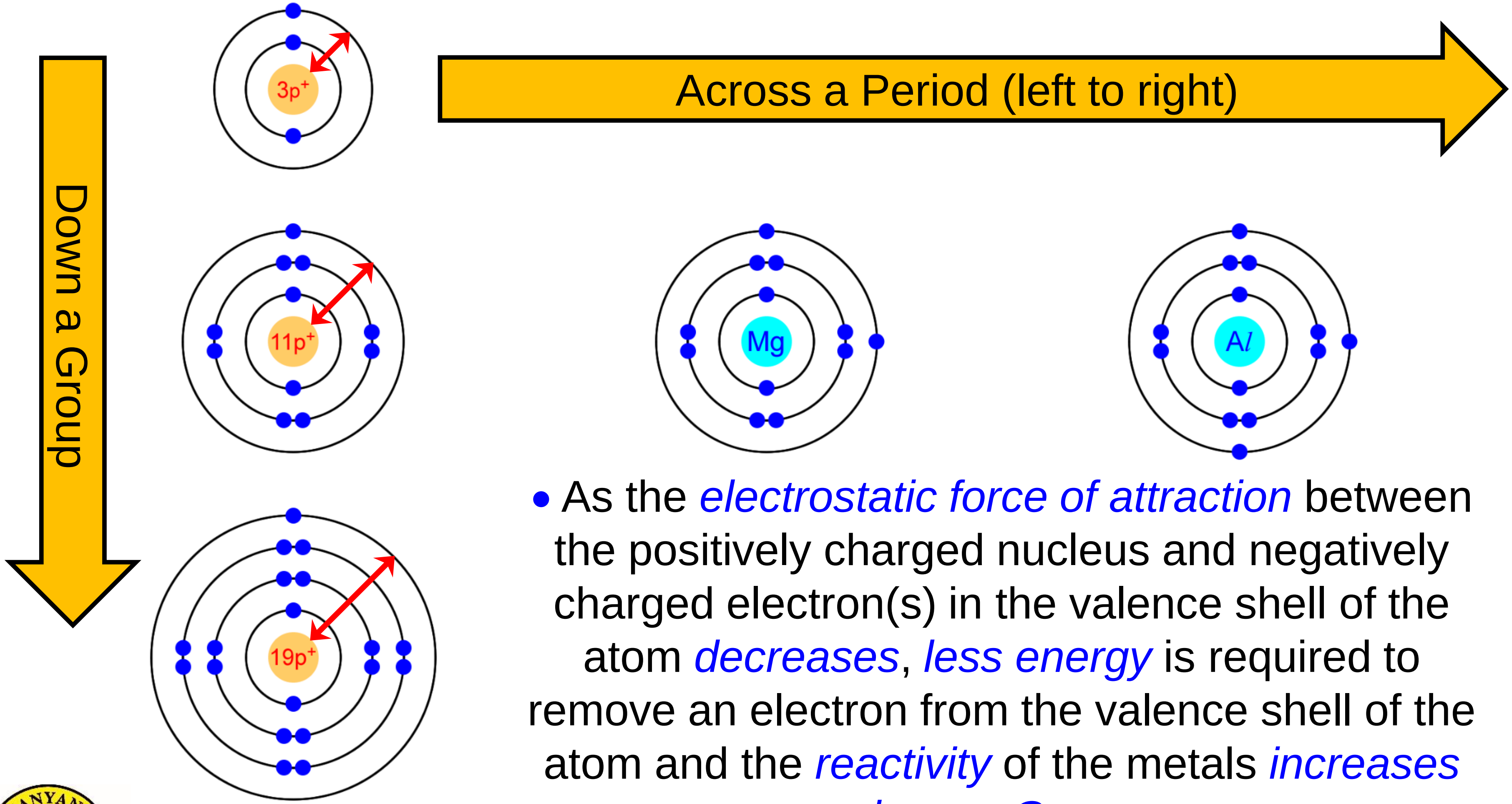

Across a Period (left to right)

Down a Group

- As the *electrostatic force of attraction* between the positively charged nucleus and negatively charged electron(s) in the valence shell of the atom *decreases*, *less energy* is required to remove an electron from the valence shell of the atom and the *reactivity* of the metals *increases down a Group*.

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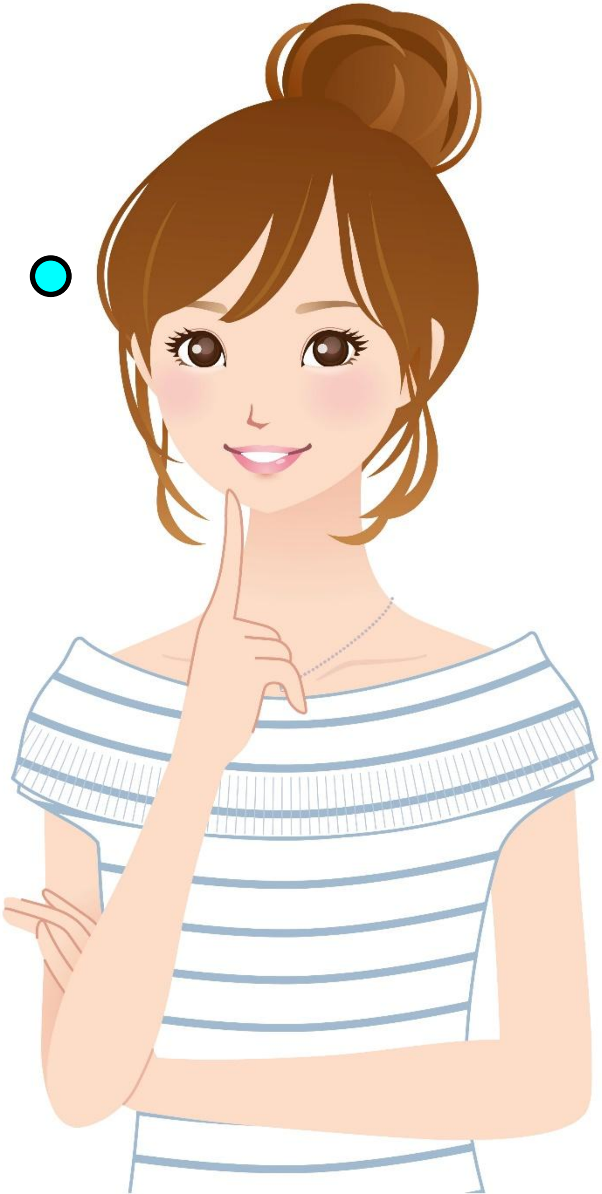
What is the *reactivity series* of metals? How is it determined and why is it useful?



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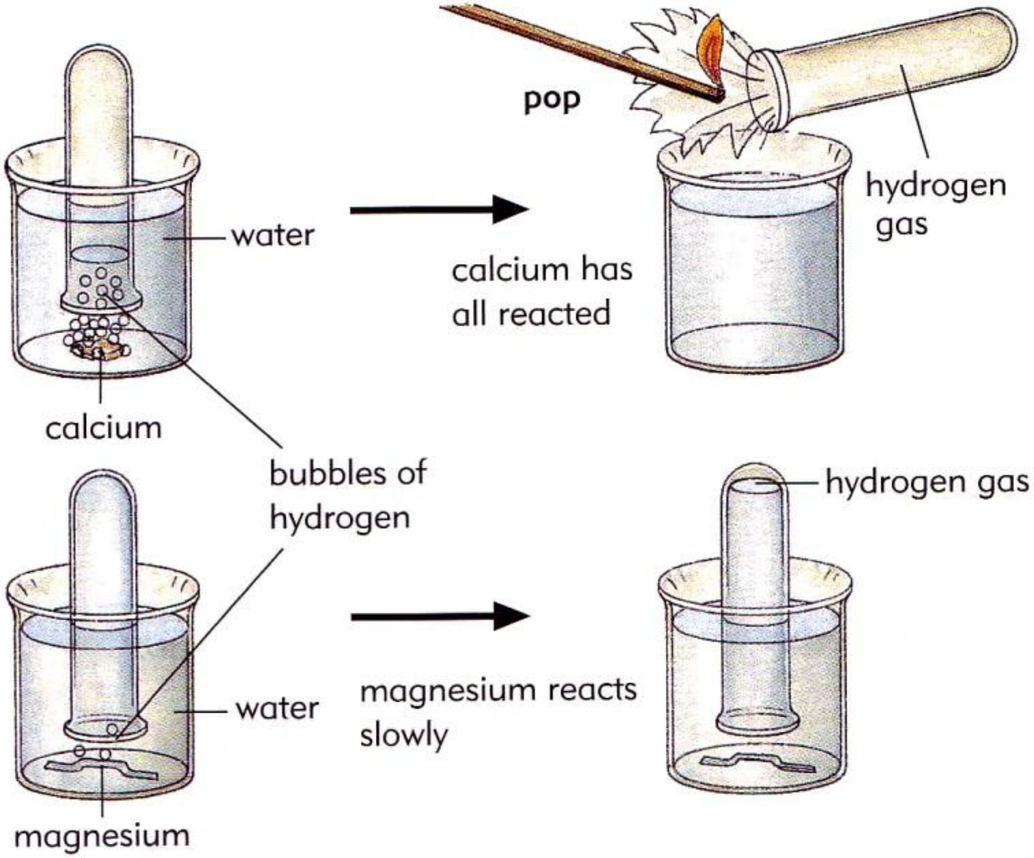

- Metals can be placed in order of reactivity based upon several criteria. One example is to observe how vigorously different metals react with *cold water*.

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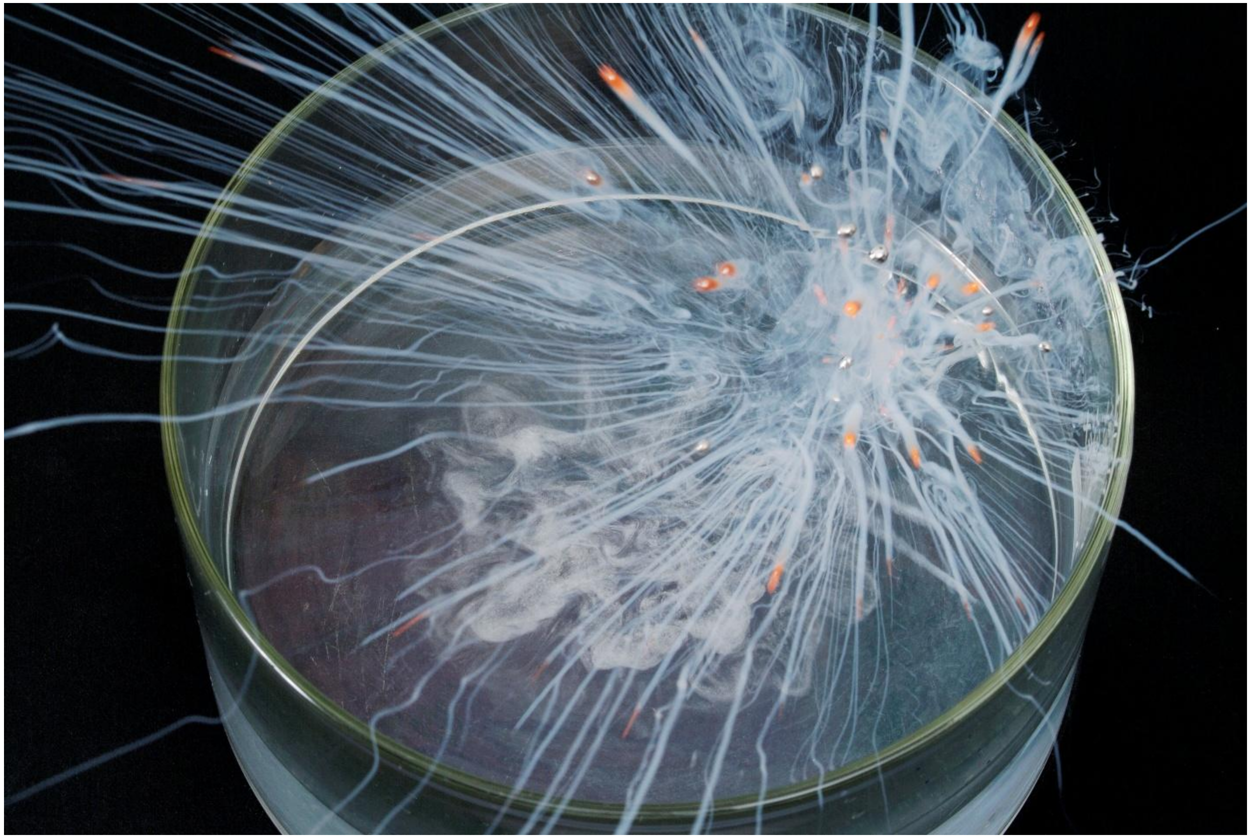

potassium + water → potassium hydroxide + hydrogen

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

Potassium

METALS

sodium + water → sodium hydroxide + hydrogen

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

Potassium → Sodium

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calcium + water → calcium hydroxide + hydrogen

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$$

Potassium → Sodium → Calcium

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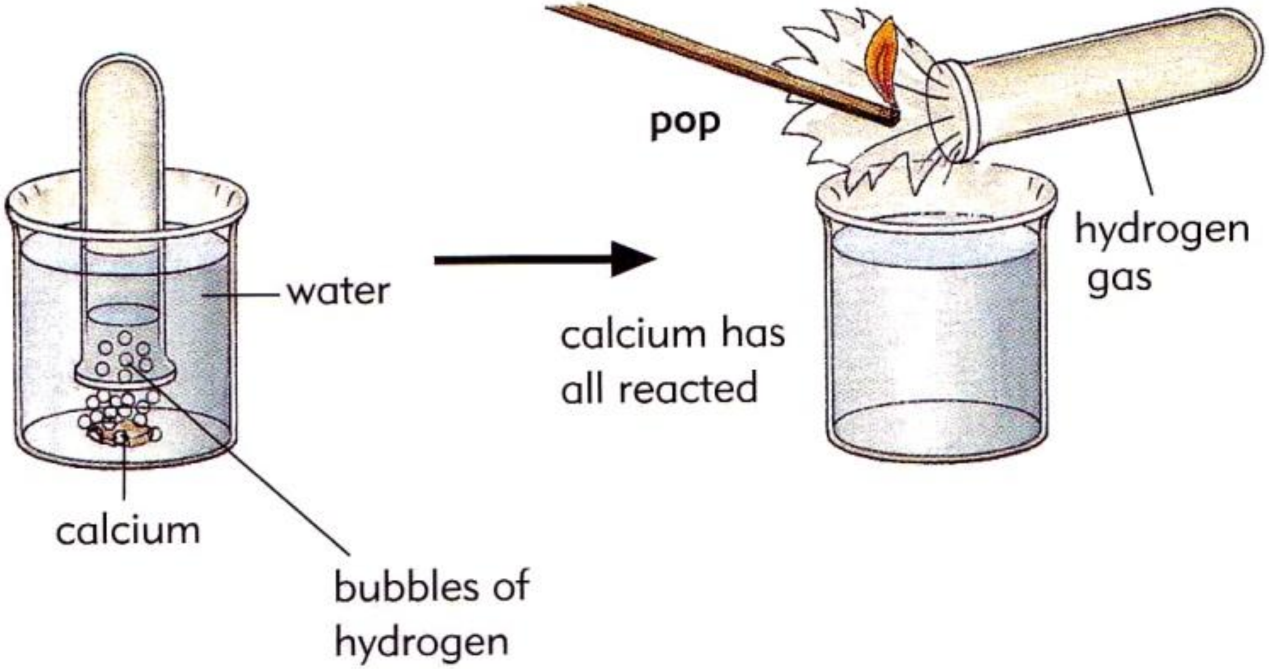

calcium + water → calcium hydroxide + hydrogen

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$

Potassium → Sodium → Calcium

METALS

- *Potassium*, *sodium* and *calcium* all react vigorously with cold water to produce an alkaline solution and hydrogen gas.

potassium + water → potassium hydroxide + hydrogen

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

sodium + water → sodium hydroxide + hydrogen

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

calcium + water → calcium hydroxide + hydrogen

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$$

Potassium → Sodium → Calcium

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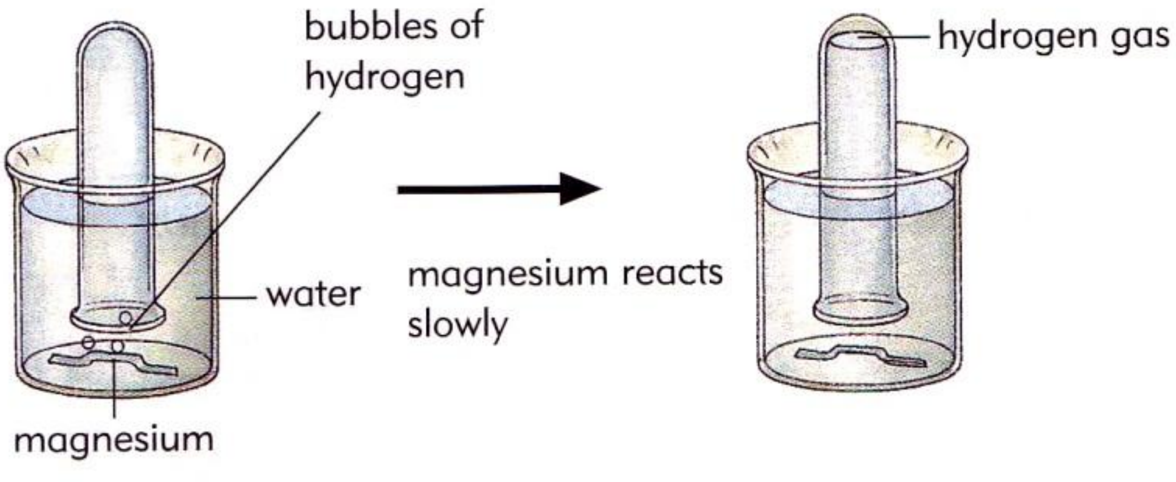
- *Magnesium* reacts very slowly with *cold water*.



- *Magnesium* reacts very slowly with *cold water* to produce magnesium hydroxide and hydrogen gas.

magnesium + water → magnesium hydroxide + hydrogen

$$Mg(s) + 2H_2O(g) \rightarrow Mg(OH)_2(s) + H_2(g)$$

Potassium → Sodium → Calcium → Magnesium

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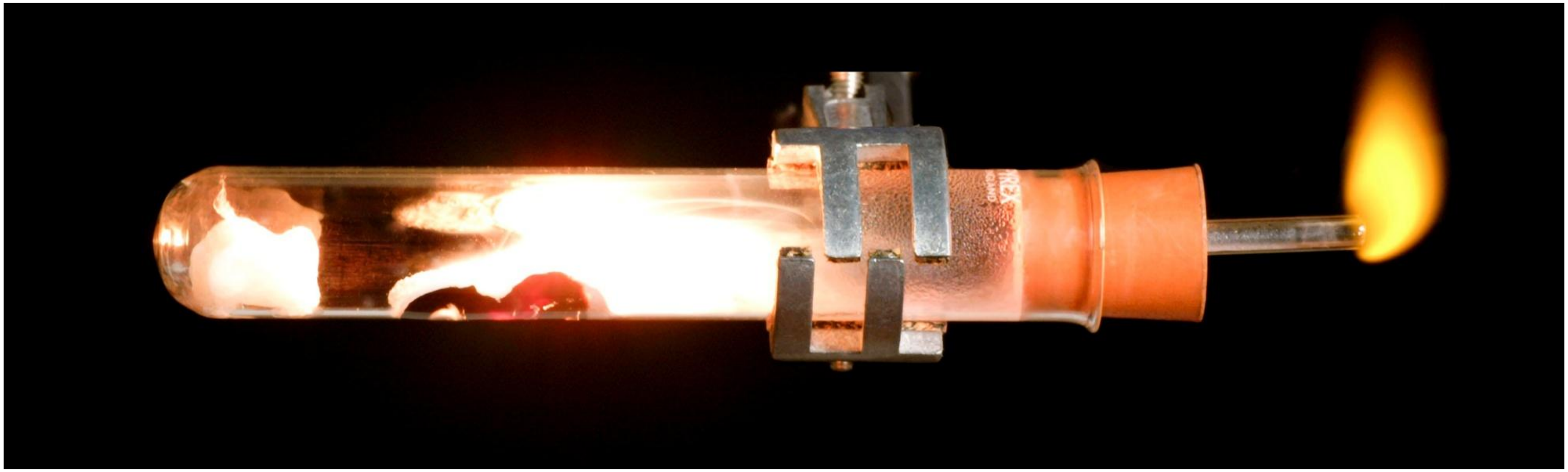
- *Magnesium* reacts very vigorously with *steam*.



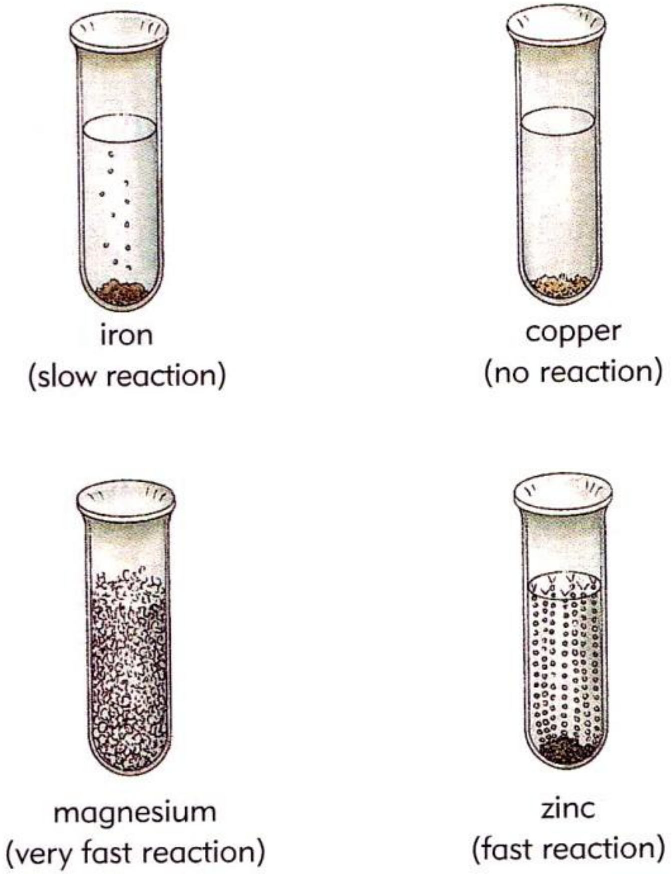
- *Magnesium* reacts very vigorously with *steam* to produce magnesium oxide and hydrogen gas.

magnesium + steam → magnesium oxide + hydrogen

$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$

Potassium → Sodium → Calcium → Magnesium

METALS



- Another way of arranging the metals in order of reactivity is to observe how vigorously different metals react with *dilute acids*.

Potassium → Sodium → Calcium → Magnesium

METALS

magnesium
(very fast reaction)

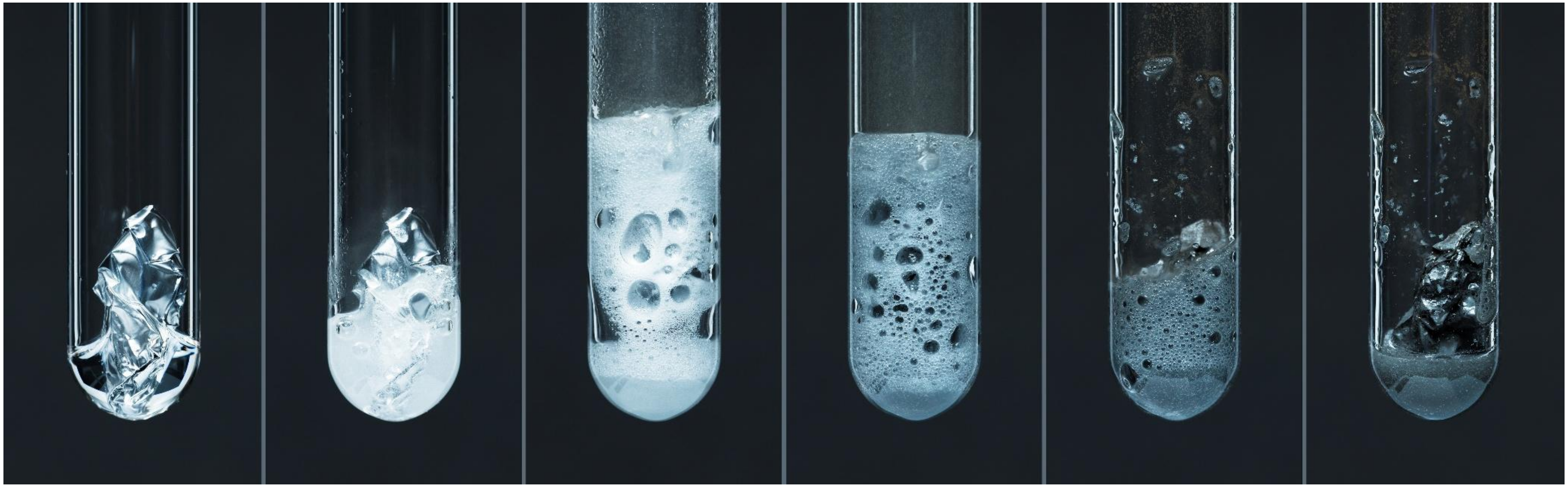
- Magnesium reacts vigorously with *dilute strong acids*.

magnesium + sulfuric acid → magnesium sulfate + hydrogen

$$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$$

Potassium → Sodium → Calcium → Magnesium

METALS



- Aluminium reacts vigorously with *dilute strong acids*.

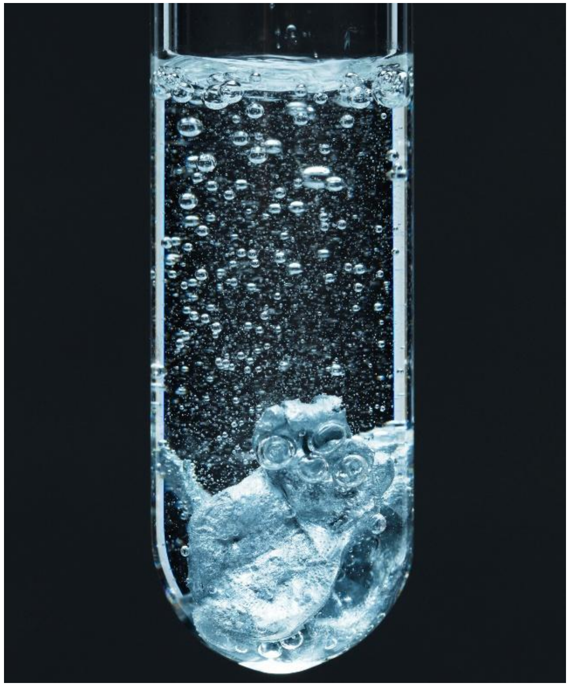
Note: The protective layer of Al_2O_3 must first be removed.

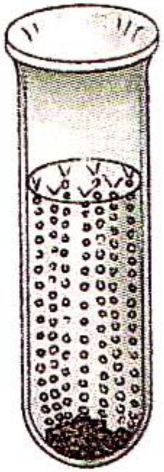
aluminium + hydrochloric acid → aluminium chloride + hydrogen

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium

METALS





zinc
(fast reaction)

- Zinc gives a fast reaction with *dilute strong acids*.

zinc + nitric acid → zinc nitrate + hydrogen

$$Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc

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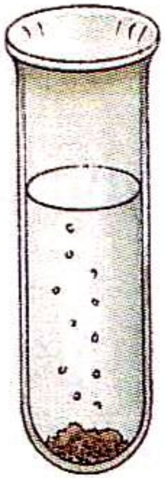

iron
(slow reaction)

- Iron gives a slow reaction with *dilute strong acids*.

iron + hydrochloric acid → iron(III) chloride + hydrogen

$$2Fe(s) + 6HCl(aq) \rightarrow 2FeCl_3(aq) + 3H_2(g)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron

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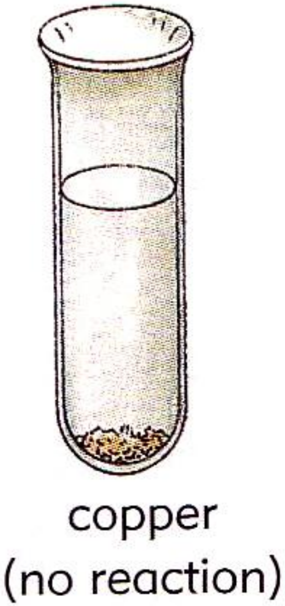


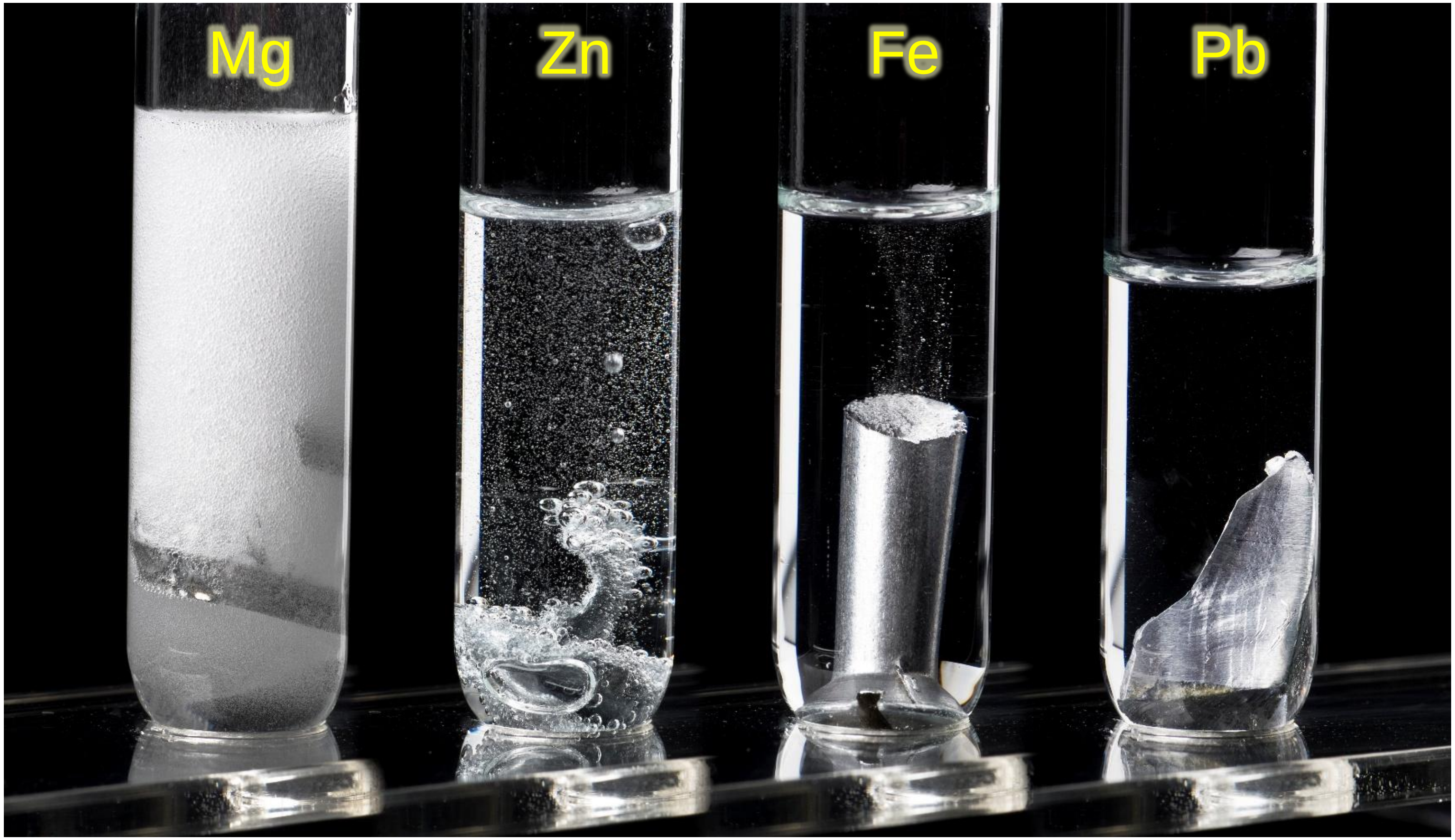
- Copper does not react with *dilute strong acids*.

copper + sulfuric acid → no observed reaction

$$Cu(s) + H_2SO_4(aq) \rightarrow \text{no observed reaction}$$

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS



- From left-to-right: magnesium, zinc, iron and lead in test tubes of *dilute strong acid*.

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Magnesium, aluminium, zinc, iron and lead all react with acids to produce a salt and hydrogen gas.

magnesium + sulfuric acid → magnesium sulfate + hydrogen

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

zinc + nitric acid → zinc nitrate + hydrogen

$Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$

iron + hydrochloric acid → iron(III) chloride + hydrogen

$2Fe(s) + 6HCl(aq) \rightarrow 2FeCl_3(aq) + 3H_2(g)$

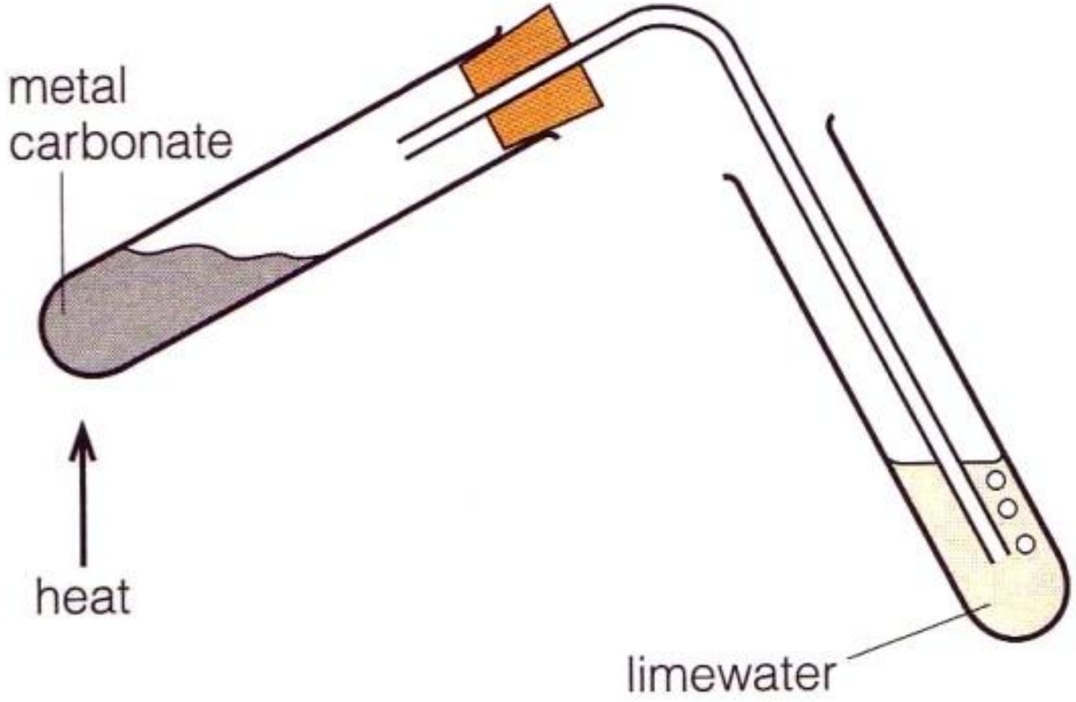
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

	Element
← ← less reactive more reactive → →	Potassium – K
	Sodium – Na
	Calcium – Ca
	Magnesium – Mg
	Aluminium – A*l*
	Zinc – Zn
	Iron – Fe
	Lead – Pb
	Copper – Cu
	Silver – Ag

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS



- Another example is to observe how easily the metal carbonates *decomposes on heating*. The more reactive the metal, the more stable its carbonate, and the less likely it is to undergo *thermal decomposition*.

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- With the exception of lithium carbonate, Li_2CO_3, *Group 1 metal carbonates* do *not* decompose on heating in a non-luminous Bunsen burner flame. Note that lithium is the least reactive of the Group 1 metals.

- Carbonates of other metals (including lithium carbonate) *decompose* on heating in a non-luminous Bunsen burner flame to produce a metal oxide and carbon dioxide gas as the products. In general, the temperature at which metal carbonates decompose *decreases* upon *descending* the reactivity series of metals.

METALS

- Calcium carbonate (white) decomposes into calcium oxide (white) and carbon dioxide at 890 °C.

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

- Zinc carbonate (white) decomposes into zinc oxide (yellow when hot, white when cold) and carbon dioxide at 400 °C.

$$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$$

- Lead(II) carbonate (white) decomposes into lead(II) oxide (yellow) and carbon dioxide at 350 °C.

$$PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$$

- Reversible Decomposition of Zinc Oxide

- Zinc oxide is white when cold, but yellow when hot. When heated, the white zinc oxide loses oxygen to form a yellow *non-stoichiometric** oxide. On cooling, the compound absorbs oxygen from the atmosphere to form white zinc oxide once again.

*A compound whose proportions cannot be written using integers.

METALS

- Reversible Decomposition of Zinc Oxide

- Zinc oxide is white when cold, but yellow when hot. When heated, the white zinc oxide loses oxygen to form a yellow *non-stoichiometric** oxide. On cooling, the compound absorbs oxygen from the atmosphere to form white zinc oxide once again.

*A compound whose proportions cannot be written using integers.

METALS

- Copper(II) carbonate (green) decomposes into copper(II) oxide (black) and carbon dioxide at 300 °C.

$$CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$$

- Silver carbonate (yellow) decomposes into silver oxide (black) and carbon dioxide at 210 °C.

$$Ag_2CO_3(s) \rightarrow Ag_2O(s) + CO_2(g)$$

The silver oxide (black) undergoes further thermal decomposition at 280 °C to form elemental silver and oxygen.

$$2Ag_2O(s) \rightarrow 4Ag(s) + O_2(g)$$

METALS

- Thermal Decomposition of Copper(II) Carbonate

- When heated by a non-luminous Bunsen burner flame, green copper(II) carbonate decomposes to form black copper(II) oxide and carbon dioxide gas. The carbon dioxide gas produces a white precipitate when bubbled through limewater.

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

- Thermal Decomposition of Copper(II) Carbonate

- When heated by a non-luminous Bunsen burner flame, green copper(II) carbonate decomposes to form black copper(II) oxide and carbon dioxide gas. The carbon dioxide gas produces a white precipitate when bubbled through limewater.

METALS

- Thermal Decomposition of Copper(II) Carbonate

- When heated by a non-luminous Bunsen burner flame, green copper(II) carbonate decomposes to form black copper(II) oxide and carbon dioxide gas. The carbon dioxide gas produces a white precipitate when bubbled through limewater.

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

	Element	Thermal Stability
← ← less reactive more reactive → →	Potassium – K	Stable to heating with Bunsen burner (does not decompose). Decomposes when the temperature is very high.
	Sodium – Na	
	Calcium – Ca	Decomposes to the metal oxide and carbon dioxide gas with increasing ease down the reactivity series. Example: $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
	Magnesium – Mg	
	Aluminium – A*l*	
	Zinc – Zn	
	Iron – Fe	
	Lead – Pb	
	Copper – Cu	
	Silver – Ag	

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

- The temperature at which a reaction takes place can affect the products that are formed.

- This means that a reaction between the *same two chemicals*, but at *different temperatures*, can produce *different products*.

- This happens because the product of the reaction formed at the *lower temperature* is *thermally unstable* and *decomposes* at the *higher temperature* to form a different reaction products.

METALS

- For example, if calcium reacts with *water* at *room temperature*, the main product is *calcium hydroxide*.

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$$

- But if calcium reacts with *steam* at a *high temperature*, the main product is *calcium oxide*, because any calcium hydroxide that is initially formed undergoes *thermal decomposition*.

to start: $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$

then decomposes: $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(l)$

overall: $Ca(s) + H_2O(g) \rightarrow CaO(s) + H_2(g)$

METALS

- At room temperature, calcium hydroxide is formed.

calcium + water → calcium hydroxide + hydrogen

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

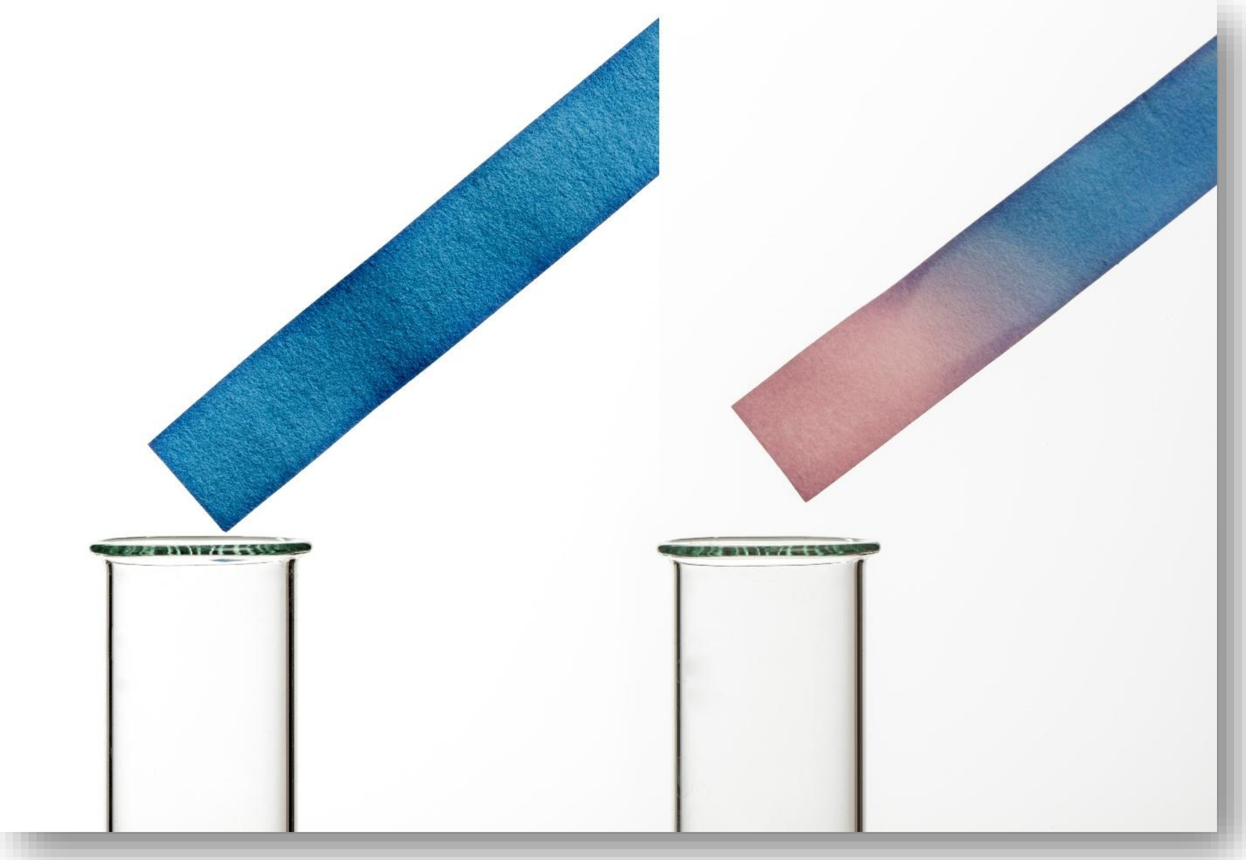
- At high temperatures, calcium oxide is formed.

calcium hydroxide → calcium oxide + water

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

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS



- Qualitative test for water vapour: Anhydrous cobalt(II) chloride paper changes from blue to pink.

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

Summary

- Calcium reacts with water (*room temperature*) to form *calcium hydroxide* and hydrogen:

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g)$$

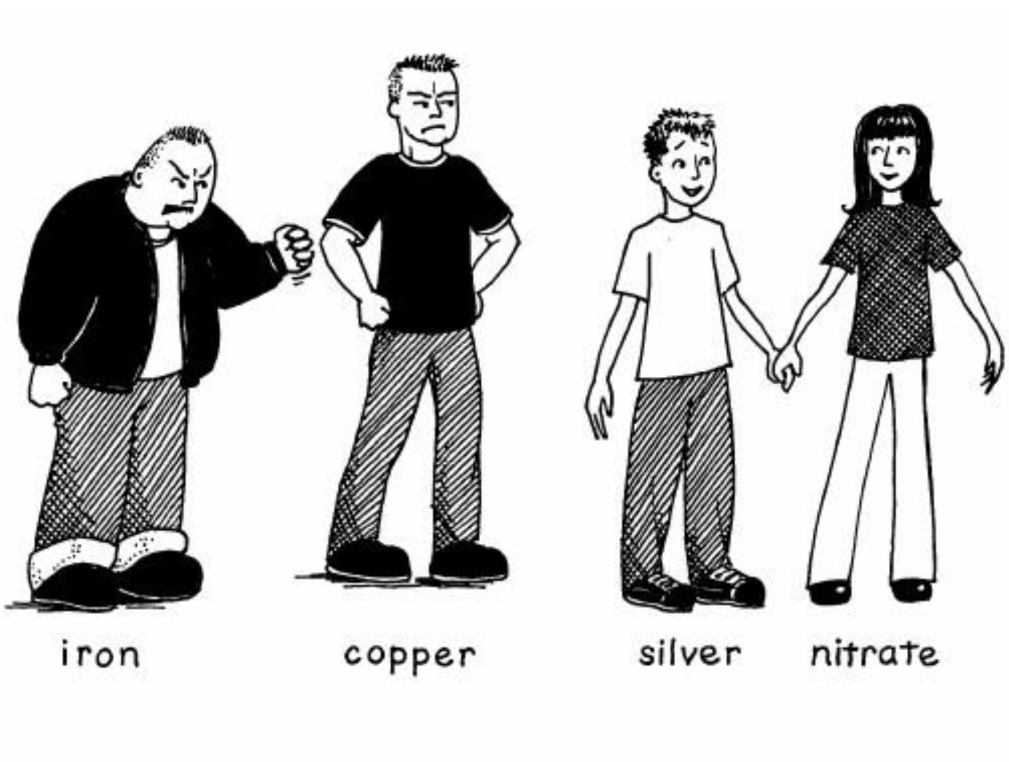
- Calcium reacts with steam (*high temperature*) to form *calcium oxide* and hydrogen:

$$Ca(s) + H_2O(g) \rightarrow CaO(s) + H_2(g)$$

METALS

Please tell me about *displacement reactions*.

METALS



- A more reactive metal will displace a less reactive metal from its compounds.

Carbon ↓ (between Aluminium and Zinc) — *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\text{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\text{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → **Silver**

METALS

- The reaction between elemental copper and aqueous silver nitrate.
- Copper is more reactive than silver and will therefore displace silver from its compounds.

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

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

- Observations:

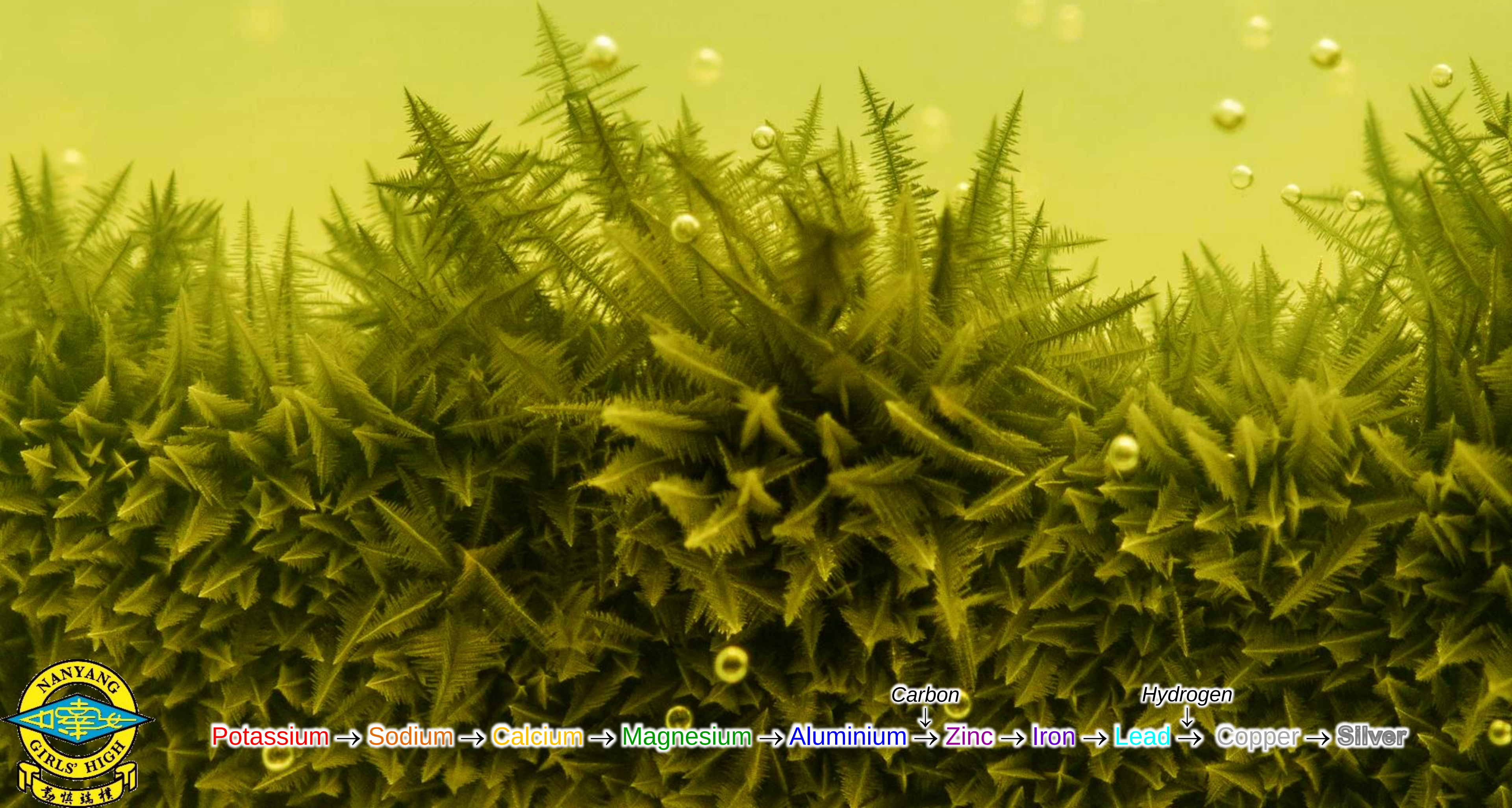
→ Solid copper metal dissolves into the solution.

→ The appearance of the solution changes from a colourless solution of $AgNO_3$ to a blue solution of $Cu(NO_3)_2$.

→ Silver-grey crystals of elemental silver deposit over the surface of the copper.

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS



Carbon ↓ *Hydrogen* ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

- The reaction between elemental iron and aqueous lead(II) nitrate.
- Iron is more reactive than lead and will therefore displace lead from its compounds.

$$2Fe(s) + 3Pb(NO_3)_2(aq) \rightarrow 2Fe(NO_3)_2(aq) + 3Pb(s)$$

$$2Fe(s) + 3Pb^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + 3Pb(s)$$

- Observations:

→ Solid iron metal dissolves into the solution.

→ The appearance of the solution changes from a colourless solution of $Pb(NO_3)_2$ to a yellow solution of $Fe(NO_3)_3$.

→ Grey crystals of elemental lead deposit over the surface of the iron.

$$\text{Potassium} \rightarrow \text{Sodium} \rightarrow \text{Calcium} \rightarrow \text{Magnesium} \rightarrow \text{Aluminium} \overset{\overset{\textit{Carbon}}{\downarrow}}{\rightarrow} \text{Zinc} \rightarrow \text{Iron} \rightarrow \text{Lead} \overset{\overset{\textit{Hydrogen}}{\downarrow}}{\rightarrow} \text{Copper} \rightarrow \text{Silver}$$

METALS

Carbon Hydrogen
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

NANYANG GIRLS' HIGH
勤慎端樸

METALS

- The reaction between elemental zinc and aqueous copper(II) sulfate.
- Zinc is more reactive than copper and will therefore displace copper from its compounds.

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

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

- Observations:

→ Solid zinc metal dissolves into the solution.

→ The appearance of the solution changes from a blue solution of $CuSO_4$ to a colourless solution of $ZnSO_4$.

→ Reddish-brown (pink) crystals of elemental copper deposit over the surface of the zinc.

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\text{Carbon}\,\downarrow}{\rightarrow}$ Zinc → Iron → Lead $\overset{\text{Hydrogen}\,\downarrow}{\rightarrow}$ Copper → **Silver**

METALS

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- The reaction between elemental aluminium and aqueous copper(II) sulfate.
- Aluminium is more reactive than copper and will therefore displace copper from its compounds.

$$2Al(s) + 3CuSO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3Cu(s)$$

$$2Al(s) + 3Cu^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Cu(s)$$

- Observations:

→ Solid aluminium metal dissolves into the solution.

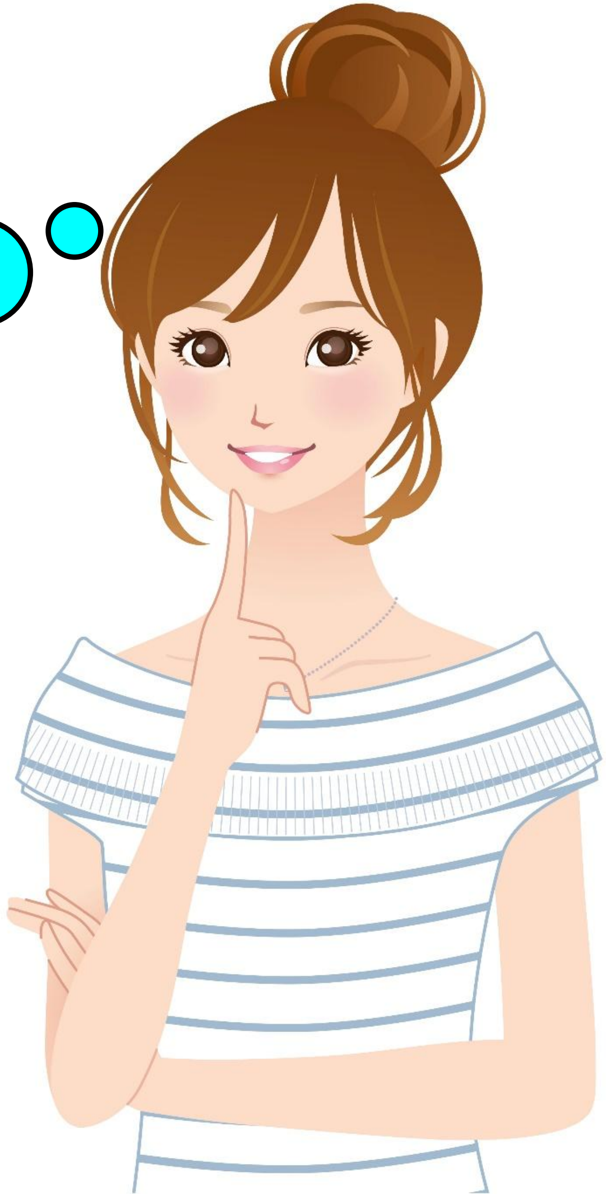
→ The appearance of the solution changes from a blue solution of $CuSO_4$ to a colourless solution of $Al_2(SO_4)_3$.

→ Reddish-brown (pink) crystals of elemental copper deposit over the surface of the aluminium foil.

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\text{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\text{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → **Silver**

METALS

For displacement reactions, it is important to describe any changes in the appearance of the *solid* and any changes in the appearance of the *solution*.



METALS

- Study the following reactions. Use the data to place the metals in order of reactivity, from most to least reactive.

$$Cu(s) + Pt(NO_3)_2(aq) \rightarrow Cu(NO_3)_2(aq) + Pt(s)$$

$$Fe(s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$$

$$Sn(s) + Cu(NO_3)_2(aq) \rightarrow Sn(NO_3)_2(aq) + Cu(s)$$

$$Fe(s) + Sn(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Sn(s)$$

Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Study the following reactions. Use the data to place the metals in order of reactivity, from most to least reactive.

$$Cu(s) + Pt(NO_3)_2(aq) \rightarrow Cu(NO_3)_2(aq) + Pt(s)$$

$$Fe(s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$$

$$Sn(s) + Cu(NO_3)_2(aq) \rightarrow Sn(NO_3)_2(aq) + Cu(s)$$

$$Fe(s) + Sn(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Sn(s)$$

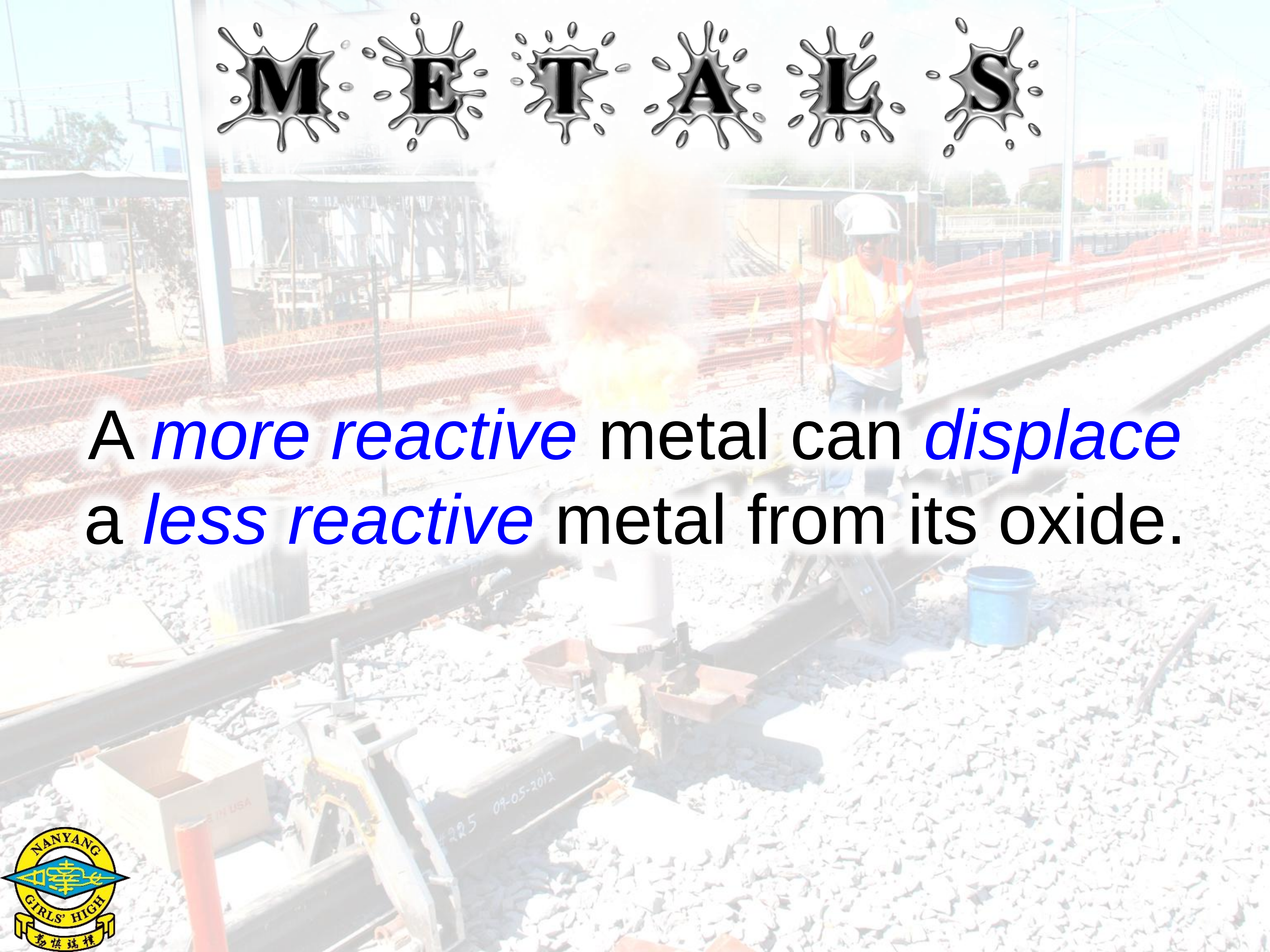
Most Reactive → Least Reactive

Iron (Fe) → Tin (Sn) → Copper (Cu) → Platinum (Pt)

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

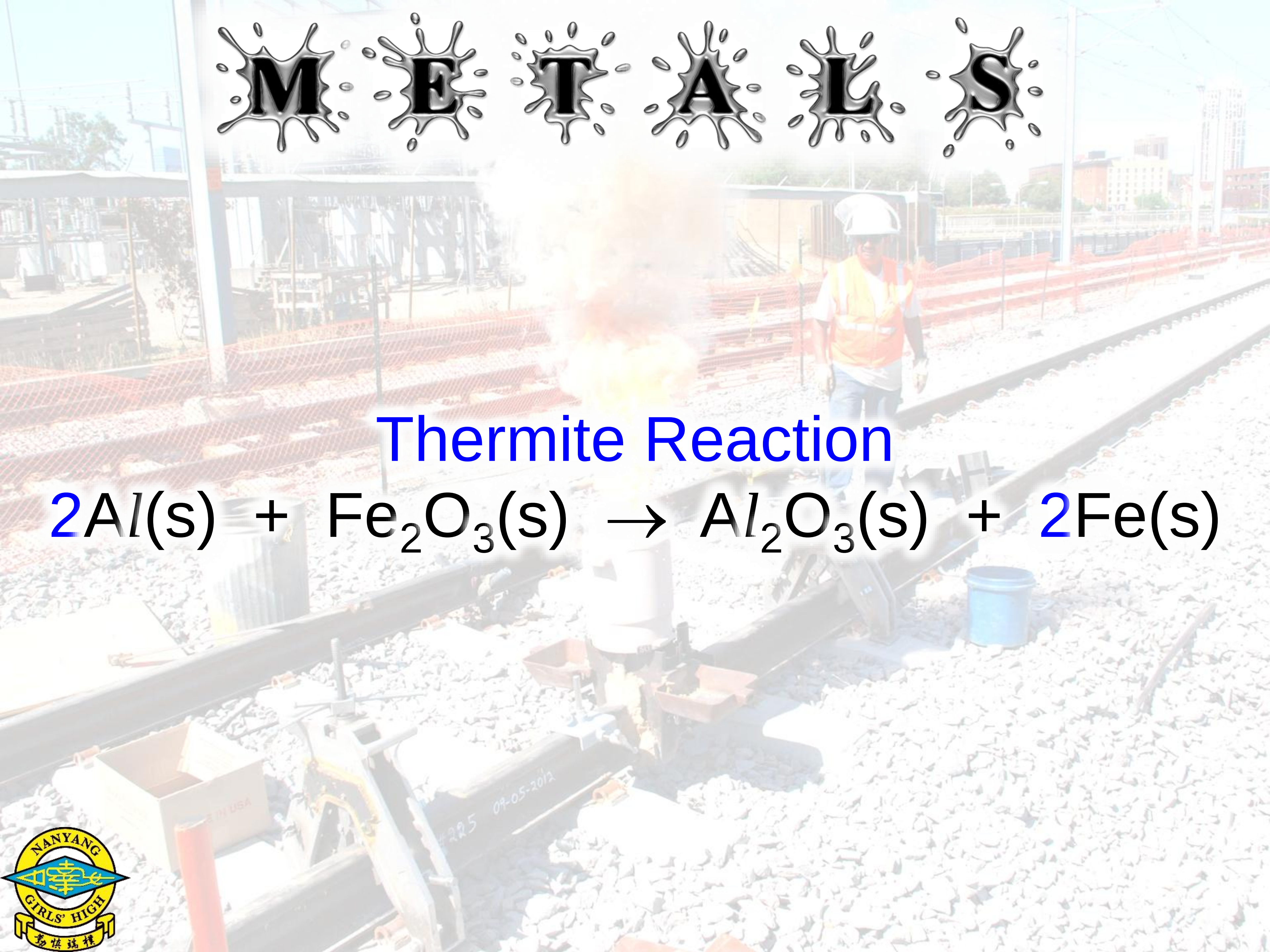
METALS

METALS



A *more reactive* metal can *displace* a *less reactive* metal from its oxide.

METALS

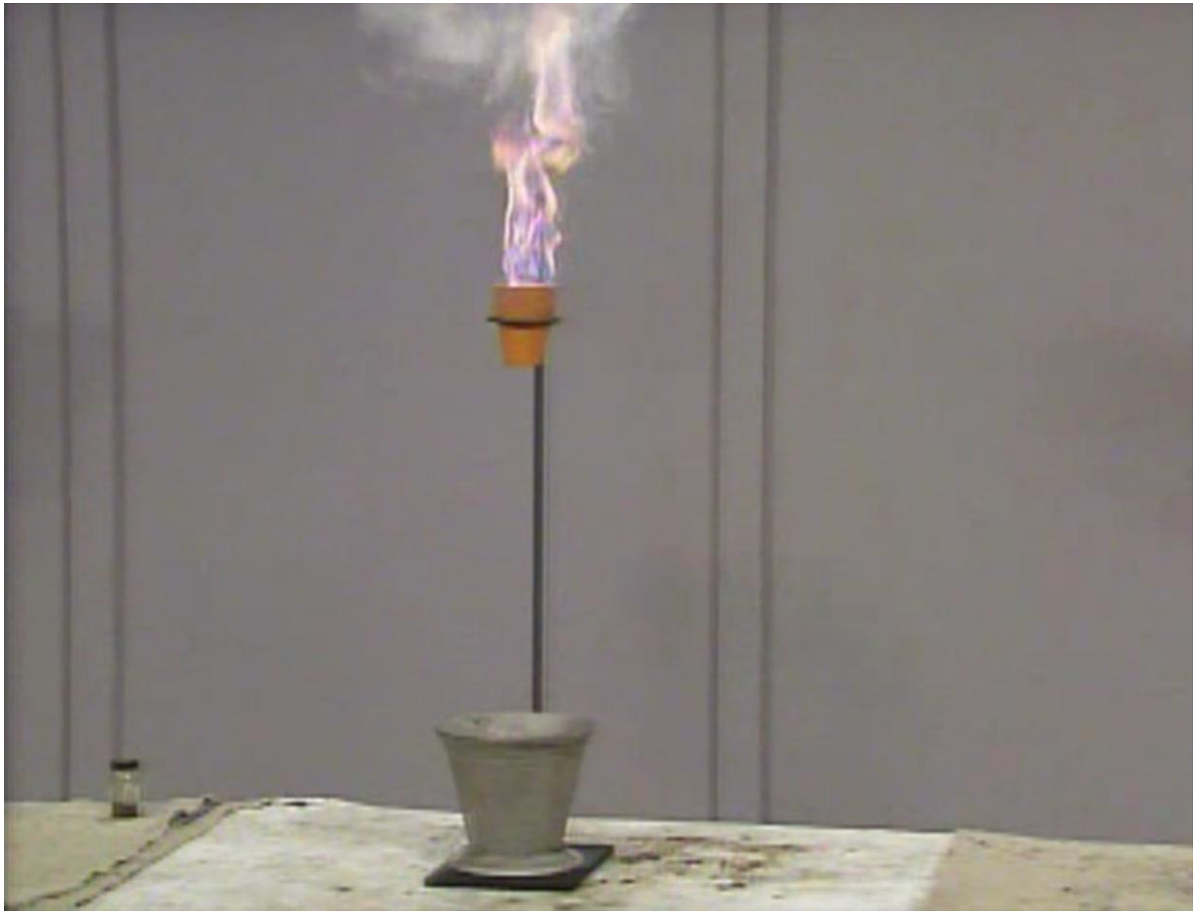


Thermite Reaction

$$2\mathrm{A}l(s) + \mathrm{Fe_2O_3}(s) \rightarrow \mathrm{A}l_2\mathrm{O_3}(s) + 2\mathrm{Fe}(s)$$

METALS

- Thermite Reaction



$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\substack{\textit{Carbon}\\ \downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\substack{\textit{Hydrogen}\\ \downarrow}}{\rightarrow}$ Copper → Silver

METALS

- Thermite Reaction

$$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\textit{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\textit{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → Silver

METALS

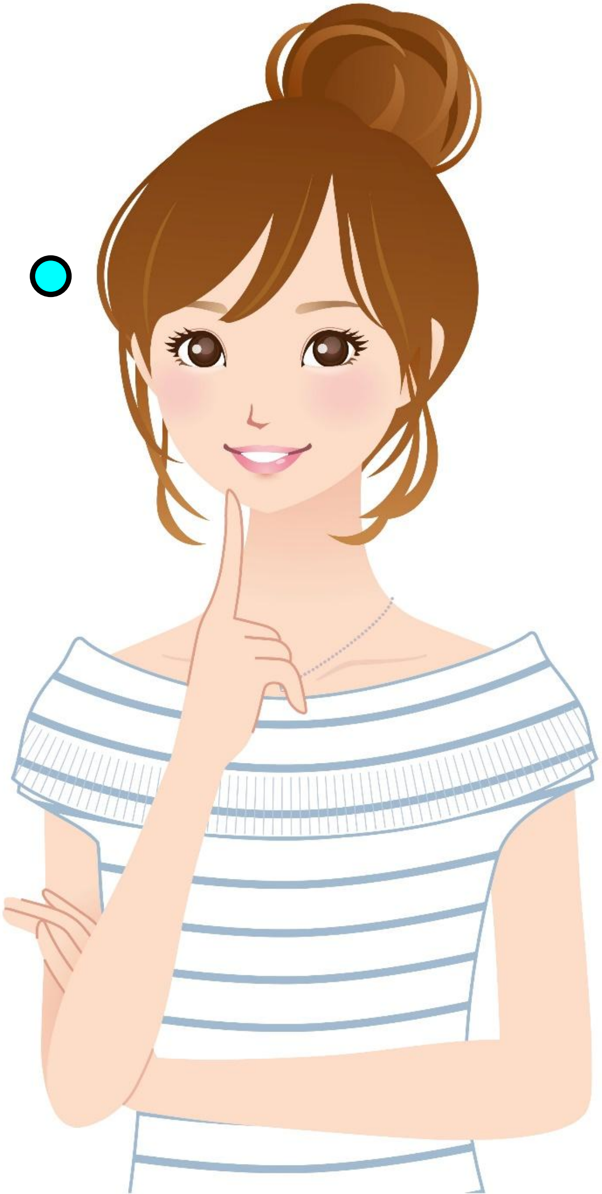
	Element	Reactions	Stability of CO_3^{2-}	Ease with which the metal loses electrons
↑ ↑ more reactive less reactive ↓ ↓	Potassium – K	Cold water	↓ ↓ ↓ ↓ ↓ • Metal carbonates become less stable and decompose more readily on heating, *i.e.* the temperature at which the metal carbonate decomposes decreases. $MCO_3 \rightarrow MO + CO_2$ ↓ ↓ ↓ ↓ ↓	↑ ↑ ↑ ↑ ↑ • Metals form positive ions more readily. • Metals lose valence electrons more easily. ↑ ↑ ↑ ↑ ↑
	Sodium – Na	Cold water		
	Calcium – Ca	Cold water		
	Magnesium – Mg	Steam		
	Aluminium – A*l*	Dilute acid		
	(Carbon – C)	(Not applicable)		
	Zinc – Zn	Dilute acid		
	Iron – Fe	Dilute acid		
	Lead – Pb	Dilute acid		
	(Hydrogen – H)	(Not applicable)		
	Copper – Cu	Displacement		
	Silver – Ag	Displacement		

Carbon ↓ *Hydrogen* ↓

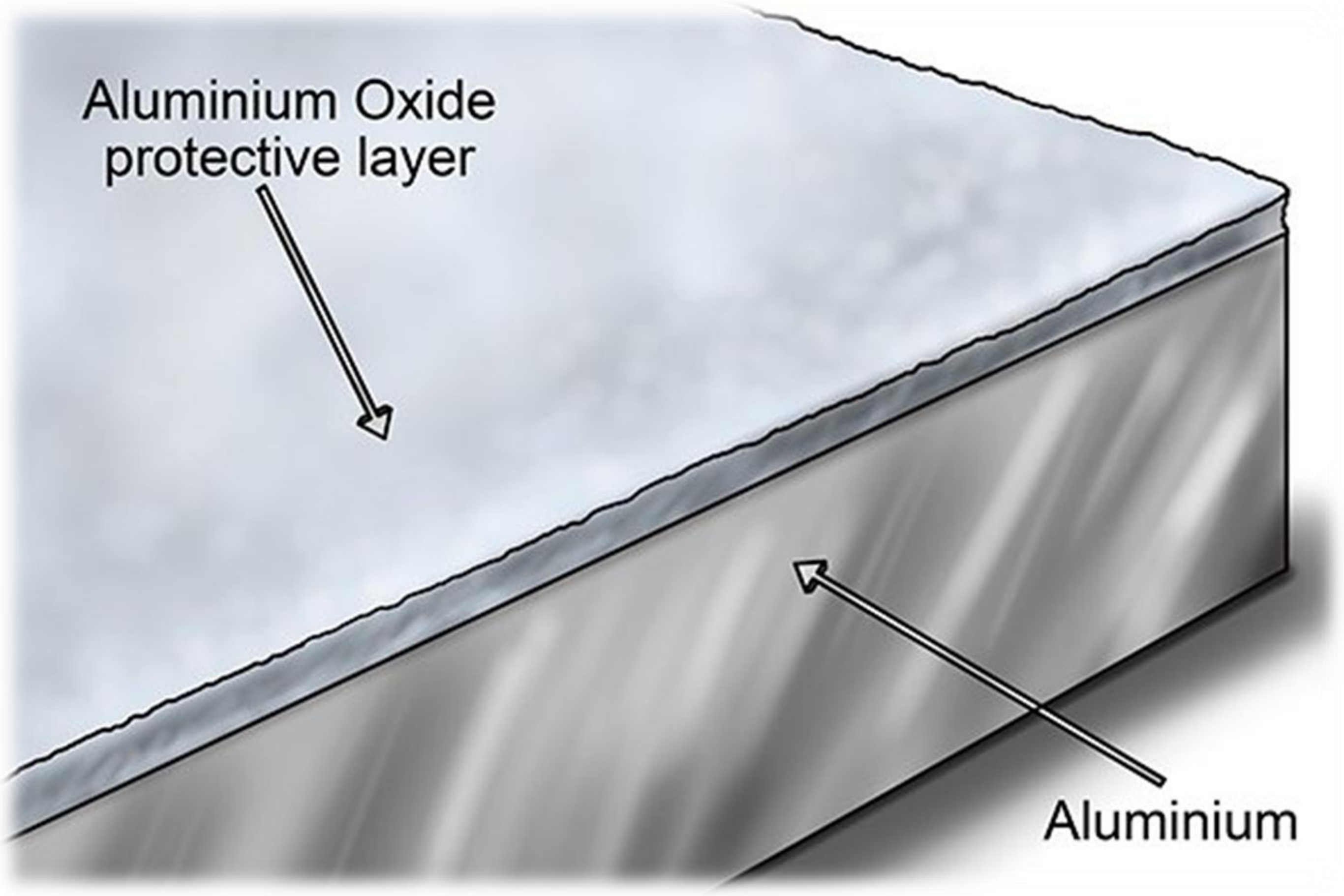
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

Why do *aluminium* and *lead* sometimes appear to be less reactive than they really are?



METALS



Carbon ↓ *Hydrogen* ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Aluminium is a reactive metal, placed between magnesium and zinc in the reactivity series. It reacts readily with oxygen in the air to form solid aluminium oxide, Al_2O_3.

- The surface layer of aluminium oxide is firmly bonded to the aluminium. This durable and non-porous layer of aluminium oxide protects the aluminium underneath from any further reactions.

- Before the aluminium can react any further, the protective layer of aluminium oxide must first be removed, either chemically or physically, in order to expose the fresh aluminium lying underneath.

Carbon ↓ *Hydrogen* ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Many lead(II) salts are insoluble in water, for example lead(II) chloride, $PbCl_2$, and lead(II) sulfate $PbSO_4$.

- When lead is added to sulfuric acid, there will be an initial reaction forming lead(II) sulfate, but the reaction will quickly stop as an insoluble layer of lead(II) sulfate forms over the surface of the lead, preventing the acid and metal from coming into further contact with each other:

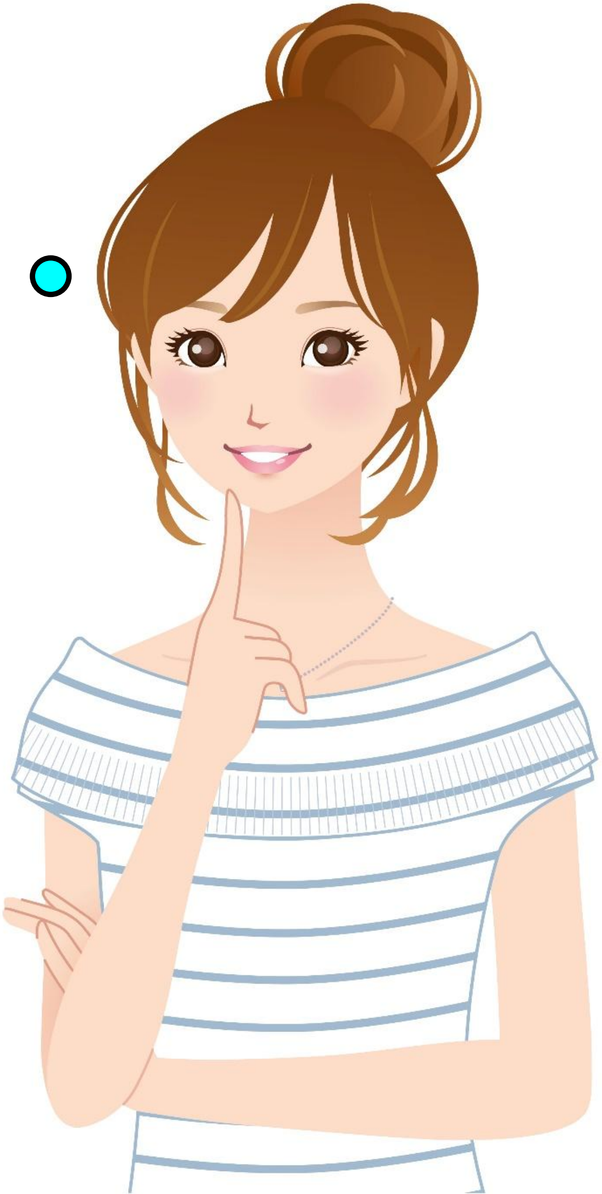
$$Pb(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + H_2(g)$$

- Because there maybe little or no observed reaction, this may lead to the erroneous conclusion that lead is unable to displace hydrogen, hence lead is less reactive than hydrogen.

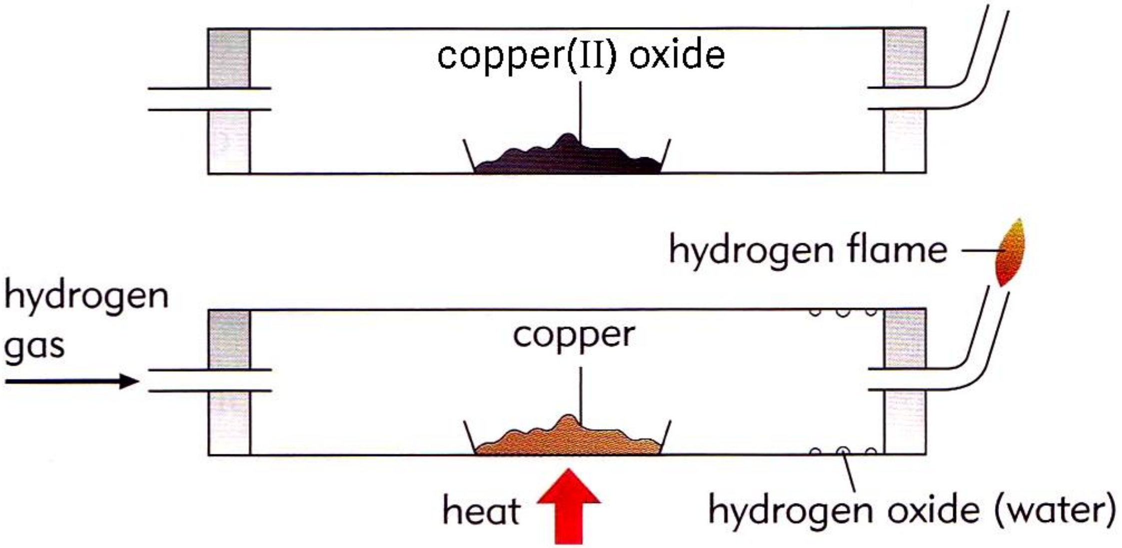
Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

How does a knowledge of the *reactivity series* of metals help me determine the best way of *extracting* a metal from its ore?



METALS

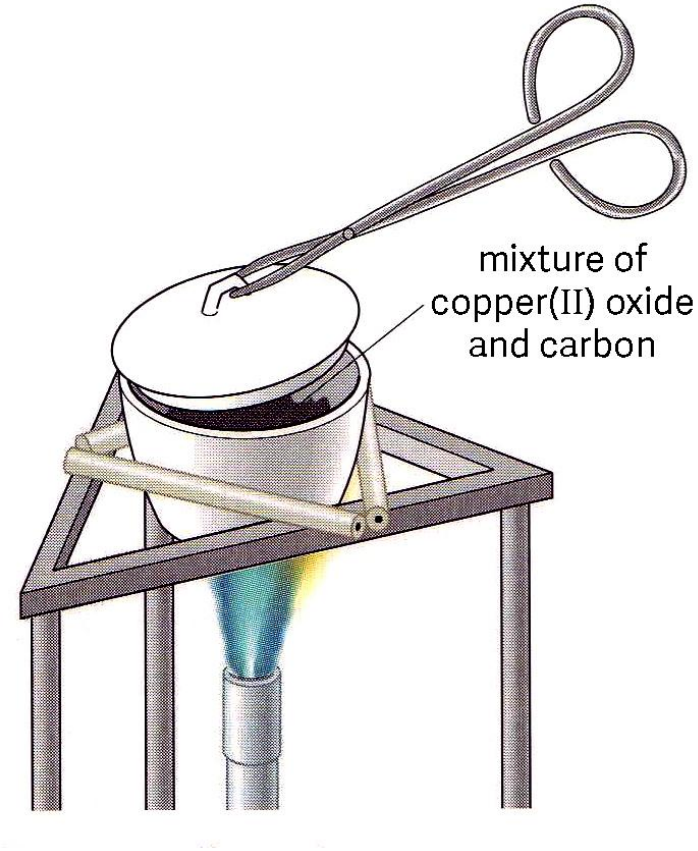


- Metals *below hydrogen* in the reactivity series can be displaced from their oxides by heating them with *hydrogen*, for example:

$$CuO(s) + H_2(s) \rightarrow Cu(s) + H_2O(l)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\textit{Carbon}}{\overset{\downarrow}{\rightarrow}}$ Zinc → Iron → Lead $\overset{\textit{Hydrogen}}{\overset{\downarrow}{\rightarrow}}$ Copper → **Silver**

METALS



- Metals that are placed *below carbon* in the reactivity series, (*i.e.* metals that are *less reactive* than carbon) can be displaced from their oxides by heating them with carbon. For example:

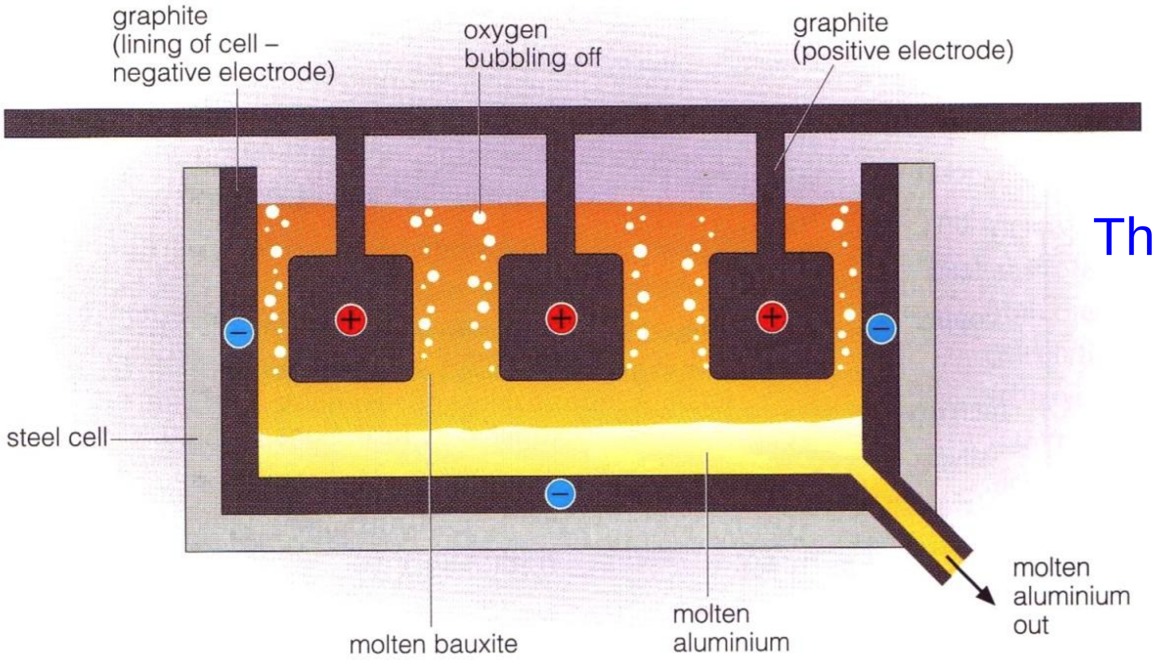
$$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$$

In this example, the copper has been *reduced* by carbon.

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\text{Carbon}}{\overset{\downarrow}{\rightarrow}}$ Zinc → Iron → Lead $\overset{\text{Hydrogen}}{\overset{\downarrow}{\rightarrow}}$ Copper → **Silver**

METALS

- Metals placed *above* carbon in the reactivity series (*i.e.* metals that are *more reactive* than carbon) are extracted from their oxides by *electrolysis*, e.g. the extraction of Al from $Al_2O_3(l)$.



The Electrolysis of Molten Aluminium Oxide

Carbon ↓ Hydrogen ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Pure anhydrous aluminium oxide can only be electrolysed in its *molten state* (in order for it to contain *mobile ions*).

- Pure anhydrous aluminium oxide melts at the very high temperature of 2072°C. As a consequence, the electrolysis of aluminium oxide on an industrial scale would consume a very large amount of energy in order to keep it molten. This would be both expensive and potentially very polluting.

- In order to perform the electrolysis at a more manageable temperature, aluminium oxide is dissolved in a mixture of molten cryolite (sodium hexafluoroaluminate – Na_3AlF_6), calcium fluoride (CaF_2) and aluminium fluoride (AlF_3). This allows the aluminium oxide to be electrolysed at the much lower temperature of 950°C (impurities lower melting points).

Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- An electric current of between 40 000 to 100 000 A is used for the electrolysis of molten aluminium oxide.

- Aluminium ions are reduced to metallic aluminium at the negative cathode:

$$Al^{3+}(l) + 3e^{-} \rightarrow Al(l)$$

- Oxide ions are oxidised to molecular oxygen at the positive anode:

$$2O^{2-}(l) \rightarrow O_2(g) + 4e^{-}$$

- Because there is a high concentration of fluoride ions in the electrolyte, fluorine gas may also be produced at the anode:

$$2F^{-}(l) \rightarrow F_2(g) + 2e^{-}$$

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\text{Carbon}}{\downarrow}$ → Zinc → Iron → Lead $\overset{\text{Hydrogen}}{\downarrow}$ → Copper → Silver

METALS

- The graphite anode reacts with oxygen to produce carbon dioxide gas:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

- Because the graphite anode is oxidised, it needs to be replaced at regular intervals.

- This may seem inefficient. However, oxidation of the graphite anode is an *exothermic* process ($\Delta H = -394$ kJ/mol) and the energy that is released into the surroundings is used to *heat the electrolyte* so that less external energy is required in order to keep it molten.

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

Carbon ↓ Hydrogen ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

Aluminium is now commonplace but was considered to be a precious metal until the late 1800s. Although aluminium is the third most abundant element and most abundant metal in the Earth's crust, it was at first found to be exceedingly difficult to extract the metal from its various compounds. The great expense of refining the metal made the small available quantity of pure aluminium *more valuable than gold*. Bars of aluminium were exhibited at the Exposition Universelle of 1855, and Napoleon III's most important guests were given aluminium cutlery, while those less worthy dined with mere silver!

https://en.wikipedia.org/wiki/precious_metal

Carbon ↓ (between Aluminium and Zinc)

Hydrogen ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

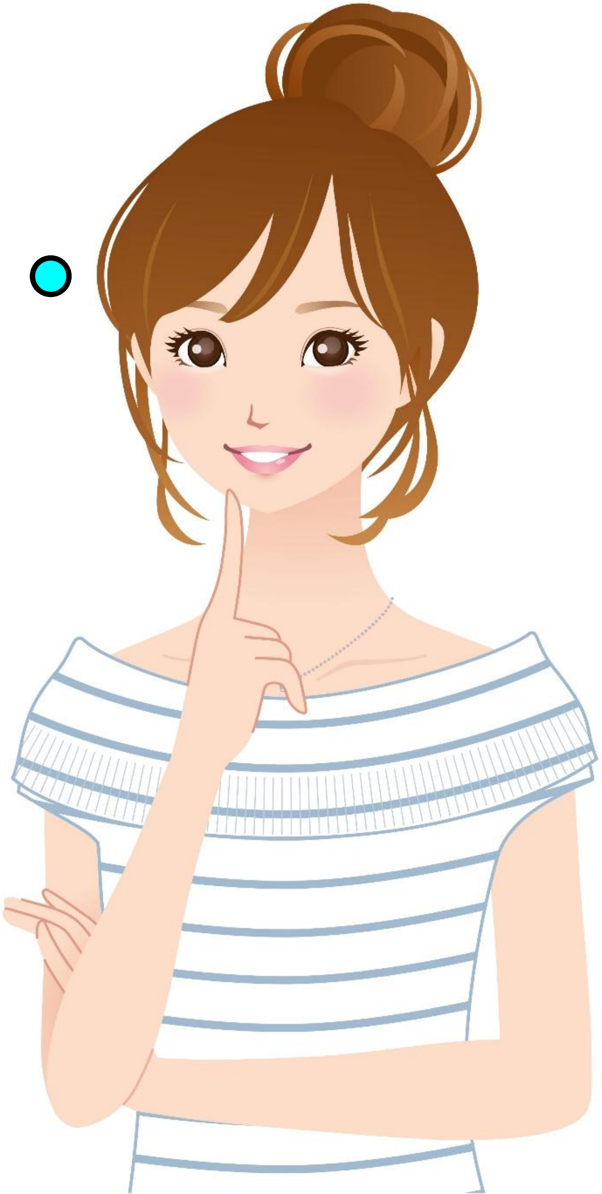
METALS

	Element	Extraction from Ore	
more reactive ↑ ↑	Potassium – K	Electrolysis	
	Sodium – Na	Electrolysis	
	Calcium – Ca	Electrolysis	
	Magnesium – Mg	Electrolysis	
	Aluminium – A*l*	Electrolysis	
	(Carbon – C)	(Not applicable)	← for reference
↓ ↓ less reactive	Zinc – Zn	Reduced by carbon	
	Iron – Fe	Reduced by carbon	
	Lead – Pb	Reduced by carbon	
	(Hydrogen – H)	(Not applicable)	← for reference
	Copper – Cu	Reduced by C or H_2	
	Silver – Ag	Reduced by C or H_2	

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\text{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\text{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → Silver

METALS

How is *iron* extracted from its ore in a *blast furnace*?



METALS

Iron occurs naturally in an ore called *haematite*. The main component of haematite is iron(III) oxide, Fe_2O_3.

Carbon ↓ *Hydrogen* ↓

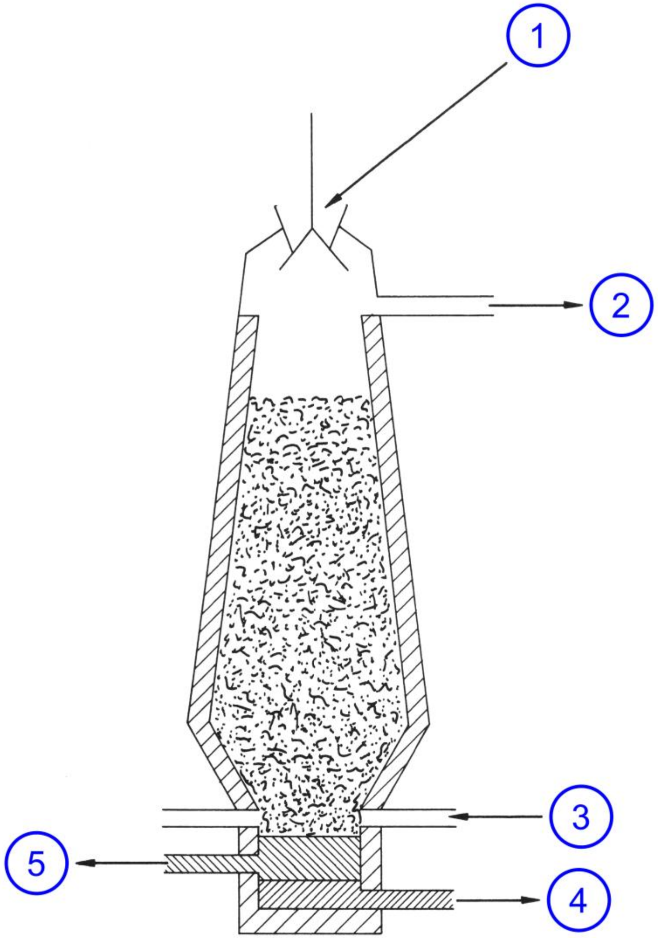
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

METALS

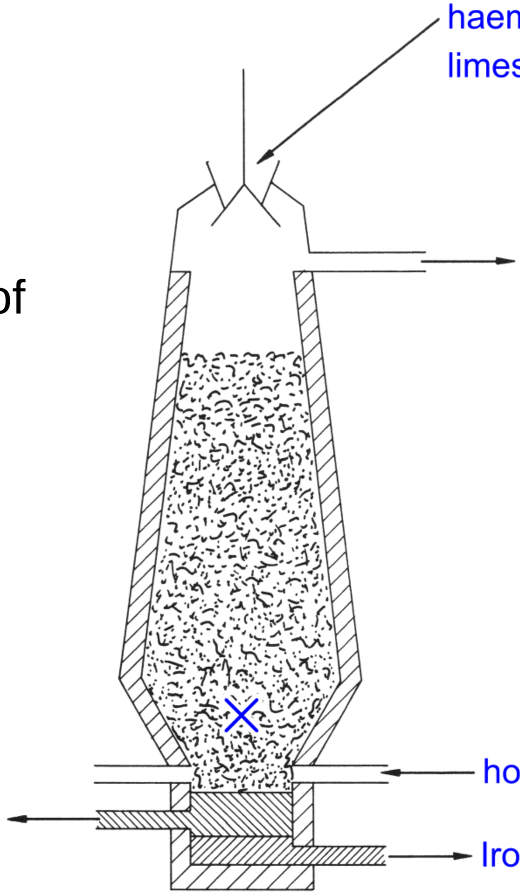
METALS

METALS



Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\text{Carbon}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\text{Hydrogen}}{\rightarrow}$ Copper → Silver

METALS



coke (carbon, C)
haematite (iron(III) oxide, Fe_2O_3)
limestone (calcium carbonate, $CaCO_3$)

waste gases:
carbon monoxide, CO
carbon dioxide, CO_2

hot air (oxygen, O_2)

slag (calcium silicate, $CaSiO_3$)

Iron, Fe

× Marks the hottest part of the blast furnace where temperatures can reach up to 1700 °C.

Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- There are *five* important reactions that take place in the blast furnace:

1) The coke (carbon) reacts with oxygen in the hot air to form carbon dioxide. This is an *exothermic reaction* which raises the temperature inside the blast furnace:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

2) The limestone (calcium carbonate) undergoes *thermal decomposition* to form calcium oxide and carbon dioxide:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium → *Carbon* ↓ → Zinc → Iron → Lead → *Hydrogen* ↓ → Copper → Silver

METALS

3) The carbon dioxide reacts with more coke (carbon) to form carbon monoxide:

$$CO_2(g) + C(s) \rightarrow 2CO(g)$$

4) The carbon monoxide reacts with the haematite (iron(III) oxide) to form iron and carbon dioxide:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(\mathit{l}) + 3CO_2(g)$$

5) The haematite contains sand (silica, SiO_2) as an impurity. The calcium oxide formed in reaction 2) reacts with the sand to form slag (calcium silicate, $CaSiCO_3$) which can be easily separated from the molten iron:

$$CaO(s) + SiO_2(s) \rightarrow CaSiO_3(\mathit{l})$$

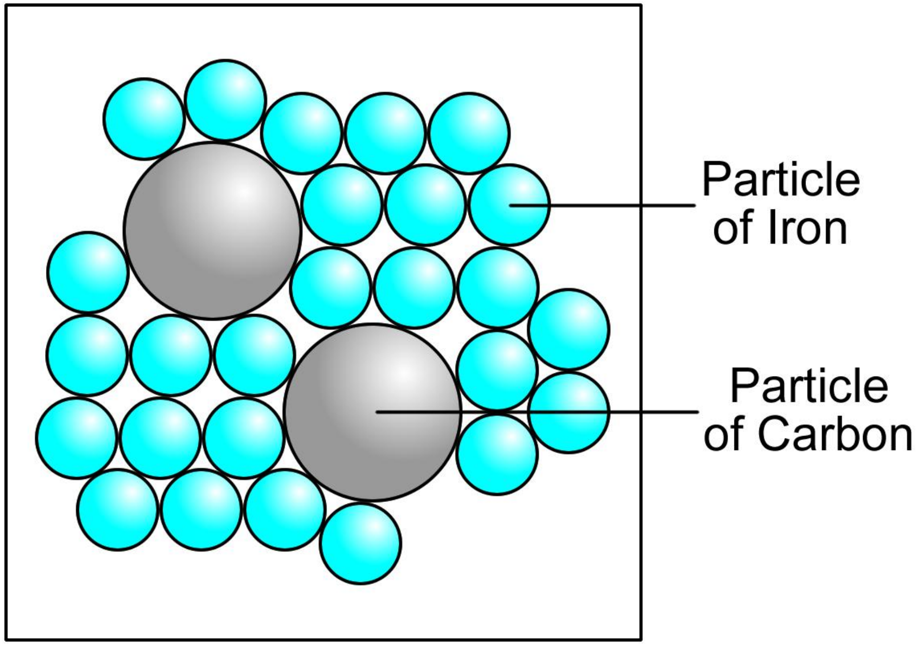
Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

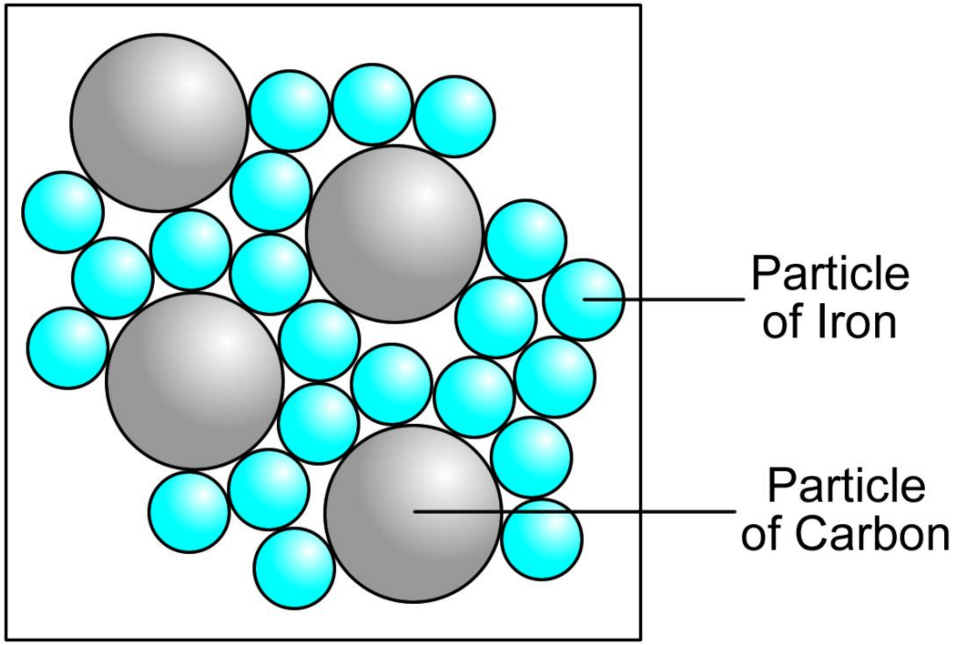
What are the properties of the different *alloys* of iron?

METALS



- *Low carbon steel* – compared to high carbon steel – is relatively soft, and therefore more easily shaped (*i.e.* more malleable and ductile). Low carbon steel is used to manufacture the bodies of motorcars.

METALS

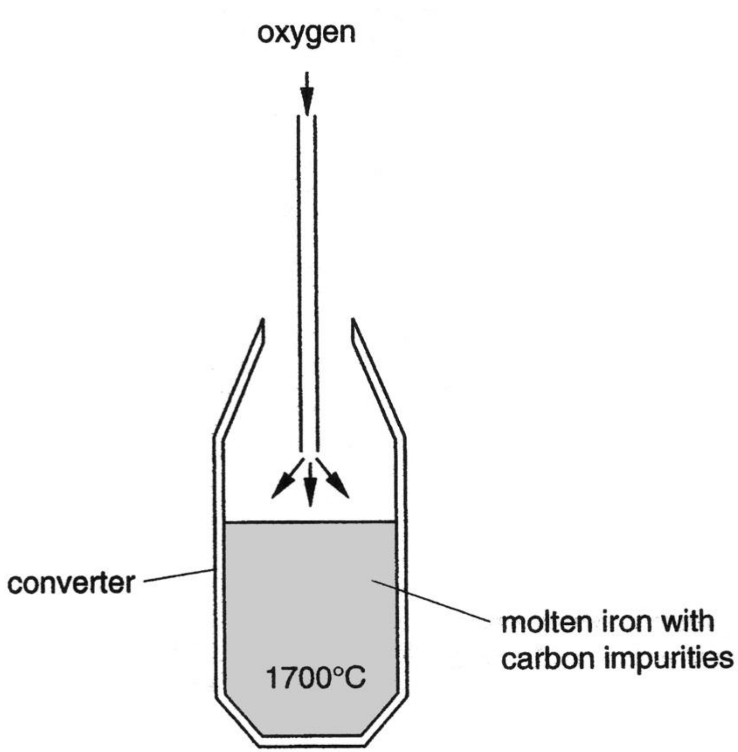


- *High carbon steel* – compared to low carbon steel – is very hard (less malleable and ductile), but also quite brittle. High carbon steel is used to make springs, knife blades, and masonry drills.

Carbon ↓ *Hydrogen* ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS



- Steel with different levels of carbon can be made by blowing oxygen through molten iron in a *converter*.

- Carbon reacts with the oxygen and removed as carbon dioxide, thus lowering the carbon content of the iron.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

- *Stainless steel* is resistant to corrosion. It is an alloy made by adding chromium to iron.

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\text{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\text{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → **Silver**

Summary – Iron and its Alloys

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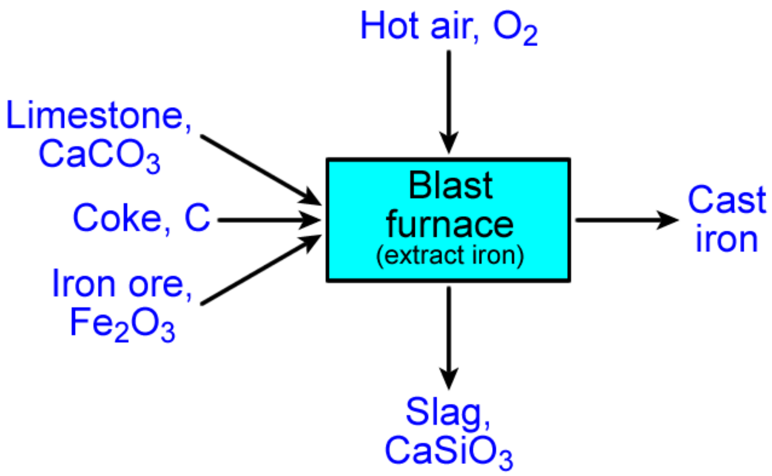
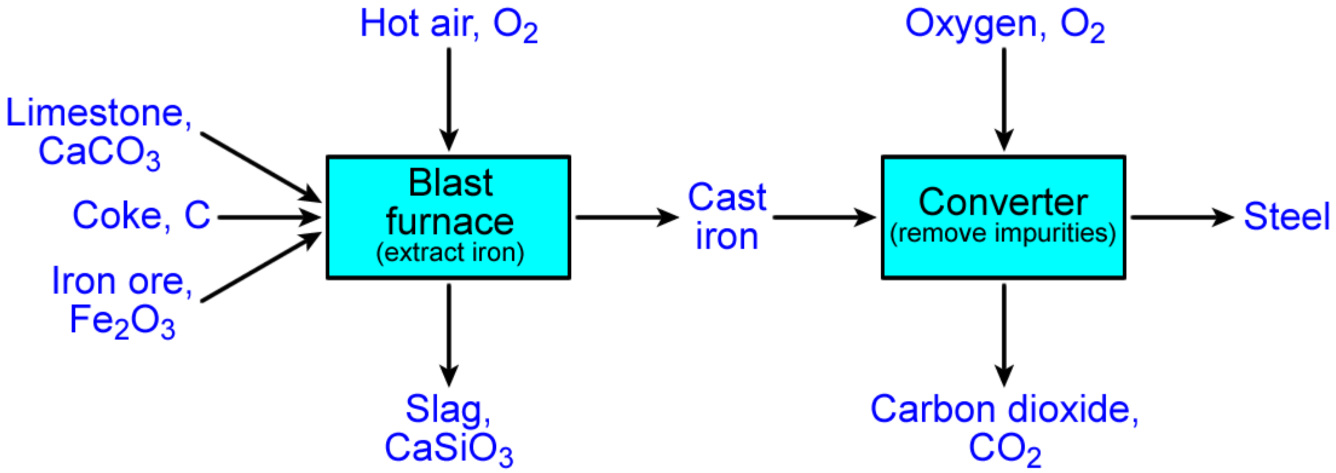
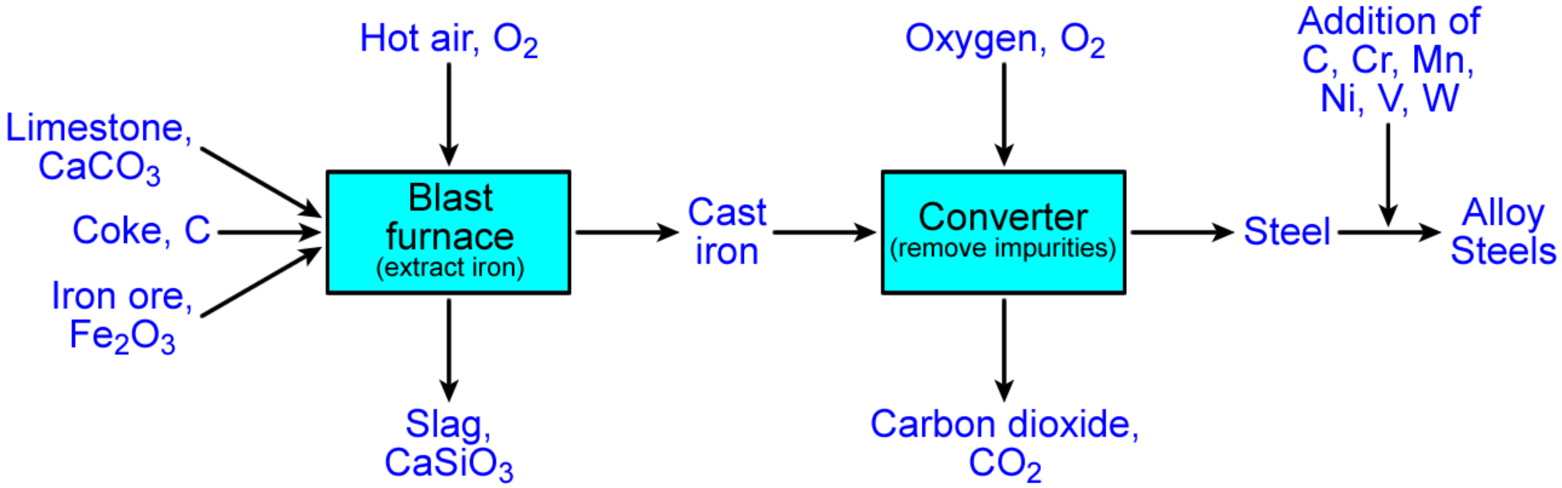


Summary – Iron and its Alloys

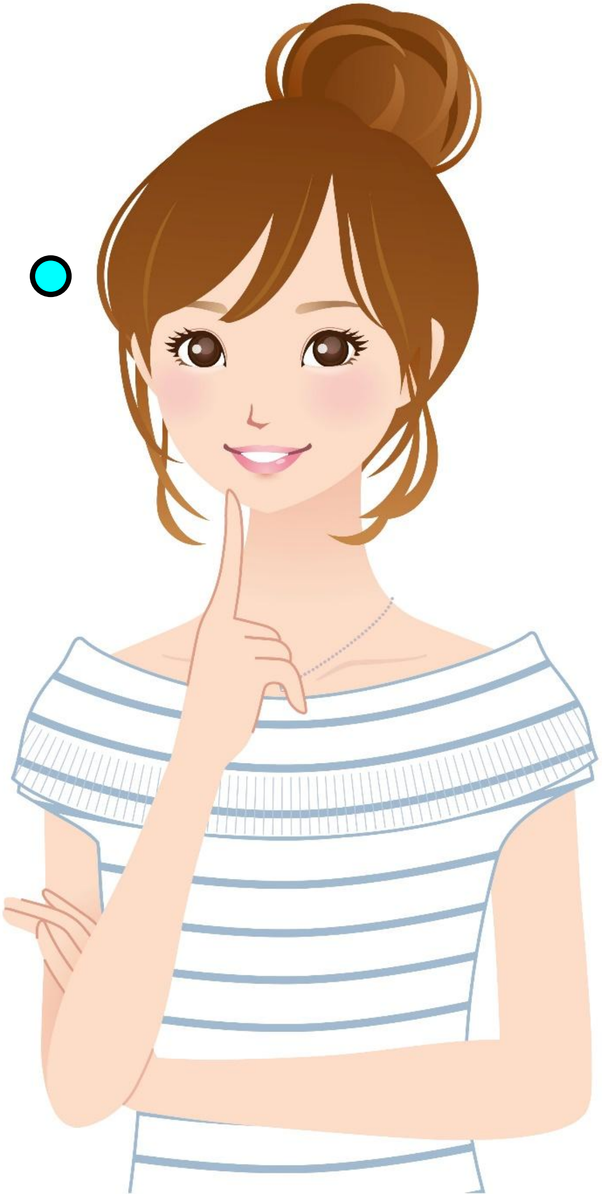


Summary – Iron and its Alloys



METALS

What are the different ways of preventing iron from *corroding*?

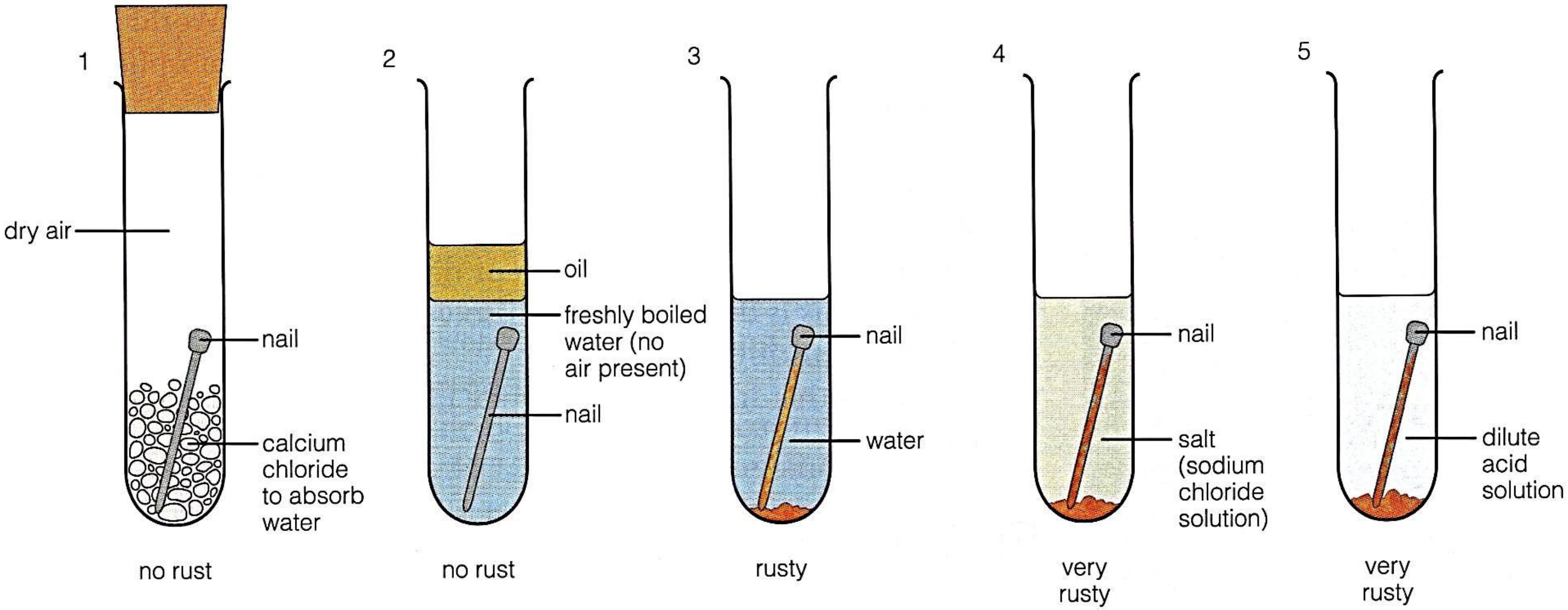


METALS

Corrosion

METALS

- Study the reactions shown below. What are the important conditions that cause iron to corrode / rust?



- *Air* (specifically, oxygen) must be present.
- *Water* must be present.

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

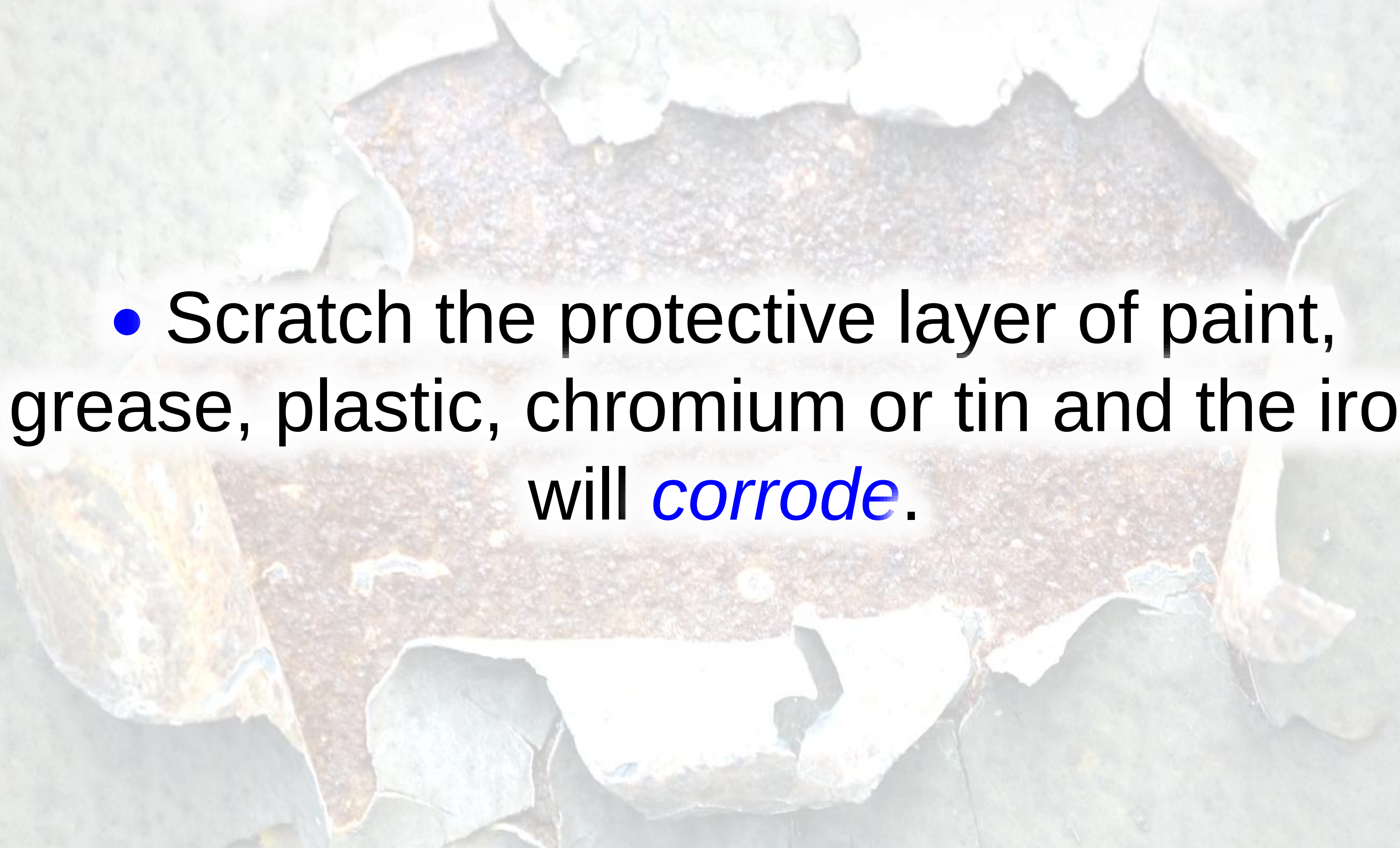
METALS

- Iron can be protected against corrosion / rusting by preventing air and water coming into contact with its surface.
- This can be easily achieved by covering the surface of the iron with a thin layer of:

→ Paint
→ Grease
→ Plastic
→ Chromium (chrome plating)
→ Tin (tin plating)
→ Zinc (galvanizing)

METALS

METALS



- Scratch the protective layer of paint, grease, plastic, chromium or tin and the iron will *corrode*.

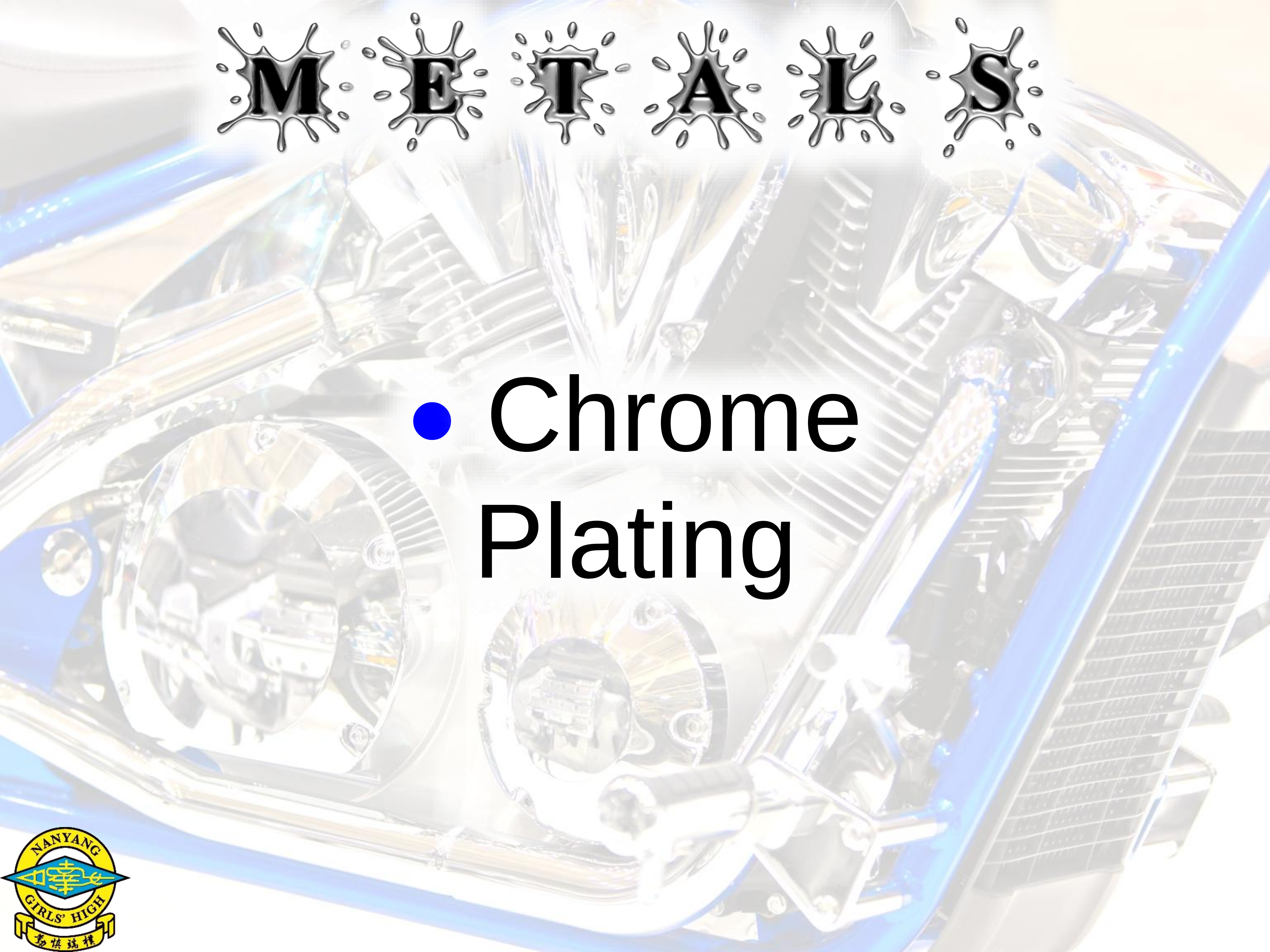
METALS

METALS

- Painting

METALS

METALS



- Chrome Plating

METALS

- Tin Plating

METALS

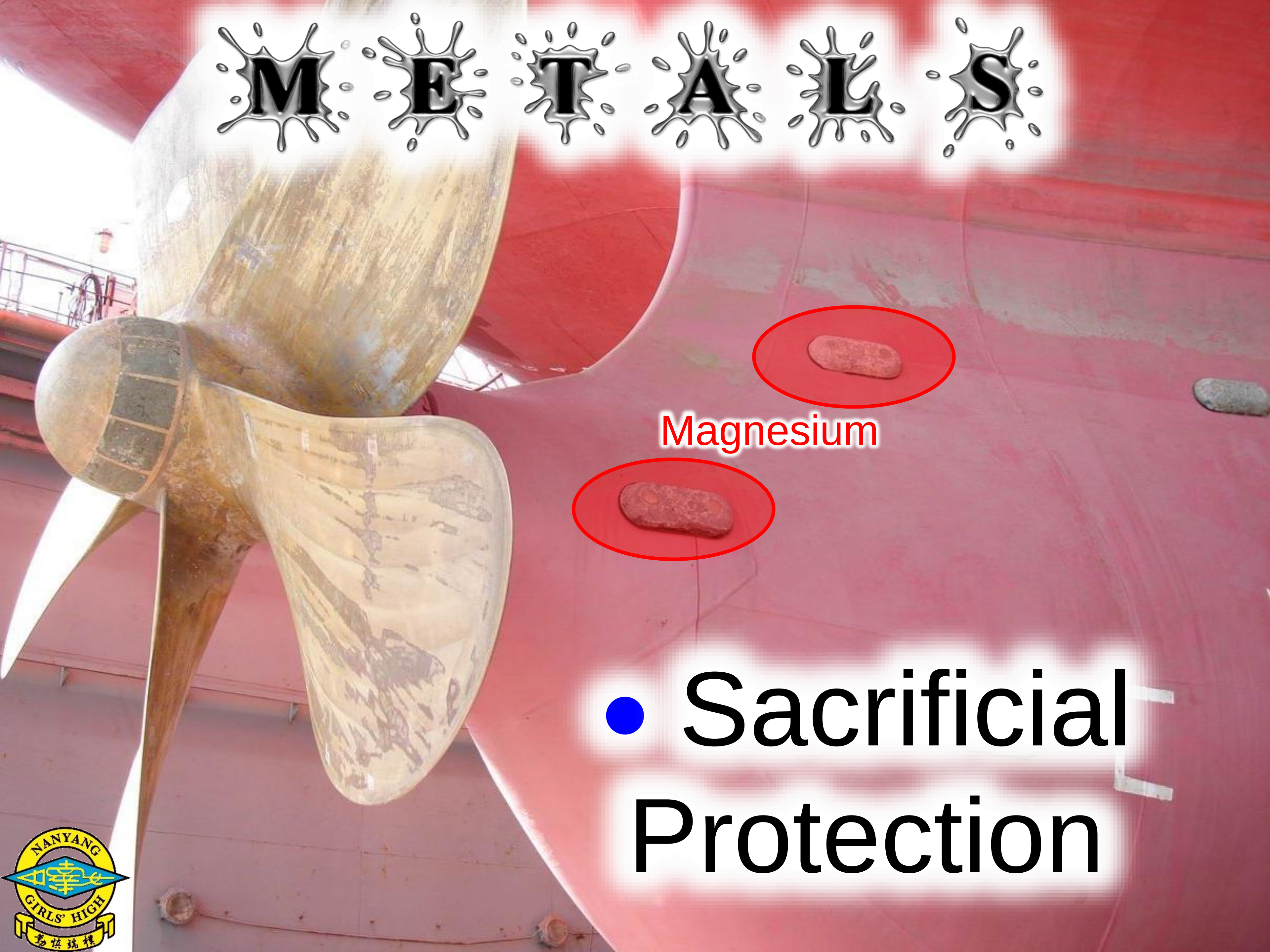
METALS

- Zinc Plating (Galvanizing)

METALS

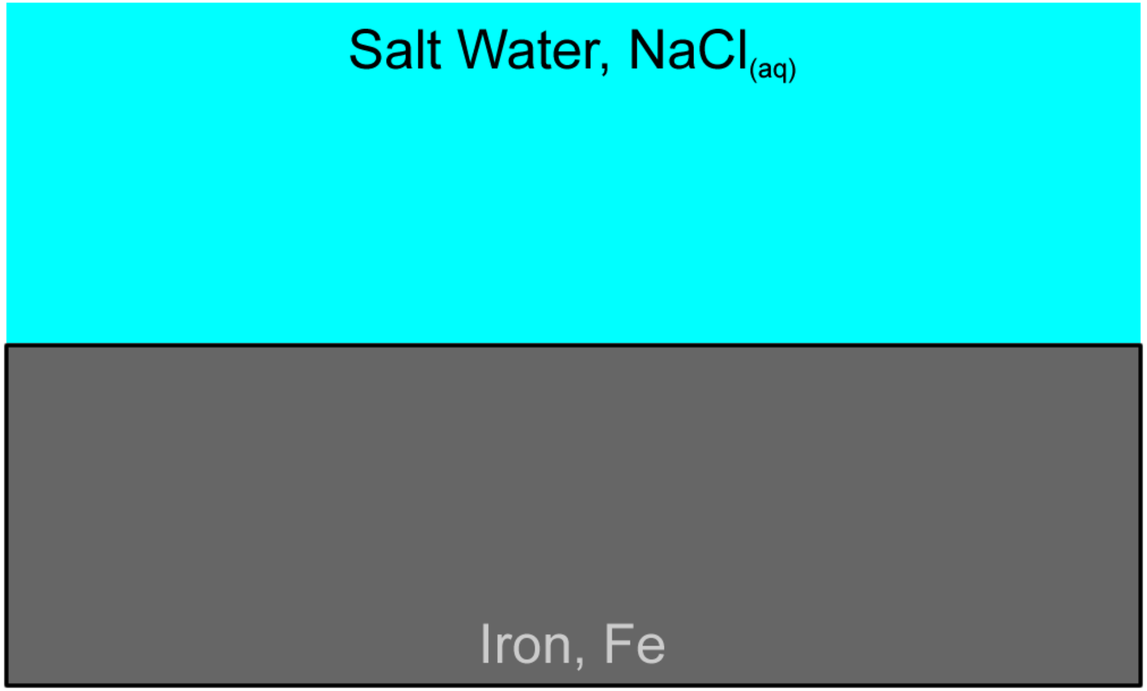
- Sacrificial Protection

METALS



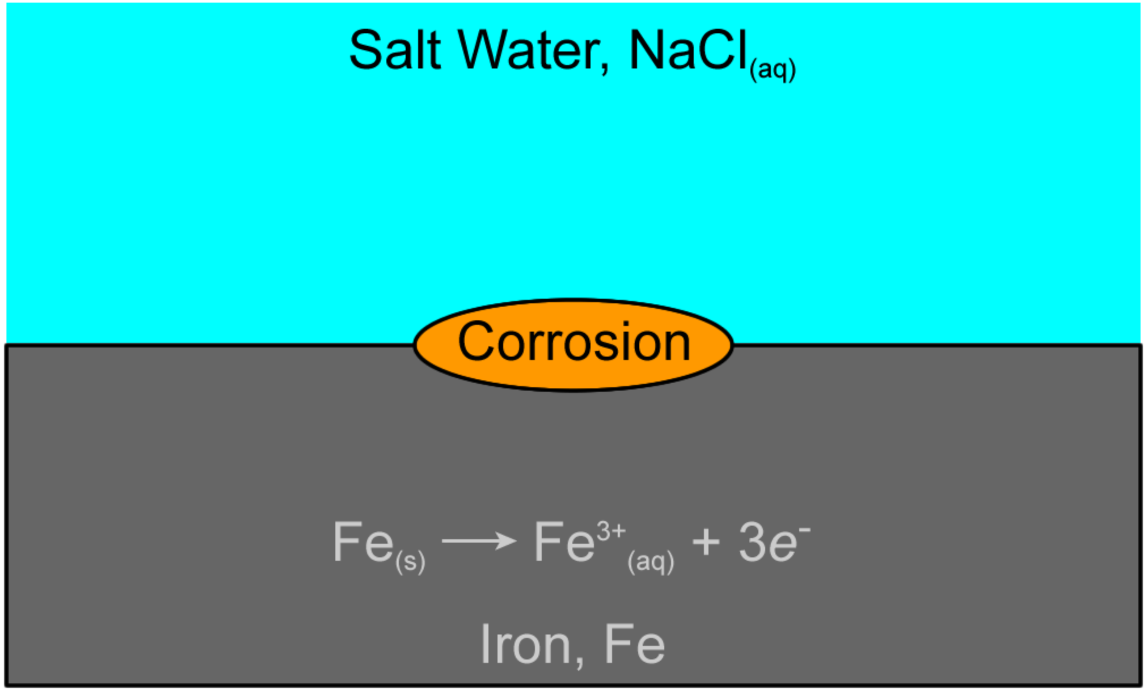
- Sacrificial Protection

Sacrificial Protection



- Iron in contact with air / oxygen and water will corrode.

Sacrificial Protection



Salt Water, $NaCl_{(aq)}$

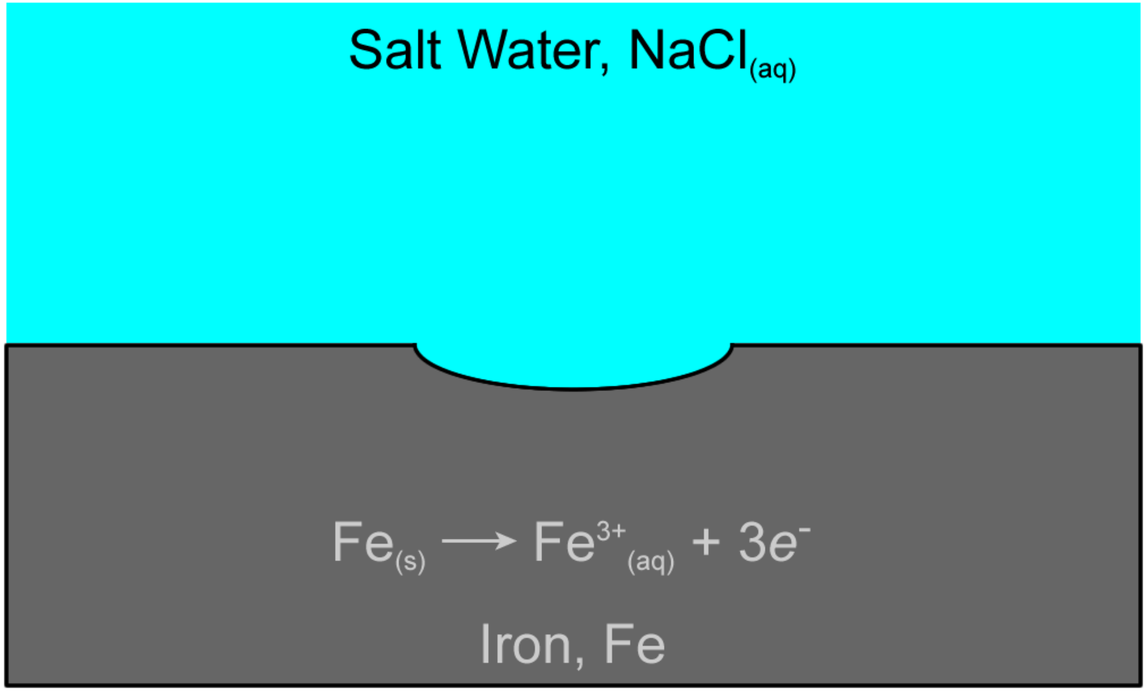
Corrosion

$Fe_{(s)} \longrightarrow Fe^{3+}_{(aq)} + 3e^-$

Iron, Fe

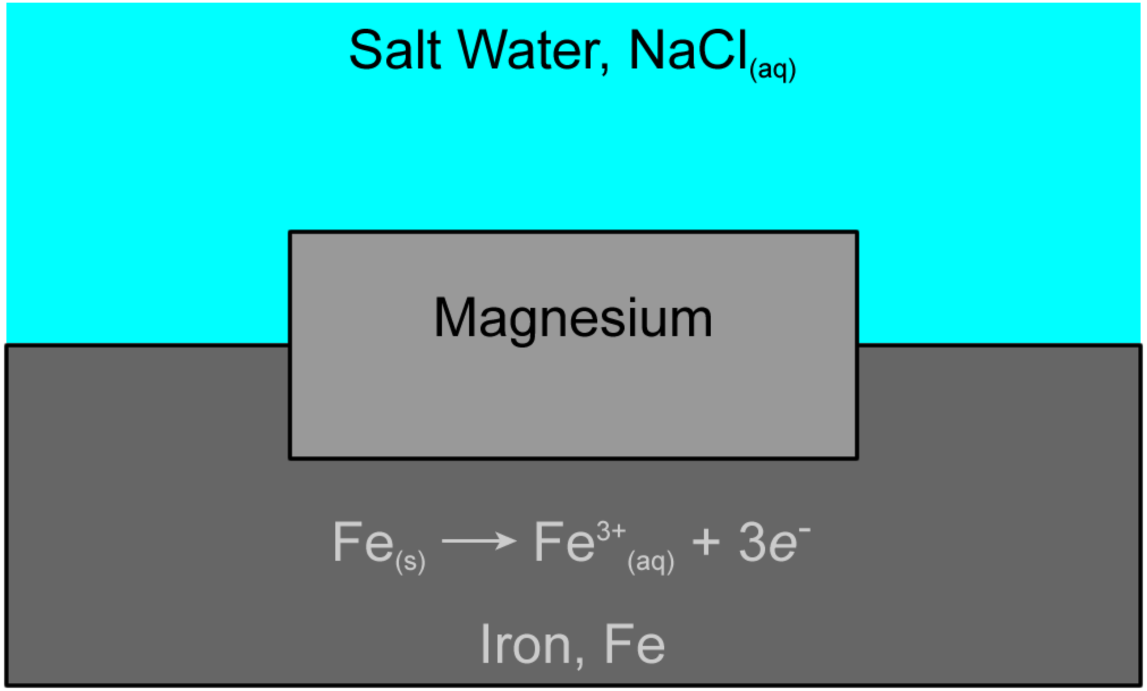
- Iron in contact with air / oxygen and water will corrode.

Sacrificial Protection



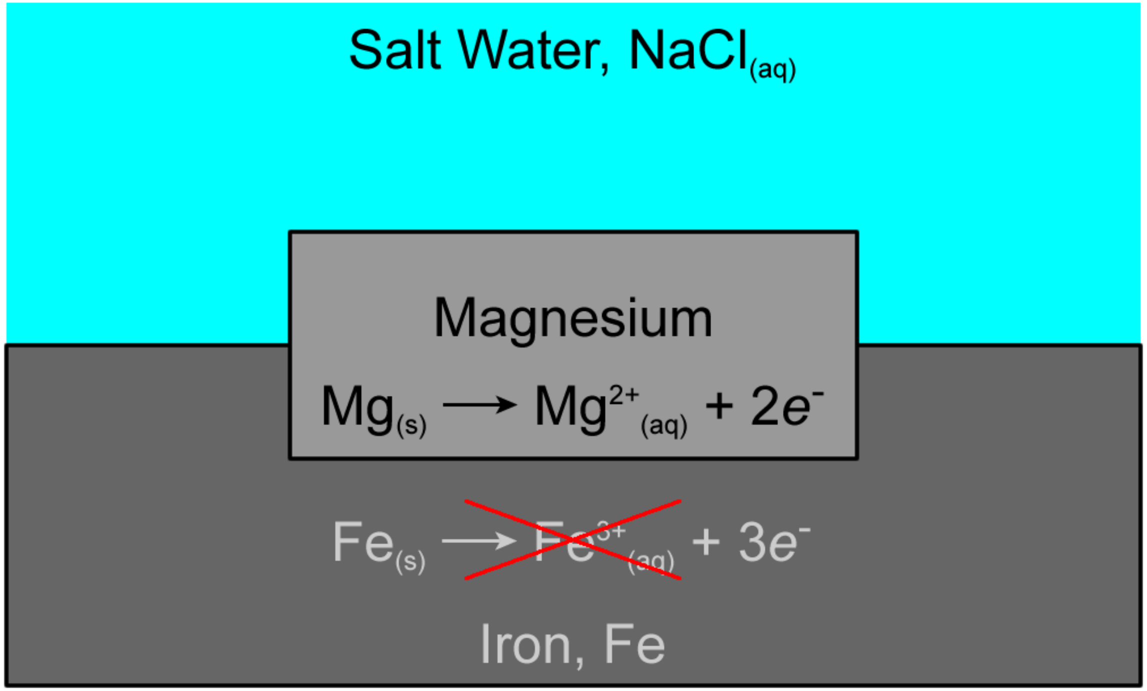
- Iron in contact with air / oxygen and water will corrode.

Sacrificial Protection



- Magnesium is more reactive than iron.
- The more reactive magnesium corrodes preferentially.

Sacrificial Protection



- Magnesium is more reactive than iron.
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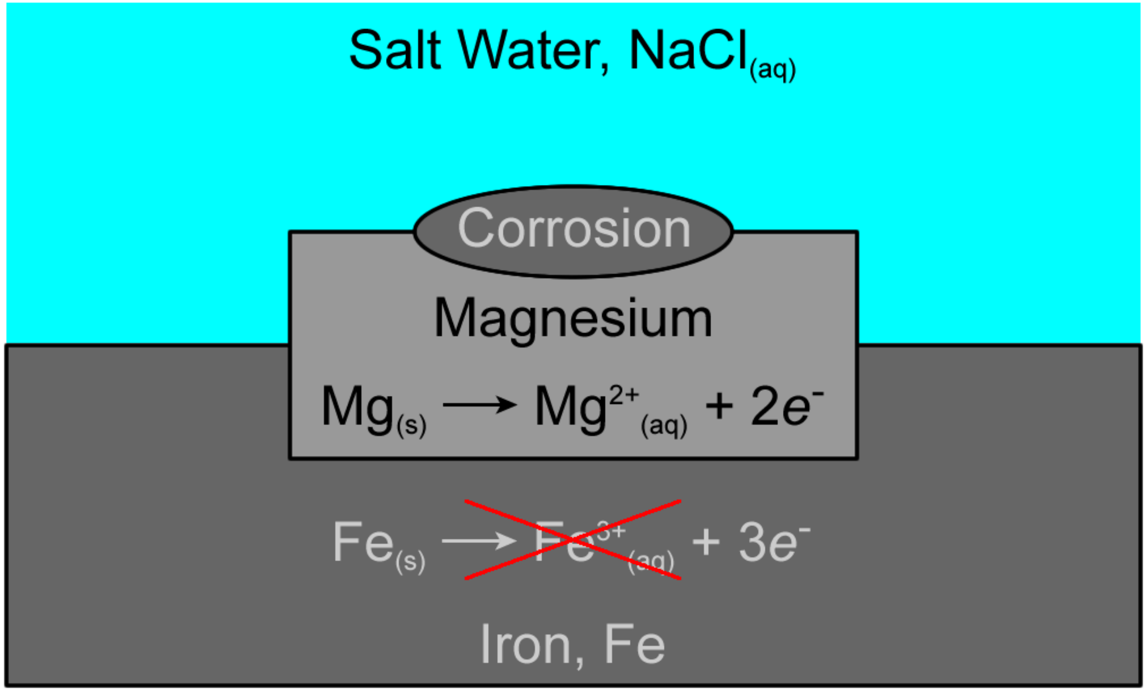
Potassium → Sodium → Calcium → Magnesium → Aluminium → (Carbon ↓) Zinc → Iron → Lead → (Hydrogen ↓) Copper → Silver

Sacrificial Protection



Salt Water, $NaCl_{(aq)}$

Corrosion

Magnesium

$Mg_{(s)} \longrightarrow Mg^{2+}_{(aq)} + 2e^-$

$Fe_{(s)} \longrightarrow \cancel{Fe^{3+}_{(aq)}} + 3e^-$

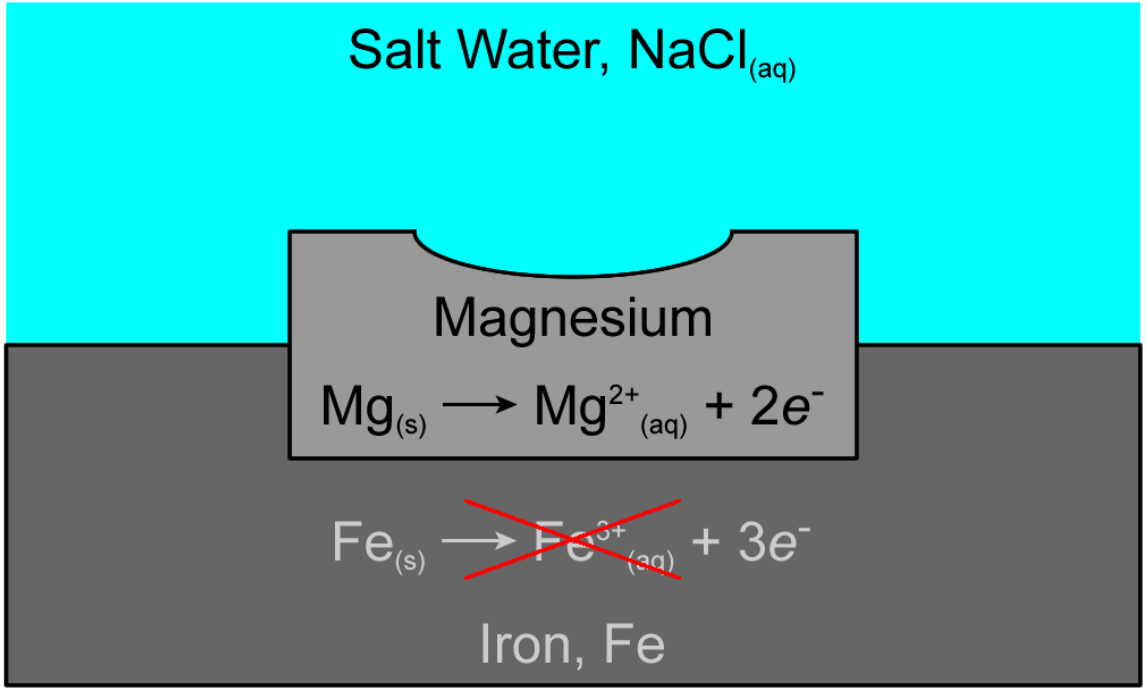
Iron, Fe

- Magnesium is more reactive than iron.
- The more reactive magnesium corrodes preferentially.

Carbon ↓ (between Aluminium and Zinc); Hydrogen ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

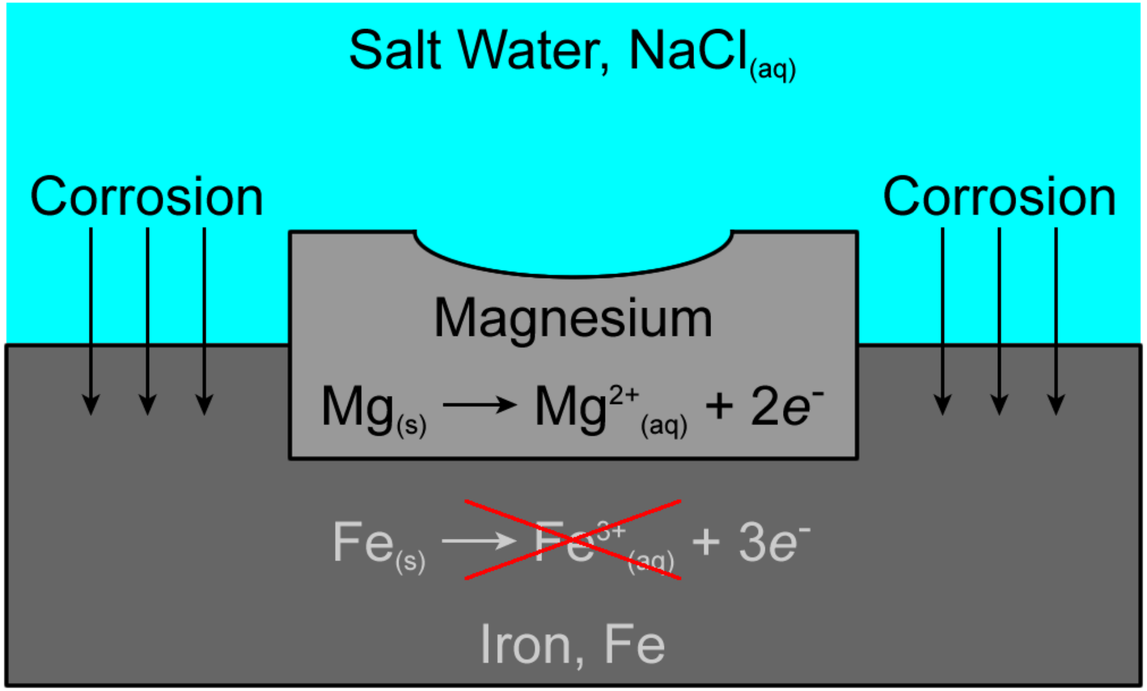
Sacrificial Protection



- Magnesium is more reactive than iron.
- The more reactive magnesium corrodes preferentially.

Potassium → Sodium → Calcium → Magnesium → Aluminium → *Carbon* ↓ → Zinc → Iron → Lead → *Hydrogen* ↓ → Copper → **Silver**

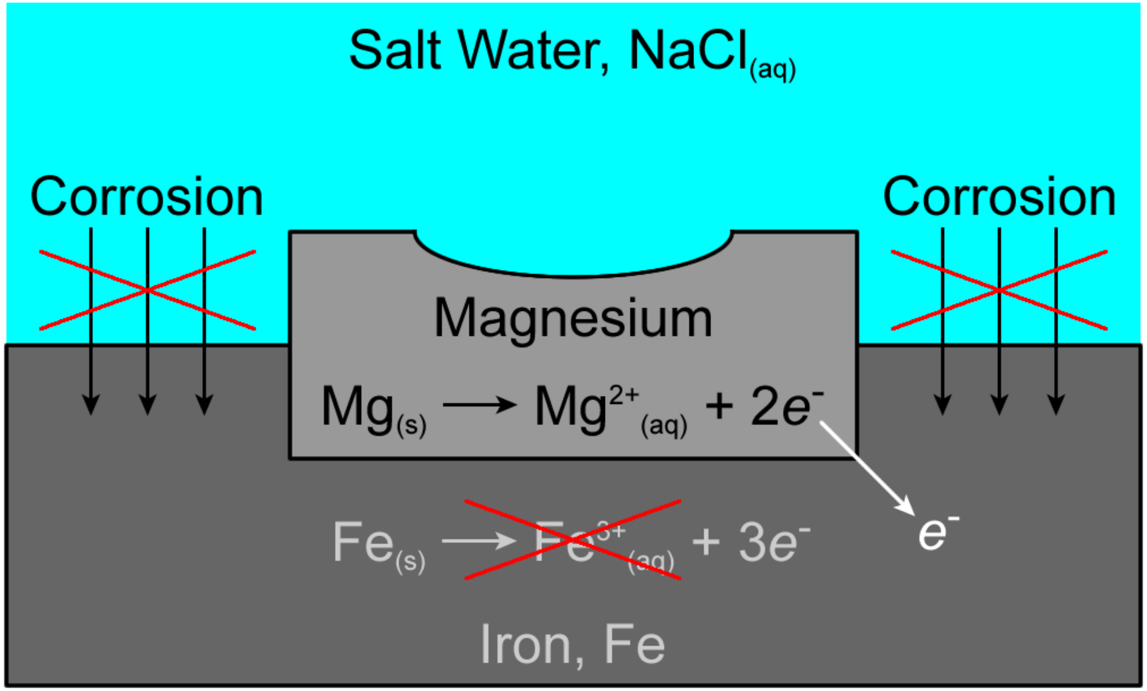
Sacrificial Protection



- Magnesium is more reactive than iron.
- The more reactive magnesium corrodes preferentially.

Potassium → Sodium → Calcium → Magnesium → Aluminium → (Carbon) Zinc → Iron → Lead → (Hydrogen) Copper → Silver

Sacrificial Protection



- Magnesium is more reactive than iron.
- The more reactive magnesium corrodes preferentially.

Carbon ↓ Hydrogen ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

Sacrificial Protection

- Magnesium is more reactive that iron. This means that the magnesium will oxidise / corrode in preference to the iron, therefore preventing the iron from corroding.

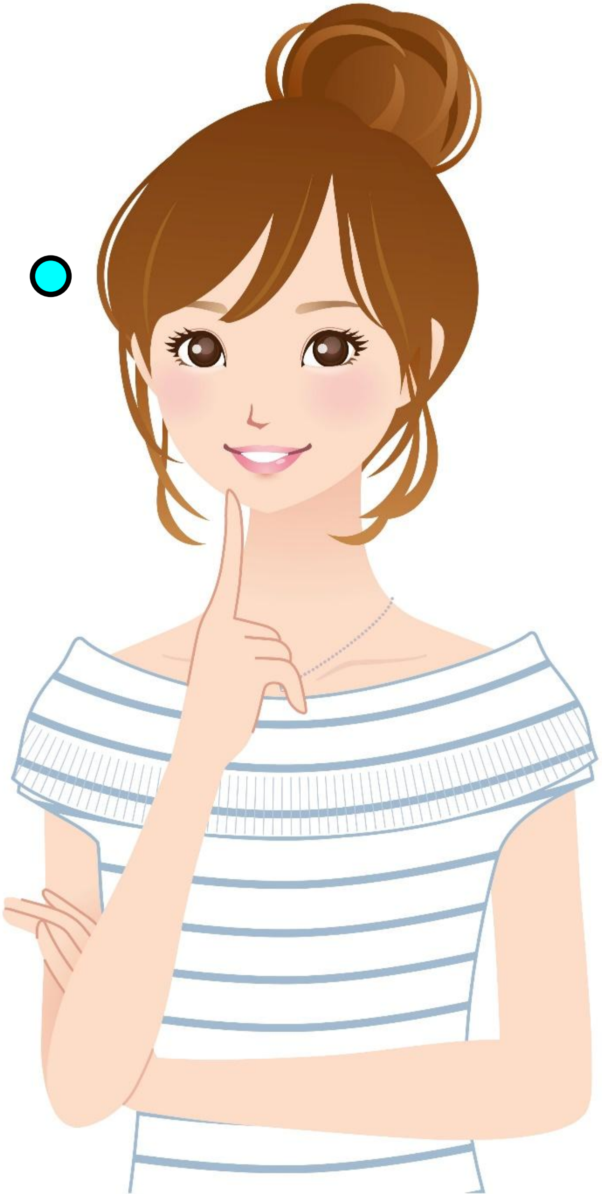
- Note that, unlike many of the other ways of protecting iron, it is not necessary for a thin layer of magnesium to cover the entire surface of the iron *i.e.* the iron can be exposed to air / oxygen and water and – as long as the iron is in contact with a block of magnesium – the iron will not corrode.

METALS

- How is this bridge protected from corrosion?

METALS

What are the advantages and disadvantages of *recycling* metals?



METALS

Carbon ↓ (between Aluminium and Zinc) — Hydrogen ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- The Earth’s resources are finite, meaning that there is a limited supply of metal ores, and a limited supply of the chemicals that are necessary to extract the metals from their ores.

- Recycling therefore means that metals will be available in large quantities for a longer time.

- Recycling metals also saves energy and reduces the volume of greenhouse gases (*e.g.* carbon dioxide) that are released, reducing the effects of climate change.

METALS

- Recycling metals conserves the Earth’s natural resources, saves energy and reduces pollution. It also makes good economic sense to recycle metals, as producing 1000 kg of iron through recycling is much cheaper than extracting the same mass of iron from its ore. The recycling of metals also produces jobs for people.

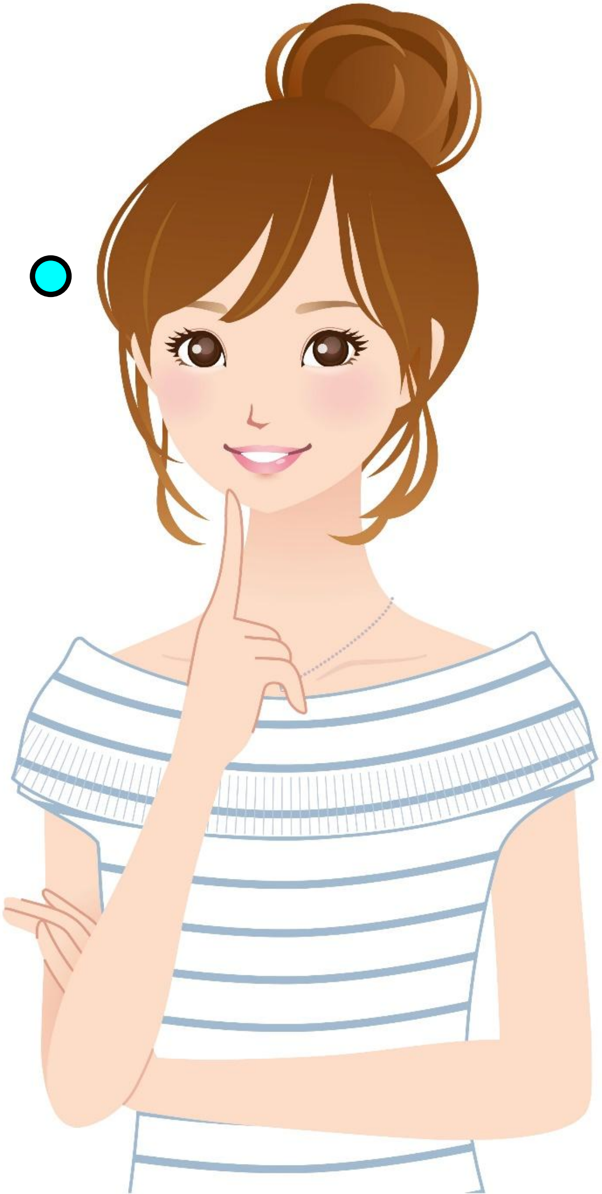
- The main disadvantage or problem with recycling metals is public apathy. Only a small percentage of the metals that could be recycled are recycled – people need to be educated and encouraged to recycle metals. The recycling of metals requires purpose built factories, and the collection of metals for recycling from households and industry poses another problem.

METALS

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\textit{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\textit{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → Silver

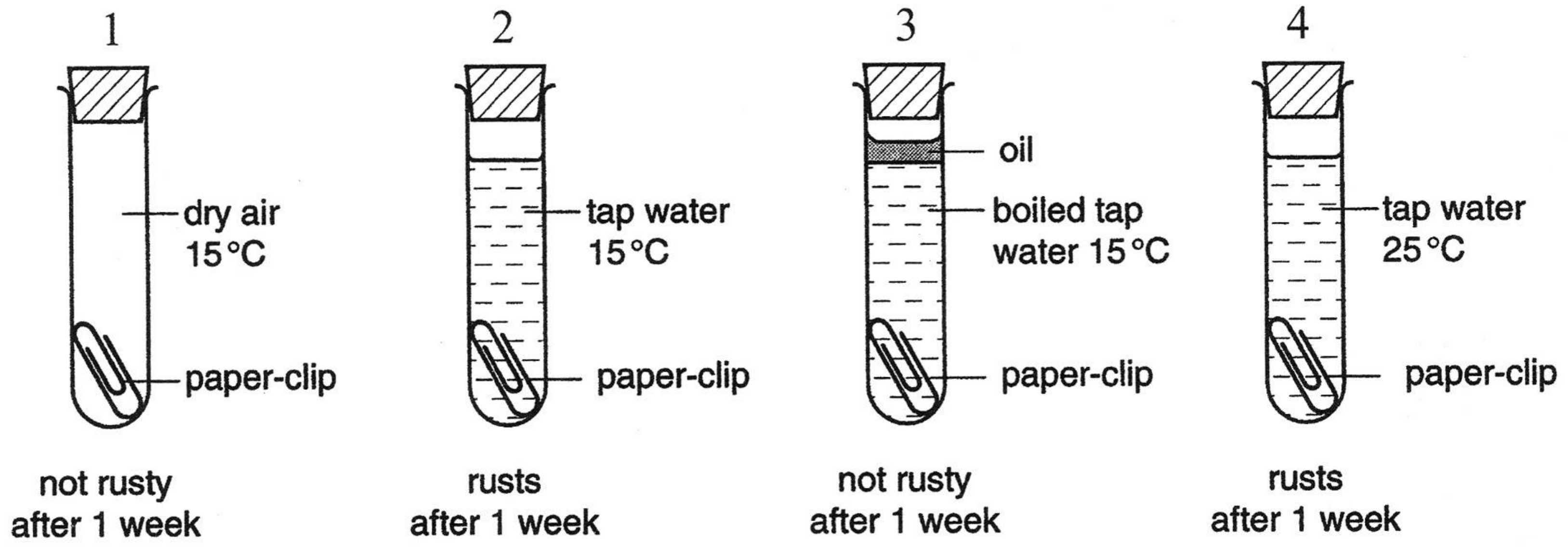
METALS

Could I please have some *questions* to test my understanding of what I have learnt?



METALS

1. Four experiments on rusting are shown.



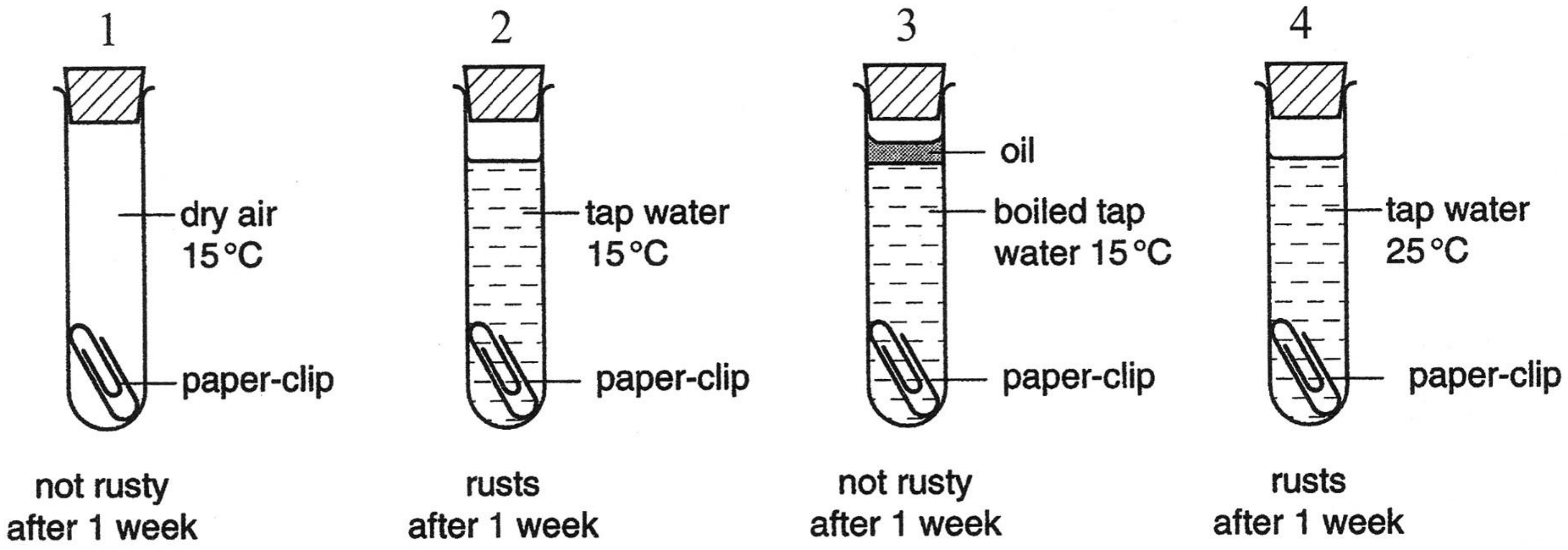
- Which two experiments can be used to show that air is needed for iron to rust?

A 1 and 3 **B** 1 and 4 **C** 2 and 3 **D** 2 and 4

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\text{Carbon}}{\overset{\downarrow}{\rightarrow}}$ Zinc → Iron → Lead $\overset{\text{Hydrogen}}{\overset{\downarrow}{\rightarrow}}$ Copper → Silver

METALS

1. Four experiments on rusting are shown.



- Which two experiments can be used to show that air is needed for iron to rust?

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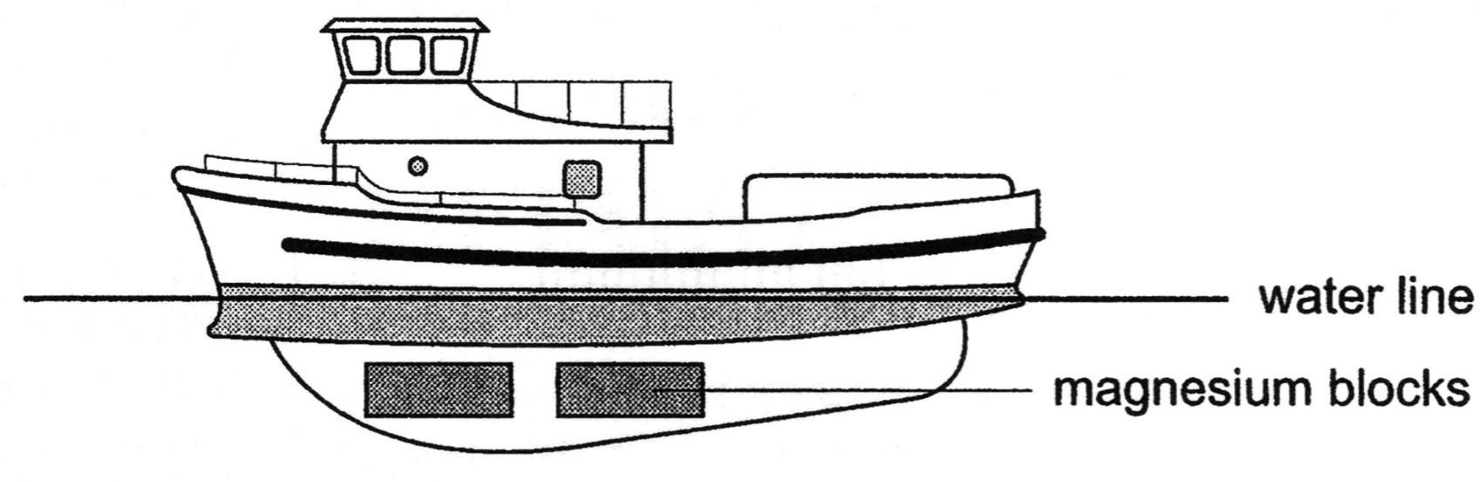
Carbon ↓ between Aluminium and Zinc; *Hydrogen* ↓ between Lead and Copper

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

2. The diagram shows a boat made from iron. Some magnesium blocks are attached to the iron below the water line.



- Why does the magnesium stop the iron from rusting?

A The magnesium reacts in preference to the iron.
B The magnesium reacts to form a protective coating over the iron.
C The magnesium forms an alloy with the iron.
D The magnesium stops oxygen in the water from getting to the iron.

Carbon ↓ (between Aluminium and Zinc)

Hydrogen ↓ (between Lead and Copper)

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METALS

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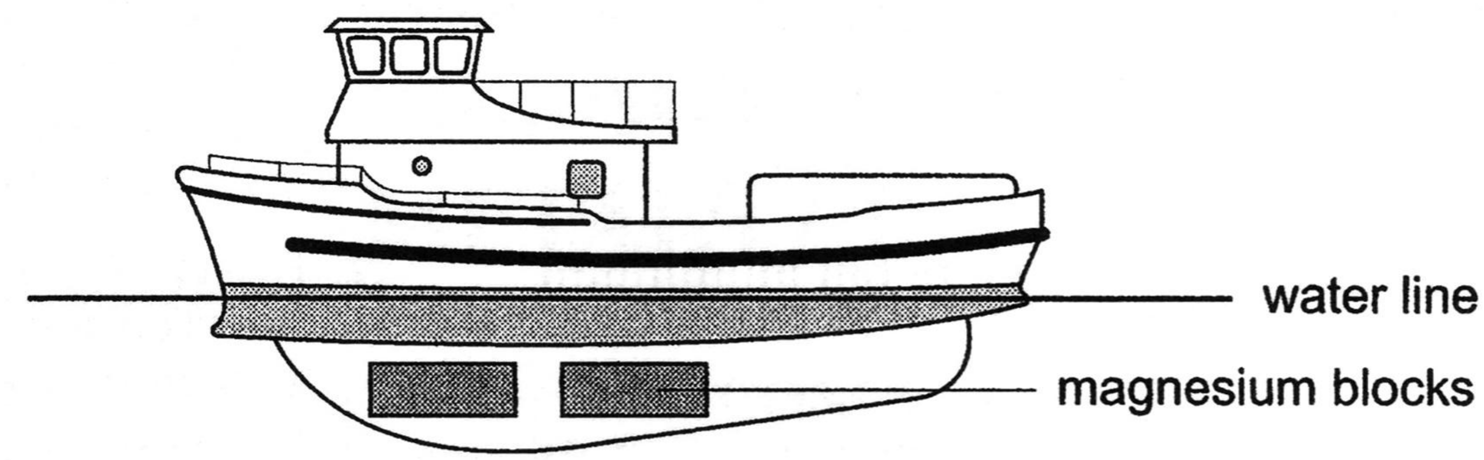

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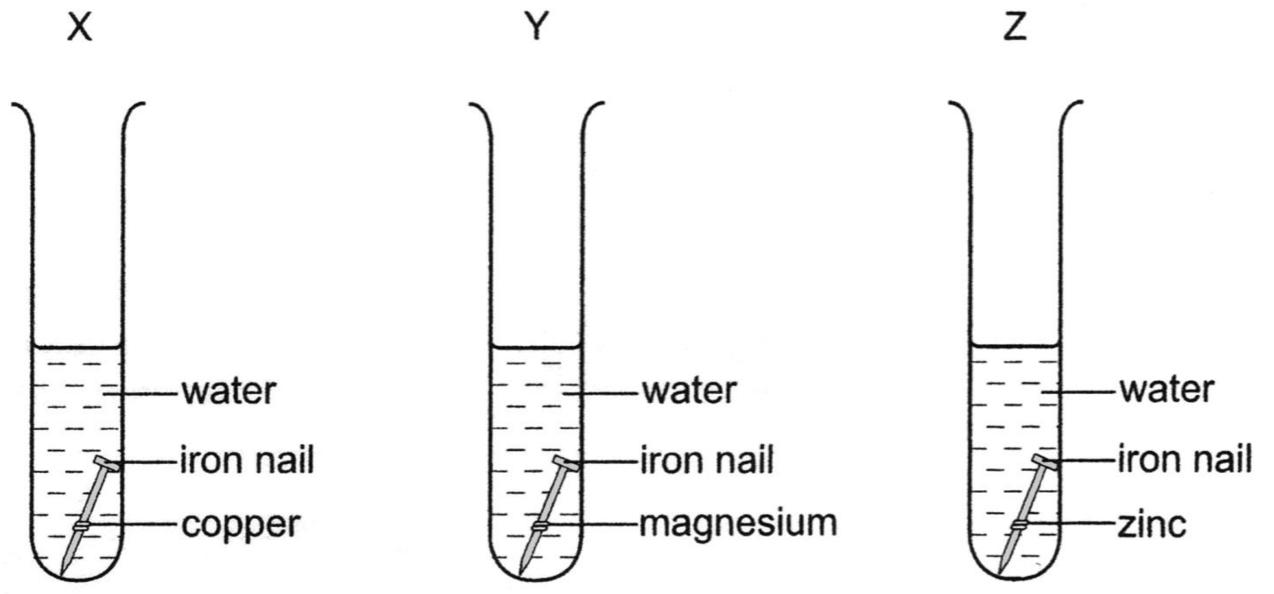
Carbon ↓ *Hydrogen* ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

3. Experiments are set up to investigate the sacrificial protection of iron.



- In which test-tubes will the iron rust?

A X only **B** Y only **C** X and Z only **D** Y and Z only

Carbon ↓ (between Aluminium and Zinc)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

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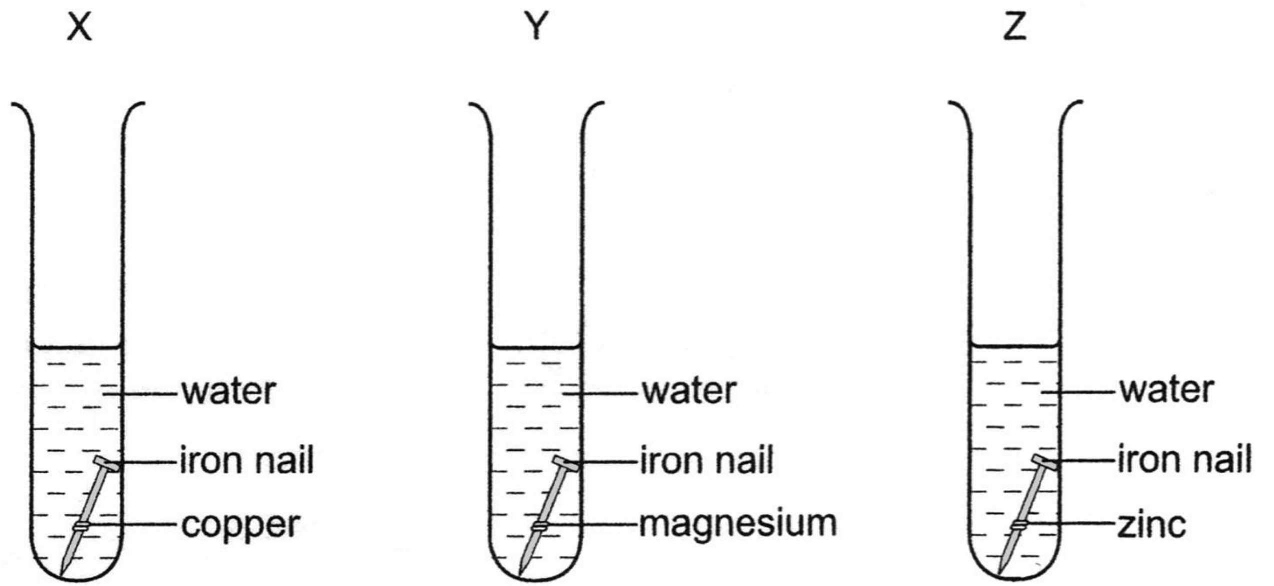


- In which test-tubes will the iron rust?

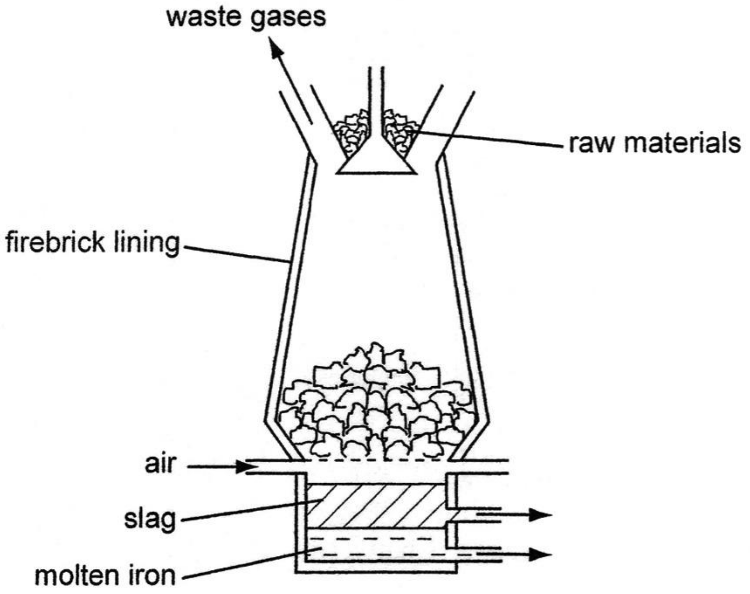
[✓ **A**] X only **B** Y only **C** X and Z only **D** Y and Z only

Carbon ↓ (between Aluminium and Zinc)
Hydrogen ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

4. Iron is extracted in the blast furnace using the raw materials haematite, coke and limestone.



- Which substance undergoes thermal decomposition?

A Limestone **B** Carbon dioxide

C Haematite **D** Slag

Carbon ↓ *Hydrogen* ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

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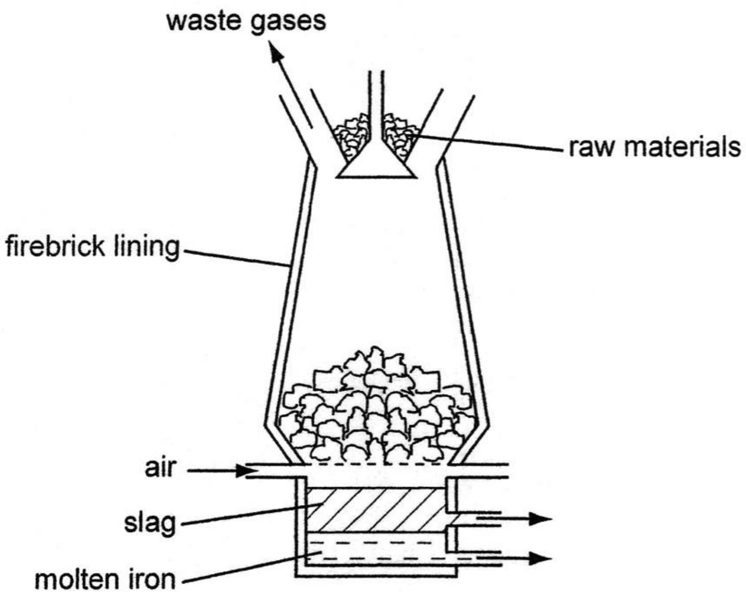

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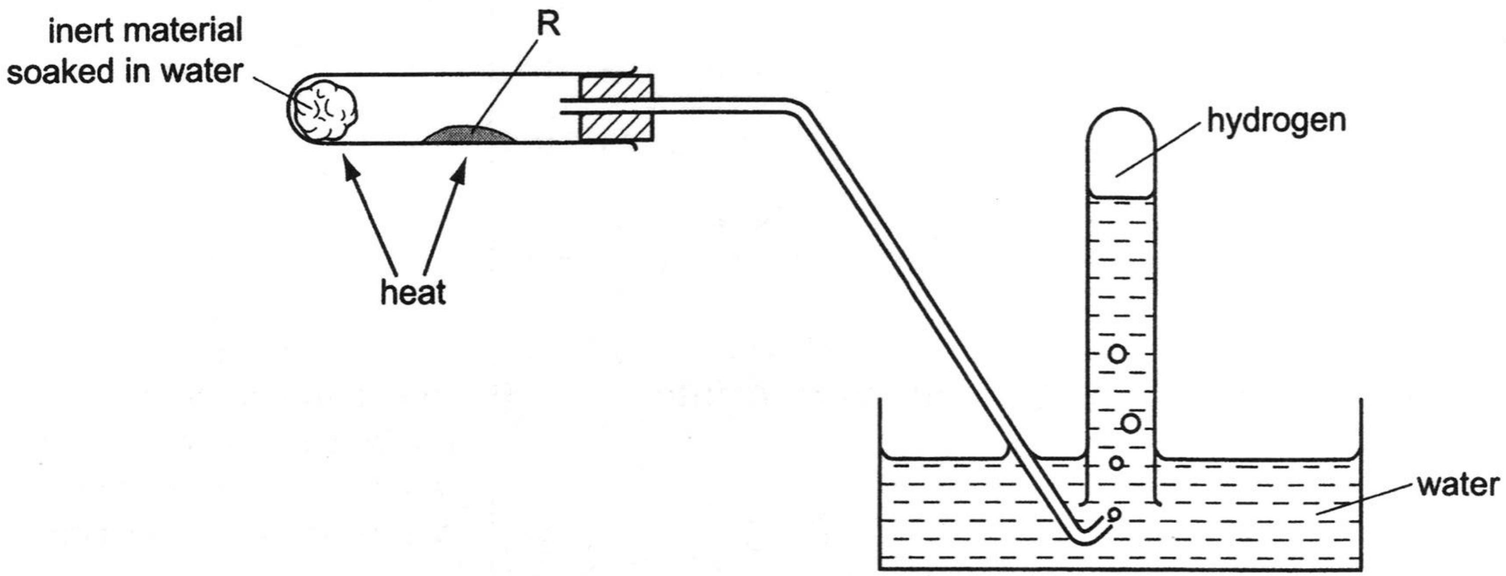
Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\text{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\text{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → Silver

METALS

5. The diagram shows an experiment to produce and collect hydrogen.



- What is **R**?

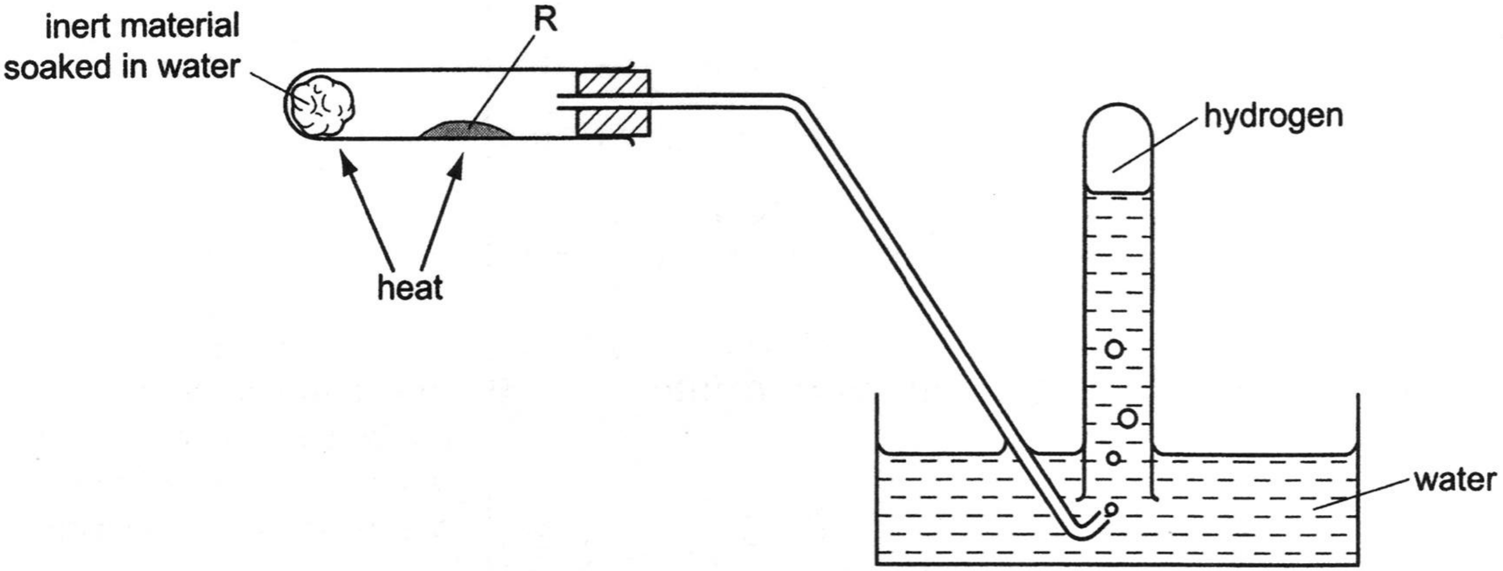
A Copper(II) oxide **B** Iron **C** Lead **D** Lead(II) oxide

Carbon ↓ Hydrogen ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

5. The diagram shows an experiment to produce and collect hydrogen.



- What is **R**?

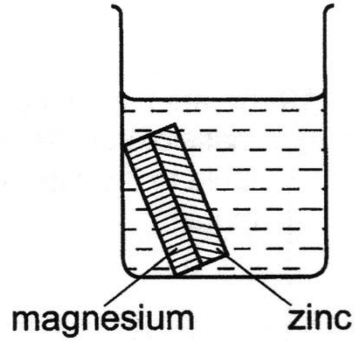
A Copper(II) oxide [✓ **B**] Iron **C** Lead **D** Lead(II) oxide

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

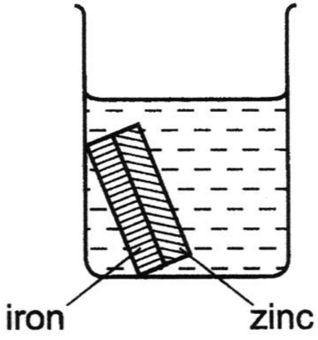
METALS

6. Each beaker contains two strips of metal fastened together and immersed in hydrochloric acid. All the strips are of the same size. After 5 minutes, which beaker contains the **least** amount of zinc ions?

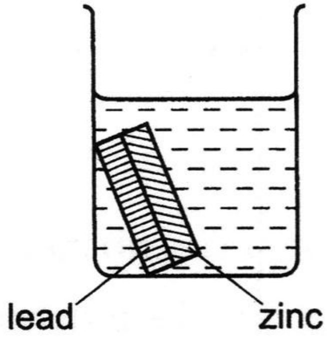
A



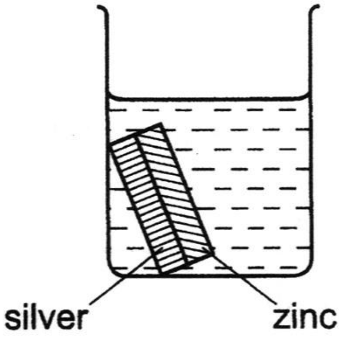
C



B



D



Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

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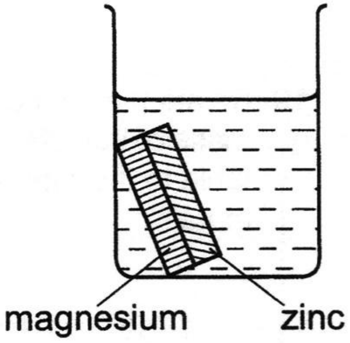
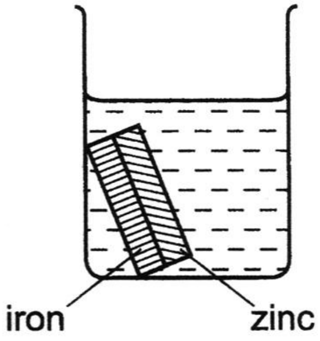
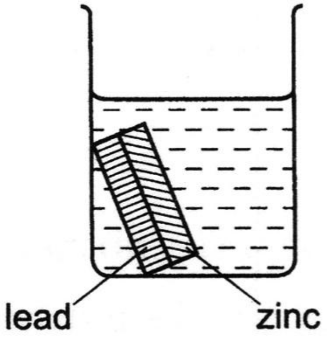
[✓ **A**]



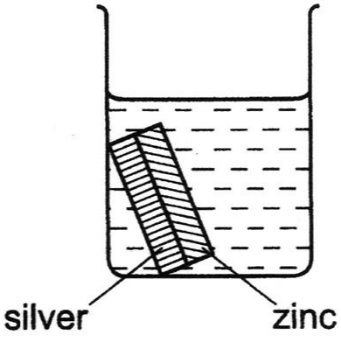
C



B



D

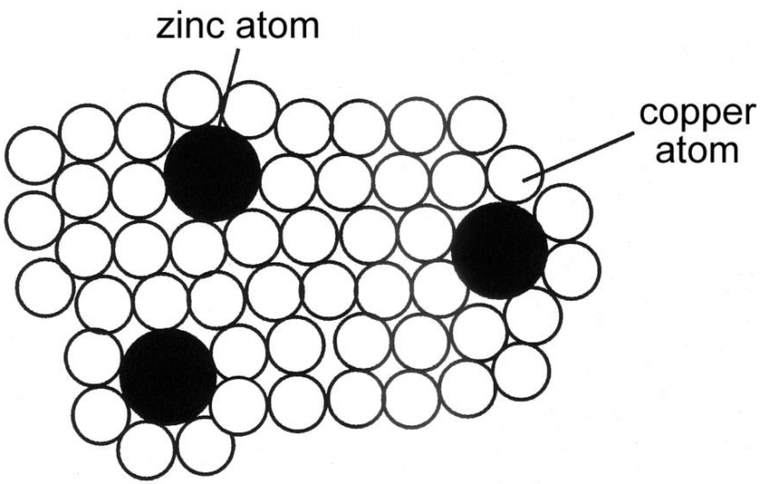


Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

7. The diagram shows the structure of brass.



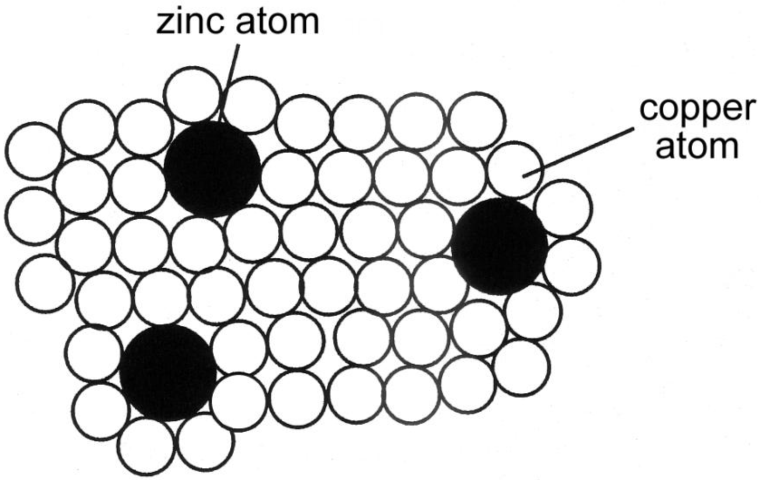
- Why is brass harder than pure copper?

A The zinc atoms form strong covalent bonds with the copper atoms.
B The zinc atoms have more electrons than the copper atoms.
C The zinc atoms prevent the “sea” of electrons moving freely.
D The zinc atoms prevent the layers of copper atoms from sliding over each other easily.

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\overset{\textit{Carbon}}{\downarrow}}{\rightarrow}$ Zinc → Iron → Lead $\overset{\overset{\textit{Hydrogen}}{\downarrow}}{\rightarrow}$ Copper → **Silver**

METALS

7. The diagram shows the structure of brass.



- Why is brass harder than pure copper?

A The zinc atoms form strong covalent bonds with the copper atoms.

B The zinc atoms have more electrons than the copper atoms.

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[✓ **D**] The zinc atoms prevent the layers of copper atoms from sliding over each other easily.

Potassium → Sodium → Calcium → Magnesium → Aluminium → (*Carbon*) → Zinc → Iron → Lead → (*Hydrogen*) → Copper → Silver

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

1) copper + silver nitrate → copper(II) nitrate + silver

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

1) copper + silver nitrate → copper(II) nitrate + silver

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

$$Cu(s) + 2Ag^+(aq) + \cancel{2NO}_3^-(aq) \rightarrow Cu^{2+}(aq) + \cancel{2NO}_3^-(aq) + 2Ag(s)$$

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

Carbon ↓ (between Aluminium and Zinc)
Hydrogen ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

2) zinc + copper(II) sulfate → zinc sulfate + copper

Carbon ↓ Hydrogen ↓

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

2) zinc + copper(II) sulfate → zinc sulfate + copper

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

$$Zn(s) + Cu^{2+}(aq) + \cancel{SO}_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + \cancel{SO}_4^{2-}(aq) + Cu(s)$$

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Potassium → Sodium → Calcium → Magnesium → Aluminium $\overset{\textit{Carbon}}{\downarrow}$ → Zinc → Iron → Lead $\overset{\textit{Hydrogen}}{\downarrow}$ → Copper → Silver

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

3) iron + lead(II) nitrate → iron(II) nitrate + lead

Carbon ↓ (between Aluminium and Zinc) *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

3) iron + lead(II) nitrate → iron(II) nitrate + lead

$$Fe(s) + Pb(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Pb(s)$$

$$Fe(s) + Pb^{2+}(aq) + \cancel{2NO}_3^{-}(aq) \rightarrow Fe^{2+}(aq) + \cancel{2NO}_3^{-}(aq) + Pb(s)$$

$$Fe(s) + Pb^{2+}(aq) \rightarrow Fe^{2+}(aq) + Pb(s)$$

Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

4) zinc + hydrochloric acid → zinc chloride + hydrogen

Carbon *Hydrogen*
↓ ↓
Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

4) zinc + hydrochloric acid → zinc chloride + hydrogen

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

$$Zn(s) + 2H^+(aq) + \cancel{2Cl^-}(aq) \rightarrow Zn^{2+}(aq) + \cancel{2Cl^-}(aq) + H_2(g)$$

$$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$$

Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → Silver

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

5) magnesium + chromium(III) chloride → magnesium chloride + chromium

Carbon ↓ (between Aluminium and Zinc); *Hydrogen* ↓ (between Lead and Copper)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

METALS

- Write a balanced chemical equation and an ionic equation to describe each one of the following displacement reactions.

5) magnesium + chromium(III) chloride → magnesium chloride + chromium

$$3Mg(s) + 2CrCl_3(aq) \rightarrow 3MgCl_2(aq) + 2Cr(s)$$

$$3Mg(s) + 2Cr^{3+}(aq) + \cancel{6Cl^-}(aq) \rightarrow 3Mg^{2+}(aq) + \cancel{6Cl^-}(aq) + 2Cr(s)$$

$$3Mg(s) + 2Cr^{3+}(aq) \rightarrow 3Mg^{2+}(aq) + 2Cr(s)$$

Carbon ↓ (between Aluminium and Zinc)

Potassium → Sodium → Calcium → Magnesium → Aluminium → Zinc → Iron → Lead → Copper → **Silver**

Hydrogen ↓ (between Lead and Copper)

METALS

Presentation on
Metals and Reactivity Series
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