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# 2022 FE Revision – Metals and Electrolysis ANSWERS

## **Reactivity Series of Metals**

Metals can be ranked in a series based on their reactivity with water, steam and dilute acid. This series also indicates whether a metal can be extracted by reduction of its ore using carbon or hydrogen, and whether the compound (carbonate) of a metal is stable to heat. The table below ranks 12 (out of about 100) metals, from most reactive to least reactive.

The more reactive metal can reduce. the ion or compound of a less reactive metal to form that less reactive metal. Alternatively, we say, the more reactive metal can displace the less reactive metal from its salt solution or oxide.

Examples:

- $\overline{Zn}(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$  or  $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$ Observations: the **blue** solution fades in colour and a pink or brown **solid** is formed Explanation: the blue of the solution is due to Cu<sup>2+</sup> (aq) and as its concentration falls the blue colour fades off. The solid copper metal is formed which appears as pink or brown or red-brown in colour.
- 2  $2AI(s) + Fe_2O_3(s) \rightarrow AI_2O_3(s) + 2Fe(s)$

## **Extraction of Metals**

Many metals are extracted from their oxides (ores) by reduction with carbon or carbon monoxide. Those that cannot be extracted by such a method (e.g. aluminium) would usually be extracted by electrolysis. Iron is extracted in a blast furnace at above 1000°C. The inputs into the furnace are haematite (Fe<sub>2</sub>O<sub>3</sub>), coke (C), limestone (CaCO<sub>3</sub>) and hot air. The outputs are molten iron, molten slag (CaSiO<sub>3</sub>) and waste gases such as CO<sub>2</sub>.

The main reaction is  $Fe_2O_3$  (s) + 3CO (g)  $\rightarrow$  2Fe (l) + 3CO<sub>2</sub> (g)

The side reactions are

- 1. Generate CO (the reducing agent):  $C + O_2 \rightarrow CO_2$ , then  $CO_2 + C \rightarrow 2CO_2$
- 2. Remove sand (acidic impurity): CaCO<sub>3</sub>  $\rightarrow$  CaO + CO<sub>2</sub>, then CaO + SiO<sub>2</sub>  $\rightarrow$  CaSiO<sub>3</sub> (slag)

## Alloying

Metals are soft and malleable due to the regular layers of atoms that can slide over each other easily. So they are usually mixed with other metals or with carbon to form alloys which are harder and stronger - since the atoms are of different sizes they no longer form regular layers and thus cannot slide over each other.

Iron is alloyed with carbon to make steel. Mild steel (or low-carbon steel, with ~0.25% carbon) is the most common steel, used to make car bodies, bridges and railway lines. Hard steel (or high-carbon steel, with ~1% carbon) is harder and less malleable or more brittle than mild steel, and is used to make tools. Stainless steel has additives other than carbon; it is used to make cutlery and surgical instruments.

## Corrosion of Metal

The corrosion (or rusting) of iron requires air(oxygen) and water Rusting can be prevented by

- 1. Surface protection: cover the iron with a barrier (paint or oil or plastic or metal such as tin or chromium) to block air and water from reaching the iron underneath.
- 2. Sacrificial protection: coat the iron with zinc (or have the iron in contact with magnesium or zinc). The more reactive metal reacts with oxygen and water in place of the iron thus protecting it. The advantage is that even if the layer of magnesium or zinc is scratched to expose the iron underneath the iron is still protected (until all the magnesium or zinc is corroded off).
- 3. Stainless steel: the iron is mixed with carbon and chromium and nickel. The chromium and nickel prevents the iron from reacting with water and oxvgen.

Aluminium, even though it is more reactive than iron, does not corrode easily. This is because it is covered with a thin layer of non-porous aluminium oxide that protects the metal from reacting.

## **Recycling Metals**

Benefits		Disadvantages	
1	Helps conserve metals and ores	1	Recycling involves collecting transporting washing
2.	Saves cost and reduces pollution during extraction		sorting, etc, thus recycling may be more expensive
	of new metals		than extracting
3.	Reduces use of unsightly landfills and open mines	2.	Smelting during recycling can produce pollutants

#### **Practice Questions**

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(a)	Choosing your answer from the list only, name			
(i)	two metals which have to be stored under oil	barium	and	sodium
(ii)	two metals which will not react with steam	copper	and	silver
(iii)	two metals that react with water to give hydrogen and an alkali	barium	and	sodium
(iv)	two metals that could be produced by the reduction of their oxides with carbon	zinc	and	iron
(v)	two metals that do <b>not</b> react with dilute acids	copper	and	silver
(vi)	one metal whose carbonate does not decompose on heating	sodium		

(b) Aluminium is near the top of the reactivity series, but tends to react slowly. State the reason for this.

Aluminium has a tenacious non-porous layer of aluminium oxide on its surface. This layer protects the metal underneath from reacting with other reactants such as water and oxygen, thus aluminium reacts slowly.

- 2 chromium and iron (a)
  - (b) platinum
  - (c) aluminium, zinc, chromium and iron
  - Aluminum has low density so more aluminium cans can be packaged per ton, for delivery (d)(i)
  - (ii) Steel is very strong so the cans can withstand knocks and rough handling during delivery
  - $2AI + 6H^+ \rightarrow 2AI^{3+} + 3H_2$  or  $Fe + 2H^+ \rightarrow Fe^{2+} + H_2$ (iii)

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	Silver	Barium	Copper	Chromium
chromium(III) nitrate solution (green)	no reaction	grey deposit in colourless solution	no reaction	
barium nitrate solution (colourless)	no reaction		no reaction	no reaction
silver nitrate solution (colourless)		silvery deposit in colourless solution	silvery deposit in blue solution	silvery deposit in green solution
copper(II) nitrate solution (blue)	no reaction	reddish-brown deposit in colourless solution		reddish-brown deposit in green solution

- (b)  $2Cr + 3Cu^{2+} \rightarrow 2Cr^{3+} + 3Cu$
- barium (c)
- (d)  $Cr_2(CO_3)_3 \rightarrow Cr_2O_3 + 3CO_2$

## Electric Cell vs Electrolytic Cell

	Electric Cell	Electrolytic Setup	
Key Ideas	<ol> <li>The 2 electrodes are 2 different metals</li> <li>Electrons flow from the more reactive metal to the less reactive metal, and are then 'captured' by the cations in the electrolyte.</li> <li>The greater the difference in the reactivity of the metals the greater the voltage produced.</li> </ol>	<ol> <li>During electrolysis, using inert electrodes, the electrolyte decomposes.</li> <li>If the electrolyte is molten and contains 1 type of cations and 1 type of anions, then the cations would be attracted to the negative electrode and become reduced, while the anions would be attracted to the positive electrode and become oxidized</li> <li>If the electrolyte is an aqueous solution, it would contain H<sup>+</sup> and OH<sup>-</sup> as well as the dissolved substance's cation and anion. At the electrodes, competition will occur and the result is based on the electro-chemical series.</li> </ol>	
Energy Conversion	Spontaneous Reaction: chemical energy converted to electrical energy System does work on load/surroundings.	Non-spontaneous Reaction: electrical energy converted to chemical energy Surrounding (power supply) does work on system.	
Nature of Electrodes	Electrodes must be metals of different reactivity, so that a potential difference can be set up between them.	Electrodes may be inert (graphite or platinum), or active (eg. copper, silver, other metals).	
Electrode Reactions	Anode is where oxidation occurs. The anode loses electrons, thus becoming smaller. Cathode is where reduction occurs. Positive ions in the electrolyte around the cathode gain the electrons.	Anode is where oxidation occurs. Negative ions (anions) migrate here and are oxidized. Cathode is where reduction occurs. Positive ions (cations) migrate here and are reduced.	
Electrode Sign Conventions	Negative Electrode is the Anode.Reason: It is the more reactive metal thus electrons are generated continuously at this electrode (oxidation occurs here). Therefore, to the user, this is the source of negative charge and is thus by convention called the 'negative' terminal.Positive Electrode is the cathode	Negative Electrode is the <u>Cathode</u> .         Reason: This electrode is connected to the negative terminal of the power source and receives electrons from it, so it is negative. Reduction occurs here, as cations in the electrolyte gain electrons here.         Positive Electrode is the anode	

## **Factors affecting Electrolysis**

1 Electro-chemical Series

2 *Concentration Effect* – our focus is only between the halides and hydroxide ions. Example:

Electrolysis of <u>dilute</u> NaCl produces O<sub>2</sub> at the anode  $(4OH^- \rightarrow 4e^- + O_2 + 2H_2O)$ 

Electrolysis of <u>concentrated</u> NaCl produces Cl<sub>2</sub> at the anode  $(2Cl^- \rightarrow 2e^- + Cl_2)$ 

3 Nature of electrode

Inert electrodes are graphite (carbon) and platinum. They do not interfere in the electrolysis (although they may undergo side reactions, e.g. graphite may react with the **oxygen** produced at the anode) Active electrodes are metals (other than platinum). An active anode would take part in the electrolysis and be preferentially **oxidized** over OH<sup>-</sup> and other anions. Thus the active anode dissolves away.

#### **Applications of Electrolysis**

1 Aluminium extraction: aluminium ore (bauxite) is a mixture of aluminium **oxide** (Al<sub>2</sub>O<sub>3</sub>) and other oxides. These other oxides are first removed and then the melted or solubilized aluminium oxide is electrolyzed.



Electroplating: a metal object can be covered with another metal via electroplating. Example: in copper plating, the object is the **cathode** and the anode is a piece of **copper**, the electrolyte is copper(II) sulfate solution. If the object is non-conductive, then it is first sprayed with a metallic paint or covered with a layer of **graphite** powder before the electroplating can take place. During the plating process the copper anode dissolves, forming Cu<sup>2+</sup>, thus replenishing the Cu<sup>2+</sup> used up for plating. Since for every Cu<sup>2+</sup> used up at the cathode one Cu<sup>2+</sup> is produced at the anode, the blue colour of the solution remains the same, as [Cu<sup>2+</sup> (aq)] is **constant**. (Note: square brackets, [], represent 'concentration')

3 Purification of copper: when copper is first manufactured it is only about 98% pure and is often contaminated with silver. To get 100% (pure) copper, the piece of impure copper is made the **anode**, a small piece of pure copper is the cathode, and the electrolyte is copper(II) sulfate solution. The anode becomes **smaller** and the cathode becomes **larger** as electrolysis progresses.

## **Practice Questions**

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	electrolyte	positive electrode	negative electrode
А.	dilute sulfuric acid	$4OH^{-} \rightarrow O_2 + 2H_2O + 4e^{-}$	$2H^+ + 2e^- \rightarrow H_2$
В.	dilute sodium chloride	$40H^{-} \rightarrow O_2 + 2H_2O + 4e^{-}$	$2H^+ + 2e^- \rightarrow H_2$
C.	concentrated sodium chloride	$2CI^{-} \rightarrow CI_2 + 2e^{-}$	$2H^+ + 2e^- \rightarrow H_2$
D.	concentrated potassium iodide	$2I^{-} \rightarrow I_2 + 2e^{-}$	$2H^+ + 2e^- \rightarrow H_2$
E.	copper(II) sulfate solution	$40H^{-} \rightarrow 0_2 + 2H_2O + 4e^{-}$	Cu <sup>2+</sup> + 2e <sup>-</sup> → Cu

- (b)(i) E and A (since it is becoming more concentrated)
- (ii) C and D
- (iii) A and B

(a) **B and D** 

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(b) C:  $2H^+ + 2e^- \rightarrow H_2$  D:  $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ 

- (c) A
- (d) concentrated copper(II) bromide solution
- (e)(i) pH value increases as the solution changes from neutral to alkaline. At electrode C, as H<sup>+</sup> is reduced, OH<sup>−</sup> is left behind, thus as OH<sup>−</sup> concentration increases, alkalinity increases
- (ii) **50 cm<sup>3</sup>**
- (iii) At electrode A, Cu<sup>2+</sup> is preferentially reduced over H<sup>+</sup>, thus the layer of copper deposit on electrode A increases as electrolysis progresses, causing its mass to increase
- 5 (a) H<sup>+</sup>, SO<sub>4</sub><sup>2-</sup>, OH<sup>-</sup>
  - (b) At positive electrode, oxygen is given off:  $4OH^- \rightarrow O_2 + 2H_2O + 4e^-$ .

At negative electrode, hydrogen is given off:  $4H^+ + 4e^- \rightarrow 2H_2$ .

The equations show that for every 4 mol of electrons that flow from positive electrode to negative electrode, 1 mol of oxygen gas and 2 mol of hydrogen gas are formed, thus volume of hydrogen collected is twice that of oxygen.

- (c) This is because oxygen gas is more soluble in water than hydrogen gas. Thus as bubbles of oxygen gas are formed some become dissolved in the solution, so less oxygen than in theory is actually collected.
- (d) Self-ionization of water is very small thus concentration of free mobile ions (of H<sup>+</sup> and OH<sup>-</sup>) is very small, hence the poor conductivity. When ammonia dissolves in water some molecules form free mobile ions of  $NH_4^+$  and  $OH^-$ , thus overall, concentration of free mobile ions increases, hence conductivity increases.
- 6 (a) This is so that free mobile ions (of Pb<sup>2+</sup> and Br<sup>-</sup>) are present
  - (b)  $2Br^{-} \rightarrow Br_2 + 2e^{-}$
  - (c) Yes, the blue solution fades in colour and may become colourless. The blue colour is due to Cu<sup>2+</sup> (aq) thus as Cu<sup>2+</sup> is preferentially reduced and its concentration decreases, the blue colour fades off
  - (d) No colour change is observed. During electrolysis the copper anode is preferentially oxidized: Cu → Cu<sup>2+</sup> + 2e<sup>-</sup>. For every Cu<sup>2+</sup> removed from the solution due to reduction at the cathode: Cu<sup>2+</sup> + 2e<sup>-</sup> → Cu, one Cu<sup>2+</sup> is produced at the anode, thus the concentration of Cu<sup>2+</sup> remain constant and there is no colour change.
  - (e) Amount of Cu collected = 1.26/64 = 0.01969 mol

Since equations are:  $Cu^{2^+} + 2e^- \rightarrow Cu$  and  $Pb^{2^+} + 2e^- \rightarrow Pb$ , amount of electrons used at both electrodes are the same, thus

Amount of Pb collected = 0.01969 mol

∴ Mass of Pb = 0.01969 x 207 = 4.075 = 4.08 g (3 sf)