St Ninian's High School

Chemistry Department

National 5 Chemistry

Unit 3: Chemistry in Society

Metals

Summary Notes

Name ____

Learning Outcomes

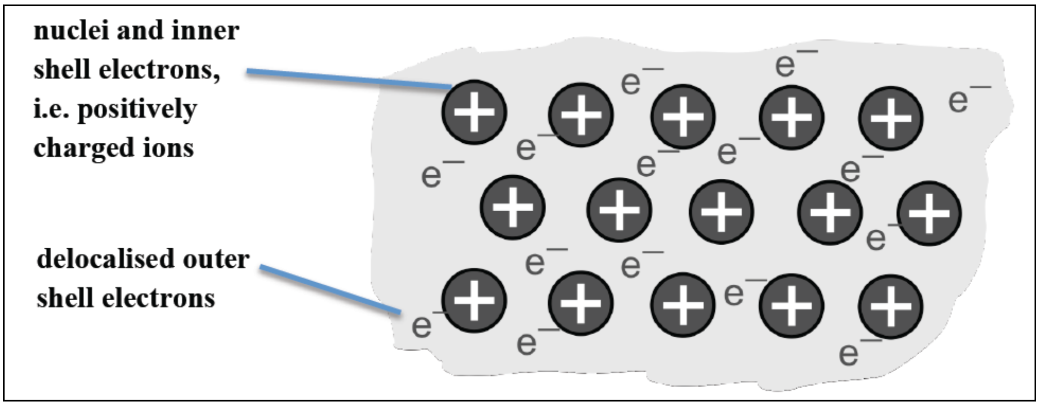
After completing this topic you should be able to :

1	state that metals have delocalised outer electrons which explains how all metals are able to conduct electricity when solid or liquid
2	metallic bonding involves the attraction of metal nuclei for delocalised outer electrons
3	relate specific properties of metals, e.g. density, thermal and electrical conductivity, malleability, strength, to their uses
4	describe the reactions of different metals with oxygen, water and dilute acid
5	state that these reactions give an indication of the reactivity of the metal
6	explain the terms reduction, oxidation and redox
7	write ionic redox equations to show the reactions of metals with oxygen, water and dilute acid
8	state that ores are naturally-occurring compounds of metals
9	state that the less reactive metals are found uncombined in the earth's crust and that the more reactive metals have to be extracted from ores
10	explain that the extraction of a metal from its ore is an example of reduction since metal ions gain electrons
11	state that some metals can be obtained from metal oxides by heat alone; some by heating with carbon or carbon monoxide; some need other methods
12	state that iron is produced from iron ore in the blast furnace
13	state the two key reactions which take place in the blast furnace are i) production of carbon monoxide and ii) reduction of iron oxide
14	state that aluminium is produced by the electrolysis of aluminium ore
15	write redox equations to show the extraction of metals
16	calculate the percentage composition of a particular metal in an ore.

Properties of Metals

Metallic Bonding

The electrons in the outer electron level of metal atoms are not tightly held in position. Chemists consider the electrons in metal atoms to be **delocalised** which helps to explain why all metals are able to conduct electricity. Overall the structure of metals (metallic bonding) is regarded as a regular array of positively charged ions surrounded by a sea of delocalised electrons. Each ion is attracted to the pool of electrons and vice versa. Since the electrons in the pool do not belong to any particular ion they are free to move from one ion to another - this explains electrical conduction.



Metallic Bonding - an array of positively charged ions surrounded by a sea of delocalised electrons.

Properties of Metals

Useful Metals

Metals are generally shiny, strong, good conductors of heat and electricity, malleable and ductile. **Malleable** means that they can be beaten out into shape. **Ductile** means that they can be drawn out to form thin wires without snapping. The last two properties of metals allows them to be made into almost any shape, for example, wires, pipes, cans, etc.

Reactions of Metals

Oxidation and Reduction

Some metals are more reactive than others. When metals take part in chemical reactions they lose electrons. For example magnesium will react with oxygen forming magnesium oxide.

$$2Mg(s) \quad + \quad O_2(g) \longrightarrow 2MgO(s)$$

If we consider what happens to the magnesium atoms:

Magnesium atoms have electron arrangement 2, 8, 2. Mg

Magnesium ions have stable electron arrangement 2, 8 Mg^{2+}

Overall $Mg \longrightarrow Mg^{2+} + 2e^-$

A chemical reaction where a species losses electrons is called **oxidation**. Equations which show the movement of electrons are called **ion-electron equations**. They are listed on page 10 of the data booklet.

If we consider what happens to the oxygen atoms:

Oxygen atoms have electron arrangement 2, 6. O

Oxygen ions have electron arrangement 2, 8 O^{2-}

Overall $O + 2e^- \longrightarrow O^{2-}$

A chemical reaction where a species gains electrons is called **reduction**.

An easy way to remember this is **OILRIG;**

O	**O**xidation
I	**I**s
L	**L**oss
R	**R**eduction
I	**I**s
G	**G**ain

When you have an oxidation and reduction reaction occurring together the overall reaction is called **redox.** When metals react with oxygen, water and dilute acids a redox reaction occurs. The metal is always undergoing oxidation. In this section you will investigate the reactivity of different metals compared with one another in reactions involving oxygen, water and acids.

Reactions of Metals

Metals and Oxygen

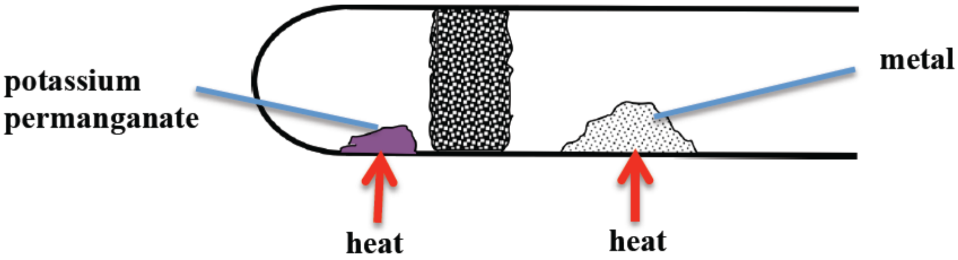
Most metals react with oxygen. When this happens, a metal oxide is formed:

In General:

Metal	**+**	**Oxygen**	**→**	**Metal Oxide**

For example magnesium and oxygen will react to produce magnesium oxide.

$$2Mg(s) \quad + \quad O_2(g) \longrightarrow 2Mg^{2+}O^{2-}(s)$$



Metals and Water

Some metals will react with water to form a metal hydroxide solution and release hydrogen gas.

In General:

Metal	**+**	**Water**	**→**	**Metal Hydroxide**	**+**	**Hydrogen**

Some metals are so reactive in water that they must be stored under oil to prevent any moisture getting in contact with the surface of the metal. Potassium will react with water to produced potassium hydroxide and hydrogen: $2K(s) \quad + \quad 2H_2O(l) \longrightarrow 2K^+(aq) \quad + \quad OH^-(aq) \quad + \quad H_2(g)$

Metals and Dilute Acids

Some metals will react with dilute acids to form a salt solution and release hydrogen gas.

In General:

Metal	**+**	**Acid**	**→**	**Salt**	**+**	**Hydrogen**

The balanced ionic equation for magnesium and hydrochloric acid is

$$Mg(s) \quad + \quad 2H^+(aq) \quad + \quad 2Cl^-(aq) \longrightarrow Mg^{2+}(aq) \quad + \quad 2Cl^-(aq) \quad + \quad H_2(g)$$

Reactions of Metals

The Reactivity Series

The reactivity series lists metals in order of reactivity with the most reactive listed first and the least reactivity at the bottom. The following table summarises the reactions of metals with oxygen, water and dilute acid.

<table>
<tr><th rowspan="2">Metal</th><th rowspan="2">Reactivity</th><th colspan="3">Reaction with</th></tr>
<tr><th>Oxygen</th><th>Water</th><th>Dilute acid</th></tr>
<tr><td>potassium</td><td>most reactive</td><td rowspan="11">metals which react with oxygen</td><td rowspan="5">metals which react with water</td><td rowspan="10">metals which react with dilute acid</td></tr>
<tr><td>sodium</td><td rowspan="12">↓</td></tr>
<tr><td>lithium</td></tr>
<tr><td>calcium</td></tr>
<tr><td>magnesium</td></tr>
<tr><td>aluminium</td><td rowspan="9">no reaction</td></tr>
<tr><td>zinc</td></tr>
<tr><td>iron</td></tr>
<tr><td>tin</td></tr>
<tr><td>lead</td></tr>
<tr><td>copper</td><td rowspan="4">no reaction</td></tr>
<tr><td>mercury</td><td rowspan="3">no reaction</td></tr>
<tr><td>silver</td></tr>
<tr><td>gold</td><td>least reactive</td></tr>
</table>

Reactions of Metals

Redox Equations

The reactions of metals with oxygen, water and dilute acids are all example of redox since oxidation and reduction take place. Redox equations are written by combining oxidation and reduction equations.

Example 1: Calcium and oxygen

Balanced equation $Ca(s) + ½O_2(g) \longrightarrow CaO(s)$

Oxidation art $Ca(s) \longrightarrow Ca^{2+}(s) + 2e^-$

Reduction part $½O_2(g) + 2e^- \longrightarrow O^{2-}(s)$

A redox equation can be written by combining the oxidation and reduction equations but it must never contain electrons. The oxidation and reduction equations may have to multiplied so that the electrons cancel. In this example the 2 electrons in each equation will cancel anyway.

Redox equation $Ca(s) + ½O_2(g) \longrightarrow Ca^{2+}O^{2-}(s)$

Example 2: Sodium and water

Balanced equation $2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$

Oxidation part $Na(s) \longrightarrow Na^+(aq) + e^-$

Reduction part $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

In this example the oxidation equation must be multiplied by two and allow the electrons to cancel.

Redox equation $2Na(s) + 2H_2O(l) \longrightarrow 2Na^+OH^-(aq) + H_2(g)$

Example 3: Magnesium and hydrochloric acid

Balanced equation $Mg(s) + 2HCl(aql) \longrightarrow MgCl_2aq) + H_2(g)$

Oxidation part $Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$

Reduction part $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

In this example the electrons are balanced and will cancel out.

Redox equation $Mg(s) + 2HCl(aq) \longrightarrow 2Mg^{2+}(Cl^-)_2(aq) + H_2(g)$

Example 4: Aluminium and hydrochloric acid

Oxidation equation $Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$

Reduction equation $2H^+(aq) + 2e^- \longrightarrow H_2(g)$

To cancel out the electrons, the oxidation must be multiplied by 2, and the reduction by 3.

Oxidation $2Al(s) \longrightarrow 2Al^{3+}(aq) + 6e^-$

Reduction $6H^+(aq) + 6e^- \longrightarrow 3H_2(g)$

Redox $2Al(s) + 6H^+(aq) \quad 2Al^{3+}(aq) + 3H_2(g)$

Extraction of Metals

Metal Ores

Metals such as gold and silver are unreactive and are found uncombined in the Earth's crust as the metallic element.

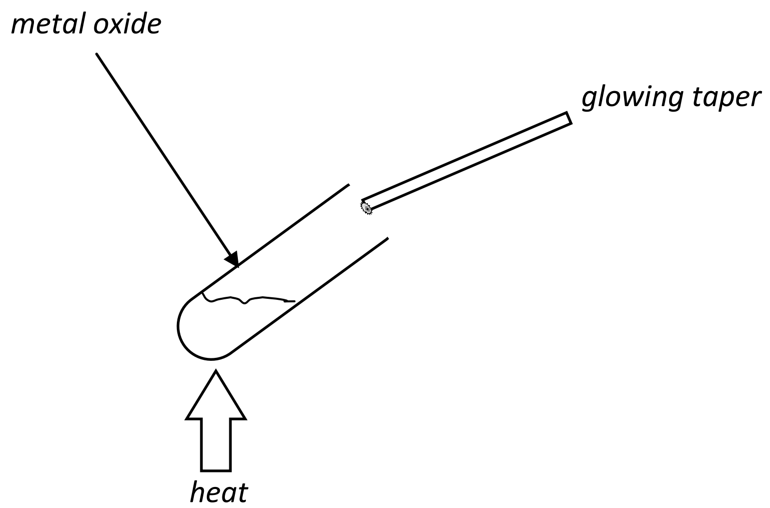
Most metals are found in the ground combined with other elements in the form of ionic compounds called ores. An **ore** is defined as a naturally occurring metal compound. Examples of ores are listed below.

Name of ore	Metal compound found in the ore
haematite	iron oxide
bauxite	aluminium oxide
iron pyrite (fool's gold)	iron sulfide

A metal can be extracted from its ore. Extraction involves separating the metal from the other elements in the ore. The method of extraction depends on the position of the metal in the reactivity series. In general the less reactive the metal the easier it is to extract the metal from its ore. **Extraction** of metals from ores is an example of a reduction reaction since the metal ions gain electrons to produce metal atoms.

Extraction by Heat Alone

In this activity your teacher will attempt to extract 3 different metals from their ores using heat alone. As your teacher carries out this demonstration record your observations in a results table.



Some metals can be extracted by heat alone such as mercury and silver, from their corresponding oxides. Most metals are too reactive to be extracted by heat alone.

In General:

$$\textbf{Metal Oxide} \xrightarrow{\textbf{Heat}} \textbf{Metal} + \textbf{Oxygen}$$

Extraction of Metals

Extraction by Heating with Carbon

In this activity you will watch your teacher demonstrate a method of heating iron oxide with carbon. You will then attempt to extract 2 for metals from their ores by heating with carbon in the same way.

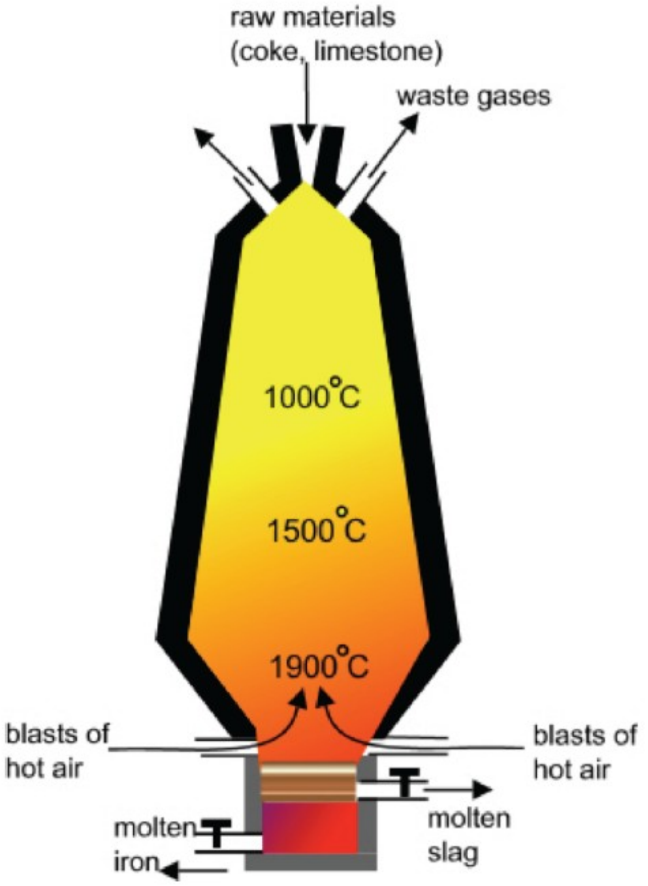
Some metals can be extracted by heating with carbon such as zinc, tin, lead, copper and iron from their corresponding oxides. Metals above zinc in the reactivity series are too reactive to be extracted by heating with carbon or carbon monoxide gas. In this extraction process the metal ions are reduced and the carbon is oxidised.

In General:

Metal Oxide	**+**	**Carbon**	$\xrightarrow{\textbf{Heat}}$	**Metal**	**+**	**Carbon Dioxide**

The Blast Furnace

Both iron and copper can be extracted from the ore by heating with carbon. However, iron is a more reactive metal than copper. It is therefore a little harder to extract it from its ore. In industry a special furnace has to be used so that the temperature is high enough. This is called a blast furnace.



There are 3 raw materials added

i) iron oxide (metal ore)

ii) coke (carbon)

iii) limestone (removes impurities)

The iron ore is called haematite which is iron(III) oxide, Fe_2O_3.

The molten slag layer can be used to produce cement. It is less dense than the molten iron.

In the blast furnace there are 3 reactions taking place are:

i) Carbon + oxygen $\longrightarrow$ carbon dioxide

ii) Carbon dioxide + carbon $\longrightarrow$ carbon monoxide

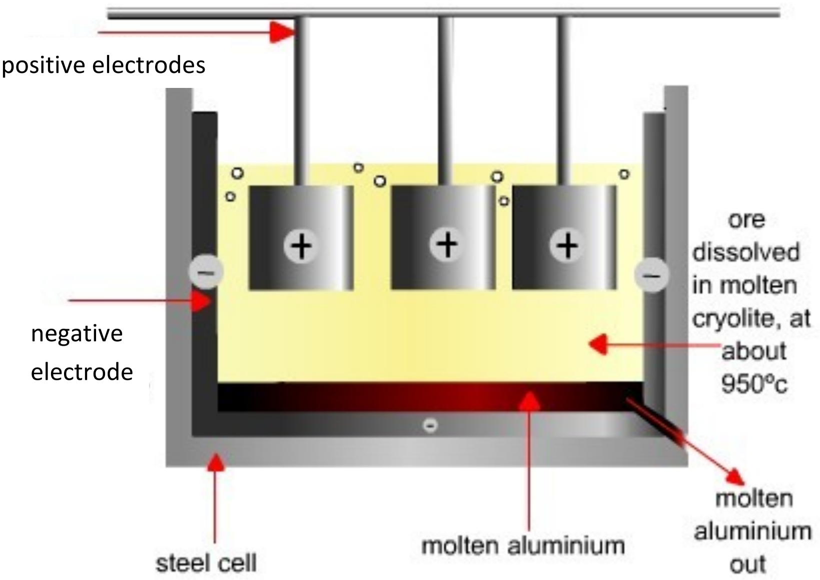
iii) Carbon monoxide + iron oxide $\longrightarrow$ iron + carbon dioxide

Extraction of Metals

Electrolysis

Very reactive metals are strongly bonded in their ores and cannot be extracted using carbon/carbon monoxide. More energy is needed and this is provided by electrolysis. Electrolysis is a process which uses electricity to break down a substance.

Metals above zinc in the reactivity series are extracted using electrolysis such as aluminium. Aluminium is extracted from its molten ore, bauxite (Al_2O_3).



When electricity is passed through the molten ore, the positive ions are attracted to the negative electrode where they gain electrons producing metal atoms. The aluminium flows out as a liquid.

Reaction at the negative electrode: $Al^{3+}(l) + 3e^- \longrightarrow Al(l)$ (reduction)

Summary of Reactions and Extractions of Metals

As elements: most reactive metals at the top				**As compounds:** least reactive compounds at the top
Metal	**with oxygen**	**with water**	**with acid**	**as oxides**
potassium	metals that are stored under oil	metals that react with cold water	metals that are too reactive to try in acid	metal oxides do not decompose on heating with carbon; electrical energy required to decompose compounds
sodium				
lithium				
calcium				
magnesium			metals that react to form hydrogen	
aluminium				
zinc				metal oxides decompose on heating with carbon
iron				
nickel				
tin				
lead				
hydrogen				
copper			metals that do not react with dilute acids	metal oxides decompose on heating to form metal
silver				
mercury				
gold				

Extraction of Metals

Percentage Composition

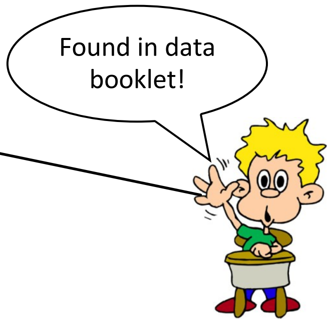
Metals are found in different ores, for example there are two common types of iron ore. When chemists try to extract a metal they want to know how much metal will be produced from a certain mass of ore. By calculating the percentage composition of an element in a compound it is possible to know what mass of pure metal can be expected.

The percentage composition of a metal in a compound is equal to:

$$\text{\% composition} = \frac{\text{mass of the metal in the chemical formula}}{\text{gram formula mass of the compound}} \times 100$$

or

$$\boxed{\%\ by\ mass = \frac{m}{GFM} \times 100}$$



Example 1: Find the % composition of titanium present in titanium ore, TiO_2.

Step 1: Find the gram formula mass using the data booklet

GRM = (1 x 48) + (2 x 16) = 48 + 32 = 80 g

Step 2: State the mass of the metal within the GFM

Mass of titanium = 48

Step 3: Carry out the calculation using the relationship in the data booklet.

% by mass = (48 ÷ 80) x 100

% by mass = 60 %

In other words 60 % of the ore is titanium.

Extraction of Metals

Example 2: Find the % composition by mass of aluminium in aluminium sulfate, $Al_2(SO_4)_3$.

Step 1: Find the gram formula mass using the data booklet

GRM = (2 x 27) + (3 x 32) + (12 x 16) = 54 + 96 + 192 = 342 g

Step 2: State the mass of the metal within the GFM

Mass of aluminium = 2 x 27 = 54

Step 3: Carry out the calculation using the relationship in the data booklet.

% by mass = (54 ÷ 342) x 100

% by mass = 15·8 %

In other words 15·8 % of the mass of a sample of aluminium sulfate is due to aluminium.

Summary Statements

- ☐ Metals have delocalised outer electrons which explains how all metals are able to conduct electricity when solid or liquid.
- ☐ Metallic bonding involves the attraction of metal nuclei for delocalised outer electrons.
- ☐ Relate specific properties of metals, e.g. density, thermal and electrical conductivity, malleability, strength, to their uses.
- ☐ Describe the reactions of different metals with oxygen, water and dilute acid.
- ☐ The reactions of metals with oxygen, water and acid gives an indication of the reactivity of the metal.
- ☐ Explain the terms reduction, oxidation and redox.
- ☐ Write ionic redox equations to show the reactions of metals with oxygen, water and dilute acid
- ☐ State that ores are naturally-occurring compounds of metals.
- ☐ State that the less reactive metals are found uncombined in the earth's crust and that the more reactive metals have to be extracted from ores.
- ☐ Explain that the extraction of a metal from its ore is an example of reduction since metal ions gain electrons.
- ☐ State that some metals can be obtained from metal oxides by heat alone (Hg & Ag) ; some by heating with carbon (Zn - Cu) or carbon monoxide; some need other methods (Al - electrolysis).
- ☐ State that iron is produced from iron ore in the blast furnace.
- ☐ State the two key reactions which take place in the blast furnace are i) production of carbon monoxide and ii) reduction of iron oxide.
- ☐ State that aluminium is produced by the electrolysis of aluminium ore.
- ☐ Write redox equations from oxidation and reduction equations to show the extraction of metals.
- ☐ Calculate the percentage composition of a particular metal in an ore using the following relationship:

$$\boxed{\%\ by\ mass = \frac{m}{GFM} \times 100}$$