Carbon and its compounds

Chemical properties of carbon

All allotropes of carbon have similar chemical properties but different physical properties.

a) Reaction with oxygen

When carbon is burnt in excess pure oxygen, carbon dioxide is formed.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

When carbon is burnt in limited supply of oxygen, carbon monoxide is formed.

$$2C_{(s)} + O_{2(g)} \rightarrow 2CO_{(g)}$$

b) Reaction with oxides of metals

$\left.\begin{array}{l} K \\ Na \\ Ca \\ Mg \\ Al \end{array}\right\}$ Oxides of these metals are not reduced by carbon.

(C)

$\left.\begin{array}{l} Zn \\ Fe \\ Pb \\ Cu \end{array}\right\}$ Oxides of these metals are reduced by carbon to a metal.

$\left.\begin{array}{l} Hg \\ Ag \end{array}\right\}$ Oxides of these metals are decomposed on heating before they react with carbon.

Carbon reduces oxides of metals that are lower than it in the activity series. Such oxides include copper(II) oxide, zinc oxide, lead(II) oxide and iron(III) oxide. All these are reduced to free metal while carbon is oxidised to carbon dioxide. The reaction requires heating the mixture of carbon and the metallic oxide.

$$2CuO_{(s)} + C_{(s)} \rightarrow 2Cu_{(s)} + CO_{2(g)}$$

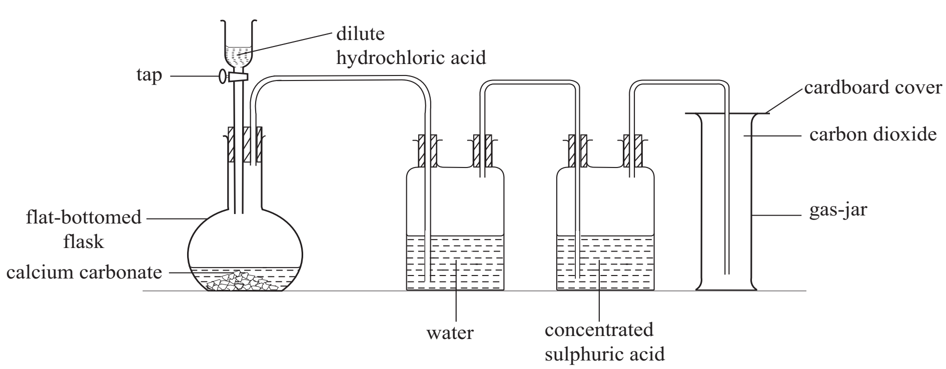
$$2Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + 3CO_{2(g)}$$

This reaction is used in extraction of the metals. Those metals higher in reactivity series than carbon have a higher affinity for oxygen and will not give it up to carbon.

$$MgO_{(s)} + C_{(s)} \rightarrow \textit{No reaction}$$

Carbon dioxide

Laboratory preparation of carbon dioxide



Using a dropping funnel, dilute hydrochloric acid is added to calcium carbonate in a flat-bottomed flask fitted with a delivery tube.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Ionically: $CO_3^{2-}(s) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

Effervescence occurs and a colourless gas, which is carbon dioxide, is passed through a bottle containing water or potassium hydrogencarbonate solution to absorb any fumes of hydrochloric acid. It is then passed through concentrated sulphuric acid to remove any moisture thus drying it. Carbon dioxide is an acidic gas and therefore it does not react with the drying agent. The gas is collected by downward delivery in a gas-jar since the gas is denser than air.

If the gas is not required dry it can be collected over water since carbon dioxide is only slightly soluble in water.

Note:
Dilute sulphuric acid is not used with calcium carbonate because the reaction produces calcium sulphate which is sparingly soluble and thus forms a coating on the calcium carbonate which stops further reaction.

$$CaCO_3(s) + H_2SO_4(aq) \rightarrow CaSO_4(s) + H_2O(l) + CO_2(g)$$

Lead(II) carbonate is also not used because when it reacts with dilute hydrochloric acid or sulphuric acid, the reaction soon slows down and then stops. This is due to the formation of lead(II) chloride or lead(II) sulphate, both of which are insoluble salts. The insoluble salt coats the carbonate preventing it from reacting with the acid.

$$PbCO_3(s) + 2HCl(aq) \rightarrow PbCl_2(s) + CO_2(g) + H_2O(l)$$
$$PbCO_3(s) + H_2SO_4(aq) \rightarrow PbSO_4(s) + CO_2(g) + H_2O(l)$$

Industrial preparation of carbon dioxide

In industries, carbon dioxide is obtained as a by-product of fermentation of sugars to alcohol.

$$C_6H_{12}O_6(aq) \rightarrow 2C_2H_5OH(aq) + 2CO_2(g)$$

Properties of carbon dioxide

Physical properties

1 It is slightly soluble in water forming carbonic acid.

$$H_2O(l) + CO_2(g) \rightleftharpoons H_2CO_3(aq)$$

2 It turns litmus pink since it is a weak acidic gas.
3 It is a colourless gas.
4 It is denser than air.

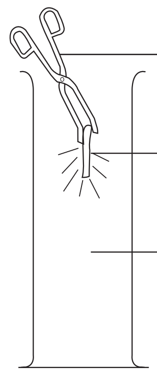
Chemical properties

(a) Effect of carbon dioxide on burning magnesium

When a piece of burning magnesium ribbon is lowered into a gas-jar containing carbon dioxide, it continues to burn for a short time with a spluttering flame. Black particles of carbon are formed on the sides of the gas-jar and white ash of magnesium oxide is also formed.

$$2Mg(s) + CO_2(g) \rightarrow 2MgO(s) + C(s)$$

The heat from the burning magnesium decomposes carbon dioxide into carbon and oxygen. The decomposition of carbon dioxide provides more oxygen which supports continued burning of magnesium to form magnesium oxide.



(b) Effect of carbon dioxide on calcium hydroxide solution (lime-water)

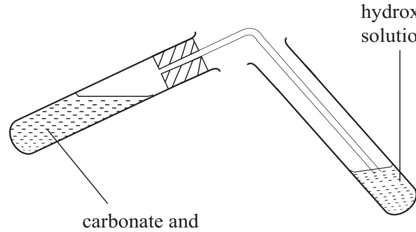
When carbon dioxide is bubbled through calcium hydroxide solution , a *white precipitate* is formed. The white precipitate is due to the formation of an insoluble substance, calcium carbonate, in water.

$$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(l)}$$

The above **test** is used to distinguish carbon dioxide from any other gas.

However, if more carbon dioxide is bubbled through the white precipitate, the precipitate dissolves to form a colourless solution due to the formation of calcium hydrogencarbonate, which is soluble in water.

$$CaCO_{3(s)} + H_2O_{(l)} + CO_{2(g)} \rightarrow Ca(HCO_3)_{2(aq)}$$



(c) Reaction with alkalis

Sodium hydroxide solution readily absorbs carbon dioxide to produce sodium carbonate.

$$2NaOH_{(aq)} + CO_{2(g)} \rightarrow Na_2CO_{3(aq)} + H_2O_{(l)}$$

With excess carbon dioxide, a white precipitate of sodium hydrogencarbonate is formed. The precipitate is sparingly soluble in cold water.

$$Na_2CO_{3(aq)} + H_2O_{(l)} + CO_{2(g)} \rightarrow 2NaHCO_{3(s)}$$

When solid sodium hydroxide is exposed to air, a colourless solution is formed and later a white crystalline solid is formed. Sodium hydroxide is deliquescent and therefore absorbs water from air to form a solution. The solution absorbs carbon dioxide from air forming a white crystalline solid of sodium carbonate.

decahydrate.

$$2NaOH_{(s)} + CO_{2(s)} + 9H_2O_{(l)} \rightarrow Na_2CO_3.10H_2O_{(s)}$$

Uses of carbon dioxide

1. Carbon dioxide is used in the manufacture of **carbonated drinks**. Solutions of carbon dioxide in water have a refreshing taste (pleasant taste).
2. Carbon dioxide is used as a **refrigerating agent**.
3. Carbon dioxide is used in **fire extinguishers**.

5 Carbon dioxide is used in **baking**.

Carbon monoxide

Carbon monoxide is a poisonous, colourless gas with practically no smell. It is formed by the partial combustion of carbon.

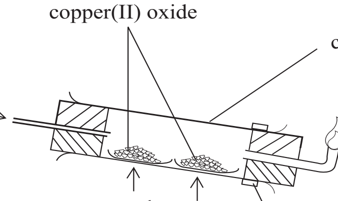
$$2C_{(s)} + O_{2(g)} \rightarrow 2CO_{(g)}$$

Properties of carbon monoxide

1 It is a colourless gas.
2 It has no effect on litmus paper, that is, it is a neutral gas.
3 It burns in air with a blue flame forming carbon dioxide.

$$2CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$$

4 It is insoluble in water.
5 It is a reducing agent. It reduces some metallic oxides of copper, lead, zinc and iron, that is, oxides of metals below carbon in activity series. The porcelain boat is heated strongly and the excess carbon monoxide is lighted at the jet.



$$\underset{\text{(black)}}{CuO_{(s)}} + CO_{(g)} \rightarrow \underset{\text{(brown)}}{Cu_{(s)}} + CO_{2(g)}$$

$$\underset{\text{(red-brown)}}{Fe_2O_{3(s)}} + 3CO_{(g)} \rightarrow \underset{\text{(grey)}}{2Fe_{(s)}} + 3CO_{2(g)}$$

$$\underset{\text{(white)}}{ZnO_{(s)}} + CO_{(g)} \rightarrow \underset{\text{(grey)}}{Zn_{(s)}} + CO_{2(g)}$$

$$PbO_{(s)} + CO_{(g)} \rightarrow Pb_{(s)} + CO_{2(g)}$$

Lead(II) oxide (yellow) is reduced to grey globules. Carbon monoxide does not, however, reduce the oxides of metals higher than carbon in the reactivity series. Such metals have a higher affinity for oxygen than carbon monoxide.

$$MgO_{(s)} + CO_{(g)} \rightarrow \textit{No reaction}$$

Uses of carbon monoxide

1 In the manufacture of synthetic petrol
2 In the reduction of ores and refining of nickel.

The effect of carbon monoxide on environment.

Carbon monoxide is poisonous because it forms a fairly stable compound with haemoglobin which reduces the oxygen-carrying capacity of the blood.

Carbonates

Effects of heat on carbonates

Carbonates of potassium and sodium are not decomposed by heat. Carbonates of calcium, magnesium, zinc, iron, lead and copper are decomposed by heat to an oxide and carbon dioxide.

When a white solid of lead(II) carbonate is heated strongly in a test-tube, a brown residue that turned to a yellow solid on cooling is formed.

$$PbCO_{3(s)} \rightarrow PbO_{(s)} + CO_{2(g)}$$

When a green solid of copper(II) carbonate is heated, a black residue of copper(II) oxide is formed.

$$CuCO_{3(s)} \rightarrow CuO_{(s)} + CO_{2(g)}$$

When a white solid of zinc carbonate is strongly heated, a yellow residue that turned to a white solid on cooling is formed.

$$ZnCO_{3(s)} \rightarrow ZnO_{(s)} + CO_{2(g)}$$

Magnesium and calcium carbonates (white) decompose to magnesium and calcium oxides (white) and carbon dioxide. Therefore, there is no observable change.

$$MgCO_{3(s)} \rightarrow MgO_{(s)} + CO_{2(g)}$$

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

Ammonium carbonate sublimes when heated. The cause of this sublimation is that ammonium carbonate dissociates on heating to ammonia, water and carbon dioxide, which recombine on cooling.

$$(NH_4)_2CO_{3(s)} \rightleftharpoons 2NH_{3(g)} + CO_{2(g)} + H_2O_{(l)}$$

Testing for carbonate (CO_3^{2-})

Add dilute nitric acid to the solution or solid to be tested. Effervescence (bubbles) of a colourless gas that forms a white precipitate with calcium hydroxide solution (lime-water) indicates the presence of a carbonate ion.

$$CO_3^{2-}{}_{(s\ or\ aq)} + 2H^+_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g)}$$

In qualitative analysis, identification tests for carbonate ions can be carried out as follows.

Test	Observations	Deductions
(a) Heat a spatula endfull of P strongly in a dry test-tube.	*Green solid. Colourelss gas which turned damp blue litmus paper red and lime-water milky. Black residue.*	*CO_2 given off $\therefore$ CO_3^{2-} present. Residue is CuO or Cu^{2+} present.*
(b) To one spatula endful of P add 4 cm^3 of dilute nitric acid.	*Green solid dissolved in acid. Bubbles or effervescence of a colourless gas which turned damp blue litmus paper red and lime-water milky. A blue solution.*	*CO_2 given off $\therefore$ CO_3^{2-} confirmed. Cu^{2+} present.*

Exercise

Complete the following table.

Test	Observations	Deductions
(a) Heat a spatula endfull of **X** strongly in a dry test-tube.	*White solid. Colourless gas which turned damp blue litmus paper red and lime-water milky. Brown residue that turned to yellow solid.*	
(b)(i) Heat a spatula endfull of **Z** strongly in a dry test-tube.	*White solid. Colourless gas which turned damp blue litmus paper red and lime-water milky. Yellow residue that turned to white solid.*	

(ii) To one spatula endful of **Z** add 4 cm^3 of dilute nitric acid.	*A white solid dissolved in acid. Bubbles or effervescence of a colourless gas which turned damp blue litmus paper red and lime-water milky. Colourless solution.*	

Sodium carbonate (soda ash)

Laboratory preparation of sodium carbonate

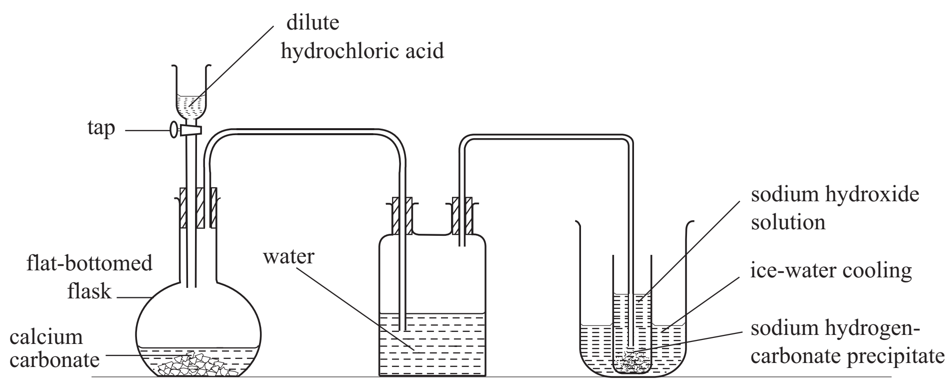
Using a dropping funnel, dilute hydrochloric acid is poured on to calcium carbonate in a flat-bottomed flask fitted with a delivery tube. Effervescence occurs and the gas (carbon dioxide) produced is passed through water to remove traces of acid.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Ionically: $CO_3^{2-}(s) + 2H^+(aq) \rightarrow H_2O(l) + CO_2(g)$

Then carbon dioxide is passed into a moderately concentrated solution of sodium hydroxide for sometime until finally a white precipitate of sodium hydrogencarbonate appears.

$$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l) \quad (1^{st}\text{ stage})$$
$$Na_2CO_3(aq) + H_2O(l) + CO_2(g) \rightarrow 2NaHCO_3(s) \ (2^{nd}\text{ stage})$$



The white precipitate is filtered off and washed two or three times with cold water. The solid is transferred into a dish and heated to a constant mass. Sodium carbonate is obtained as a fine white powder.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(l) + CO_2(g)$$

Uses of sodium carbonate

1 It is used for softening of water for domestic purpose.
2 It is used in manufacture of glass.
3 It is used to make dry soap powders.

Washing soda

Washing soda is sodium carbonate decahydrate ($Na_2CO_3.10H_2O$). When exposed to air, the crystals lose mass and become coated with a fine white powder. Each molecule of washing soda gives up, to the atmosphere, nine molecules of water of crystallization forming sodium carbonate monohydrate ($Na_2CO_3.H_2O$).

$$Na_2CO_3.10H_2O(s) \rightarrow Na_2CO_3.H_2O(s) + 9H_2O(l)$$

Such an action, that is, the giving up of water of crystallization to the atmosphere is termed as <u>efflorescence.</u>

Washing soda is used for softening water by precipitating the calcium ions from solution as calcium carbonate.

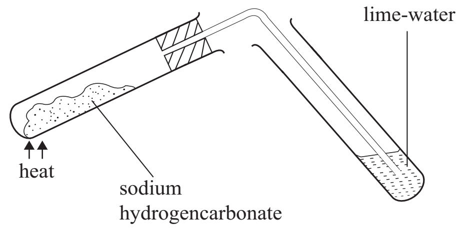
$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$$

Hydrogencarbonates

Action of heat on sodium hydrogencarbonate

When sodium hydrogencarbonate (white) is heated in a test-tube, a colourless gas that forms a white precipitate with calcium hydroxide solution (lime-water) is given off and a colourless liquid (water) is seen to condense on the cooler parts of the test-tube. A white residue is observed.

$$2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + H_2O_{(l)} + CO_{2(g)}$$



Distinguishing between a carbonate and hydrogencarbonate

A carbonate ion can be distinguished from hydrogencarbonate ion by adding magnesium sulphate or magnesium chloride solution separately to their solutions in water. A *white precipitate* indicates the presence of a carbonate ion. With hydrogencarbonate ion, there is *no observasable change*.

$$Mg^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \rightarrow MgCO_{3(s)}$$
$$Mg^{2+}_{(aq)} + 2HCO_3^{-}{}_{(aq)} \rightarrow Mg(HCO_3)_{2(aq)}$$

However, on heating the solution of magnesium hydrogencarbonate, a white precipitate is formed. This is because the magnesium hydrogencarbonate decomposes to the insoluble magnesium carbonate.

$$Mg(HCO_3)_{2(aq)} \rightarrow MgCO_{3(s)} + H_2O_{(l)} + CO_{2(g)}$$

Carbon cycle

Carbon cycle describes the processes that increase or decrease the carbon dioxide concentration in the environment (atmosphere).

Atmospheric carbon dioxide is used by green plants to synthesize glucose through a process called **photosynthesis**. Glucose in cells of both plants and animals is oxidised to obtain energy through a process called **respiration**. During this process, carbon dioxide is produced as a waste product which is released to the atmosphere.

When animals and plants die they undergo **decomposition**. During this process, the organic matter is broken down by micro-organisms and some of carbon compounds are converted to carbon dioxide. When these fuels are burnt, a process called **combustion**, the carbon compounds are oxidised to carbon dioxide.

The concentration of carbon dioxide in the atmosphere would be reduced by planting more trees, which use the gas in photosynthesis, and by burning less fossil fuels.

Iron

The main iron ores are haematite (Fe_2O_3), magnetite (Fe_3O_4), iron disulphide (pyrites, FeS_2) and spathic iron ore ($FeCO_3$).

Extraction of iron

The ore is first mixed with coke and limestone and the mixture is fed into the blast furnace from the top of the furnace. Hot air is blown into the furnace at the bottom which comes into contact with the red hot coke, producing carbon dioxide.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

Higher up the furnace, the source of oxygen is less and more coke combines with carbon dioxide produced to form carbon monoxide.

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

Carbon monoxide produced reduces the iron oxides to molten iron.

$$Fe_3O_4(s) + 4CO(g) \rightleftharpoons 3Fe(l) + 4CO_2(g)$$

$$Fe_2O_3(s) + 3CO(g) \rightleftharpoons 2Fe(l) + 3CO_2(g)$$

Molten iron runs to the bottom of the furnace and is tapped off into moulds where it is solidified. Limestone is decomposed by heat to calcium oxide and carbon dioxide.

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

The iron contains impurities such as silicon dioxide (sand), which combine with calcium oxide to form a molten slag that floats on top of the molten iron and it is tapped off.

$$CaO(s) + SiO_2(s) \rightarrow CaSiO_3(l)$$

The waste gases, mainly nitrogen and oxides of carbon, escape from the top of the furnace. Slag is used in making road foundations, phosphorus fertilizers and cement.