Paper Reference(s)

6252/01

Edexcel GCE

Chemistry (Nuffield)

Advanced Subsidiary

Unit Test 2

Thursday 17 January 2008 – Morning

Time: 1 hour 30 minutes

Passage for Section B

Do not return the insert with the question paper.

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Fluorine

Fluorine was first isolated by the French chemist Henri Moissan in 1886 after twenty-five years of continuous research. Fluorine is the most reactive element, reacting with nearly all organic and inorganic chemicals. It started to be produced on an industrial scale in the 1940s. It was needed to make uranium(VI) fluoride, UF_6, a chemical essential for the enrichment of uranium which was needed initially for the first atomic bombs and is now used in some nuclear power stations.

Fluorine is produced industrially by the electrolysis of an electrolyte made by dissolving anhydrous potassium fluoride in anhydrous hydrogen fluoride. Anhydrous hydrogen fluoride is a poor conductor of electricity so it cannot be electrolysed on its own. A 2:1 mixture of hydrogen fluoride and potassium fluoride has high electrical conductivity and a melting point of about 70 °C. An aqueous solution of hydrofluoric acid cannot be used because hydroxide ions would be preferentially discharged at the anode.

The electrolysis cell uses a carbon anode and a steel cathode. The anode is made of hard carbon, not graphite, since graphite would rapidly disintegrate due to infiltration of the small fluorine atoms between the carbon layers. A typical cell has up to 40 anodes and contains 1,250 kg of electrolyte. Such a cell can work at 12 V with a current of 6,000 A producing 4 kg of fluorine per hour. Because of the hydrogen fluoride impurities, about 2.08 mol of hydrogen fluoride are needed to make each mole of fluorine, F_2. Cells operate at about 90 °C. Cooling is necessary to maintain this temperature. This is achieved by a cooling jacket around the cell through which water is pumped at 80 °C.

Hydrogen gas is collected from the cathode. It is contaminated with hydrogen fluoride gas which is removed by passing the mixture through sodium or potassium hydroxide solution. Fluorine gas, also contaminated with hydrogen fluoride, is collected from the anode. The hydrogen fluoride is removed by reaction with sodium fluoride to form sodium hydrogen difluoride, $NaHF_2$. The electrolyte has to be continually replenished by the addition of hydrogen fluoride. Fluorine is used immediately or liquefied and stored. Fluorine is transported as a gas in steel containers.

Fluorine is still used to make uranium(VI) fluoride in a two-step process. First, uranium(IV) fluoride is made from uranium(IV) oxide and hydrofluoric acid, then uranium(IV) fluoride is reacted with fluorine directly. An alternative second step is to react uranium(IV) fluoride with chlorine trifluoride (made by reacting fluorine with chlorine).

Fluorine is also used to make sulphur hexafluoride, by direct fluorination of sulphur. The reaction is highly exothermic and is self-sustaining without the application of heat. Crude gas from the reactor is heated at 500 °C to disproportionate any disulphur decafluoride present.

$$S_2F_{10}(g) \rightarrow SF_6(g) + SF_4(g)$$

The mixture is then washed with sodium hydroxide to remove sulphur tetrafluoride.

Sulphur hexafluoride is an important gaseous electrical insulator in high voltage transformers or X-ray equipment.

Fully saturated fluorocarbons made from fluorine have many uses, including use as refrigerants, lubricants, and in artificial blood. Many anaesthetics, such as enflurane and isoflurane, are made indirectly from fluorine. Important inorganic compounds made from fluorine include nitrogen trifluoride and tungsten hexafluoride used in electronics manufacture and 'graphite fluoride' used in battery manufacture.

(538 words)

(Source: adapted from *Blue John and family: F, HF and Fluorides* by Harold Fielding and Brian Lee, Chemistry in Britain, April 1978)