Anaesthetics

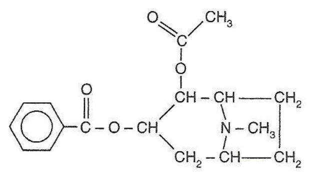
LOCAL ANAESTHETICS

Local anaesthetics (sometimes called local analgesics) block pain in a specific area when injected under the skin or applied topically, but do not affect the overall level of consciousness.

COCAINE

The first anaesthetic to be used in medicine was cocaine. It comes from a South American plant called *Erythoxylum coco*, and its leaves have been chewed for centuries by Indians from Peru and Ecuador to deaden pain and increase endurance. It was first used clinically in 1884. However, as well as being very addictive it can have unpleasant side effects, such as anxiety, nausea, and headaches. In a few cases it can cause breathing difficulties, convulsions, coma, and even death.

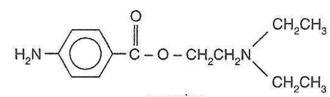
Cocaine also acts as a stimulant. It used to be legal and was used socially in Victorian times and recommended by Sigmund Freud to his patients. Even though it is now illegal cocaine abuse has risen rapidly during the past few decades. Cocaine and its synthetic derivatives work by suppressing nerve transmissions. They do this by blocking the action of acetylcholine, a neurotransmitter which allows repetitive impulses to travel along nerves. They also decrease the blood supply to the area by constricting the blood vessels.



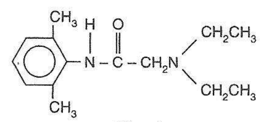
cocaine

PROCAINE AND LIDOCAINE

Due to its undesirable side effects, derivatives of cocaine were synthesized which retain the anaesthetic properties of cocaine but which do not also affect the brain or act as stimulants. Two of the most important compounds which fulfil these criteria are procaine and lidocaine. Both have similar structures to cocaine and both are used in dentistry and for minor surgery.



procaine



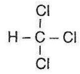
lidocaine

GENERAL ANAESTHETICS

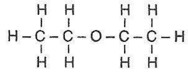
A general anaesthetic renders the patient unconscious so that they can feel no pain. Before 1840 major surgery, which often meant amputation, was carried out without any anaesthetic except perhaps alcohol and a lead bullet or leather belt to bite on. The first anaesthetic to be used was ethoxyethane (ether) in 1846. Ether is highly inflammable and very strong smelling so it was rapidly superseded by trichloromethane (chloroform). For the next hundred years chloroform or nitrogen(I) oxide (N_2O) were widely used in surgery, and a nitrogen(I) oxide and air mixture is still used by some dentists today. During that period various other compounds, such as cyclopropane, were tried, however, all have disadvantages. Like ether, cyclopropane is prone to ignite and explode, nitrogen(I) oxide is not very efficient and trichloromethane can lead to liver damage.

In the 1950s a research programme was initiated to find a safer alternative. The objectives were that the anaesthetic should be stable, safe, and non-inflammable. In addition of course all anaesthetics must have a low boiling point and must be capable of being inhaled easily. Several fluorinated hydrocarbons fulfilled these criteria and the most successful one was 2-bromo-2-chloro-1,1,1-trifluoroethane (halothane), which is now widely used all over the world.

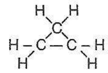
However, halothane is a CFC (chlorofluorocarbon) and can initiate damage to the ozone layer so research is still continuing to find more environmentally friendly alternatives. There are still risks associated with using general anaesthetics and advances in medicine, such as 'key hole surgery' and the use of lasers, are diminishing the need for them.



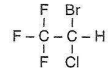
trichloromethane (chloroform)



ethoxyethane (ether)



cyclopropane

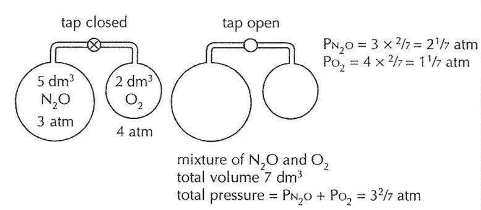


2-bromo-2-chloro-1,1,1-trifluoroethane (halothane)

PARTIAL PRESSURES OF COMPONENT GASES

It is important that anaesthetists administer the correct gas and air mixture. Too much anaesthetic and the patient may suffer oxygen starvation or even death, too little and the patient may recover consciousness prematurely and feel pain. The partial pressure of a component gas in a mixture is equal to the pressure it would exert if it occupied the total volume on its own. Dalton's law states that the total pressure exerted by a mixture of ideal gases is equal to the sum of the partial pressures of all the component gases.

e.g. consider a 5 dm^3 flask containing N_2O at 3 atm pressure connected by a tap to a 2 dm^3 flask containing O_2 at 4 atm pressure. When the tap is opened, the total pressure of the mixture will be $3\frac{2}{7}$ atm.



Partial pressures can also be expressed in terms of mole fractions. In a mixture of two gases A and B where the number of moles of A is n_A and the number of moles of B is n_B then:

$$P_A = \frac{n_A}{n_A + n_B} \times P_{total}$$

$$P_B = \frac{n_B}{n_A + n_B} \times P_{total} \quad \text{and} \quad P_{total} = P_A + P_B$$