Oxford Cambridge and RSA

For issue on or after: 13 March 2024

A Level Chemistry B (Salters)

H433/02 Scientific literacy in chemistry

Advance Notice Article

To prepare candidates for the examination taken on
Tuesday 18 June 2024 – Morning

INSTRUCTIONS
- **Before** the exam, read this article carefully and study the content of the learning outcomes for A Level Chemistry B (Salters).
- You can ask your teacher for advice and discuss this article with others in your class.
- You can investigate the topic of this article yourself using any resources available to you.
- Do **not** take this copy of the article or any notes into the exam.

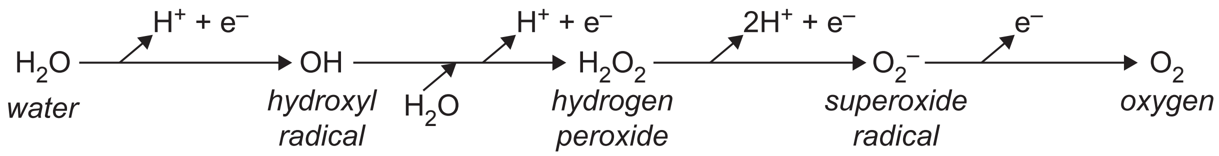
INFORMATION
- In the exam you will answer questions on this article. The questions are worth 20–25 marks.
- A clean copy of this article will be given to you with the question paper.
- This document has **4** pages.

ADVICE
- In the exam you won't have time to read this article in full but you should refer to it in your answers.

Reactive Oxygen Species

Adapted from *Oxygen – the molecule that made the world* by Nick Lane, Oxford 2002, pages 115–119.

Radiation is present all around us. It can interact directly with all kinds of molecules, but in our bodies it is most likely to interact with water, knocking out electrons and forming oxygen. The three intermediates formed by irradiating water, the hydroxyl radicals, hydrogen peroxide and superoxide radicals are called reactive oxygen species and react in very different ways. However, because all three are linked and can be formed from each other, they might be considered equally dangerous. Indeed the three work together as part of an insidious catalyst system. We will consider each in turn, in the order they are produced by the radiation on route from water to oxygen.



Hydroxyl radicals (OH) are the first to be formed. These are extremely reactive fragments or random muggers. They can react with all biological molecules at speeds approaching their rate of diffusion. This means that they react with the first molecules in their path, and it is virtually impossible to stop them from doing so. When a hydroxyl radical reacts with a protein, lipid (fat) or DNA, it snatches a proton and an electron to itself and sinks back into the sublime chemical stability of water. But of course, the act of snatching an electron leaves the reactant short of an electron. So, another radical is formed, this time part of the protein, lipid or DNA. This is a fundamental feature of all free radical reactions – one radical creates another, and if this radical is also reactive, then a chain reaction will ensue. Thus, the essential feature of a radical is an unpaired electron, while the essential feature of free-radical chemistry is the chain reaction.

We are all familiar with radical chain reactions when they happen in fatty foods such as butter: they are responsible for rancidity. The fats in the butter oxidise and taste disgusting. The same type of reaction also takes place in cell membranes, which are made mostly of lipids. The process is then called lipid peroxidation. Radical damage is less obvious when it affects proteins or DNA, but radical damage to DNA is one of the main causes of genetic mutation and accounts for the high rates of cancer suffered by radiation victims.

A dramatic non-biological example of the power of radical chain reactions is the hole in the ozone layer. The devastation that can be caused by chlorofluorocarbons (CFCs) such as freon is a result of the formation of radicals in the upper atmosphere. CFCs are quite robust molecules that can survive buffeting by the weather in the lower atmosphere. However, they are shredded by ultraviolet rays in the upper atmosphere and disintegrate to release chlorine atoms. Being one electron short of a full pack, chlorine atoms are dangerously reactive radicals. They can steal electrons from almost anything, setting in motion a chain reaction. According to the US Environmental Protection Agency, a single gram of freon will often destroy as much as 70 kg of ozone. The radical chain reaction ends when two radicals react with each other, and their unpaired electrons conjoin in blissful chemical union.

If radiation strips a second electron from water, the next fleeting intermediate is hydrogen peroxide (H_2O_2) – whose bleaching properties give its name to the peroxide blonde. Bleaching is caused by the oxidation of organic pigments as hydrogen peroxide strips electrons from them. The oxidising properties of hydrogen peroxide can kill bacteria. Most industrial uses of hydrogen peroxide also draw on its power as an oxidising agent.

Despite its widespread use as an oxidising agent, hydrogen peroxide is unusual in that it lies chemically half-way between oxygen and water. This gives the molecule something of a split personality. It can go either way in its reactions (losing or gaining electrons) depending on the chemical company it keeps. It can even go both ways at once, when reacting with another hydrogen peroxide molecule. In this case, one of the molecules gains two electrons to become water, while the other loses two electrons to become oxygen.

A far more dangerous and significant reaction, however, takes place in the presence of iron, which can pass on electrons one at a time to hydrogen peroxide to generate hydroxyl radicals. If dissolved iron is present, hydrogen peroxide is a real hazard. Organisms go to great lengths to avoid contamination with dissolved iron. The reaction between hydrogen peroxide and iron is called the Fenton reaction, after the Cambridge chemist Henry Fenton who first discovered it in 1894.

$H_2O_2 + Fe^{2+} \rightarrow OH^- + OH + Fe^{3+}$

He later showed that the reaction could damage almost any organic molecule. Thus, the main reason that hydrogen peroxide is toxic is that it produces hydroxyl radicals in the presence of dissolved iron. Ironically, the greatest danger lies in its **slow** reactivity in the absence of iron. It has time to diffuse throughout the cell. Hydrogen peroxide may diffuse into the cell nucleus, for example, and there mix with the DNA before it encounters iron, which transforms it into a brutish hydroxyl radical. The insidious infiltration of hydrogen peroxide means that it is more dangerous than the hydroxyl radicals produced outside the nucleus. Some proteins, such as haemoglobin, also contain iron. If they happen to run into hydrogen peroxide they can be mutilated on the spot.

What, then, of the third of our intermediates, the superoxide radical, O_2^-? Like hydrogen peroxide, the superoxide radical is not terribly reactive. However, it too has an affinity for iron, dissolving it from proteins and storage depots. To understand why this is harmful, we need to think again about the Fenton reaction. The Fenton reaction is dangerous because it produces hydroxyl radicals but it grinds to a halt when all the accessible iron is used up. Any chemical that regenerates dissolved iron is capable of re-starting the reaction. Because the superoxide ion is one electron away from molecular oxygen, it is more likely to lose that electron to form oxygen than it is to gain three electrons to form water. Only a few molecules are able to **accept** a single electron, however. One of the best places for the superoxide to jettison its spare electron is iron. This converts iron back into the form where it can participate in the Fenton reaction:

$O_2^- + Fe^{3+} \rightarrow O_2 + Fe^{2+}$

In summary, then, the three intermediates between water and oxygen operate as an insidious catalytic system that damages biological molecules in the presence of iron.

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