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The Open University

Water pollution control

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1.1

Contents

Section 1: Water basics

1.1 Introduction

We can all relate to water. We know we need it to survive – indeed, the early great civilisations of Egypt and Mesopotamia were centred on river valleys where there was a plentiful supply of fresh, clean water.

When we take water into our bodies, it is used in several ways. For example:

- for cooling – it helps keep our bodies at around 37 °C
- as a waste disposal medium
- as a conductor for nerve impulses
- as a component in the digestion of food
- as a solvent in which vital chemical reactions take place.

You can see from the above that even if you didn't move an inch, your body would still need water to keep you alive.

Water is a fascinating subject, encompassing chemistry, biology and physics. Apart from keeping us alive, water is used extensively in industrial processes, for recreation and for transport. It is something we can't do without.

The water we use for domestic purposes ought to be free from contaminants, yet water pollution is a major problem in many countries. According to the World Health Organization (WHO, 2002), about 1.7 million people die each year due to unsafe water, sanitation and hygiene. This text endeavours to outline the need for monitoring of the aquatic environment, leading to effective means of protection being put in place. It details various sources of water pollution, and describes the effects that different pollutants have on water. Sewage treatment is then considered in detail, and various treatment methods are presented; this is followed by sludge treatment and disposal. Next, the important subject of water quality testing is addressed, with details of the different tests. Industrial wastewater treatment is introduced, and the final section looks at river quality modelling.

The self-assessment questions (SAQs) located throughout the text will help you to review and remember what you have read.

1.2 The hydrological cycle

The *hydrological cycle* – the continuous cycling of water between land, open water surfaces and the sea, either directly or indirectly – is a complex process that has been known about for a long time (Figure 1). Probably the oldest reference to the hydrological cycle is found in the *Chandogya*, one of the principal Upanishads, which says 'rivers … lead from sea to sea'. It reveals that as early as 1000 BCE, attempts were being made to interpret and explain recurrent phenomena on the basis of direct experience.

Figure 1 Early understanding of the water cycle?

View description

The identifiable mechanisms of the cycle are complicated not only by the characteristics of air–water–land interfaces across which the cycle operates, but also by climatic factors that vary in both time and space. The various operations and mechanisms within the cycle are illustrated in Figure 2.

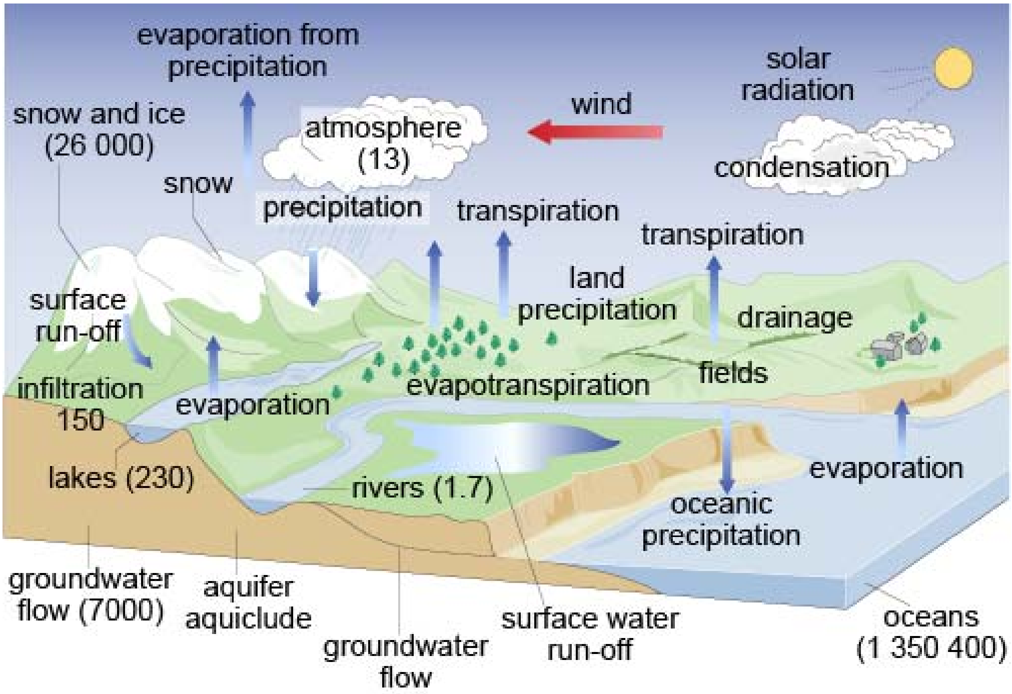

Figure 2 The hydrological cycle (volumes are in $\mathrm{Tm}^3 = 10^{12}\ \mathrm{m}^3$)

View description

1.3 The natural aquatic environment

Now that you have been introduced to the basic hydrological cycle, this subsection will consider the importance of water and how crucial dissolved oxygen is to aquatic life. The physical, chemical and biological characteristics of natural waters will then be explored. Importantly, how the parameters vary will be considered. Seasonal effects are important, as you might imagine.

1.3.1 Water, the medium of life

Water is an excellent *solvent*, so it is never pure – even in its ‘natural’ state, it contains a variety of soluble inorganic and organic compounds. Water can also carry large amounts of insoluble material in *suspension*. The amounts and types of impurities vary with location and time of year, and determine some of the characteristics of a particular watercourse.

One of the most important determining factors is the presence of organic material in *solution* or in suspension. Organic material can be used as food by the organisms living in natural water, provided the material is *biodegradable*. The basis of a *trophic system* in a river is the inorganic and organic materials it contains, their biodegradation by decomposer organisms, and the products of the photosynthetic activities of the primary producers (green plants and algae).

In water, as on land, the primary producers are eaten by herbivores (primary consumers) and these in turn are devoured by the secondary consumers (carnivores). The interdependence of these organisms gives a complex food web within which there are many food chains, the successive links in the chains being composed of different species in a predator–prey relationship. For a river, a typical food chain could be:

alga → protozoan → mayfly nymph → small fish (e.g. minnow) → large fish (e.g. pike)

Scavengers eat bottom debris, including dead organisms. Any uneaten dead organisms are broken down by decomposers (mostly bacteria and fungi), releasing nutrients that can be taken up by plants.

Through this cyclic movement of nutrients, the water environment achieves an ecological equilibrium. In theory, in any given stretch of water a balance occurs between the production of living material and the death and decomposition of organisms over a period of time. The river neither becomes choked with living organisms nor is devoid of them – although, depending on location and geological conditions, the numbers and varieties of organisms in the *biota* vary enormously. The maintenance of equilibrium is dependent on the complexity of biota and the interlinking of food chains and webs.

If the water contains low levels of plant nutrients then the conditions are said to be *oligotrophic*. This may occur when the physical and chemical characteristics of the land through which the water passes are such that nutrients are sparse or are not dissolved out of the soil and rocks.

The opposite condition, with high levels of nutrients in the water, is described as eutrophic; the gradual increase with time of plant nutrients in a body of water is called *eutrophication*.

Flowing and standing water

A typical river has several sources in high ground that are characterised by steep gradients, swift current velocities, and erosion of the surrounding rocks and soil. As the gradient lessens, the current velocity decreases and the river deepens and widens. The river then tends to deposit stones, gravel and sand. This variation in the flow downhill has a direct influence on the types of organisms and *substratum* to be found at different points along the river. The whole length of the river can be subdivided into different zones, each characterised by its own typical fauna and flora.

In contrast to rivers, standing bodies of deep water such as lakes and reservoirs may be affected by thermal stratification. Figure 3 illustrates this effect for a typical lake. In the summer, there is very little mixing between the cooler, denser water at the bottom of the lake (hypolimnion) and the warmer, less dense water at the lake surface (epilimnion). Thus, stream and river water running into the lake will tend to stay in the upper layer. This water carries nutrients, so organisms flourish in the epilimnion and there is a high rate of *primary production*. In the hypolimnion, the dead remains of primary production settle out, forming a layer of bottom sediment.

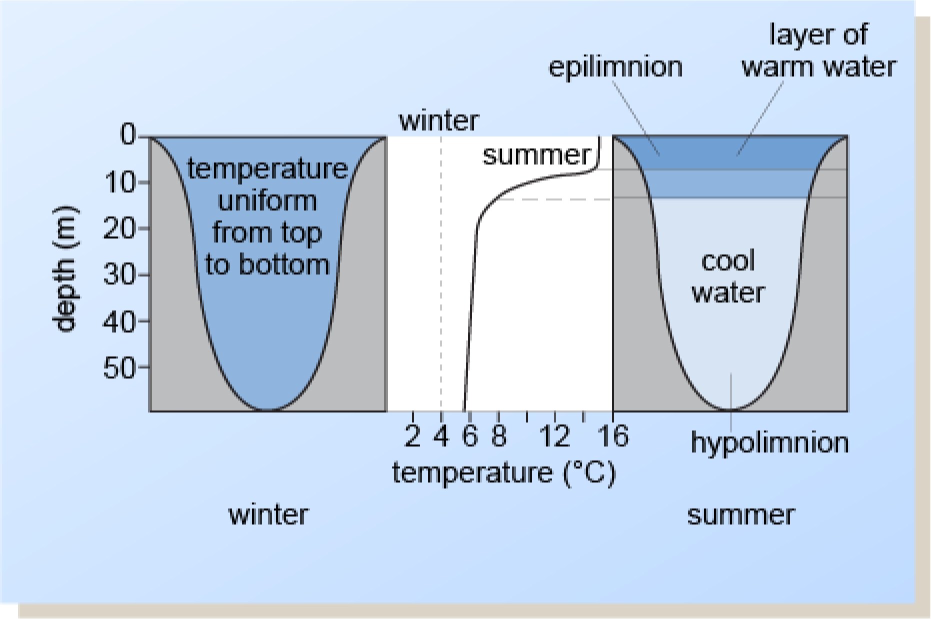

Figure 3 Thermal stratification of a lake

View description

The lack of mixing between the layers (stratification), together with the absence of light penetration to the bottom of the lake, determines the ecological characteristics of a deep lake or reservoir. In a deep lake, the absence of light prevents the growth of plant life in the bottom layers, although decomposer and scavenger microorganisms can live on and in the sediment.

Figure 3 shows that, in contrast to summer conditions, thermal stratification is absent in winter. This is because the density of fresh water is greatest at about 4 °C. Thus when the temperature of the surface layer drops to this temperature, the layer will fall to the bottom of the lake, displacing any colder (but less dense) water – which will now rise to the surface. The lake ‘turns over’ and mixing occurs at all levels, leading to uniform temperature and uniform conditions throughout. This mixing process can bring partially decomposed bottom sediments to the surface, where further biodegradation can occur. This can also cause a significant deterioration in water quality.

SAQ 1

Why is eutrophication more likely in a shallow lake than in a river? Describe the conditions it can bring about.

View answer

1.3.2 Dissolved oxygen – measurement

Organic and inorganic nutrients are the basic food supply essential for maintaining the plants and animals in natural watercourses. Equally essential to aquatic life is a supply of oxygen, needed for respiration. Oxygen dissolved in the water is also needed in the biodegradation of organic matter by aerobic bacteria. The more organic matter there is in a river or effluent, the greater the amount of dissolved oxygen that will be ‘demanded’ by the biodegrading bacteria. A measure of this oxygen demand, the *biochemical oxygen demand* (BOD), can therefore be used as a measure of the polluting capacity of an effluent. BOD can be measured experimentally, and the procedure for its determination will be given in Section 6.

Oxygen dissolved in natural waters arises from two main sources – the atmosphere and photosynthesis. Atmospheric air, which contains 21% oxygen by volume, can dissolve in water up to a limit. Green plants in the presence of sunlight generate oxygen by photosynthesis. These two sources replenish the oxygen used up in aerobic processes by aquatic organisms. The solubility of oxygen in water depends on the temperature, the pressure and the amount of dissolved solids present.

C_s is the maximum amount of oxygen in grams that can be held in one cubic metre of solution – called the *saturation concentration*. It is therefore expressed in $\mathrm{g\ m^{-3}}$ (grams per cubic metre). You may also find C_s expressed in units of $\mathrm{mg\ l^{-1}}$ (milligrams per litre) or ppm (parts per million); these are the same as $\mathrm{g\ m^{-3}}$, so a solubility of $20\ \mathrm{g\ m^{-3}}$ is the same as a solubility of 20 ppm, which is the same as $20\ \mathrm{mg\ l^{-1}}$. (You may like to verify this for yourself.)

Table 1 shows the solubility of oxygen from air at atmospheric pressure in pure water at various temperatures. This is calculated using the following expression (Baca and Arnett, 1976):

$$C_s = 14.65 - 0.410\,22\,T + 0.007\,91\,T^2 - 0.000\,077\,74\,T^3$$

where C_s is the solubility (in g m^{-3}) of oxygen in water at 1 atmosphere pressure, and T is the temperature in °C (we will return to this expression in Section 8). As can be seen, the solubility decreases with an increase in water temperature.

Table 1 Saturation concentration of oxygen in water at 1 atmosphere at different temperatures

T (°C)	C_s (g m^{-3})
0	14.65
5	12.79
10	11.26
15	10.01
20	8.99
25	8.12
30	7.36

The minimum concentration of dissolved oxygen required to support a balanced population of desirable aquatic flora and fauna is 5 g m^{-3}. The figures in Table 1 are for water at normal atmospheric pressure of 1 atmosphere. Decreasing the atmospheric pressure on the water decreases the saturation concentration. Therefore streams at high altitude are not able to dissolve as much oxygen as those at the same temperature nearer sea level.

The presence of dissolved solids in the water also affects the solubility of oxygen. Electrical conductivity can be used as a measure of the total dissolved solids in a water sample, i.e. its *salinity* (Table 2). Waters with dissolved salts, also called saline waters, are sometimes treated by desalination to provide drinking water.

Table 2 Relationship between electrical conductivity measured at 25 ºC and salinity

Electrical conductivity (μS cm^{-1})	Salinity (mg l^{-1})
100	51
500	241
1000	493
2000	1016
3000	1558
4000	2113
5000	2678
6000	3253
7000	3836
8000	4426
9000	5022
10 000	5625
12 000	6846
14 000	8087
16 000	9345
18 000	10 621
20 000	11 911

(Adapted from Hoare, 2010)

A correction factor can be used to calculate the saturation concentration of dissolved oxygen in saline waters. Table 3 gives a set of such factors. For a given conductivity and temperature, the factor should be multiplied by the appropriate saturation concentration from Table 1.

Table 3 Correction factors for dissolved oxygen in water, based on conductivity measured at 25 ºC

	Conductivity at 25 ºC (μS cm^{-1})							
	1000	5000	10 000	12 000	14 000	16 000	18 000	20 000
Temperature (ºC)								
5	0.996	0.981	0.963	0.955	0.947	0.940	0.932	0.924
10	0.996	0.982	0.964	0.957	0.949	0.942	0.934	0.927
15	0.997	0.983	0.965	0.958	0.951	0.944	0.937	0.929
20	0.997	0.983	0.966	0.960	0.953	0.946	0.939	0.932
25	0.997	0.984	0.968	0.961	0.954	0.948	0.941	0.934

(Adapted from USGS, 2006)

Exercise 1

A water supply company contemplates building a desalination plant near an estuary, with a view to producing drinking water from the plentiful *brackish* water in the estuary. When measured at 25 ºC, the conductivity of the water in the estuary is about 16 000 μS cm^{-1} during the ebb tide (which is when water will be drawn for desalination, as the salinity level will be at its lowest). Having dissolved oxygen in the water would be beneficial for the pre-treatment stages of the desalination system. Calculate the dissolved oxygen level in the water when it is at 10 ºC.

View answer

1.3.3 Dissolved oxygen – rate

The *rate* at which oxygen dissolves in water is dependent on several factors. One of these, the oxygen deficit (D), is the difference between the saturation concentration of oxygen (C_s) and the concentration of oxygen actually present (C), i.e.

$$D = C_s - C \quad (1)$$

The oxygen deficit is the driving force for the replenishment of oxygen used up in polluted water. The greater the oxygen deficit, the greater the transfer rate of oxygen into the water. Other factors important in the dissolution of oxygen in water include the turbulence of the water, its ratio of surface area to volume, the presence of animals and plants in the water, and any other dissolved substances. These factors are described further below.

Exercise 2

A river at a certain location has a dissolved oxygen content of 8.1 g m^{-3}. Using the data given in Table 1, calculate the oxygen deficit if the river water has a temperature of 10 °C.

View answer

Figure 4 illustrates how the dissolved oxygen (DO) concentration varies between the water surface and the interior of a water body when oxygen is consumed in the water. The resultant oxygen deficit causes oxygen to be transferred from the surface into the water body. As mentioned earlier, the greater the deficit, the greater the rate of oxygen transfer into the water.

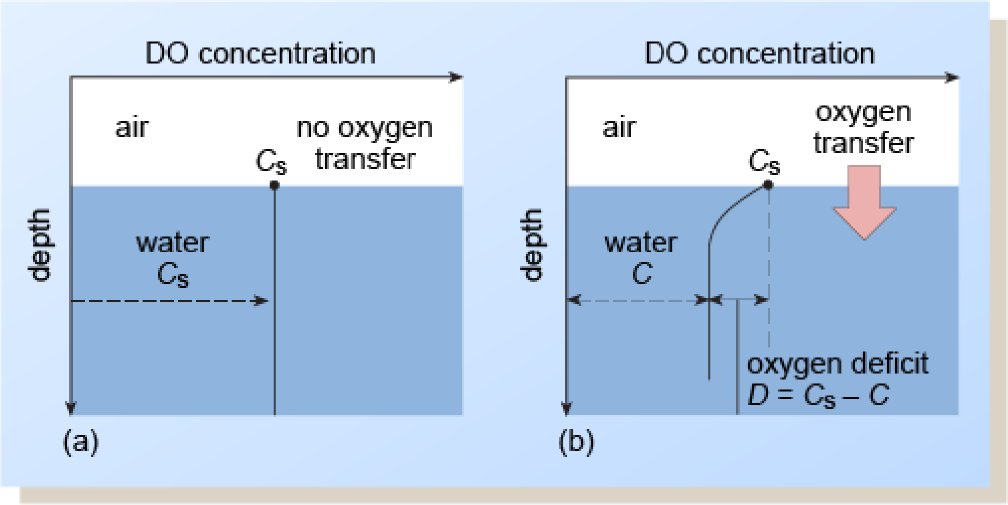

Figure 4 (a) Water body at equilibrium, with no consumption of oxygen; (b) consumption of oxygen in a water body resulting in an oxygen deficit, and oxygen consequently being transferred into the water body

View description

Other factors

The rate of oxygen transfer into a water body also depends on the turbulence of the water, since this helps transport oxygen from the surface layers to the main body of the water. Rapidly flowing, turbulent streams are therefore able to take up oxygen more rapidly than smooth -flowing slow ones.

Another factor governing the transfer of oxygen into a watercourse is the ratio of surface area to volume. A large surface area permits a greater diffusion of oxygen into the water. Hence shallow, wide rivers are reoxygenated more rapidly than deep, narrow ones. Agitation increases the ratio of surface area to volume – as, for example, when water flows over dams and weirs, and when waves are produced by strong winds. A further advantage of agitation is the entrainment of air bubbles as air is drawn into the water body.

The amount of oxygen in a water body at any given time depends not only on the characteristics mentioned above but also on biological and other factors. Almost all aquatic animals and plants use oxygen in carrying out their metabolic processes and so constantly tend to increase the oxygen deficit. If organic pollutants are present, the oxygen deficit is increased further as biodegradation takes place. At the same time as oxygen is being consumed, oxygen replenishment via photosynthesis and natural aeration takes place.

Figure 5 shows graphically the processes of oxygen demand and replenishment.

- Curve (a) shows the oxygen demand of a polluted water sample.
- Curve (b) shows the reaeration process observed when oxygen is forced to dissolve in the water due to the oxygen deficit created by the biodegradation taking place.
- Curve (c) shows the net result of the oxygen demand and replenishment processes. This is called the dissolved oxygen sag

curve. This is, in effect, the difference between the demand and replenishment curves.

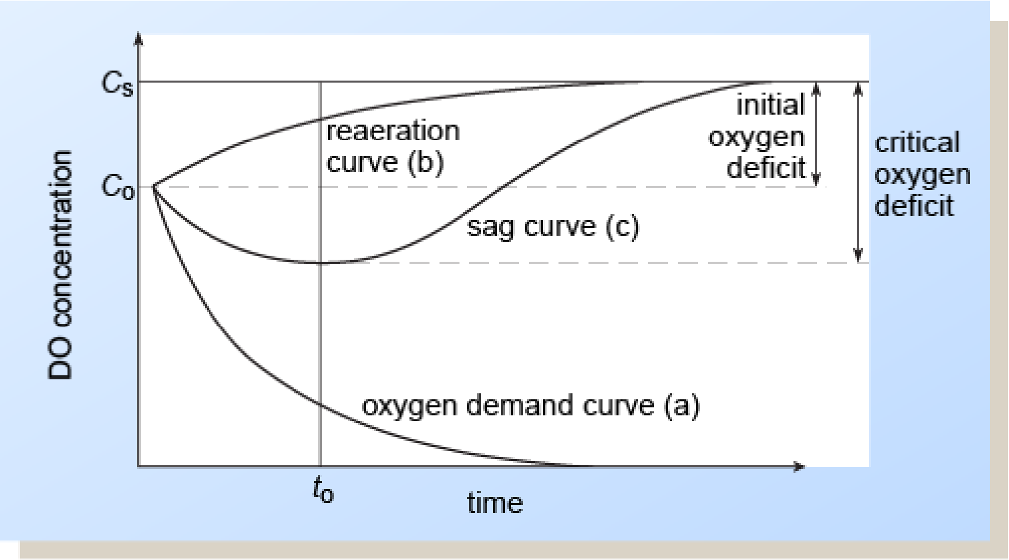

Figure 5 The dissolved oxygen sag curve

View description

SAQ 2

Which of the following events would not affect the rate of oxygen transfer from the atmosphere to a body of water?

- Doubling the oxygen deficit
- Large amounts of salts being discharged into the water
- A slight breeze blowing over the water
- The water flowing over a weir
- Raising the temperature of the water by 10 °C

View answer

1.3.4 Dissolved oxygen – variation

There are diurnal and (in temperate countries) seasonal differences in oxygen concentration in water. Figure 6 illustrates the diurnal variation that may occur. This variation is related to plant growth, light intensity and temperature. Variations of up to 10 g m^{-3} have been recorded in 24 hours.

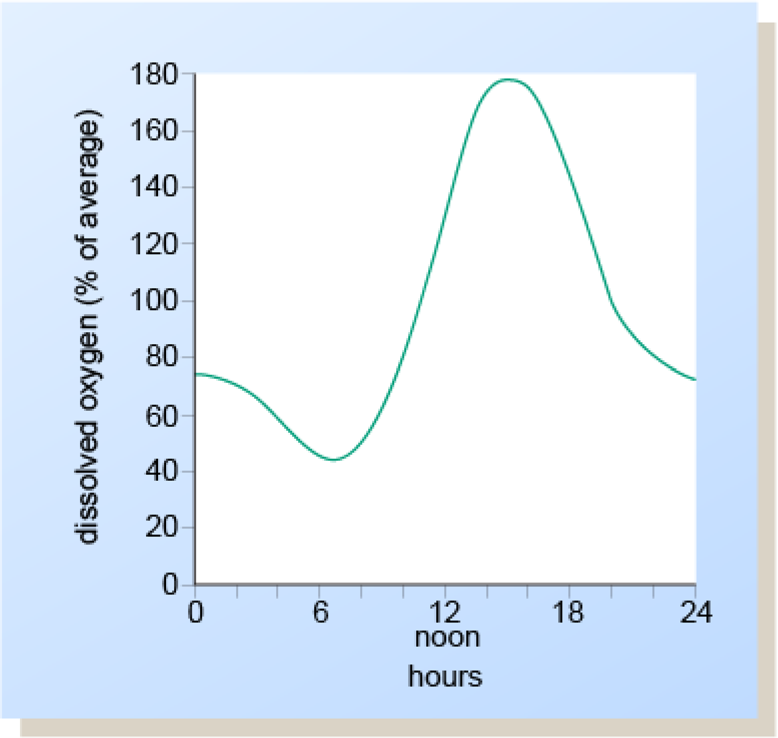

Figure 6 Hourly variation of dissolved oxygen in a water body

View description

The amount of dissolved oxygen rises to a maximum during the day because of photosynthesis occurring in daylight. It decreases through the night because none is produced by photosynthesis, but respiration continues using oxygen as it does during the daylight hours. In temperate countries in the northern hemisphere, this extreme diurnal variation occurs mainly between April and October, because the lower temperatures during the rest of the year tend to slow down or inhibit metabolic processes and plants become dormant; in temperate countries in the southern hemisphere, it is between October and April. Tropical countries exhibit diurnal variation all year round.

Figure 7 illustrates seasonal changes in dissolved oxygen in a temperate country in the northern hemisphere. An increase occurs in the summer months because of longer days (more daylight) and therefore increased photosynthetic activity.

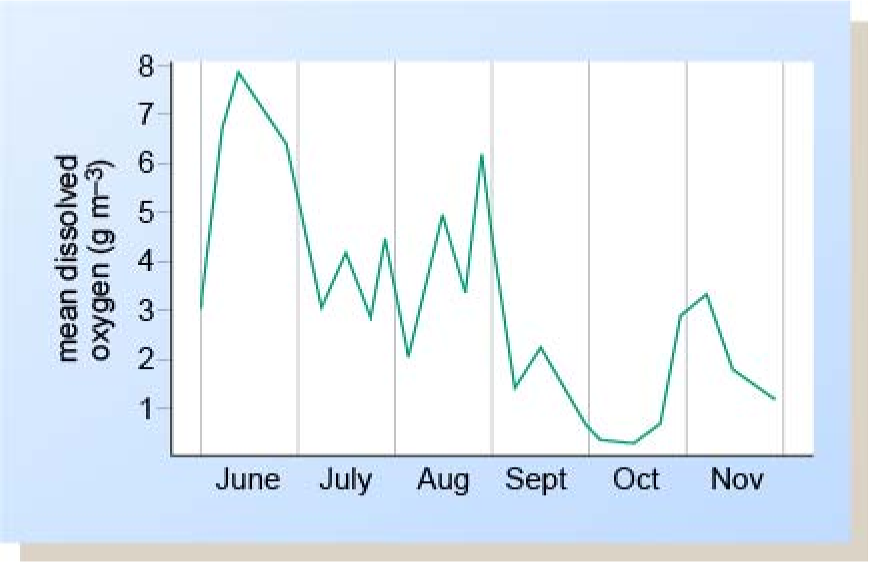

Figure 7 Seasonal variation of dissolved oxygen in a water body in a temperate country in the northern hemisphere (The Open University, 2007)

View description

In some circumstances oxygen supersaturation can occur, i.e. more oxygen is dissolved in the water than the saturation concentration allows. This occurs because plants produce pure oxygen (whereas air contains 21% oxygen). Therefore, when photosynthesis rather than atmospheric aeration is responsible for the oxygenation of the water, up to five times the saturation concentration is theoretically possible at the same temperature and pressure. In practice 500% is never attained, but up to 200% has been recorded in a shallow river with profuse plant growth on bright sunny days (YSI Environmental, 2005).

SAQ 3

When is the level of dissolved oxygen in a river likely to be at its highest and at its lowest?

View answer

1.3.5 Physical characteristics of natural water

A river’s physical characteristics include:

- clarity/*turbidity*
- colour
- speed of flow
- turbulence
- odour
- the presence of plants and macroscopic animal life.

These physical characteristics are determined by the location, geology and climate of the catchment area. In turn, they influence the chemical and biological characteristics of the watercourse.

Figure 8 shows how the aquatic ecosystem is a complex set of physical, chemical and biological interrelationships. Nutrients and dissolved oxygen may be essential to aquatic life, but other conditions must also be satisfactory.

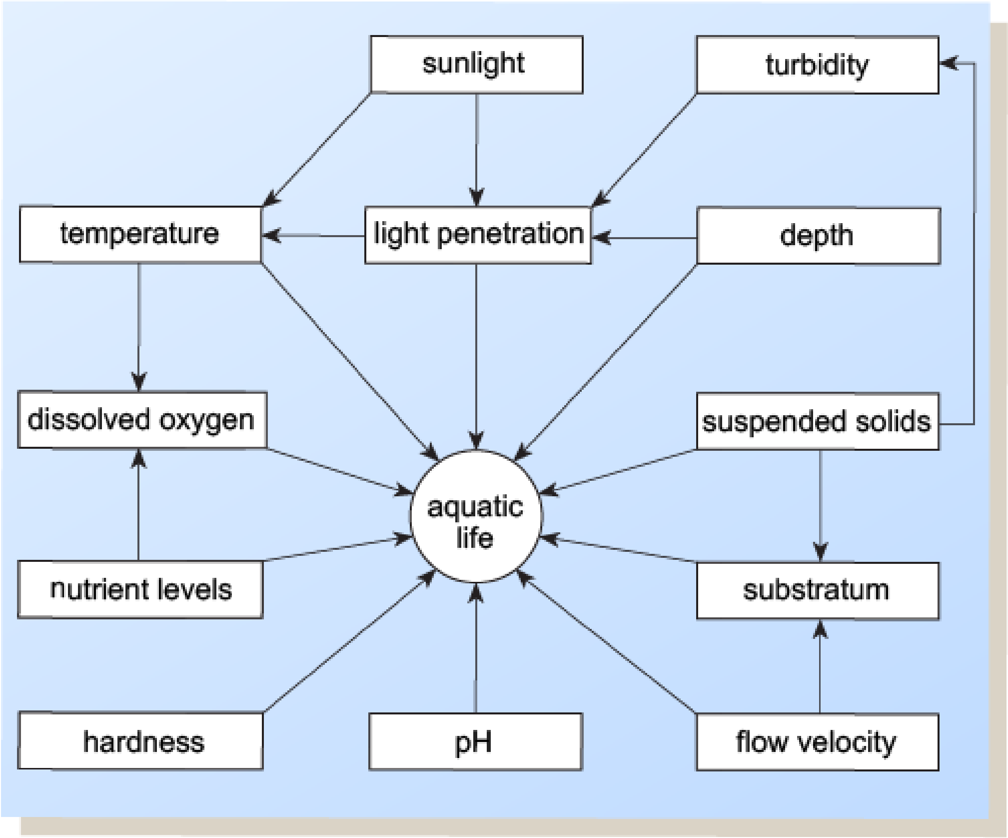

Figure 8 Influences acting on aquatic life

View description

Turbidity, colour and suspended solids

As water runs off the land, some insoluble substances are carried along as suspended solids. Many organisms (down to sub-micron level) present on the suspended solids are also transported. Depending on their sizes and the velocity of the river flow, the solid particles may settle out at a certain point or be carried on further. Quantities are affected by seasonal changes and, in temperate countries, tend to be higher in winter because of increased storm run-off due to higher rainfall and melting snow.

The quantity of suspended solids (measured in $g\ m^{-3}$) affects the turbidity or cloudiness of the water. Suspended solids may also contribute colour to the water. Turbidity is measured in *nephelometric* turbidity units (NTU).

Particles of all sizes tend to reduce light penetration; this reduces the rate of photosynthesis and therefore causes a reduction in the growth of plant life. Very small particles that settle out on the bottom of the stream may have a blanketing effect, thus preventing certain bottom dwellers from living there and green plants from photosynthesising.

Speed of flow and turbulence

To get nutrients and dissolved oxygen to all parts of a body of water, good mixing is important. You have seen one example in the form of thermal stratification in deep lakes in summer, where poor mixing and reduced sunlight led to the bottom layers not being supplied with the necessary conditions for plant growth.

The same principle applies to rivers and streams. Fast-moving, turbulent streams mix and agitate the water, aiding the transfer of oxygen from the atmosphere to the river, and carrying nutrients more efficiently to the plants and animals in the river. Small, rapidly flowing streams are nearly always saturated with dissolved oxygen; large, sluggish rivers may have oxygen concentrations well below saturation levels.

Temperature

All aquatic organisms have a fairly well-defined temperature tolerance range that determines their distribution. As you have already seen, temperature affects the saturation concentration of dissolved oxygen (see Table 1). An increase in water temperature will reduce the oxygen solubility and increase the metabolic activity of aquatic organisms. The combination of these two effects means that oxygen demand by organisms increases just when oxygen supply is being reduced.

Coarse fish such as yellow perch can tolerate water temperatures up to 26 °C (Hinshaw, 2007) and dissolved oxygen levels as low as 3.5 g m^{-3} (Hart et al., 2006). However, game fish such as the Atlantic salmon (Figure 9) die if the oxygen concentration drops below 5 g m^{-3} (Hendry and Cragg-Hine, 2003) or if the water temperature rises above 20 °C (Fisheries and Oceans Canada, 2012).

Figure 9 An Atlantic salmon

View description

1.3.6 Chemical characteristics of natural water

Since water is such a good solvent, it is not surprising to find many different chemical substances present in it. Water reaching a river will contain inorganic and organic compounds dissolved by rainwater percolating through the soil and rocks. In addition, some gases will dissolve in rainwater during its passage through the air.

The substances present in water may be conveniently grouped into:

- dissolved gases such as oxygen and carbon dioxide
- those involved in the bicarbonate–carbonate equilibrium derived from carbon dioxide (described below)
- other dissolved substances.

All gases will dissolve to a greater or lesser extent in water. As you have seen, oxygen has a low solubility in water, as do nitrogen, argon and some of the other gases present in the atmosphere.

Carbon dioxide, by contrast, is very soluble in water: 1 m^3 of water at 20 °C will hold 878 g of pure carbon dioxide. However, carbon dioxide is also special for another reason: when it dissolves, it reacts with the water to form bicarbonate and carbonate ions. The chemical equation describing this process is the bicarbonate–carbonate equilibrium (Equation 2).

$$\underset{\substack{\text{dissolved}\\\text{carbon}\\\text{dioxide}}}{CO_2} + H_2O \rightleftharpoons \underset{\substack{\text{hydrogen}\\\text{ion}}}{H^+} + \underset{\substack{\text{bicarbonate}\\\text{ion}}}{HCO_3^-} \rightleftharpoons \underset{\substack{\text{hydrogen}\\\text{ions}}}{2H^+} + \underset{\substack{\text{carbonate}\\\text{ion}}}{CO_3^{2-}} \tag{2}$$

View description

Because all the reactions are reversible, the whole system reaches *equilibrium* – so natural waters will contain various proportions of carbon dioxide, bicarbonate and carbonate.

How the acidity of the water affects the equilibrium in Equation 2

The acidity or alkalinity of water is expressed in terms of pH. This is an index on a scale from 0 to 14 and is derived from the concentration of hydrogen ions (H^+) in a solution. Low values of pH (lower than 7) correspond to ‘acid’ (high H^+) conditions, whereas high values of pH (greater than 7) correspond to ‘alkaline’ (low H^+) conditions.

For a reversible reaction, a change in the concentration of one of the chemical species in the reaction will produce a corresponding shift in the concentrations of the other species in order to ‘compensate’ for the change. So in Equation 2, if the concentration of hydrogen ions (H^+) increases, the reactions move towards the left to compensate. A new equilibrium is reached with higher concentrations of CO_2 and bicarbonate, and a lower concentration of carbonate. Conversely, a decrease in hydrogen ions shifts the reactions in Equation 2 to the right.

Figure 10 illustrates the relationship between pH and concentrations of CO_2, HCO_3^- and CO_3^{2-}. It shows that in high-pH (alkaline) water, most of the carbon dioxide ends up as bicarbonate and carbonate, whereas in low-pH (acidic) water, the carbon dioxide stays in solution without reacting further.

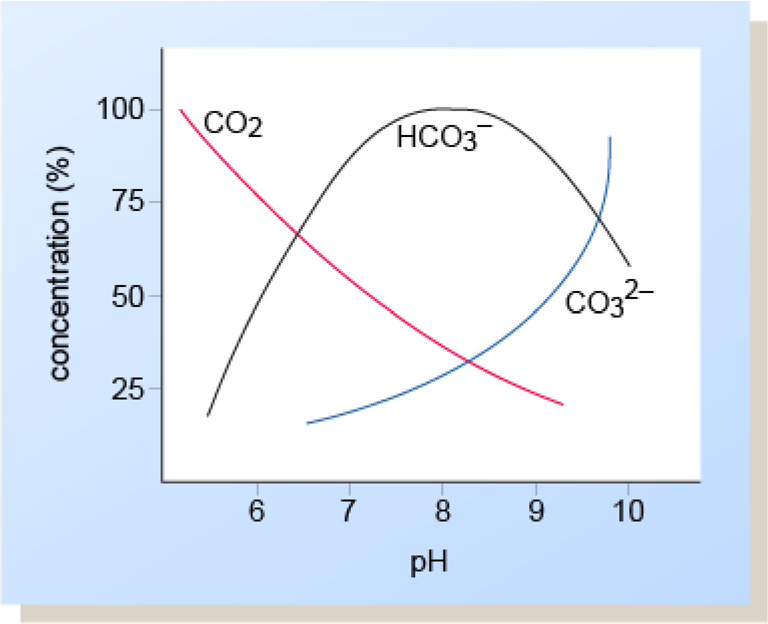

Figure 10 The bicarbonate–carbonate equilibrium

View description

The bicarbonate–carbonate equilibrium is important in photosynthesis, in which aquatic plants take up the inorganic carbon in carbon dioxide in the presence of sunlight for synthesising new cell material. All plants can use dissolved CO_2 for this purpose, but none apparently can use carbonate directly (though blue-green algae can use bicarbonate for photosynthesis). Thus, up to a point, low-pH waters – with available carbon dioxide – are more favourable for photosynthesis. However, if the pH falls much below 6 then the acidity of the water becomes limiting on plant growth.

The supply of carbon dioxide to the aquatic environment comes from the atmosphere through diffusion, and as the product of aerobic and anaerobic metabolism (respiration).

Apart from substances derived from the atmosphere, there are usually other substances dissolved in natural waters. ‘Salinity’ is a general term that means the concentration of *ionic* constituents (salts) dissolved in water. These include the naturally available carbonates, sulfates and chlorides of sodium, calcium, potassium and magnesium. However, ‘salinity’ may also mean specifically the sodium chloride content from either sewage or seawater intrusion. High chloride content in watercourses can also be caused by run-off from salted roads in winter. Some groundwaters may have high mineral content due to the geology of the area.

The pH of natural water usually varies from approximately 6 to 8 depending on the types of rocks and substrate surrounding the watercourse, although in some drainage areas it can be as low as 4.

Water hardness

You might be familiar with hard water from seeing scale deposited in kettles. As well as causing scale formation, both *temporary hardness* and *permanent hardness* make lathering with ordinary soap difficult. The result is the formation of scum that floats on the surface of washing water. On the benefit side, the dissolved solids or minerals often give hard water a pleasant taste; they are of nutritional importance to plants and microorganisms, and may have various medicinal functions for humans.

Hardness in water is mainly due to the presence of *ions* of the metals calcium (Ca^{2+}), magnesium (Mg^{2+}) and iron (Fe^{2+}). Rivers and lakes fed by water that has run from areas of chalk and limestone ($CaCO_3$) contain an abundance of calcium. Calcium and magnesium account for at least 70% of the *cations* in water.

When calcium, magnesium and iron are present in water as bicarbonate salts, e.g. $Ca(HCO_3)_2$, and the water is boiled or heated above 70 °C, carbonate salts of the metals are precipitated. Such water is said to possess temporary or carbonate hardness because the carbonate salts (e.g. calcium carbonate) are largely insoluble, and are thus removed from the water and deposited as scale (see Equation 3).

$$\underset{\substack{\text{calcium}\\\text{bicarbonate}}}{Ca(HCO_3)_2} \xrightarrow{\text{heat}} H_2O + CO_2 + \underset{\substack{\text{calcium carbonate}\\\text{(scale)}}}{CaCO_3} \tag{3}$$

View description

If scale is deposited on the heating elements of kettles and the surfaces of industrial boilers, it shortens their life and makes them less efficient.

When calcium, magnesium and iron are present as chloride or sulfate salts (e.g. $CaCl_2$), the hardness is called permanent or non-carbonate hardness. Although this type of hardness also contributes to scaling, in this case the precipitate is due to the decreased solubility of these metal salts at higher temperatures and not to the formation of new insoluble compounds.

Plant nutrients

Of the major nutrients of plants, nitrogen and phosphorus are important growth-limiting factors in primary production (i.e. if levels of either of these nutrients fall below a certain level, plants will not grow even if other nutrients are in plentiful supply). Increased levels of nitrogen and phosphorus are the cause of eutrophication, as described earlier. Both nitrogen and phosphorus enter watercourses through natural leaching by water of the soluble nitrates and phosphates found in soils and rocks, as well as from sewage effluent and agricultural sources. Phosphorus is the limiting growth factor for certain types of *blue-green algae*, and both phosphorus and nitrogen limit the growth of *green algae*.

Radiological aspects

Environmental radiation comes from a range of naturally occurring and anthropogenic sources, and varies a great deal with location. Globally, natural sources are estimated to contribute on average about 80% of the radiation dose experienced by people – the remainder is from medical exposure (WHO, 2013a). Any exposure to radiation can lead to cancer and the greater the exposure, the greater the risk.

The contribution that drinking water makes to radiation intake is very small, and is due largely to naturally occurring *radionuclides* in the uranium and thorium decay series.