1

Introducing Fuel Cells

1.1 Historical Perspective

This book is an introduction to fuel-cell systems; it aims to provide an understanding of the technology — what it is, how it works and what are its applications. Essentially, a fuel cell can be defined as a device that produces electrical power directly from a fuel via an electrochemical process. In some respects, this operation is similar to that of a conventional battery except that the reactants are stored outside the cell. Therefore, the performance of the device is limited only by the availability of the fuel and oxidant supply and not by the cell design. For this reason, fuel cells are rated by their power output (kW) rather than by their capacity (kWh).

Before addressing the technology in depth, it is necessary to understand that by virtue of being electrochemical, fuel cells have both chemical and electrical characteristics. Accordingly, their development has been inextricably linked with the development of electrochemistry as a distinct branch of physical chemistry.

At the beginning of the 19th century, it was recognized that an ‘electrochemical cell’ (nowadays, commonly called a ‘battery’) could be made by placing two dissimilar metals in an aqueous salt solution. This discovery was made by Alessandro Volta, the professor of experimental physics at Pavia University, who constructed a pile of alternating discs of copper (or silver or brass) and zinc (or tin) that were separated by pasteboard discs (or ‘any other spongy matter’) soaked in brine. When the top and bottom of the pile were connected by a wire, the assembly delivered, for the first time in history, a more or less steady flow of electricity. Volta introduced the terms ‘electric current’ and ‘electromotive force’, the latter to denote the physical phenomenon that causes the current to flow. In due course, he conveyed his findings in a letter dated 20 March 1800 to Joseph Banks, the then president of the Royal Society. Known as the ‘Volta (or Voltaic) pile’, this was the first ‘primary’ (or non-rechargeable) power source, as opposed to a ‘secondary’ (or rechargeable) power source.

Sir Humphry Davy, who was working at the Royal Institution in London, soon recognized that the Volta pile produces electricity via chemical reactions at the metal–solution interfaces — hydrogen is evolved on the ‘positive’ copper disc, and zinc is consumed at the ‘negative’ disc. Indeed, this recognition of the relationship between chemical and electrical effects prompted Davy to coin the word ‘electrochemical’, from which sprang the science of ‘electrochemistry’. He gave warning that Volta’s work was ‘an alarm bell to experimenters all over Europe’. His prediction was soon to be verified.
Volta had sent his letter to the Royal Society in two parts because he anticipated problems with its delivery given that correspondence from Italy had to pass through France, which was then at war with Britain. While waiting for the second part to arrive, Joseph Banks had shown the first few pages to Anthony Carlisle (a fashionable London surgeon) who, in turn, with the assistance William Nicholson (a competent amateur scientist) assembled on 30 April 1800 the first pile to be constructed in England. Almost immediately, on 2 May 1800, the two investigators found that the current from their device when passed through a dilute salt solution via two platinum wires was capable of decomposing water into its constituents — hydrogen at one wire and oxygen at the other. Details of the discovery were published in Nicholson’s own journal in July of the same year. Thus, the new technique of ‘molecular splitting’ — to be coined ‘electrolysis’ by Michael Faraday much later in 1834 and derived from the Greek ‘lysis’ = separation — was demonstrated before Volta’s own account of the pile was made public in September 1800. A schematic representation of the process is shown in Figure 1.1a.

It was left to Michael Faraday, Davy’s brilliant student, to identify the mechanisms of the processes that take place within ‘electrolytic’ cells and to give them a quantitative basis. In addition, he was also the guiding force behind the nomenclature that is still in use today. First, Faraday with the assistance of Whitlock Nicholl (his personal physician and accomplished linguist) devised the name ‘electrode’ to describe a solid substance at which an electrochemical reaction occurs and ‘electrolyte’ to describe the chemical compound that provides an electrically conductive medium between electrodes. (Note that in the case of dissolved materials, it is fundamentally incorrect to refer to the ‘electrolyte solution’ as the ‘electrolyte’; nevertheless, the latter terminology has become common practice.) To distinguish between the electrode by which conventional current (i.e., the reverse flow of electrons) enters an electrolytic cell and the electrode by which

![Figure 1.1](image-url)
Introducing Fuel Cells

it leaves, Faraday sought the assistance of the polymath William Whewell, the Master of Trinity College at the University of Cambridge. In a letter dated 24 April 1834, he asked Whewell:

‘Can you help me out to two good names not depending upon the idea of a current in one direction only or upon positive or negative?’

In other words, he wanted terms that would be unaffected by any later change in the convention adopted for the direction of current. Eventually, they settled on calling the positive electrode an ‘anode’ and the negative electrode a ‘cathode’, which were coined from Greek ‘ano‐dos’ (‘upwards’–‘a way’) to represent the path of electrons from the positive electrode to the negative and ‘katho‐dos’ (‘downwards’–‘a way’) to represent the counter direction. For an electrolytic cell, then, the anode is where the current enters the electrolyte and the cathode is where the current leaves the electrolyte. Thus the positive electrode sustains an oxidation (or ‘anodic’) reaction with the liberation of electrons, while a reduction (or ‘cathodic’) reaction takes place at the negative electrode with the uptake of electrons.

With use of the Greek neutral present participle ‘ion’ — ‘a moving thing’ — to describe the migrating particles in electrolysis, two further terms were obtained, namely, ‘anion’, i.e., the negatively charged species that goes to the anode against the current (or with the flow of negative charge), and ‘cation’, i.e., the positively charged species that goes to the cathode with the current (or against the flow of negative charge). The operation of an electrolysis cell is shown in Figure 1.1a. It should be noted that the anode–cathode terminology for an ‘electrolytic cell’ applies to a ‘battery under charge’ (secondary system).

A fuel cell operates in the reverse manner to an electrolysis cell, i.e., it is a ‘galvanic’ cell that spontaneously produces a voltage (similar to a ‘battery under discharge’). The anode of the electrolysis cell now becomes the cathode and the cathode becomes the anode; see Figure 1.1b. Nevertheless, the directions of the migration of anions and cations with respect to current flow are unchanged such that the positive electrode remains a positive electrode and the negative electrode remains a negative electrode. Thus, in a fuel cell, the fuel is always oxidized at the anode (positive electrode), and the oxidant is reduced at the cathode (negative electrode).

There is some debate over who discovered the principle of the fuel cell. In a letter written in December 1838 and published on page 43 of the January issue of the January–June 1839 Volume XIV of The London and Edinburgh Philosophical Magazine and Journal of Science, the German scientist Christian Friedrich Schönbein described his investigations on fluids that were separated from each other by a membrane and connected to a galvanometer by means of platina wires. In the 10th of 14 reported tests, one compartment contained dilute sulfuric acid that held some hydrogen, whereas the other compartment contained dilute sulfuric acid that was exposed to air. Schönbein detected a current and concluded that this was caused ‘by the combination of hydrogen with (the) oxygen (contained dissolved in water)’. This discovery was largely overlooked, however, after the publication of a letter from William Robert Grove, a Welsh lawyer and a scientist at the Royal Institution; see Figure 1.2a. The letter, which was dated 14 December 1838, appeared on page 127 of the February issue of the aforementioned Volume XIV and described his evaluation of electrode and electrolyte materials for use
in batteries. Unfortunately, the order in which these two letters had been written is unknown as Schönbein did not date his letter in full — he gave the month, but not the day. In fact, this chronology is of little importance given the following postscript that Grove had added to his letter in January 1839:

‘I should have pursued these experiments further, and with other metals, but was led aside by some experiments with different solutions separated by a diaphragm and connected by platinum plates; in many of these I have been anticipated.’

In the same postscript, Grove went on to speculate that by connecting such cells in series sufficient voltage could be created to decompose water (by electrolysis).

Grove carried out many experiments that demonstrated the principle of the fuel cell. In 1842, he realized that the reaction at the electrodes was dependent on an area of contact between the gas reactant and a layer of liquid that was sufficiently thin to allow the gas to diffuse to the solid electrode (today, this requirement is commonly related to the formation of a ‘three-phase boundary’ or ‘triple-point junction’ where gas, electrolyte and electrocatalyst come into simultaneous contact, *v.i.*). At that time, Grove was the professor of experimental chemistry at the London Institution in Finsbury Circus, and in the same communication he reported the invention of a ‘gaseous voltaic battery’. The device employed two platinized platinum electrodes (to increase the real surface area), and a series of fifty such pairs when semi-immersed in dilute sulfuric acid solution was found ‘to whirl round’ the needle of a galvanometer, to give a painful shock to five persons joining hands, to give a brilliant spark between charcoal points, and to decompose hydrochloric acid, potassium iodide and acidulated water. An original sketch of four such cells is reproduced in Figure 1.2b. It was also found that 26 cells were the minimum number required to electrolyse water. Grove had indeed realized

![Figure 1.2](https://commons.wikimedia.org/w/index.php?curid=20390734)
the desire expressed in his 1839 postscript in that he had achieved the beautiful symmetry inherent in the ‘decomposition of water by means of its composition.’

The aforementioned apparatus became widely recognized as the first fuel cell and Grove was designated as the ‘Father of the Fuel Cell.’ Historically, this title is not fully justified. More accurately, Schönbein should be credited with the discovery of the fuel-cell effect in 1838 and Grove with the invention of the first working prototype in 1842. Happily, such accreditations were of little concern to the two scientists and they became close friends. For almost 30 years, they exchanged ideas and developments via a dynamic correspondence and visited each other frequently.

It is interesting to note that many latter-day authors have attributed the introduction of the term ‘fuel cell’ to Ludwig Mond and Charles Langer in their description of a new form of gas battery in 1889. Remarkably, however, there is no mention of ‘fuel cell’ in this communication. Other claims that William W. Jacques, in reporting his experiments to produce electricity from coal, coined the name are equally ill-founded. A. J. Allmand in his book *The Principles of Applied Electrochemistry*, published in 1912, appears to attribute the appellation ‘fuel cell’ to the Nobel Laureate Friedrich Wilhelm Ostwald in 1894.

Grove concluded his short paper in 1842 with the following modest entreaty:

‘Many other notions crowd upon my mind, but I have occupied sufficient space and must leave them for the present, hoping that other experimenters will think the subject worth pursuing.’

Unfortunately, however, the invention of the first internal combustion engine to become commercially successful by Jean Joseph Étienne Lenoir in 1859, coupled ironically with Faraday’s earlier discovery of electromagnetic force, diverted the course of electricity generation from electrochemical to electromagnetic methods. As a result, the fuel cell became merely an object of scientific curiosity during much of the next half-century. Meanwhile, knowledge of electrochemical conversion and storage of energy progressed largely through the development of battery technologies.

In 1894, a well-documented criticism against heat engines was expressed by Friedrich Ostwald, who drew attention to the poor efficiency and polluting problems associated with producing electrical power via the combustion of fossil fuels rather than by direct electrochemical oxidation. A fuel cell is inherently a more thermodynamically efficient device since, unlike an engine in which heat is converted to mechanical work, the cell is not subject to the rules of the Carnot cycle. By virtue of this cycle, the efficiency of a thermal engine is always lowered to a value far below 100%, as determined by the difference between the temperature at which heat is taken in by the working fluid and the temperature at which it is rejected. On this basis, Ostwald advocated that:

‘The path which will help to solve this biggest technical problem of all, this path must be found by the electrochemistry. If we have a galvanic element which directly delivers electrical power from coal and oxygen, [...] we are facing a technical revolution that must push back the one of the invention of the steam engine. Imagine how [...] the appearance of our industrial places will change! No more smoke, no more soot, no more steam engine, even no more fire, [...] since fire will now only be needed for the few processes that cannot be accomplished electrically, and those will daily diminish. [...] Until this task shall be tackled, some time will pass by.’
Regrettably, Ostwald was proven to be correct as regards his closing prediction for although attempts were made at the turn of the century to develop fuel cells that could convert coal or carbon into electricity (for instance, the work of William W. Jacques in the United States), the need for an expensive platinum catalyst and its poisoning by carbon monoxide formed during the coal gasification limited cell affordability, usefulness and lifetime. Consequently, interest in such ‘direct carbon fuel cells’ dwindled.

In the 1930s, Emil Bauer and H. Preis in Switzerland experimented with solid oxide fuel cells (SOFCs). Given the limitations of solid oxides at that time (i.e., poor electrical conductivity and chemical stability), G.H.J. Broers and J.A.A. Ketelaar in the late 1950s turned to the use of fused salts as electrolytes. The work gave birth to the molten carbonate fuel cell (MCFC), which eventually became one of the main types of fuel cell in commercial production.

The renaissance of the fuel‐cell concept in the 20th century can be attributed largely to the work of Englishman F.T. (Tom) Bacon. He was an engineer by profession and thus appreciated the many potential advantages of the fuel cell over both the internal combustion engine and the steam turbine as a source of electrical power. His interest in fuel cells dated as far back as 1932, and he ploughed a lone furrow, with little support or backing, but showed enormous dedication to the challenge of developing practical cells. Early in his career, Bacon elected to study the alkaline‐electrolyte fuel cell (AFC), which used nickel‐based electrodes, in the belief that platinum‐group electrocatalysts would never become commercially viable. In addition, it was known that the oxygen electrode is more readily reversible in alkaline solution than in acid. This choice of electrolyte and electrodes necessitated operating the cell at moderate temperatures (100–200°C) and high gas pressures. Bacon restricted himself to the use of pure hydrogen and oxygen as reactants. Finally, in August 1959, he demonstrated the first workable fuel cell — a 40‐cell system that could produce about 6 kW of power, which was sufficient to run a forklift truck and to operate a welding machine as well as a circular saw.

A major opportunity to apply fuel cells arose in the early 1960s with the advent of space exploration. In the United States, fuel cells were first employed to provide spacecraft power during the fifth mission of Project Gemini. Batteries had been employed for this purpose in the four earlier flights, as well as in those conducted in the preceding Project Mercury. This switch in technology was undertaken because payload mass is a critical parameter for rocket‐launched satellites, and it was judged that fuel cells, complete with gas supplies, would weigh less than batteries. Moreover, the objective of Project Gemini was to evolve techniques for advanced space travel — notably, the extravehicular activity and the orbital manoeuvres (rendezvous, docking, etc.) required for the moon landing planned in the following Project Apollo. Thus, lunar flights demand a source of power of longer duration than that available from batteries.

A proton‐exchange membrane fuel cell (PEMFC) system manufactured by the General Electric Company was adopted for the Gemini missions (two modules, each with a maximum power of about 1 kW), but this was replaced in Project Apollo by an AFC of circulating electrolyte design, as pioneered by Bacon and developed by the Pratt and Whitney Aircraft Company (later the United Technologies Corporation). Both
types of system were fuelled by hydrogen and oxygen from cryogenic tanks. The AFC could supply 1.5 kW of continuous power, and its in-flight performance during all 18 Apollo missions was exemplary. In the 1970s, International Fuel Cells (a division of United Technologies Corporation) produced an improved AFC for the Space Shuttle orbiter that delivered eight times more power than the Apollo version and weighed 18 kg less. The system provided all of the electricity, as well as drinking water, when the Space Shuttle was in flight.

The successful exploitation of fuel cells in the space programme drove research activity worldwide during the 1970s to develop systems that would generate power with high efficiency and low emissions for terrestrial applications. Research was stimulated further by the hiatus in the global oil supply in 1974. What followed was the emergence of various national initiatives on fuel-cell development. In the United States, demonstrations of phosphoric acid fuel cell (PAFC) technology by the American Gas Association led to a Notice of Market Opportunities (NOMO) initiative. This activity, in turn, renewed interest in the MCFC by US researchers, and in the mid-1980s, national research and development programmes were established in Japan and Europe. Renewed interest in the PEMFC was championed in the late 1980s by Geoffrey Ballard, a Canadian pioneer, who saw the potential for the technology to replace internal combustion engines. Since then, this system has been the subject of much advancement for a variety of applications, so much so that it merits two chapters in this book.

### 1.2 Fuel-Cell Basics

To understand how the reaction between hydrogen and oxygen produces an electric current, and where the electrons are released, it is necessary to consider the reaction that takes place at each electrode. The reactions vary for different types of fuel cell, but it is convenient to start with a cell based around an acid electrolyte, not only because this system was used by Grove but also because it is the simplest and still the most chosen for commercial applications.

At the anode of an acid fuel cell, hydrogen is oxidized and thereby releases electrons and creates $H^+$ ions, as expressed by:

$$2H_2 \rightarrow 4H^+ + 4e^- \quad (1.1)$$

This reaction also releases energy in the form of heat.

At the cathode, oxygen reacts with electrons taken from the electrode, and $H^+$ ions from the electrolyte, to form water, i.e.,

$$O_2 + 4e^- + 4H^+ \rightarrow 2H_2O \quad (1.2)$$

Thus the overall cell reaction is:

$$2H_2 + O_2 \rightarrow 2H_2O + \text{heat} \quad (1.3)$$
Clearly, for both the electrode reactions to proceed continuously, electrons produced at the negative electrode must pass through an electrical circuit to the positive. Also, H\(^+\) ions must pass through the electrolyte solution — an acid is a fluid with free H\(^+\) ions and so serves this purpose very well. Certain polymers and ceramic materials can also be made to contain mobile H\(^+\) ions. These materials are commonly called ‘proton-exchange membranes’, as an H\(^+\) ion is also known as a proton. The PEMFC is examined in detail in Chapter 4.

The cell reaction (1.3) shows that two hydrogen molecules will be needed for each oxygen molecule if the system is to be kept in balance. The operating principle is illustrated in Figure 1.3.

In a fuel cell with an alkaline electrolyte (AFC), the overall reaction of hydrogen oxidation is the same, but the reactions at each electrode are different. In an alkaline solution, hydroxyl (OH\(^-\)) ions are available and mobile. At the anode, these ions react with hydrogen to release electrons and energy (heat) together with the production of water:

\[
2\text{H}_2 + 4\text{OH}^- \rightarrow 4\text{H}_2\text{O} + 4e^- \tag{1.4}
\]

At the cathode, oxygen reacts with electrons taken from the electrode, and water in the electrolyte and thereby forms new OH\(^-\) ions:

\[
\text{O}_2 + 4e^- + 2\text{H}_2\text{O} \rightarrow 4\text{OH}^- \tag{1.5}
\]

Comparing equations (1.4) and (1.5) shows that, as with an acid electrolyte, twice as much hydrogen is required compared with oxygen. The operating principle of the AFC is presented in Figure 1.4.

There are many other types of fuel cell, each distinguished by its electrolyte and the reactions that take place on the electrodes. The different systems are described in detail in the following chapters.

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**Figure 1.3** Electrode reactions and charge flow for fuel cell with an acid electrolyte. Note that although the negative electrons flow from the anode to cathode, the ‘conventional positive current’ flows from cathode to anode.
1.3 Electrode Reaction Rates

The oxidation of hydrogen at the negative electrode liberates chemical energy. It does not follow, however, that the reaction proceeds at an unlimited rate; rather, it has the 'classical' energy form of most chemical reactions, as shown in Figure 1.5. The schematic represents the fact that some energy must be used to excite the atoms or molecules sufficiently to start the chemical reaction — the so-called 'activation energy'. This energy can be in the form of heat, electromagnetic radiation or electrical energy. In visual terms, the activation energy helps the reactant to overcome an 'energy hill',

![Diagram of fuel cell](image-url)

*Figure 1.4* Electrode reactions and charge flow for a fuel cell with an alkaline electrolyte. Electrons flow from negative anode to positive cathode, but 'conventional positive current' flows from cathode to anode.

![Energy diagram](image-url)

*Figure 1.5* Classical energy diagram for a simple exothermic chemical reaction.
and once the reaction starts, everything rolls downhill. Thus, if the probability of an atom or molecule having sufficient energy is low, then the reaction will only proceed slowly. This is indeed the case for fuel-cell reactions, unless very high temperatures are employed.

The three main ways of dealing with the slow reaction rates are to (i) use catalysts, (ii) raise the temperature and (iii) increase the electrode area. Whereas the first two options can be applied to any chemical reaction, the electrode area has a special significance for electrochemical cells. The electrochemical reactions take place at the location where the gas molecules (hydrogen or oxygen) meet the solid electrode and the electrolyte (whether solid or liquid). The point at which this occurs is often referred to as the ‘three-phase boundary/junction’ or the ‘triple-phase boundary/junction’ (v.s.).

Clearly, the rate at which either electrode reaction proceeds will be proportional to the area of the respective electrode. Indeed, electrode area is such an important issue that the performance of fuel cells is usually quoted in terms of the current per cm$^2$. Nevertheless, the geometric area (length × width) is not the only issue. The electrode is made highly porous so as to provide a great increase in the ‘effective’ surface area for the electrochemical reactions. The surface area of electrodes in modern fuel cells, such as that shown in Figure 1.6, can be two to three orders of magnitude greater than the geometric area. The electrodes may also have to incorporate a catalyst and endure high temperatures in a corrosive environment; catalysts are discussed in Chapter 3.

**Figure 1.6** Transmission electron microscope image of a fuel-cell catalyst. The black spots are the catalyst particles that are finely divided over a carbon support. The structure clearly has a large surface area. *(Source: Courtesy of Johnson Matthey Plc.)*
1.4 Stack Design

Because a fuel cell functions at a low voltage (i.e., well below 1 V), it is customary to build up the voltage to the desired level by electrically connecting cells in series to form a ‘stack’. There are a number of different designs of fuel cell, but in each case the unit cell has certain components in common. These are as follows:

- An electrolyte medium that conducts ions. This may be a porous solid that contains a liquid electrolyte (acid, alkali or fused salt) or a thin solid membrane that may be a polymer or a ceramic. The membrane must be an electronic insulator as well as a good ionic conductor and must be stable under both strong oxidizing and strong reducing conditions.

- A negative fuel electrode (anode) that incorporates an electrocatalyst, which is dispersed on an electronically conducting material. The electrode is fabricated so that the electrocatalyst, the electrolyte and the fuel come into simultaneous contact at a three-phase boundary (v.s.).

- A positive electrode (cathode), also with a triple-point electrocatalyst, at which the incoming oxygen (either alone or in air) is reduced by uptake of electrons from the external circuit.

- A means of electrically connecting individual cells together. The design of interconnector depends on the geometry adopted for the cells.

- Seals that keep the gases apart and also prevent cell-to-cell seepage of liquid electrolyte, which otherwise would give rise to partial short-circuits.

A stack also has current-collectors that are located at the two ends of the stack and are connected by end-plate assemblies.

Historically, the flat plate is by far the preferred geometry for fuel cells, and one way of assembling such cells in series is to connect the edge of each negative electrode to the positive of the next cell through the string, as illustrated in Figure 1.7. (For simplicity, the diagram ignores the difficulty of supplying gas to the electrodes.) The problem with this method, however, is that the electrons have to flow across the face of the electrode to the current collection point at the edge. The electrodes might be quite good conductors, but if each cell is only operating at about 0.7 V, even a small voltage drop can be significant. Consequently, this type of stack design is not used unless the current flows are very low, the electrodes are particularly good conductors and/or the dimensions of the stack are small.

A much better method of cell interconnection for planar fuel cells is to use a ‘bipolar plate’. This is an electrically conducting plate that contacts the surfaces of the positive electrode of one cell and the negative electrode of the next cell (hence the term ‘bipolar’). At the same time, the bipolar plate serves as a means of feeding oxygen to the negative anode and fuel gas to the positive cathode of the adjacent cells. This is achieved by having channels machined or moulded on either side of the plate along which the gases can flow and the products, i.e., pure water in the case of hydrogen fuel, can exit. Various designs of channel geometry have been proposed to maximize the access of gases and the removal of water, e.g., pin-type, series–parallel, serpentine, integrated and interdigitated flow-fields. The different types are described in later chapters when considering the stacking arrangement of each type of fuel cell. The arrangement of the
channels (also known as the ‘flow-field’) leads the bipolar plate to be also known as the flow-field plate. Bipolar plates must also be relatively impermeable to gases, sufficiently strong to withstand stack assembly and easily mass produced. They are made of a good electronic conductor such as graphite or stainless steel. For transport applications, low weight and low volume are essential. The method of connecting two plates to a single cell is illustrated in Figure 1.8; the respective gases are supplied orthogonally.

To connect several cells in series, anode–electrolyte–cathode assemblies have to be prepared. These are then ‘stacked’ together with bipolar plates placed between each pair of cells. In the particular arrangement shown in Figure 1.9, the stack has vertical channels for feeding hydrogen over the anodes and horizontal channels for feeding oxygen (or air) over the cathodes. The result is a solid block, in which the electric current passes efficiently more or less straight through the cells, rather than over the surface of each electrode one after the other.

The electrodes and electrolytes are also well supported, and the whole structure is clamped together to give a strong and robust device. Although simple in principle, the design of the bipolar plate has a significant effect on fuel-cell performance. If the electrical connection between cells is to be optimized, then the area of contact points should be as large as possible, but this would mitigate good gas flow over the electrodes. If the contact points have to be small, at least they should be frequent. This may render the plate more complex, difficult and expensive to manufacture, as well as fragile. Ideally, bipolar plates should be as thin as possible so as to minimize both the electrical resistance between individual cells and the stack size. On the other hand, such an
approach would narrow the gas channels and thereby place greater demands on the pumps for supplying gases. High rates of flow are sometimes required, especially when using air instead of pure oxygen at the positive electrode. For low-temperature fuel cells, the circulating air has to evaporate and carry away the product water. Moreover, in many

*Figure 1.8* Single cell with end-plates for collecting current from the whole face of the adjacent electrode and applying gases to each electrode.

*Figure 1.9* A three-cell stack showing how bipolar plates connect the anode of one cell to the cathode of its neighbour.
cases, additional channels have to pass through the bipolar plate to carry a cooling fluid. Some further challenges for the bipolar plate are considered in the next section.

1.5 Gas Supply and Cooling

The arrangement given in Figure 1.9 has been simplified to show the basic principle of the bipolar plate. In practice, however, the twin problems of gas supply and preventing leaks mean that the design is somewhat more complex.

Because the electrodes must be porous (to permit the access of gas), they allow leakage of the gas through their edges. Consequently, the edges must be sealed. Sometimes this is done by making the electrolyte compartment slightly larger than one, or both, of the electrodes and fitting a gasket around each electrode, as presented in Figure 1.10. Such assemblies can then be made into a stack in which the fuel and oxygen can then be supplied to the electrodes using the external manifolds as shown disassembled in Figure 1.11. With this arrangement, the hydrogen should only come into contact with the anodes as it is fed vertically through the fuel-cell stack. Similarly, the oxygen (or air) fed horizontally through the stack should only contact the cathodes and certainly not the edges of the anodes. Such would not be the case for the basic design illustrated in Figure 1.9.

The externally manifolled design suffers from two major disadvantages. The first is that it is difficult to cool the stack. Fuel cells are far from 100% efficiency, and considerable quantities of heat are generated, as well as electrical power. In practice, the cells in this type of stack have to be cooled by the reactant air passing over the positive electrodes. This means that air has to be supplied at a higher rate than that demanded by the cell chemistry — sometimes the flow is sufficient to cool the cell, but it is wasteful of energy. The second disadvantage of external manifolding is that there is uneven pressure over the gasket round the edge of the electrodes, i.e., at the points where there

![Figure 1.10](image_url) The construction of cathode–electrolyte–anode units with edge seals that prevent the gases leaking in or out through the edges of the porous electrodes.
Introducing Fuel Cells

is a channel and the gasket is not pressed firmly onto the electrode. This increases the probability of leakage of the reactant gases.

‘Internal manifolding’ is a more common stack arrangement and requires a more complex design of bipolar plate, such as that displayed schematically in Figure 1.12. In this arrangement, the plates are made larger relative to the electrodes and have extra channels running through the stack for the delivery of fuel and oxygen to the electrodes. Holes are carefully positioned to feed the reactants into the channels that run over the surface of the electrodes. Reactant gases are fed in at the ends of the stack where the respective positive and negative electrical connections are also made. An example of a commercial fuel‐cell stack is shown in Figure 1.13.

A stack with internal manifolding can be cooled in various ways. The most practical method is to circulate a liquid coolant through electrically conductive metal plates that are inserted between groups of cells. In this passive approach, the heat within the plane of the plate must be conducted out to one or more of the edges of the plate for transfer to a heat-exchanger external to the fuel‐cell stack. Alternatively the bipolar plates themselves can be made thicker and machined to incorporate extra channels that allow passage of cooling air or water. The preferred cooling method varies greatly with the type of fuel cell and is addressed in later chapters.

From the foregoing discussion, it should be apparent that the bipolar plate is a key component of a fuel‐cell stack. As well as being a fairly intricate item to manufacture, the choice of material for its construction raises issues. For low‐temperature fuel
cells, graphite was one of the first materials to be employed, but it is difficult to work and brittle and, consequently, has now largely been replaced by various carbon composite materials. Stainless steel can also be used, but it will corrode in some types of fuel cell. Ceramic materials have found application in fuel cells that operate at high temperatures. The bipolar plate nearly always is a major contributor to the capital cost of a fuel cell.
1.6 Principal Technologies

Setting aside practical issues such as manufacturing and materials costs, the two fundamental technical problems with fuel cells are:

- The slow reaction rates, particularly for the oxygen reduction reaction, which lead to low levels of current and power.
- The fact that hydrogen is not a readily available fuel.

To address these problems, many different types of fuel cell have been developed and tested. The systems are usually distinguished by the electrolyte that is used and the operating temperature, though there are always other important differences as well. There are six principal types of fuel cell, namely:

- Low temperature (50–150°C): alkaline electrolyte (AFC), proton-exchange membrane (PEMFC), direct methanol (DMFC) and other liquid-fed fuel cells.
- Medium temperature (around 200°C): PAFC.
- High temperature (600–1000°C): molten carbonate (MCFC) and SOFC.

Some operational data on each type are given in Table 1.1. There are other less well-known types such as the direct borohydride (DBFC) and direct carbon fuel cells (DCFC); the former operates at low temperatures and the latter at high temperatures.

### Table 1.1 Principal types of fuel cell.

<table>
<thead>
<tr>
<th>Fuel cell type</th>
<th>Mobile ion</th>
<th>Operating temperature (°C)</th>
<th>Fuel</th>
<th>Applications and notes</th>
</tr>
</thead>
<tbody>
<tr>
<td>Alkaline (AFC)</td>
<td>OH⁻</td>
<td>50–200</td>
<td>Pure H₂</td>
<td>Space vehicles, e.g., Apollo, Shuttle</td>
</tr>
<tr>
<td>Proton-exchange</td>
<td>H⁺</td>
<td>30–100 + a</td>
<td>Pure H₂</td>
<td>Vehicles and mobile applications, and for lower power CHP systems</td>
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<td></td>
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<td>Direct methanol</td>
<td>H⁺</td>
<td>20–90</td>
<td>Methanol</td>
<td>Portable electronic systems of low power, running for long times</td>
</tr>
<tr>
<td>(DMFC)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H⁺</td>
<td>~220</td>
<td>H₂, (low S, low CO, tolerant to CO₂)</td>
<td>Large numbers of 200-kW CHP systems in use</td>
</tr>
<tr>
<td>(PAFC)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Molten carbonate</td>
<td>CO₃²⁻</td>
<td>~650</td>
<td>H₂, various hydrocarbon fuels (no S)</td>
<td>Medium- to large-scale CHP systems, up to MW capacity</td>
</tr>
<tr>
<td>(MCFC)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Solid oxide</td>
<td>O²⁻</td>
<td>500–1000</td>
<td>Impure H₂, variety of hydrocarbon fuels</td>
<td>All sizes of CHP systems, 2kW to multi MW</td>
</tr>
<tr>
<td>(SOFC)</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

CHP, combined heat and power.

a) New electrolyte materials as described in Chapter 4 are enabling higher operating temperatures for the PEMFC.

1 Although hydrogen is preferred for most types of fuel cell, other fuels can be used for some technologies. For example, methanol is employed in the direct methanol fuel cell (DMFC) and carbon as the fuel in the direct carbon fuel cell (DCFC).
To date, the PEMFC has proved to be the most successful commercially. The electrolyte is a solid polymer, in which protons are mobile. The chemistry is the same as that shown Figure 1.3 for an acid-electrolyte system. The PEMFC runs at relatively low temperatures, so the problem of slow reaction rates is addressed by using sophisticated catalysts and electrodes. Platinum has been the preferred catalyst. It is an expensive metal but, through improvements in materials, only minute amounts are now required. Consequently, in modern PEMFC designs, the platinum makes a relatively small contribution to the total cost of the fuel-cell system. More recent research suggests that in some cases platinum can be eliminated from the catalyst. Further discussion of the PEMFC is given in Chapter 4. The PEMFC has to be fuelled with hydrogen of high purity, and methods for meeting this requirement are discussed in Chapter 10.

The DMFC is a variant of the PEMFC. The technology differs from the PEMFC only in that methanol in its native liquid form is used as fuel. Other liquid fuels such as ethanol and formic acid may also be viable for some applications. Unfortunately, most of these liquid-fuelled cells produce very low levels of power, but, even with this limitation, there are many potential applications for such devices in the rapidly growing area of portable electronics devices. Such cells, for the foreseeable future at least, will remain low-power units and will therefore suit applications that require slow and steady consumption of electricity over long periods.

As mentioned earlier, an AFC system was chosen for the Apollo and Space Shuttle orbiter craft. The problem of slow reaction rate was overcome by using highly porous electrodes, with a platinum catalyst, and sometimes by operating at quite high pressures. Although some historically important AFCs have been run at about 200°C, the systems usually operate below 100°C. Unfortunately, the AFC is susceptible to poisoning by the carbon dioxide in the atmosphere. Thus the air and fuel supplies must be free from this gas, or else pure oxygen and hydrogen must be supplied.

The PAFC was the first type of fuel cell to reach commercialization and the technology enjoyed a reasonable degree of widespread terrestrial use in the period 1980–2000. Many 200-kW systems, manufactured by the International Fuel Cells Corporation, were installed in the United States and Europe. Other systems were produced by Japanese companies. In the PAFC, porous electrodes, platinum catalysts and a moderately high temperature (~220°C) help to boost the reaction rate to a reasonable level. Such PAFC systems were fuelled with natural gas, which is converted to hydrogen within the fuel-cell system by steam reforming. The required equipment for steam reforming unfortunately adds considerably to the costs, complexity and size of the fuel-cell system. Nevertheless, PAFC systems have demonstrated good performance in the field, for instance, units have run for periods in excess of 12 months without any maintenance that has required shutdown or human intervention. A typical installation of a 400 kW PAFC system is shown in Figure 1.14.

The most common form of SOFC operates in the region of 600–1000°C. These high temperatures permit high reaction rates to be achieved without the need for expensive platinum catalysts. At these elevated temperatures, fuels such as natural gas can be used directly (internally reformed) within the fuel cell without the need for a separate processing unit. The SOFC thus addresses the aforementioned key problems (viz. slow reaction rates and hydrogen supply) and takes full advantage of the inherent simplicity of the fuel-cell concept. Nevertheless, SOFCs are made from thin ceramic materials that are difficult to handle and therefore are expensive to manufacture. In addition,
a large amount of extra equipment is needed to make a full SOFC system, e.g., air and fuel preheaters, heat-exchangers and pumps. Also the cooling system is more complex than for low-temperature fuel cells. Care also has to be taken during start-up and shutdown of SOFC systems, on account of the intrinsic fragile nature of the ceramic materials in the stacks.

The MCFC has an interesting and distinguishing feature in that it requires carbon dioxide to be fed to the positive electrode, as well as oxygen. This is usually achieved by recycling some of the exhaust gas from the anode to the cathode inlet. The high temperature means that a good reaction rate is achieved with a comparatively inexpensive catalyst — nickel. Like the SOFC, an MCFC system can be fuelled directly with gases, such as methane and coal gas (a mixture of hydrogen and carbon monoxide), without the need for an external reformer. This advantage for the MCFC is somewhat offset, however, by the nature of the electrolyte, namely, a hot and corrosive molten mixture of lithium, potassium and sodium carbonates.

### 1.7 Mechanically Rechargeable Batteries and Other Fuel Cells

At the start of this book, a fuel cell was defined as an electrochemical device that converts a fuel to electrical energy (and heat) continuously, as long as reactants are supplied to its electrodes. The implication is that neither the electrodes nor the electrolyte is consumed by operation of the cell. Of course, in all fuel cells the electrodes and
electrolytes are degraded and subject to ‘wear and tear’ during service. The first two technologies under consideration in this section are often misleadingly described as fuel cells and employ electrodes that are entirely consumed during use.

1.7.1 Metal–Air Cells

The most common type of cell of this category is the zinc–air battery, though aluminium–air and magnesium–air cells have been produced commercially. In all cases, the basic operation is the same.

At the negative electrode, the metal reacts with hydroxyl ions in an alkaline electrolyte to form the metal oxide or hydroxide. For example, the reaction with a zinc fuel is given by:

\[ \text{Zn} + 2\text{OH}^- \rightarrow \text{ZnO} + \text{H}_2\text{O} + 2e^- \]  

The electrons thus released pass round the external electric circuit to the air electrode where they are available for the reaction between water and oxygen to form more hydroxyl ions. Thus at the air electrode the reaction is exactly the same as equation (1.5) for the AFC. Cells using a salt solution (e.g., seawater) as the electrolyte solution also work reasonably well when using aluminium or magnesium as the fuel.

Metal–air cells have a very high specific energy (Wh kg\(^{-1}\)). Zinc–air batteries are employed widely in devices that require long running times at low currents, such as hearing aids. Some interest has also been shown in the development of units with higher power for application in electric vehicles. Such systems can also be ‘refuelled’ by replenishing the metal consumed at the negative electrode — which is why the technology is sometimes promoted as a ‘fuel cell’. This claim is also supported by the fact that the reaction at the positive electrode is exactly the same as for a fuel cell, and indeed the same electrodes can be used. It should be noted, however, that removal of the metal oxide will also necessitate renewal of the electrolyte solution. Thus, the metal–air systems cannot properly be described as fuel cells and are best classified as ‘mechanically rechargeable batteries’.

1.7.2 Redox Flow Cells

Another type of electrochemical power source that is sometimes taken to be a fuel cell is the ‘redox flow cell’ (or ‘flow cell’); a multicell unit is usually referred to as a ‘flow battery’. It is useful at this point to define two types of flow cell, as several different chemistries are under development:

1) Flow batteries, in which there is a decoupling of cell power and cell capacity, e.g., the bromine–polysulfide cell and the vanadium redox cell.
2) Hybrid flow batteries, in which there is no decoupling of cell power and cell capacity, e.g., the zinc–bromine battery.

The first category is different from all other fuel cells in that the oxidant is not air, and therefore it cannot be said that the fuel is ‘combusted’. In this type of cell, there is one reactant (which can be called the fuel) that is oxidized and a complementary reactant that serves as the oxidant. These are removed from the electrode compartments when the cell is being charged and stored in tanks. The capacity of such cells can thus be very large. Discharge is undertaken by resupplying the reactants to the electrodes.
Two flow batteries have been the subject of much research, namely, the sodium-bromide–sodium-polysulfide cell and the vanadium redox cell. The former cell was introduced in the 1990s by Regenesys Technologies Limited in the United Kingdom. After a utility-scale demonstration at a power station by the National Power in Cambridgeshire, United Kingdom, the development was taken over by RWE and subsequently by Prudent Energy to complement its own work on a vanadium battery. No further studies or trials of the Regenesys have been reported.

The vanadium redox battery was pioneered in the 1980s at the University of New South Wales in Sydney, Australia, and the Japanese Electrotechnical Laboratory. The operating principle of the system is illustrated in Figure 1.15. The two reactants are flowing aqueous solutions of vanadium sulfate and the electrode reactions are as follows.

At the positive electrode:

\[
\text{VO}_2^+ + 2\text{H}^+ + e^- \overset{\text{Discharge}}{\leftrightarrow} \text{VO}^{2+} + \text{H}_2\text{O}
\]

At the negative electrode:

\[
\text{V}^{2+} \overset{\text{Discharge}}{\leftrightarrow} \text{V}^{3+} + e^- \tag{1.8}
\]

Thus, in the charged state, the positive-electrolyte loop contains a solution of \(\text{V}^{5+}\) and the negative loop contains a solution of \(\text{V}^{2+}\). On discharging, the former solution is reduced to \(\text{V}^{4+}\) and the latter is oxidized to \(\text{V}^{3+}\). The difference in the oxidation state of

![Figure 1.15](image-url)  
*Figure 1.15* Operating principle of the vanadium redox battery.
vanadium in the two reactant solutions produces 1.2–1.6 V across the membrane, as determined by the electrolyte solution, temperature and state-of-charge. Regeneration takes place by reversing the flow of the solutions and applying a potential across the cell to restore the original oxidation states in the solutions.

It can easily be seen that (i) this is a reversible cell and (ii) the capacity of the cell (e.g. as measured in kWh) is determined by the amount of liquid pumped, i.e., the size of the storage tanks, and not by the dimensions of the electrodes as would be the case in a normal battery. Furthermore, the more the cells and the faster the flow of electrolyte solutions, the higher is the power rating. This approach enables economies of scale in both manufacturing and energy–power capacity.

The vanadium redox cell shares many characteristics with now-abandoned Regenesys. Numerous companies and organizations have been involved in funding and developing the vanadium technology, and several large field trials have been conducted around the world. Research and development is continuing.

In the hybrid form of flow cell, one or more of the electroactive components are deposited as a solid layer. Consequently, the system may be viewed as a combination of one battery electrode and one fuel-cell electrode. The zinc–bromine system is the best-known example of such technology. A modern version developed by Redflow Limited, an Australian-based company, is shown in Figure 1.16. As with the vanadium redox cell, the zinc–bromine cell is comprised of two fluids that pass carbon-plastic electrodes that are each placed in a half-cell either side of a microporous polyolefin membrane. During discharge, zinc and bromine combine into zinc bromide and thereby generate 1.8 V across each cell. During charge, metallic zinc will be drawn out of solution and deposited (plated) as a thin film on one side of the negative electrode. Meanwhile, bromine evolves as a dilute solution at the positive electrode on the other side of the membrane. Because bromine is a highly volatile and reactive liquid, it is complexed with an organic reagent to form a poly-bromo compound, which is an oil and is immiscible with the aqueous electrolyte solution. The oil sinks down to the bottom of the electrolytic

![Figure 1.16](Redox zinc–bromine battery. (Source: Courtesy of Redflow Pty Ltd.)
tank and is separated and stored in a special compartment in the external reservoir of the positive electrode until required again for discharge. The capacity of the cell is limited by the amount of zinc that can be plated on the negative electrode.

1.7.3 Biological Fuel Cells

Finally, it should be noted that, although not yet a principal technology, the biological fuel cell is attracting interest as a long-term prospect. The cell would normally operate with an organic fuel, such as methanol or ethanol. The distinctive ‘biological’ aspect is that the electrode reactions are promoted by enzymes present in microbes, rather than by conventional ‘chemical’ catalysts such as platinum. Hence, these systems — also known as ‘microbial fuel cells (MFCs)’ — replicate nature in the way that energy is derived from organic fuels. Biological or microbial fuel cells should be distinguished from biological methods for generating hydrogen, which is then used in a conventional fuel cell. Such methods of hydrogen production are discussed in Chapter 10. Research into advanced microfluidics, new bacterial strains, more robust separator membranes and efficient electrodes is the key to unlocking the potential of MFCs.

1.8 Balance-of-Plant Components

It should be evident that a practical fuel-cell system requires not only a readily available fuel but also a means of cooling the stack, an ability to employ the heat produced to do useful work and an application for the direct current (dc) power that is produced by the stack(s). For a fuel-cell stack to function effectively, various other components are necessary. The exact composition of this so-called balance-of-plant depends on the type of fuel cell, the available fuel and its purity and the desired outputs of electricity and heat. Typical auxiliary subsystems are:— (i) fuel clean-up processor, e.g., for sulfur removal — so-called desulfurization; (ii) steam reformer and shift reactor for the fuel; (iii) carbon dioxide separator; (iv) humidifier; (v) fuel and air delivery units; (vi) power-conditioning equipment, e.g., for inverting dc to alternating current (ac) and then transforming to line voltage; (vii) facilities for the management of heat and water; (viii) overall control and safety systems; and (ix) thermal insulation and packaging. Individual components include fuel storage tanks and pumps, compressors, pressure regulators and control valves, fuel and/or air pre-heaters, heat-exchangers and radiators, voltage regulators, motors and batteries (to provide power for pumps on start-up). These important subsystem issues are described in much more detail in Chapter 12.

The requirements for a fuel-cell system for a stationary power application and a vehicle are very different. In a stationary power plant system, such as shown in Figure 1.14, the fuel-cell stack is, in terms of size, a small part of the installation that is dominated by the fuel and heat-processing systems and the power-conditioning equipment. This will nearly always be the case for combined heat and power (CHP) facilities that run on a conventional fuel such as natural gas.

By contrast, a fuel-cell power source for a car is shown in Figure 1.17. The unit operates on gaseous hydrogen fuel that is stored on the vehicle, and the waste heat is only used to warm the car interior. The fuel-cell stack occupies the bulk of the compartment
that would normally be filled with an internal combustion engine (ICE). Other components of a hydrogen fuel-cell ‘engine’ in a vehicle, i.e., pumps, humidifier, power electronics and compressor, are generally much less bulky than those of a CHP system.

1.9 Fuel-Cell Systems: Key Parameters

To compare the performance of fuel-cell systems with each other and with other electric power generators, some key operating parameters must be considered. For electrodes and electrolytes, the key criterion is the current per unit area, which is always known as the ‘current density’ and usually expressed in terms of mA cm\(^{-2}\), except in the United States where A ft\(^{-2}\) is frequently adopted (the two units are quite similar, i.e., 1.0 mA cm\(^{-2}\) = 0.8 A ft\(^{-2}\)). The current density should be reported at a specific operating voltage, typically about 0.6 or 0.7 V. The values for current density and selected voltage can then be multiplied to give the power per unit area, in mW cm\(^{-2}\). A note of caution should be made here, namely, that electrodes frequently do not ‘scale up’ properly. That is, if the area is doubled the current will often not double. The reasons for this are varied but generally relate to issues such as the even delivery of reactants to, and removal of products from, the entire face of the electrode.

Specific power (kW kg\(^{-1}\)) and power density (kW m\(^{-3}\) or kW L\(^{-1}\)) are key ‘figures of merit’ for comparing electrical generators. Note that whereas power is measured in kW, energy is simply power delivered over a certain period of time and is measured in kWh. The capital cost of a fuel-cell system is obviously an important parameter and is usually quoted in US$ per kW for ease of comparison.

The lifetime of a fuel cell is rather difficult to specify. Standard engineering measures such as ‘mean time between failures’ (MTBF) are not entirely applicable given that the performance of a fuel cell always gradually deteriorates and the power drops fairly
steadily with time as the electrodes and the electrolyte solution both age. The degradation of a fuel cell is sometimes reported as a decline in cell voltage, given in units of mV per 1000 h. Formally, the life of a fuel cell is considered to be over when it can no longer deliver the rated power, e.g., when a 10-kW fuel cell can no longer deliver 10 kW. It should be noted that, when new, a fuel cell may be capable of providing more than the rated power, e.g., an extra 25% is not unusual.

The remaining fuel-cell characteristic of key importance is the efficiency, i.e., the electrical energy delivered by the system compared with the energy supplied as fuel. When making comparisons between systems in terms of efficiency, care should be taken that the data are expressed on the same basis. Efficiency is addressed in Chapter 2.

In the automotive industry, primary issues are the cost per kW and the power density. In round figures, current ICE technology costs US$10 per kW and delivers 1 kW L−1. Such a power source should last at least 4000 h, i.e., about 1 h of duty each day for over 10 years. For CHP plant, the capital cost is still important, but a much higher target of US$1000 per kW is generally accepted. The higher cost is due to the extra balance of plant that is required and to the fact that the system must have a substantially longer lifetime. A period of 40 000 h would be a minimum. For stationary power-generation systems, the levelized cost of electricity (LCOE) is often used as a measure of performance. The LCOE is the price at which electricity must be generated from a specific source to break even over the lifetime of the project. It is an economic assessment of the cost of the generating system and includes all the costs over its lifetime, namely, capital cost, operations and maintenance, and cost of fuel. The LCOE enables analysts to compare the costs of fuel-cell systems with other forms of power generation.

1.10 Advantages and Applications

For all types of fuel cell, a significant disadvantage or barrier to commercialization is the capital cost. There are, however, various advantages that feature more or less strongly for the different systems and lead to fuel cells being attractive for different applications. These include the following:

- **Efficiency.** As explained in Chapter 2, fuel cells are generally more efficient than piston- or turbine-based combustion engines. A further benefit is that small fuel-cell systems can be just as efficient as large ones. This capability opens up a market opportunity for small-scale cogeneration (CHP) that cannot be satisfied with turbine- or engine-based systems.
- **Simplicity.** The essentials of a fuel cell involve few, if any, moving parts. This can lead to highly reliable and long-lasting systems.
- **Low emissions.** When hydrogen is the fuel, pure water is the by-product of the main reaction of the fuel cell. Consequently, the power source is essentially ‘zero emission.’ This is a particularly attractive for vehicle applications, as there is a requirement to reduce emissions and even eliminate them within cities. Nevertheless, it should be noted that, at present, emissions of carbon dioxide are nearly always involved in the production of the hydrogen.
- **Silence.** Fuel cells are very quiet — even those with extensive extra fuel-processing equipment. Quietness is very important in both portable-power applications and for local power generation via CHP schemes.
Ironically, the fact that hydrogen is the preferred fuel is, in the main, one of the principal disadvantages of fuel cells. On the other hand, many envisage that as fossil fuels run out, hydrogen will become a major fuel and energy vector throughout the world. It could be generated, for example, by electrolysing water using electricity provided by massive arrays of photovoltaic (solar) cells. Indeed, the so-called hydrogen economy may emerge in future decades. In the meantime, it is more likely that 'hydrogen energy' will have only a very small impact globally as it is most economically produced by the steam reforming of natural gas (see Chapter 10).

In summary, the advantages of fuel cells impact particularly strongly on CHP systems (both large and small scales) and on mobile power systems — especially for vehicles and electronic equipment such as portable computers, mobile telephones and military communications equipment. A notable feature of the technology is the very wide range in system sizes, i.e., from a few watts up to several megawatts. In this respect, fuel cells are unique as energy converters.

**Further Reading**
