

(d) Building and Testing

Observations

Part 1 Cell Potentials of Various Fruit Tested with Zn and Cu Electrodes

Type of Fruit	Cell potential (V) of Zn/Cu cell
lemon	0.82
apple	0.80
orange	0.83
banana	0.80
strawberry	0.79
lime	0.87
grapefruit	0.80

Part 2 Cell Potentials of Various Metal Combinations in a Lime

Metal combinations	Cell potential (V)
Mg/Cu	1.87
Zn/Cu	0.87
Mg/Zn	0.35

Part 3 Total Potentials of Mg/Cu–Lime Cells in Series

Number of cells	LED brightness	Cell potential (V)
1	not lit	1.87
2	lit	3.60
3	lit slightly brighter	5.21

Marketing

- (e) Students should devise a variety of creative formats to market their inventions: poster, magazine advertisement, radio commercial, TV infomercial, sports team sponsorship, etc.

UNIT 5 REVIEW

(Pages 444–447)

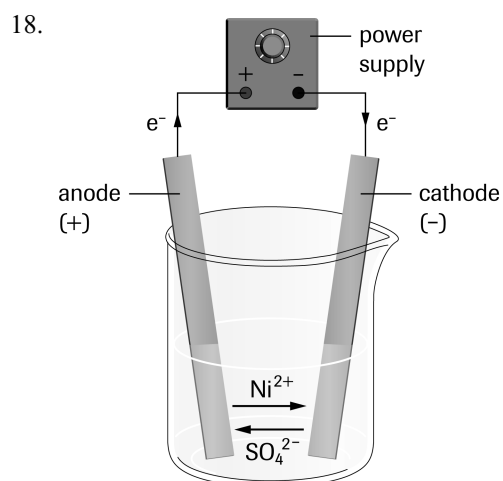
Understanding Concepts

- The gain of electrons is called *reduction*.
 - The loss of electrons is called *oxidation*.
 - An element's oxidation number *increases* when the element is oxidized.
 - An element's oxidation number *decreases* when the element is reduced.
- | | |
|--|---|
| Reactant oxidized | Reactant reduced |
| (a) $\text{Zn}_{(s)}$ | $\text{Cl}_{2(g)}$ |
| (b) $\text{Ca}_{(s)}$ | $\text{O}_{2(g)}$ |
| (c) $\text{Zn}_{(s)}$ | $\text{Cu}^{2+}_{(aq)}$ |
| (d) $\text{Zn}_{(s)}$ | $\text{Ag}^{+}_{(aq)}$ (in $\text{AgNO}_{3(aq)}$) |
| (e) $\text{Al}_{(s)}$ | $\text{Fe}^{3+}_{(s)}$ (in $\text{Fe}_2\text{O}_{3(s)}$) |
| (f) $\text{H}_{2(g)}$ | $\text{O}_{2(g)}$ |
| (g) $\text{Mg}_{(s)}$ | $\text{CO}_{2(g)}$ |
| (h) $\text{I}^{-}_{(aq)}$ (in $\text{KI}_{(aq)}$) | $\text{Cl}_{2(g)}$ |

3. (a) +2
(b) -2
(c) +4
(d) +2
(e) +2
(f) +3
(g) +4
(h) +3
4. No. For any spontaneous chemical reaction, reduction and oxidation must occur together. Electrons are lost by one reactant because they are “pulled away” by the other reactant.
5. $\begin{array}{ccccccc} +1 & -2 & 0 & & 0 & & +1 & -2 \\ \text{(a)} & \text{H}_2\text{S}_{(\text{g})} & + \text{O}_{2(\text{g})} & \rightarrow & 2 \text{S}_{(\text{s})} & + & 2 \text{H}_2\text{O}_{(\text{g})} \end{array}$
The oxidation number of sulfur changed from -2 to 0, and the oxidation number of oxygen changed from 0 to -2. Therefore, this is a redox reaction.
 $\begin{array}{ccccccc} -4 & +1 & 0 & & +4 & -2 & +1 & -2 \\ \text{(b)} & \text{CH}_{4(\text{g})} & + 2 \text{O}_{2(\text{g})} & \rightarrow & \text{CO}_{2(\text{g})} & + & 2 \text{H}_2\text{O}_{(\text{g})} \end{array}$
The oxidation numbers of carbon and oxygen changed from -4 and 0 to +4 and -2, respectively. Therefore, this is a redox reaction.
 $\begin{array}{ccccccc} +2 & -2 & +1 & -1 & +2 & -1 & +1 & -2 \\ \text{(c)} & \text{ZnO}_{(\text{s})} & + 2 \text{HCl}_{(\text{aq})} & \rightarrow & \text{ZnCl}_{2(\text{aq})} & + & \text{H}_2\text{O}_{(\text{l})} \end{array}$
Since there is no change in oxidation numbers for any of the atoms or ions, the given reaction is not a redox reaction.
 $\begin{array}{ccccccc} +2 & -2 & +2 & -2 & 0 & & +4 & -2 \\ \text{(d)} & 2 \text{PbO}_{(\text{s})} & + \text{PbS}_{(\text{s})} & \rightarrow & 3 \text{Pb}_{(\text{s})} & + & \text{SO}_{2(\text{g})} \end{array}$
The oxidation numbers of lead and sulfur changed from +2 and -2 to 0 and +4, respectively. Therefore, this is a redox reaction.
6. (a) Magnesium is higher than copper; $\text{Mg}_{(\text{s})} + \text{Cu}(\text{NO}_3)_{2(\text{aq})} \rightarrow \text{Cu}_{(\text{s})} + \text{Mg}(\text{NO}_3)_{2(\text{aq})}$
(b) Iron is lower than sodium; $\text{Fe}_{(\text{s})} + \text{Na}_2\text{SO}_{4(\text{aq})} \rightarrow \text{no reaction}$
(c) Nickel is higher than silver; $\text{Ni}_{(\text{s})} + 2 \text{AgNO}_{3(\text{aq})} \rightarrow 2 \text{Ag}_{(\text{s})} + \text{Ni}(\text{NO}_3)_{2(\text{aq})}$
(d) Calcium is higher than hydrogen; $\text{Ca}_{(\text{s})} + 2 \text{HCl}_{(\text{aq})} \rightarrow \text{H}_{2(\text{g})} + \text{CaCl}_{2(\text{aq})}$
7. Since iron is above copper on the activity series, the iron in the spoon will be oxidized. The chemical equation for this reaction is
 $\text{Fe}_{(\text{s})} + \text{CuSO}_{4(\text{aq})} \rightarrow \text{Cu}_{(\text{s})} + \text{FeSO}_{4(\text{aq})}$
8. As aluminum oxidizes, it forms a protective layer of aluminum oxide that adheres well to the underlying metal. The oxide layer effectively isolates the aluminum metal from its environment, preventing further corrosion (oxidation) from occurring.
9. (a) Separating the reactants forces the electrons that are lost at the anode to travel through a conductor, where the energy of electrons can be used to power an electrical device. Otherwise, if the reactants were in direct contact, electrons would transfer directly from one electrode to the other, and not pass through the external circuit.
(b) The salt bridge supplies anions and cations that are necessary to prevent the build-up of positive and negative charges around the anode and cathode, respectively. A flow of charges from one half-cell to the other is necessary to complete the circuit in the galvanic cell.
(c) The cell potential drops to zero because the flow of charges through the wire between the anode and cathode has stopped. The accumulation of positive charge around the anode prevents further oxidation. Similarly, accumulation of negative charge at the cathode prevents further reduction from occurring.
10. (a) $\text{Cu}_{(\text{aq})}^{2+} + \text{Zn}_{(\text{s})} \rightarrow \text{Zn}_{(\text{aq})}^{2+} + \text{Cu}_{(\text{s})}$
(b) $2 \text{Ag}_{(\text{aq})}^{+} + \text{Ni}_{(\text{s})} \rightarrow \text{Ni}_{(\text{aq})}^{2+} + 2 \text{Ag}_{(\text{s})}$
11. The porous fabric prevents the anode and cathode chemicals from reacting directly with each other, while still allowing electrolyte to flow between the anode and the cathode.
12. Both cells produce electrical energy as a result of redox reactions. The total energy a dry cell can produce is limited by the amount of chemicals it contains. The fuel cell, however, can produce electrical energy indefinitely because it has a continual fresh supply of reactants pumped into it.
13. (a) The voltage of a cell is determined by the chemicals that are reacting in the cell (electrodes and electrolytes). The voltage of a battery is determined by the number of cells inside the battery, and how they are arranged.
(b) The amount of electrical energy a battery or cell can produce is determined by the amount of chemicals contained within each cell.

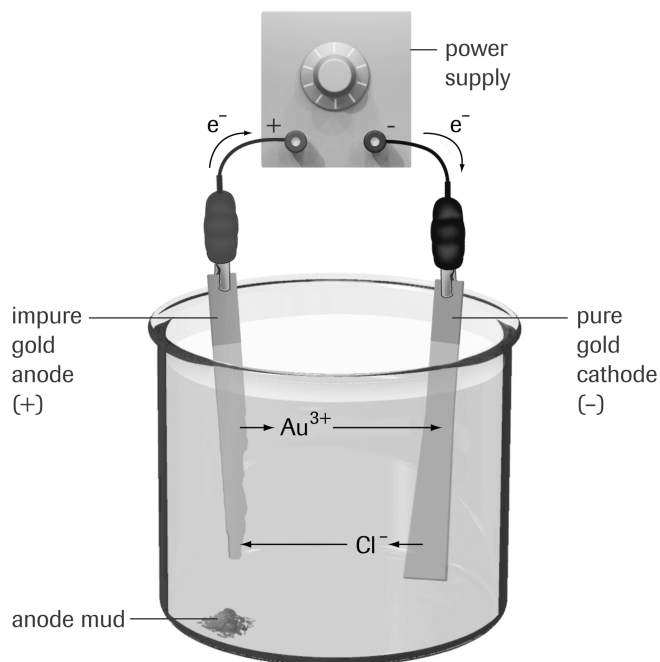
14. (a) anode: $\text{Fe}_{(s)} \rightarrow \text{Fe}_{(aq)}^{2+} + 2 e^{-}$
 cathode: $\frac{1}{2} \text{O}_{2(g)} + \text{H}_2\text{O}_{(l)} + 2 e^{-} \rightarrow 2 \text{OH}_{(aq)}^{-}$
- (b) Since water is a reactant in the reduction of oxygen, the corrosion of iron objects in the desert is extremely slow or nonexistent.
- (c) Water is not only a reactant in the corrosion process, but, like the salt bridge, it provides a medium through which ions can flow.
- (d) Painting is an effective means of rust protection only while the entire surface of the metal is coated. Galvanizing, however, provides protection even if the metal to be protected is exposed to the environment. Zinc acts as a sacrificial anode to less reactive metals (such as steel), preventing the steel from corroding.
15. (a) Since aluminum is higher on the activity series than iron, aluminum should be more reactive than iron.
- (b) The corrosion of iron is a far more serious problem than the corrosion of aluminum. As aluminum corrodes, it surrounds itself with a tightly bound protective coating of aluminum oxide. Conversely, the oxide of iron adheres poorly to the metal beneath it, flaking off readily to expose fresh iron to the environment, leading to further corrosion.
16. (a) Cathodic protection protects a metal against corrosion by supplying the metal with electrons. The inflow of electrons prevents the metal from losing electrons of its own to become oxidized.
- (b) Two forms of cathodic protection are the use of sacrificial anodes and impressed currents. A sacrificial anode is a piece of more reactive metal that is attached to the object to be protected. The object draws electrons from the oxidation of the more reactive metal, so the object is protected from corrosion. Impressed current provides electrons to the object to be protected by means of an external supply of electricity.
17. Differences Between Galvanic and Electrolytic Cells

	Galvanic cell	Electrolytic cell
Cell reaction	spontaneous	nonspontaneous
Energy	produces electrical energy	requires electrical energy
Energy changes	converts chemical energy into electrical energy	converts electrical energy into chemical energy
Electrodes	requires electrodes made of different metals	does not require electrodes made of different metals (i.e., both electrodes can be graphite); potential difference between the electrodes is created by an external source of electrical energy



19. (a) Electrorefining is an electrolytic process used to increase the purity of certain metals.

(b)



20. (a) In order for the reaction to occur as written, magnesium must be more reactive. If not, the reaction would not occur.
 (b) Based only on this chemical equation, titanium should be below magnesium on the activity series. *Note:* this prediction, however, does not take into account other factors that can affect reactivity, such as whether or not titanium forms a protective oxide layer, and the temperature of the reaction.
21. (a) Cu, Fe, Zn, Mg.
 least reactive \longrightarrow most reactive
 (b) $\text{Fe}_{(s)}/\text{Cu}_{(s)}$, $\text{Zn}_{(s)}/\text{Cu}_{(s)}$, $\text{Mg}_{(s)}/\text{Cu}_{(s)}$
 smallest cell potential \longrightarrow largest cell potential
 The magnesium/copper cell should have the largest cell potential since these metals are farthest apart on the activity series. Magnesium will oxidize more readily than any of the other metals. Conversely, the iron/copper cell should have the smallest cell potential since these metals are close together on the activity series.
22. A lower temperature generally decreases the rate of chemical reactions. Therefore, the lower temperature at 1000 m should slow the rate of corrosion. Furthermore, since the concentration of oxygen is less at 1000 m than at 100 m, and oxygen is required for corrosion to occur, the rate of corrosion at a depth of 1000 m should be slower.
23. (a) $\text{Sn}_{(\text{aq})}^{2+}$ was reduced. $\text{Cl}_{(\text{aq})}^{-}$ was oxidized.
 (b) cathode half-reaction: $\text{Sn}_{(\text{aq})}^{2+} + 2\text{e}^{-} \rightarrow \text{Sn}_{(\text{s})}$
 anode half-reaction: $2\text{Cl}_{(\text{aq})}^{-} \rightarrow \text{Cl}_{2(\text{g})} + 2\text{e}^{-}$
 (c) overall reaction: $\text{Sn}_{(\text{aq})}^{2+} + 2\text{Cl}_{(\text{aq})}^{-} \rightarrow \text{Cl}_{2(\text{g})} + \text{Sn}_{(\text{s})}$

Applying Inquiry Skills

24. (a) **Table 1 Predicted Reactions**

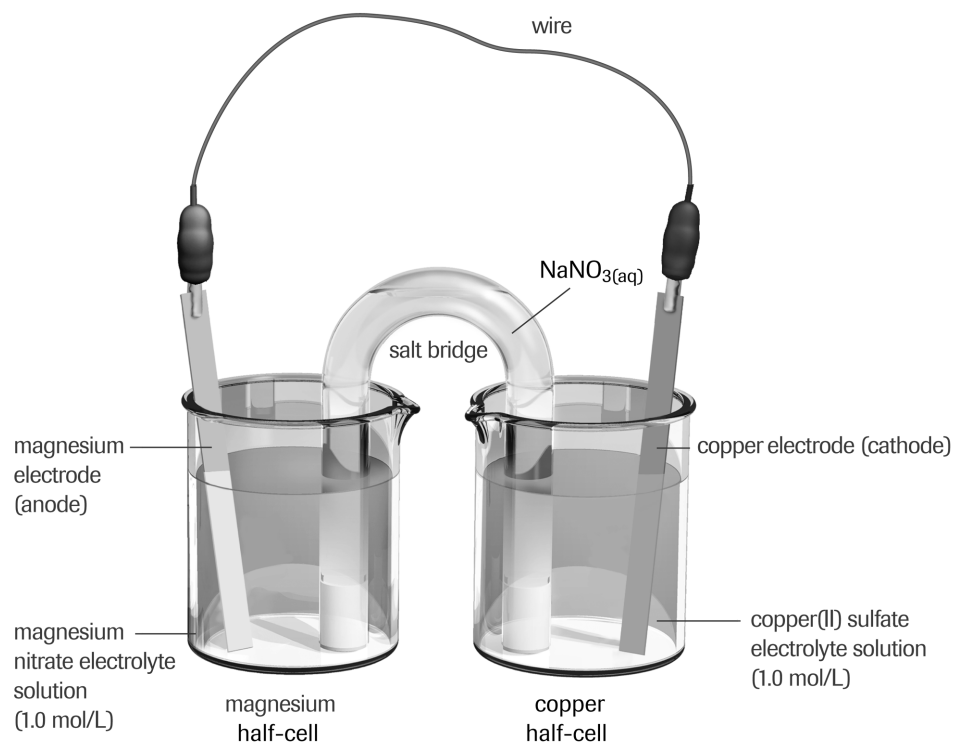
Metal \ Ion	$\text{Ag}_{(\text{aq})}^{+}$	$\text{Cu}_{(\text{aq})}^{2+}$	$\text{Mg}_{(\text{aq})}^{2+}$
$\text{Ag}_{(\text{s})}$	NR	NR	NR
$\text{Cu}_{(\text{s})}$	R	NR	NR
$\text{Mg}_{(\text{s})}$	R	R	NR

R: a spontaneous reaction occurs

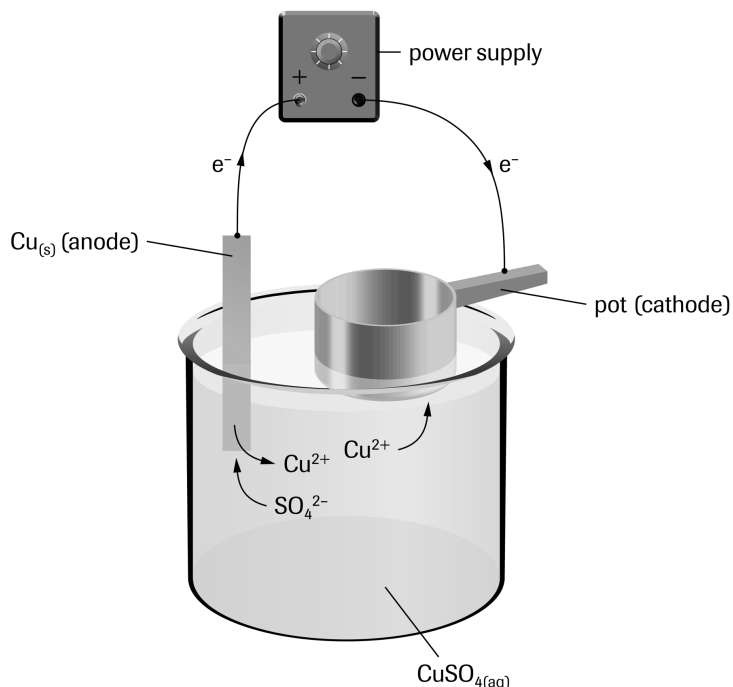
NR: a spontaneous reaction does not occur

- (b)
- $$\text{Cu}_{(\text{s})} + 2\text{Ag}_{(\text{aq})}^{+} \rightarrow 2\text{Ag}_{(\text{s})} + \text{Cu}_{(\text{aq})}^{2+}$$
- $$\text{Mg}_{(\text{s})} + 2\text{Ag}_{(\text{aq})}^{+} \rightarrow 2\text{Ag}_{(\text{s})} + \text{Mg}_{(\text{aq})}^{2+}$$
- $$\text{Mg}_{(\text{s})} + \text{Cu}_{(\text{aq})}^{2+} \rightarrow \text{Cu}_{(\text{s})} + \text{Mg}_{(\text{aq})}^{2+}$$

25. (a) Bubbles of hydrogen gas would be observed forming on the magnesium surface. Over time, the size of the magnesium strip would decrease.
 (b) The rate of the bubbling would decrease because zinc is a less reactive metal than magnesium, as given by the activity series.
 (c) Bubbling would stop since copper is below hydrogen on the activity series.
26. (a)



- (b) anode: Mg cathode: Cu
 (c) Electrons flow through the wire from Mg to Cu. Anions (NO_3^- and SO_4^{2-}) flow through the electrolyte towards Mg. Cations (Mg^{2+} , Cu^{2+} and Na^+) flow towards the copper cathode.
 (d) Laboratory galvanic cells such as this one generate (put out) insufficient energy to power devices, they are large and cumbersome to work with, and they are not portable.
27. $\text{HCl}_{(\text{aq})}$, $\text{NaCl}_{(\text{aq})}$, $\text{NaOH}_{(\text{aq})}$
 most corrosion \longrightarrow least corrosion
28. (a) The use of a sacrificial anode is a form of cathodic protection in which a reactive metal (the sacrificial anode) is attached to a metal object to be protected. The electrons released from the oxidation of the sacrificial anode prevent the metal from oxidizing.
 (b) Wrap one nail tightly with magnesium ribbon. Place the wrapped nail in one beaker containing 20 mL salt water. Place the other nail in the second beaker containing the same volume of salt water.
 (c) Reddish-orange rust should be evident on the unwrapped nail. The magnesium-wrapped nail should remain rust-free.



Any soluble copper-containing compound could theoretically be used as the plating solution. The copper anode replenishes the solution with copper ions. An inert anode, such as graphite, could also be used. However, the amount of copper that can be plated is then limited to that which is present as ions in the solution.

Making Connections

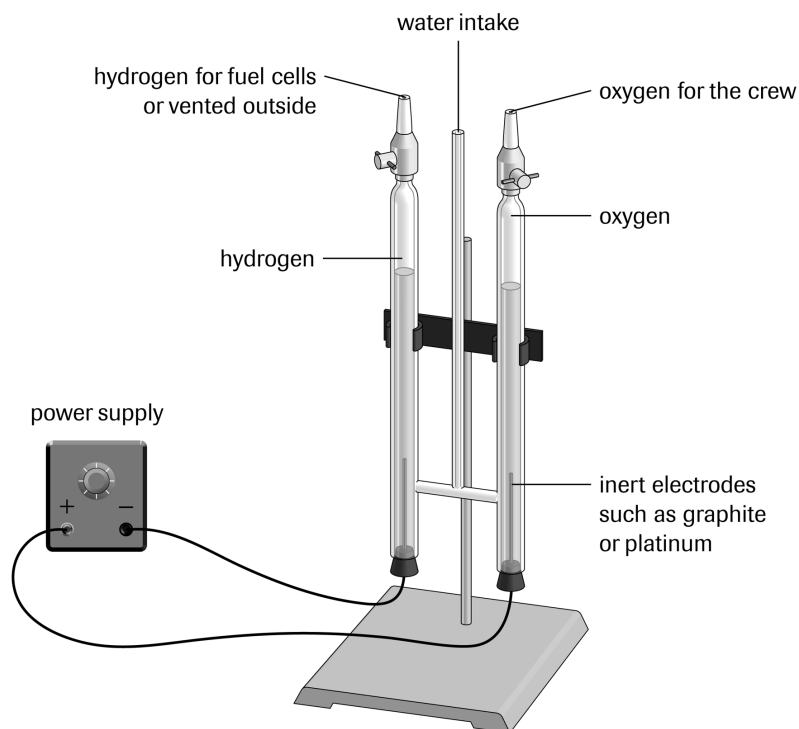
30. (a) Since salt is an electrolyte, it improves the electrical conductivity of water. Much like a salt bridge in a galvanic cell, the presence of salt in the slush that accumulates under our cars provides the ions required to prevent charge build-up at the anode and cathode sites of a corroding metal.
- (b) Salt is potentially harmful to aquatic organisms and may also contaminate ground water. Increased salt levels in ground water can stunt the growth of plants. This has been a particular problem on farms situated near major highways.
- (c) There are several alternatives to the current practice of using large quantities of solid sodium chloride. One alternative is calcium magnesium acetate, which has less ability to promote corrosion and is less harmful to aquatic organisms, but it is relatively expensive. One of the most promising alternative treatments is a brine (salt) solution that is applied before significant icing occurs. This solution uses a lot less salt but requires very good weather forecasting to predict likely trouble areas.
31. (a) Electricity generated from the combustion of fossil fuels releases sulfur dioxide. Once in the atmosphere, sulfur dioxide is further oxidized by oxygen to sulfur trioxide that combines with water vapour to form sulfuric acid. As a result, rain becomes a dilute solution of sulfuric acid. Since metal refining requires a great deal of electricity, it can be argued that metal refining indirectly promotes the production of acid rain.
- (b) Metals corrode faster in acidified rain because acid rain contains more ions than “normal” rain, and hydrogen ions in acid rain can also oxidize steel.
- (c) Combustion of fossil fuels and the production of the carbon anodes used in aluminum electrolysis produce carbon dioxide, a greenhouse gas. Aluminum production also emits polyaromatic hydrocarbons, or PAHs, into the environment. Some of these compounds have been shown to be dangerous to wildlife and carcinogenic to humans.
- (d) Possible public actions include recycling as much used metal as possible, particularly aluminum, reducing our electricity consumption, particularly in peak demand periods such as in the summer, and using political pressure to encourage further research into decreasing industrial sulfur and PAH emissions.
32. (a) reactant oxidized: $\text{Mg}_{(s)}$ reactant reduced: $\text{C (in CO}_{2(g)})$
- (b) Magnesium produces a very bright flame as it burns. As the reaction subsides, a black deposit of carbon is evident.
- (c) Carbon dioxide supports the combustion of magnesium.
- (d) Class D fires are fires that involve combustible metals such as magnesium, titanium, sodium, and potassium. Fire extinguishers specifically designed for class D fires spray copper or sodium chloride onto the fire. Copper

is particularly useful when dealing with lithium fires. Sodium chloride is effective at extinguishing sodium, potassium, and magnesium fires. The heat from the fire causes sodium chloride to form a hard crust around the burning metal, isolating the metal from oxygen. The crust also helps to dissipate heat generated in the fire. Dry sand can also be used to smother fires that involve small amounts of combustible materials.

Extension

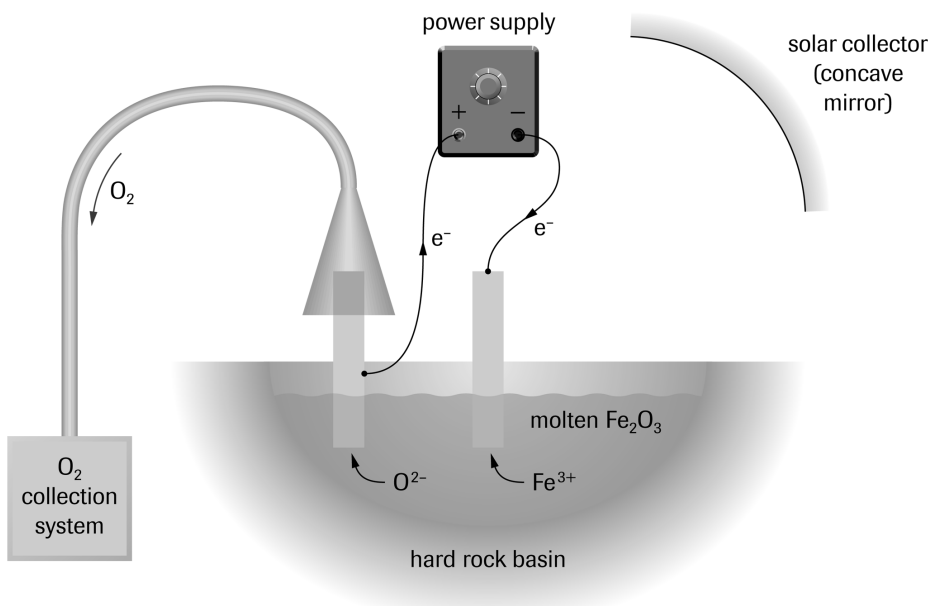
33. (a) Metals such as nickel, copper, platinum, silver, and zinc are alloyed with gold. Grades of gold include:
- 18 K gold (75% gold): 18 parts gold and 6 parts other metals
 - 14 K gold (58% gold): 14 parts gold and 10 parts other metals
 - 12 K gold (50% gold): 12 parts gold and 12 parts other metals
 - 10 K gold (42% gold): 10 parts gold and 14 parts other metals
- (b) 24 K (pure) gold is too soft for jewellery. Adding other metals to gold produces an alloy with added strength and hardness, as well as, in some cases, a desired colour.
- (c) To reduce the risk of nickel allergies, grades of gold below 18 K should be avoided, as should white gold. These types of gold jewellery are likely to contain significant proportions of nickel.

34. (a)



- (b) Electricity generated using the nuclear reactor on the submarine could be used to power the electrolysis process.
- (c) Both oxygen and hydrogen are highly explosive, so great care must be taken to keep the apparatus intact and away from sparks.

35.



Martian soil could be loaded into a reactor container that is heated using large solar collectors. The intense heat of the Sun can be used to raise the temperature of the soil to beyond the melting point of iron(III) oxide. Inert electrodes are then inserted into this molten mixture. Electrical energy generated using solar cells can be used to create a potential difference between the electrodes, and electrolyze the molten mixture. It is essential that the electrodes used not melt or react with the contents of the reactor. Oxygen gas released at the anode is collected, purified, and stored in large pressurized containers.

Another possible source of oxygen is water that has been discovered underground on Mars. Using solar energy, water could be electrolyzed into hydrogen and oxygen gas.