

Chapter 16 Redox chemistry and corrosion

Q1.

Identify each of the following half equations as involving either oxidation or reduction:

- a $\text{Na(s)} \rightarrow \text{Na}^+(\text{aq}) + \text{e}^-$
- b $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$
- c $\text{S(s)} + 2\text{e}^- \rightarrow \text{S}^{2-}(\text{aq})$
- d $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

A1.

- a oxidation
- b reduction
- c reduction
- d oxidation

Q2.

Balance the following half equations and then identify each as representing either an oxidation or a reduction reaction:

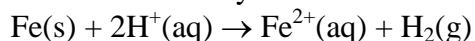
- a $\text{Fe(s)} \rightarrow \text{Fe}^{3+}(\text{aq})$
- b $\text{K(s)} \rightarrow \text{K}^+(\text{aq})$
- c $\text{F}_2(\text{g}) \rightarrow \text{F}^-(\text{aq})$
- d $\text{O}_2(\text{g}) \rightarrow \text{O}^{2-}(\text{aq})$

A2.

- a $\text{Fe(s)} \rightarrow \text{Fe}^{3+}(\text{aq}) + 3\text{e}^-$ oxidation
- b $\text{K(s)} \rightarrow \text{K}^+(\text{aq}) + \text{e}^-$ oxidation
- c $\text{F}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{F}^-(\text{aq})$ reduction
- d $\text{O}_2(\text{g}) + 4\text{e}^- \rightarrow 2\text{O}^{2-}(\text{aq})$ reduction

Q3.

Iron reacts with hydrochloric acid according to the ionic equation:



- a What has been oxidised in this reaction? What is the product?
- b Write a half equation for the oxidation reaction.
- c Identify the oxidant.
- d What has been reduced in this reaction? What is the product?
- e Write a half equation for the reduction reaction.
- f Identify the reductant.

A3.

- a Fe(s) has been oxidised to $\text{Fe}^{2+}(\text{aq})$
- b $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$
- c $\text{H}^+(\text{aq})$
- d $\text{H}^+(\text{aq})$ has been reduced to H_2
- e $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$
- f Fe(s)

Q4.

When a strip of magnesium metal is placed in a blue solution containing copper(II) ions ($\text{Cu}^{2+}(\text{aq})$), crystals of copper appear and the solution soon becomes paler in colour.

- a Show that this reaction is a redox reaction by identifying the substance that is oxidised and the one that is reduced.
- b Write a half equation for the oxidation reaction.
- c Write a half equation for the reduction reaction.
- d Write an overall redox equation.
- e Identify the oxidant and the reductant.
- f Explain why the solution loses some of its colour as a result of the reaction.

A4.

- a Magnesium is oxidised, copper ions are reduced.
- b $\text{Mg}(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^{-}$
- c $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu}(\text{s})$
- d $\text{Mg}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Cu}(\text{s})$
- e oxidant Cu^{2+} ; reductant Mg
- f The solution loses some of its blue colour due to the loss of $\text{Cu}^{2+}(\text{aq})$, which react to form $\text{Cu}(\text{s})$.

Q5.

Some ions, such as the Cu^{2+} ion, can be either oxidised or reduced.

- a Write the formula for the product of the oxidation of the Cu^{2+} ion.
- b Write the formula for the product of the reduction of the Cu^{2+} ion.
- c Name one other ion that can be either oxidised or reduced.

A5.

- a Cu^{2+}
- b Cu
- c Fe^{2+}

Q6.

Calcium that is exposed to the air forms an oxide coating.

- a What is the formula of calcium oxide?
- b What has been oxidised in this reaction?
- c Write a half equation for the oxidation reaction.
- d What has been reduced in this reaction?
- e Write a balanced half equation for the reduction reaction.
- f Write an overall equation for this redox reaction.
- g Copy the following statement and fill in the blank spaces with the appropriate words:

Calcium has been _____ by _____ to calcium ions.
The _____ has gained electrons from the _____. The
oxygen has been _____ by _____ to oxide ions.
The _____ has lost electrons to the _____.

A6.

- a CaO
- b $\text{Ca}(\text{s})$

- c $\text{Ca(s)} \rightarrow \text{Ca}^{2+}(\text{s}) + 2\text{e}^{-}$
d $\text{O}_2(\text{g})$
e $\text{O}_2(\text{g}) + 4\text{e}^{-} \rightarrow 2\text{O}^{2-}(\text{s})$
f $2\text{Ca(s)} + \text{O}_2(\text{g}) \rightarrow 2\text{CaO(s)}$
g oxidised, oxygen, oxygen, calcium, reduced, calcium, calcium, oxygen

Q7.

Assign an oxidation number to the underlined element in each the following molecules or ions:

- a $\underline{\text{Ni}}\text{O}_2$
b $\underline{\text{C}}\text{O}$
c $\underline{\text{C}}\text{O}_2$
d $\underline{\text{Br}}_2$
e $\underline{\text{N}}_2\text{H}_4$
f $\text{H}_2\underline{\text{S}}\text{O}_4$
g $\underline{\text{N}}\text{O}_3^{-}$
h $\underline{\text{C}}\text{H}_4$
i $\underline{\text{O}}_2$
j $\underline{\text{P}}\text{O}_4^{3-}$
k $\text{H}_2\underline{\text{S}}$
l $\underline{\text{Cr}}_2\text{O}_3$

A7.

- a Ni +4
b C +2
c C +4
d Br 0
e N -2
f S +6
g N +5
h C -4
i O 0
j P +5
k S -2
l Cr +3

Q8.

In each of the following redox reactions, use oxidation numbers to identify the element that has been oxidised and the one that has been reduced.

- a $\text{Pb}^{2+}(\text{aq}) + \text{Mg(s)} \rightarrow \text{Pb(s)} + \text{Mg}^{2+}(\text{aq})$
b $2\text{HgO(s)} \rightarrow 2\text{Hg(l)} + \text{O}_2(\text{g})$
c $2\text{H}_2\text{O(l)} + 2\text{F}_2(\text{g}) \rightarrow 4\text{HF(aq)} + \text{O}_2(\text{g})$
d $\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)} + 5\text{Fe}^{3+}(\text{aq})$

A8.

- a lead reduced (+2 to 0), magnesium oxidised (0 to +2)
b mercury reduced (+2 to 0), oxygen oxidised (-2 to 0)
c oxygen oxidised (-2 to 0), fluorine reduced (0 to -1)
d manganese reduced (+7 to +2), iron oxidised (+2 to +3)

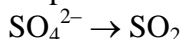
Q9.

Write half equations to represent:

- a the reduction of SO_4^{2-} to SO_2
- b the oxidation of H_2O_2 to O_2
- c the oxidation of H_2S to S
- d the reduction of MnO_4^- to MnO_2
- e the reduction of Ta_2O_5 to Ta
- f the oxidation of SO_3^{2-} to SO_4^{2-}
- g the reduction of IO_3^- to I^-

A9.

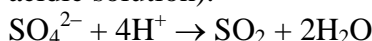
- a Step 1: Balance all atoms in the half equation except oxygen.



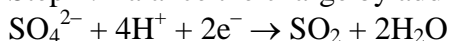
Step 2: Balance the oxygen atoms by adding water molecules.



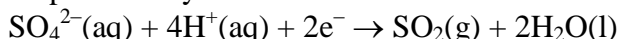
Step 3: Balance the hydrogen atoms by adding H^+ ions (which are present in acidic solution).



Step 4: Balance the charge by adding electrons to the more positive side.



Step 5: Add symbols of state.



Use the same method to balance the other half equations.

- b $\text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- (\text{aq})$
- c $\text{H}_2\text{S}(\text{g}) \rightarrow \text{S}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- (\text{aq})$
- d $\text{MnO}_4^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
- e $\text{Ta}_2\text{O}_5(\text{s}) + 10\text{H}^+(\text{aq}) + 10\text{e}^- \rightarrow 2\text{Ta}(\text{s}) + 5\text{H}_2\text{O}(\text{l})$
- f $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$
- g $\text{IO}_3^-(\text{aq}) + 6\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow \text{I}^-(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$

Q10.

Balance the following redox equations by separating them into two half equations, balancing each equation, and then combining the pair into a balanced complete redox equation.

- a $\text{H}_2\text{O}_2(\text{aq}) + \text{PbS}(\text{s}) \rightarrow \text{PbSO}_4(\text{s}) + \text{H}_2\text{O}(\text{l})$
- b $\text{I}_2(\text{aq}) + \text{H}_2\text{S}(\text{g}) \rightarrow \text{I}^-(\text{aq}) + \text{S}(\text{s})$
- c $\text{SO}_3^{2-}(\text{aq}) + \text{MnO}_4^-(\text{aq}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + \text{Mn}^{2+}(\text{aq})$
- d $\text{NO}(\text{g}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \rightarrow \text{NO}_3^-(\text{aq}) + \text{Cr}^{3+}(\text{aq})$
- e $\text{Zn}(\text{s}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cr}^{3+}(\text{aq})$
- f $\text{CuO}(\text{s}) + \text{NH}_3(\text{aq}) \rightarrow \text{Cu}(\text{s}) + \text{NO}(\text{g})$

A10.

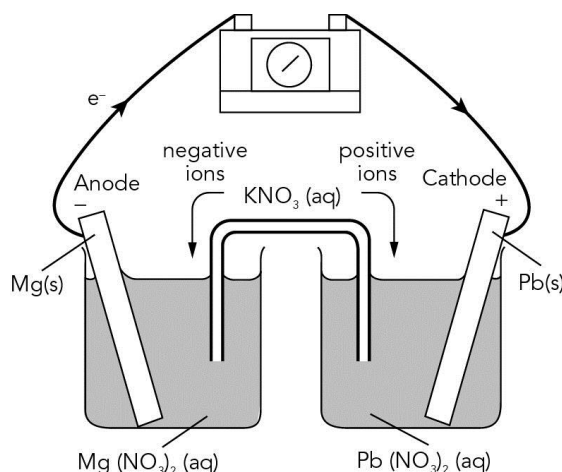
- a $\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$ $\times 4$
 $\text{PbS}(\text{s}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow \text{PbSO}_4(\text{s}) + 8\text{H}^+(\text{aq}) + 8\text{e}^-$ $\times 1$
 $4\text{H}_2\text{O}_2(\text{aq}) + \text{PbS}(\text{s}) \rightarrow \text{PbSO}_4(\text{s}) + 4\text{H}_2\text{O}(\text{l})$
- b $\text{I}_2(\text{aq}) + 2\text{e}^- \rightarrow 2\text{I}^-(\text{aq})$
 $\text{H}_2\text{S}(\text{g}) \rightarrow \text{S}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$
 $\text{I}_2(\text{aq}) + \text{H}_2\text{S}(\text{g}) \rightarrow 2\text{I}^-(\text{aq}) + \text{S}(\text{s}) + 2\text{H}^+(\text{aq})$

- c** $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$ $\times 5$
 $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$ $\times 2$
 $5\text{SO}_3^{2-}(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow 5\text{SO}_4^{2-}(\text{aq}) + 2\text{Mn}^{2+}(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
- d** $\text{NO}(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^-$ $\times 2$
 $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
 $2\text{NO}(\text{g}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow 2\text{NO}_3^-(\text{aq}) + 2\text{Cr}^{3+}(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
- e** $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ $\times 3$
 $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
 $3\text{Zn}(\text{s}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) \rightarrow 3\text{Zn}^{2+}(\text{aq}) + 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
- f** $\text{CuO}(\text{s}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s}) + \text{H}_2\text{O}(\text{l})$ $\times 5$
 $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NO}(\text{g}) + 5\text{H}^+(\text{aq}) + 5\text{e}^-$ $\times 2$
 $5\text{CuO}(\text{s}) + 2\text{NH}_3(\text{aq}) \rightarrow 5\text{Cu}(\text{s}) + 2\text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{l})$

Q11.

Two half cells are set up. One contains a solution of magnesium nitrate with a strip of magnesium as the electrode. The other contains lead nitrate with a strip of lead as the electrode. The solutions in the two half cells are connected by a piece of filter paper soaked in potassium nitrate solution. When the electrodes are connected by wires to a galvanometer, the magnesium electrode is shown to be negatively charged.

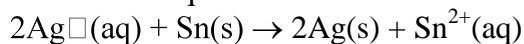
- a** Sketch the galvanic cell described. Label the positive and negative electrodes. Mark the direction of the electron flow.
- b** Write the half equations for the reactions that occur in each half cell and an equation for the overall reaction.
- c** Label the anode and cathode.
- d** Indicate the direction in which ions in the salt bridge migrate.

A11.**a**

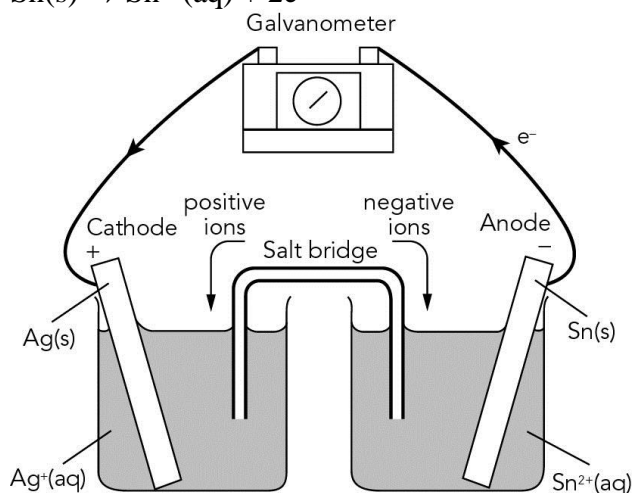
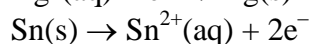
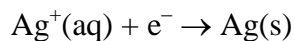
- b** $\text{Mg}(\text{s}) \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^-$
 $\text{Pb}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Pb}(\text{s})$
 Overall: $\text{Mg}(\text{s}) + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Pb}(\text{s})$
- c** lead electrode is the cathode; magnesium electrode is the anode
- d** Anions will migrate to the $\text{Mg}^{2+}(\text{aq})/\text{Mg}(\text{s})$ half cell, cations to the $\text{Pb}^{2+}(\text{aq})/\text{Pb}(\text{s})$ half cell.

Q12.

The overall equation for the redox reaction between silver ions and tin metal is:



Sketch a suitable galvanic cell to demonstrate that there is a flow of electrons between the reactants. Fully label the cell. Write half equations for the cell reactions. Show the direction of electron flow.

A12.**Q13.**

Refer to the electrochemical series (Table 16.2, page 287) and predict if the following reactions will occur spontaneously.

- Silver metal is placed in a copper nitrate solution.
- A strip of aluminium is placed in a sodium chloride solution.
- Magnesium is added to a solution of iron(II) nitrate.
- The element zinc is placed in a tin(II) solution.
- A piece of tin is placed in a silver nitrate solution.
- Lead nitrate solution is poured into a beaker containing zinc granules.
- Gold foil is added to a lead nitrate solution.

A13.

For reactions to occur spontaneously, the aqueous cation in the solution must be a stronger oxidant than the cation of the metal added.

- No
- No
- Yes
- Yes
- Yes
- Yes
- No

Q14.

Solutions of zinc nitrate, tin nitrate and copper(II) nitrate have been prepared in a laboratory, but have inadvertently been left unlabelled. Name two metals that could be used to identify each solution.

A14.

Use iron and lead to test the solutions.

The copper(II) nitrate will react with both iron and lead.

The tin nitrate will react only with the iron.

The zinc nitrate will not react with either of the metals.

Q15.

Each of these pairs of half cells combines to form a galvanic cell.

i $\text{Ag}^+(\text{aq})/\text{Ag}(\text{s})$ and $\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})$

ii $\text{Fe}^{2+}(\text{aq})/\text{Fe}(\text{s})$ and $\text{Pb}^{2+}(\text{aq})/\text{Pb}(\text{s})$

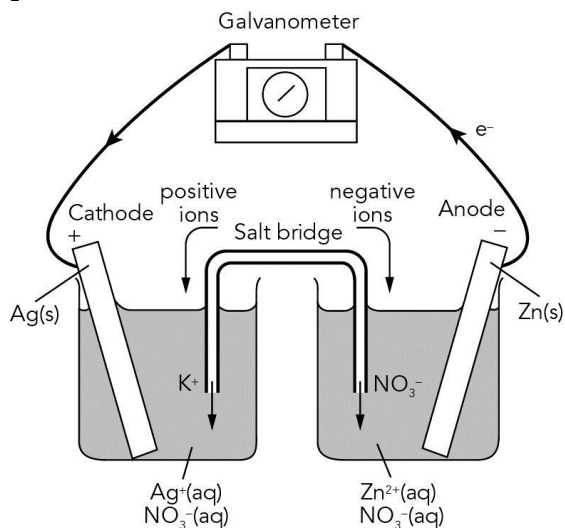
iii $\text{Ni}^{2+}(\text{aq})/\text{Ni}(\text{s})$ and $\text{Cu}^{2+}(\text{aq})/\text{Cu}(\text{s})$

Draw a diagram of each galvanic cell and on your diagram show:

- the direction of electron flow in the external circuit
- a half equation for the reaction at each electrode
- a full equation for the overall reaction in the galvanic cell
- which electrode is the anode
- which electrode is positive
- which way negative ions flow in the salt bridge

A15.

i

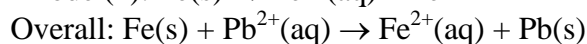
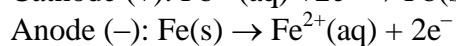
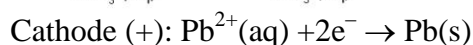
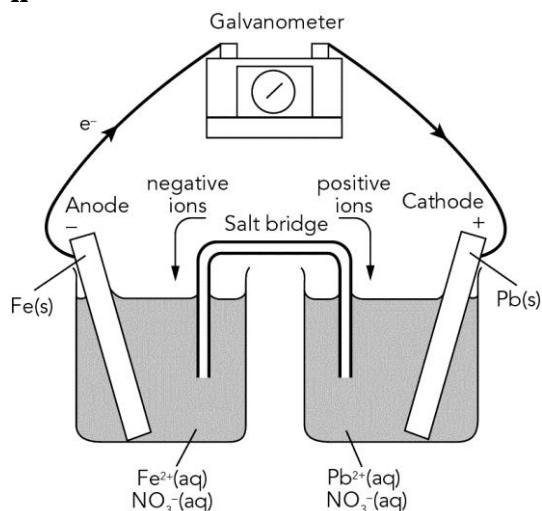


Cathode (+): $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$

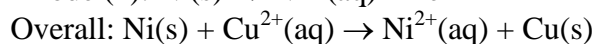
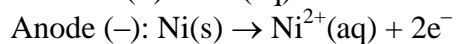
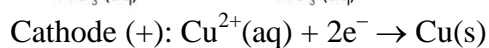
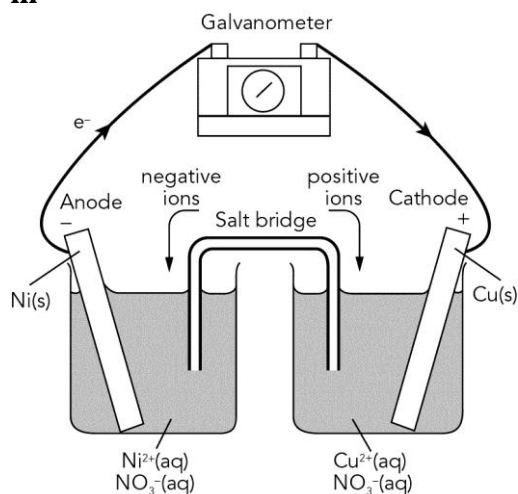
Anode (-): $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$

Overall: $\text{Zn}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$

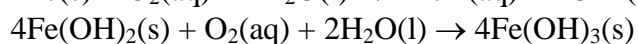
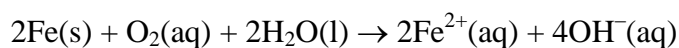
ii



iii

**Q16.**

Use equations to explain why the corrosion of iron is an electrochemical process.

A16.**Q17.**

Explain why iron corrosion occurs more rapidly near coastal environments.

A17.

Near the coast, moisture in the air contains a higher amount of dissolved salts (e.g. $\text{NaCl}(\text{aq})$). Salts are good electrolytes and accelerate the rusting process. They do this

because they facilitate the production of ions at the oxidation and reduction sites of a corroding metal.

Q18.

- a Which of the following metals could be used as a sacrificial anode in order to prevent the corrosion of iron?
magnesium, zinc, lead, copper
- b Explain your answer in terms of the electrochemical series of metals.

A18.

- a magnesium and zinc
- b Only metals more reactive (i.e. more easily oxidised) than iron will act as sacrificial anodes. These more reactive metals will be oxidised in preference to the iron. The iron will become the cathode (the site of reduction) and no oxidation of the iron will take place.

Q19.

Underwater steel pillars often corrode more rapidly just beneath the surface of the water than above. Suggest a reason for this.

A19.

Reduction of oxygen occurs when there is plenty of moisture and the oxygen concentration is high—that is, at or just above the surface of the water. A cathodic region forms here. The anodic region, then, is close to the cathode and where the oxygen concentration is less—that is, just below the surface. Oxidation, or corrosion, therefore occurs more rapidly below the surface.

Q20.

Explain why it is common for manufacturers to coat steel nails with zinc rather than copper or tin.

A20.

As zinc is more reactive than steel, it will be the anode and steel will be the cathode. Oxidation, and so corrosion, of the nails will thus be prevented.

Chapter review

Q21.

Define oxidation and reduction in terms of:

- a the transfer of oxygen
- b the transfer of electrons
- c oxidation numbers

A21.

- a Oxidation is the gain of oxygen atoms; reduction is the loss of oxygen atoms.
- b Oxidation is the loss of electrons; reduction is the gain of electrons.
- c Oxidation involves an increase in oxidation number; reduction a decrease in oxidation number

Q22.

Identify the following half equations as involving either oxidation or reduction:

- a $\text{Mg(s)} \rightarrow \text{Mg}^{2+}(\text{aq}) + 2\text{e}^{-}$
- b $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-}$
- c $\text{O}_2(\text{aq}) + 2\text{H}_2\text{O(l)} + 4\text{e}^{-} \rightarrow 4\text{OH}^{-}(\text{aq})$
- d $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{Cu(s)}$

A22.

- a oxidation
- b oxidation
- c reduction
- d reduction

Q23.

In these redox reactions, which species have been oxidised and which reduced?

- a $\text{Cu}^{2+}(\text{aq}) + \text{Mg(s)} \rightarrow \text{Cu(s)} + \text{Mg}^{2+}(\text{aq})$
- b $\text{Zn(s)} + 2\text{Ag}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag(s)}$
- c $\text{PbO(s)} + \text{CO(g)} \rightarrow \text{Pb(s)} + \text{CO}_2(\text{g})$
- d $\text{Ca(s)} + \text{HCl(aq)} \rightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2(\text{g})$
- e $\text{Pb}^{2+}(\text{aq}) + \text{Fe(s)} \rightarrow \text{Pb(s)} + \text{Fe}^{2+}(\text{aq})$

A23.

- a Mg(s) oxidised, $\text{Cu}^{2+}(\text{aq})$ reduced
- b Zn(s) oxidised, $\text{Ag}^{+}(\text{aq})$ reduced
- c CO(g) oxidised, PbO(s) reduced
- d Ca(s) oxidised, $\text{H}^{+}(\text{aq})$ reduced
- e Fe(s) oxidised, $\text{Pb}^{2+}(\text{aq})$ reduced

Q24.

Would elements from the following groups be more likely to be oxidised or reduced when they take part in chemical reactions with the other group? Give reasons for your answers.

Group A: sodium, calcium and magnesium.

Group B: sulfur, oxygen and chlorine.

A24.

Group A will be oxidised because metals tend to lose electrons to non-metals when they react. Group B will be reduced because non-metals tend to gain electrons when they react.

Q25.

a State which of the reactions below are redox reactions:

- i $\text{CO}_2(\text{g}) + \text{C(s)} \rightarrow 2\text{CO(g)}$
- ii $\text{H}_2\text{S(g)} + \text{Cl}_2(\text{g}) \rightarrow 2\text{HCl(g)} + \text{S(s)}$
- iii $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O(l)} + \text{O}_2(\text{g})$
- iv $\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O(l)} + 5\text{Fe}^{3+}(\text{aq})$
- v $\text{Cd(s)} + \text{NiO}_2(\text{s}) + 2\text{H}_2\text{O(l)} \rightarrow \text{Cd(OH)}_2(\text{s}) + \text{Ni(OH)}_2(\text{s})$
- vi $2\text{Fe(OH)}_2(\text{aq}) + \text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{Fe(OH)}_3(\text{aq})$

- b** If the above are redox reactions, state:
- i** which species has been oxidised
 - ii** which species has been reduced

A25.

- a** All are redox reactions
- b**
- i** C oxidised, CO_2 reduced
 - ii** Cl_2 reduced, H_2S oxidised
 - iii** H_2O_2 oxidised and reduced
 - iv** Fe^{2+} oxidised, MnO_4^- reduced
 - v** Cd oxidised, NiO_2 reduced
 - vi** $\text{Fe}(\text{OH})_2$ oxidised, H_2O_2 reduced

Q26.

Assign oxidation numbers to the underlined species in each of the following:

- a** V
- b** V^{2+}
- c** V_2O_5
- d** MnO_4
- e** Mn^{2+}
- f** $\text{Cr}_2\text{O}_7^{2-}$
- g** Cr^{3+}
- h** F_2O
- i** CrO_4^-
- j** O_2

A26.

- a** 0
- b** +2
- c** +5
- d** +7
- e** +2 +6
- f** +6
- g** +3
- h** +2
- i** +7
- j** 0

Q27.

Assign oxidation states for nitrogen in each of the following:

- a** HNO_3
- b** NO_3^-
- c** NO_2
- d** N_2O_4
- e** HNO_2
- f** NO
- g** N_2O
- h** N_2
- i** N_2H_4
- j** NH_3

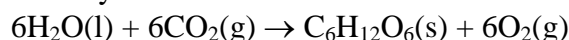
- k** NH_4^+
l LiN_3

A27.

- a** +5
b +5
c +4
d +4
e +3
f +2
g +1
h 0
i -2
j -3
k -3
l -3

Q28.

Photosynthesis is a well-known redox reaction:



Which element is oxidised and which is reduced?

A28.

Carbon in carbon dioxide is reduced, oxygen in water is oxidised.

Q29.

When zinc reacts with a solution of hydrochloric acid, hydrogen gas is produced.

- a** What has been oxidised in this reaction?
b Write a half equation for the oxidation reaction.
c What has been reduced in this reaction?
d Write a half equation for the reduction reaction.
e Identify the oxidant.
f Identify the reductant.
g Write a balanced full equation for this reaction.

A29.

- a** zinc
b $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$
c $\text{H}^+(\text{aq})$
d $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$
e $\text{Zn}(\text{s})$
f $\text{H}^+(\text{aq})$
g $\text{Zn}(\text{s}) + 2\text{HCl}(\text{aq}) \rightarrow \text{ZnCl}_2(\text{aq}) + \text{H}_2(\text{g})$

Q30.

Refer to the electrochemical series (Table 16.2, page 287) and predict whether the following mixtures would result in spontaneous reactions. If so, write overall ionic equations.

- a** Zinc metal is added to a solution of silver nitrate.
b Copper metal is placed in an aluminium chloride solution.
c Tin sulfate is placed in a copper container.

- d** Magnesium metal is added to a solution of lead nitrate.
- e** Silver metal is added to nickel chloride solution.
- f** Solutions of potassium chloride and copper chloride are mixed.
- g** Iodine is mixed with aluminium.
- h** Chlorine gas is mixed with magnesium.
- i** Potassium nitrate solution is added to silver metal.
- j** Elemental lead is placed in a solution of silver nitrate.

A30.

- a** $\text{Zn(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag(s)}$
- b** no reaction
- c** no reaction
- d** $\text{Mg(s)} + \text{Pb}^{2+}(\text{aq}) \rightarrow \text{Mg}^{2+}(\text{aq}) + \text{Pb(s)}$
- e** no reaction
- f** no reaction
- g** $\text{Al(s)} + \text{I}_2(\text{s}) \rightarrow \text{AlI}_3(\text{s})$
- h** $\text{Mg(s)} + \text{Cl}_2(\text{g}) \rightarrow \text{MgCl}_2(\text{s})$
- i** no reaction
- j** $\text{Pb(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2\text{Ag(s)}$

Q31.

You are given three colourless solutions (A, B and C) known to be sodium nitrate, silver nitrate and lead nitrate but not necessarily in this order. You also have some pieces of magnesium ribbon and copper wire. Describe how you could identify each of the solutions using only the chemicals supplied.

A31.

Place 10 mL of each solution into a clean test-tube. Add a small piece of magnesium ribbon to each solution. The magnesium should be coated by displaced metal in the silver nitrate and lead nitrate solutions, but not in the sodium nitrate. The lead coating will be black. Over time, the silver coating will change from black to silver as more metal is deposited. To confirm the identity of the silver nitrate solution, take a fresh sample and add a small piece of copper. Copper will displace silver from the solution, giving silver deposit and a blue solution. Copper will not displace lead from the solution.

Q32.

An unknown metal is placed in solutions of aluminium nitrate and iron sulfate. After a period of time, the metal is found to have reacted with the iron solution, but not the aluminium solution. Suggest the name of the unknown metal.

A32.

zinc

Q33.

If a piece of zinc metal is added to the following solutions, will a reaction be expected? If so, write the equation for the reaction.

- a** AlCl_3
- b** AgNO_3
- c** SnCl_2
- d** CuSO_4

A33.

- a no
- b $\text{Zn(s)} + \text{AgNO}_3\text{(aq)} \rightarrow \text{Zn(NO}_3)_2\text{(aq)} + \text{Ag(s)}$
- c $\text{Zn(s)} + \text{SnCl}_2\text{(aq)} \rightarrow \text{ZnCl}_2\text{(aq)} + \text{Sn(s)}$
- d $\text{Zn(s)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{ZnSO}_4\text{(aq)} + \text{Cu(s)}$

Q34.

List the following in order of increasing reducing strength:

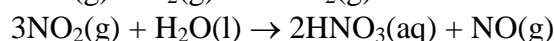
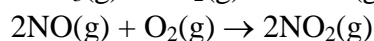
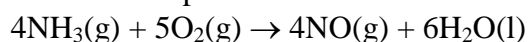
Mg, I^- , Ag, F^- , Ni, Sn, Li, Cu

A34.

F^- , Ag, I^- , Cu, Sn, Ni, Mg, Li

Q35.

The Ostwald process for the manufacture of nitric acid requires three stages:



- a Which of the stages involves redox reactions?
- b Which elements are undergoing oxidation and which are reduced?
- c Identify the oxidants and reductants.

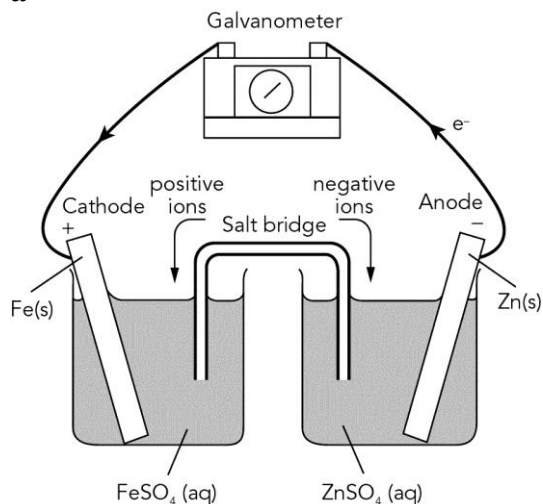
A35.

- a all
- b Nitrogen in NH_3 is oxidised in stage 1. Nitrogen in NO is oxidised in stage 2. Some nitrogen in NO_2 is oxidised and some is reduced in stage 3. Oxygen is reduced in stages 1 and 2.
- c Oxidants: oxygen and nitrogen. Reductant: nitrogen.

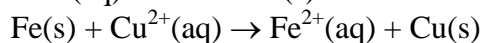
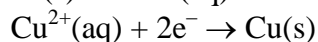
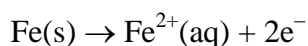
Q36.

A galvanic cell is constructed from an iron nail placed in a beaker containing a solution of iron(II) sulfate connected to another beaker containing a piece of zinc in a solution of zinc sulfate. A salt bridge connects the two half cells.

- a Sketch this cell.
- b Use the electrochemical series of metals to predict the direction of electron flow.
- c At which electrode will oxidation be occurring?
- d Write a half equation for the oxidation reaction.
- e At which electrode will reduction be occurring?
- f Write a half equation for the reduction reaction.
- g Label the anode and cathode.
- h Why is a salt bridge needed for this cell to work?

A36.**a****b** from the zinc to the iron**c** anode (zinc)**d** $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ **e** cathode (iron)**f** $\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Fe(s)}$ **h** A salt bridge is needed to complete the circuit and to maintain electrical neutrality in the half cells.**Q37.**

An iron nail becomes coated when placed in a solution of copper(II) sulfate. Use the electrochemical series to write a full equation for the oxidation/reduction reaction involved.

A37.**Q38.**

A galvanic cell is set up by combining a half cell containing a zinc rod in a solution of zinc nitrate with a half cell containing metal X in a solution of X nitrate. The zinc electrode is found to be the negative electrode. Would you expect metal X to be more or less reactive than zinc? Explain your reasoning.

A38.

Since the zinc electrode is the negative electrode, electrons must be released there. Therefore, zinc metal must be oxidised. Metal X must be less reactive than zinc.

Q39.

Two galvanic cells are constructed from different half cells.

Cell A: strip of nickel in a nickel chloride solution connected to a copper rod in a copper sulfate solution.

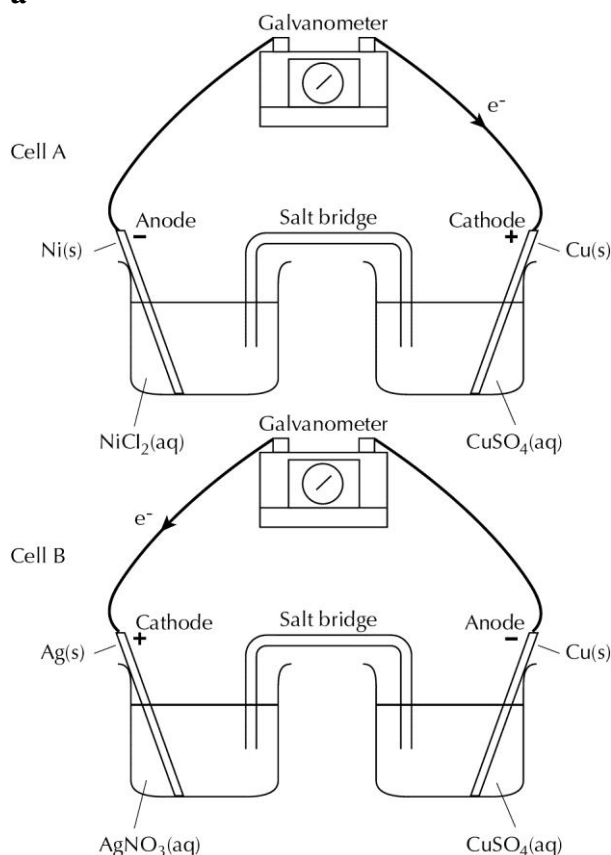
Cell B: strip of silver in a silver nitrate solution connected to a copper rod in a copper sulfate solution.

When a salt bridge is added to complete the circuit, it is found that the copper electrode is the positive electrode in cell A but is the negative electrode in cell B.

- Sketch and fully label each cell.
- Write half equations and a full equation for each cell reaction.
- Explain why the polarity of the copper electrode differs in the two cells.

A39.

a

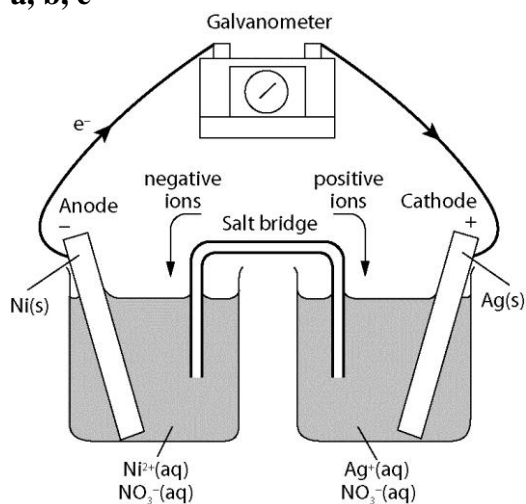


- Cell A: $\text{Ni(s)} \rightarrow \text{Ni}^{2+}(\text{aq}) + 2\text{e}^-$; $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$
Overall: $\text{Ni(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Ni}^{2+}(\text{aq}) + \text{Cu(s)}$
Cell B: $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag(s)}$; $\text{Cu(s)} \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$
Overall: $\text{Cu(s)} + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{Ag(s)}$
- The polarity of the copper electrode depends on its reactivity relative to the reactivity of the other metal to which it is connected. Copper is a weaker reductant than nickel, so the copper electrode is positive in cell A. However, copper is a stronger reductant than silver, so copper is the negative electrode in cell B.

Q40.

Draw a diagram of a galvanic cell with electrodes of silver and nickel, and solutions of their soluble ions.

- Label the anode and cathode.
- Identify the positive and negative electrodes.
- Write the equation for the oxidation and reduction reactions.
- Write the overall equation.
- Draw an arrow to indicate the direction of electron flow.
- Which electrode will lose mass?

A40.**a, b, e****c** Reduction: Cathode (+): $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$ Oxidation: Anode (-): $\text{Ni}(\text{s}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2\text{e}^-$ **d** Overall: $\text{Ni}(\text{s}) + 2\text{Ag}^+(\text{aq}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$ **f** the Ni electrode**Q41.**

Balance these half equations:

a $\text{Ag}^+(\text{aq}) \rightarrow \text{Ag}(\text{s})$ **b** $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq})$ **c** $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq})$ **d** $\text{H}^+(\text{aq}) \rightarrow \text{H}_2(\text{g})$ **A41.****a** $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$ **b** $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ **c** $\text{Zn}(\text{s}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ **d** $2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2(\text{g})$ **Q42.**

Write half equations for the following reactions:

a the reduction of NO_3^- to HNO_2 **b** the reduction of VO^{2+} to V^{3+} **c** the oxidation of NO to NO_3^- **d** the oxidation of Cl_2 to HOCl **e** the oxidation of S to SO_4^{2-} **A42.****a** $\text{NO}_3^-(\text{aq}) + 3\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{HNO}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$ **b** $\text{VO}^{2+}(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{V}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l})$ **c** $\text{NO}(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow \text{NO}_3^-(\text{aq}) + 4\text{H}^+(\text{aq}) + 3\text{e}^-$ **d** $\text{Cl}_2(\text{aq}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{HOCl}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$ **e** $\text{S}(\text{s}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 8\text{H}^+(\text{aq}) + 6\text{e}^-$

Q43.

Balance the following equations by separating them into two half equations, balancing each half equation separately, and adding them together to form the complete redox equation.

- a** $\text{I}^-(\text{aq}) + \text{ClO}^-(\text{aq}) \rightarrow \text{I}_3^-(\text{aq}) + \text{Cl}^-(\text{aq})$
b $\text{Br}^-(\text{aq}) + \text{MnO}_4^-(\text{aq}) \rightarrow \text{Br}_2(\text{l}) + \text{Mn}^{2+}(\text{aq})$
c $\text{CH}_3\text{OH}(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) \rightarrow \text{CH}_2\text{O}(\text{l}) + \text{Cr}^{3+}(\text{aq})$
d $\text{Mn}^{2+}(\text{aq}) + \text{BiO}_3^-(\text{aq}) \rightarrow \text{Bi}^{3+}(\text{aq}) + \text{MnO}_4^-(\text{aq})$

A43.

- a** $3\text{I}^-(\text{aq}) \rightarrow \text{I}_3^-(\text{aq}) + 2\text{e}^-$
 $\text{ClO}^-(\text{aq}) + 2\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \text{Cl}^-(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 $3\text{I}^-(\text{aq}) + 2\text{ClO}^-(\text{aq}) + 4\text{H}^+(\text{aq}) \rightarrow \text{I}_3^-(\text{aq}) + 2\text{Cl}^-(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
b $2\text{Br}^-(\text{aq}) \rightarrow \text{Br}_2(\text{l}) + 2\text{e}^-$
 $\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$
 $5\text{Br}^-(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) + 16\text{H}^+(\text{aq}) \rightarrow 5\text{Br}_2(\text{l}) + 2\text{Mn}^{2+}(\text{aq}) + 8\text{H}_2\text{O}(\text{l})$
c $\text{CH}_3\text{OH}(\text{aq}) \rightarrow \text{CH}_2\text{O}(\text{l}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$
 $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
 $3\text{CH}_3\text{OH}(\text{aq}) + \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 8\text{H}^+(\text{aq}) \rightarrow 3\text{CH}_2\text{O}(\text{l}) + 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$
d $\text{BiO}_3^-(\text{aq}) + 6\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{Bi}^{3+}(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
 $\text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) \rightarrow \text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^-$
 $2\text{Mn}^{2+}(\text{aq}) + 5\text{BiO}_3^-(\text{aq}) + 14\text{H}^+(\text{aq}) \rightarrow 5\text{Bi}^{3+}(\text{aq}) + 2\text{MnO}_4^-(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$

Q44.

Explain the term *cathodic protection* of iron.

A44.

Cathodic protection involves protection of a metal, usually iron, by making it the cathode of a cell. The metal can either be joined to a more reactive metal, such as zinc, or be connected to the negative terminal of a DC power source.

Q45.

Why is it not necessary to protect structures made of copper from corrosion?

A45.

Copper is not a very reactive metal and so is not easily oxidised.

Q46.

The insides of food cans are protected from corrosion by a coating of tin. Explain why:

- a** a tin layer protects the can from corrosion
b zinc would be unsuitable for coating the inside of a food can
c the contents of cans that are dented may be dangerous to use

A46.

- a** Tin is not very reactive and so does not corrode readily. It protects the steel layer under it from corroding by preventing its contact with water and oxygen.
b Zinc is too reactive to be used inside food cans. It would be oxidised easily by the contents of the can and so contaminate the food.

- c** If a can is dented, the tin layer may be cracked, exposing the steel underneath to the contents of the can. The steel forms an anode and the tin, which is less reactive, forms the cathode. The steel rapidly corrodes and contaminates the food. Rust that forms may flake off, leaving a hole in the can.

Q47.

The iron in galvanised iron nails is protected from corrosion by a coating of zinc. Explain why a coating of tin instead of zinc would not be appropriate protection for iron nails.

A47.

Any coating on a nail is likely to crack when the nail is used. If the coating is tin, the tin will form a cathode to the more reactive iron. The iron will be oxidised, that is, corrode rapidly. Zinc, however, being more reactive than iron, will be the anode and iron will be the cathode. Oxidation, and so corrosion, of the iron nail will be prevented.

Q48.

Steel garden furniture rusts much more readily near the seaside than inland. Explain why.

A48.

Near the seaside, moisture in the air contains a higher amount of dissolved salts (e.g. NaCl(aq)). Salts are good electrolytes and accelerate the rusting process. They do this because they facilitate the production of ions at the oxidation and reduction sites of a rusting metal.

Q49.

In earlier times, oxidation was defined as ‘the addition of oxygen’.

- a** Show that this definition is consistent with the definitions of oxidation used in Chapter 16.
- b** Why do you think this definition, although still useful in some situations, has been replaced?

A49.

- a** Given the high electronegativity of oxygen, the reaction of another element (fluorine excepted) with oxygen always results in the complete or partial transfer of electrons from the other element to the oxygen. This means the other element either loses one or more electrons or undergoes an increase in oxidation number as it gains oxygen—all are definitions of oxidation.
- b** This definition has been extended/replaced because it excludes many reactions that are similar to the reaction of an element with oxygen.

Q50.

The material commonly used in a salt bridge is potassium nitrate.

- a** Why is this a suitable material for this purpose?
- b** Your laboratory partner cannot find any potassium nitrate and suggests using another ionic compound, calcium carbonate, in its place. How would you respond?

A50.

- a** It is soluble in water. Neither the nitrate ion nor the potassium ion will take part in the redox reactions within the half cells and neither will form a precipitate with any other ions in solution. (All potassium salts and all nitrates are soluble in water.)
- b** Calcium carbonate has an extremely low solubility in water and so will not provide ions in solution that can migrate from the salt bridge. The circuit will not be complete.

Q51.

Construct a concept map that shows clearly the links between the following terms:

- a** oxidant, reductant, electrons, oxidation number, oxidation, reduction
- b** anode, cathode, oxidation, reduction, electrons, ions, external circuit, internal circuit, salt bridge, positive, negative

A51.

Individual students' responses required.