YOU WILL EXAMINE:
- the importance of acids and bases and their reactions in the household and the environment, and examples of common weak and strong acids and bases
- properties of acids and bases
- the Bronsted–Lowry theory of acid–base behaviour
- conjugate acid–base pairs
- polyprotic acids and amphiprotic substances
- pH as a measure of acidity of a solution
- balanced chemical equations for reactions involving acids and bases including their reactions with water
- pH calculations for strong acids and bases and for solutions that have been diluted
- the distinction between strong and weak acids and bases, and between concentrated and dilute solutions
- the use of pH measurements to determine relative strengths of acids (and bases)
- the environmental issues of acid rain and ocean acidification
- oxidation and reduction (redox) reactions
- the electron transfer theory for redox reactions
- the representation of redox reactions by half-equations involving conjugate redox pairs, and by combined overall ionic equations
- the reactivity series of metals and metal displacement reactions
- corrosion of metals and treatments to control it.

The burning, cramping stitch felt in muscles during periods of intense exercise is caused by a build-up of acidic hydrogen ions produced by chemicals in the muscles. The body does not take in enough oxygen to keep up with the demands of the increased activity so the natural acidity regulatory buffering system in the muscle is overcome and the pH drops, causing pain.

Benjamin Franklin

Write your injuries in dust, your benefits in marble.
Two new groups of reactions

Chemists make sense out of the huge number of chemical reactions by classifying them into groups. Examples of such groupings are polymerisation reactions, which you have already studied in Unit 1, and precipitation reactions, which you met in chapters 11 and 12. Often these groups are so large that they are further divided into smaller groups, as is the case with the reactions involved in organic chemistry.

Two further such groups are acid–base reactions and redox reactions. Although both of these large groups of reactions can occur in situations without water, they nevertheless occur frequently in and with water. Any study of water chemistry would not be complete without their consideration.

Introducing acids and bases

What can send a shiver down your spine, bring tears to your eyes and make your mouth pucker? Simply the taste of a freshly cut lemon or unripe pineapple. The vinegar in salad dressing, though less stringent, has the same effect. Even a cola or a glass of sparkling wine may leave a slightly sour taste in your mouth. All of these foods have a similar effect because of the presence of compounds called acids.

Hydrochloric acid is found in our stomachs, where it is used to help break down food. It is also used in industry, where it is sometimes called ‘spirit of salts’, to clean bricks and to clean off the coating of oxide on corroded iron or steel before plating the metal with a protective layer of zinc or tin. Table 13.1 (see page 239) lists common acids and their uses.

Bases are compounds with properties that in some ways complement those of acids. This means that the characteristic properties of an acid can be reduced by adding a base. If enough base is added, the acidic properties completely disappear. When this occurs, we say that the acid has been neutralised by the base. Bases all contain water-soluble bases in varying degrees of strength. Table 13.2 (see page 241) lists some common bases and their uses.

Given the corrosive nature of some of these acids and bases it is important to carefully read the labels on common supermarket cleaning products.

Safety with acids and bases

Even so-called weak acids can be dangerous. A concentrated sample of ethanoic acid (acetic acid or vinegar), which is regarded as weak, can be extremely harmful to eyes and mucous membranes. Strong acids such as nitric acid, sulfuric acid and hydrochloric acid are extremely dangerous in concentrated form and should be handled with care even in their more dilute forms. When diluting a concentrated acid, always add acid to water. Adding water directly to concentrated acids may cause them to splatter, and severe burns may result.

Strong bases such as sodium hydroxide and potassium hydroxide are also very dangerous, particularly in concentrated form and when hot. Acid and base spills on skin should be rinsed with large quantities of running water.
TABLE 13.1 Common acids and their uses

<table>
<thead>
<tr>
<th>Acid</th>
<th>Formula</th>
<th>Occurrence or use</th>
</tr>
</thead>
<tbody>
<tr>
<td>nitric</td>
<td>HNO₃</td>
<td>used to etch copper and manufacture fertilisers</td>
</tr>
<tr>
<td>hydrochloric</td>
<td>HCl</td>
<td>produced in the stomach; used to clean bricks and concrete</td>
</tr>
<tr>
<td>sulfuric</td>
<td>H₂SO₄</td>
<td>used in car batteries, plastics, insecticides, detergents and pharmaceuticals</td>
</tr>
<tr>
<td>ethanoic</td>
<td>CH₃COOH</td>
<td>found in vinegar (used in pickling and fermentation of foods)</td>
</tr>
<tr>
<td>citric</td>
<td>C₆H₈O₇</td>
<td>found in lemon juice</td>
</tr>
<tr>
<td>carbonic</td>
<td>H₂CO₃</td>
<td>found in soft drinks</td>
</tr>
<tr>
<td>lactic</td>
<td>C₃H₆O₃</td>
<td>makes milk sour; used in wrinkle-smoothing face cream</td>
</tr>
<tr>
<td>ascorbic</td>
<td>C₆H₈O₆</td>
<td>found in citrus fruits</td>
</tr>
</tbody>
</table>

The properties of acids

Acids have many common properties. They:
- usually taste sour
- are corrosive
- are molecular in structure and dissolve in water to produce an electrolyte (substance that conducts electricity)
- affect the colour of certain natural and synthetic dyes (they turn litmus, a plant dye, from blue to red)
- are neutralised by bases.

Reactions of acids

Some common reactions involving acids and their equations were introduced in chapter 11, and a brief review of these is provided here. Any of the acids H₂SO₄, HCl and HNO₃ can be substituted in the general equations that follow. For each general word equation, an example is given, showing first the molecular and then the ionic equation for the reaction.

(a) Zinc reacts with hydrochloric acid to liberate hydrogen gas.
(b) Limestone, calcium carbonate, reacts with hydrochloric acid to liberate carbon dioxide gas.
The reactions of acids are often predictable. This makes it easy to write equations for many reactions in which acids are involved.

1. acid + metal (not Cu, Hg or Ag) → salt + hydrogen gas
   
   \[
   \text{H}_2\text{SO}_4(aq) + \text{Zn}(s) \rightarrow \text{ZnSO}_4(aq) + \text{H}_2(g)
   \]
   
   \[
   2\text{H}^+(aq) + \text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + \text{H}_2(g)
   \]

2. acid + metal carbonate → salt + carbon dioxide gas + water
   
   \[
   \text{H}_2\text{SO}_4(aq) + \text{CaCO}_3(s) \rightarrow \text{CaSO}_4(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)
   \]
   
   \[
   2\text{H}^+(aq) + \text{CaCO}_3(s) \rightarrow \text{Ca}^{2+}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)
   \]

3. acid + metal hydrogen carbonate → salt + carbon dioxide gas + water
   
   \[
   \text{HCl}(aq) + \text{NaHCO}_3(s) \rightarrow \text{NaCl}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)
   \]
   
   \[
   \text{H}^+(aq) + \text{NaHCO}_3(s) \rightarrow \text{Na}^+(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(l)
   \]

4. acid + metal sulfite → salt + sulfur dioxide gas + water
   
   \[
   2\text{HCl}(aq) + \text{FeSO}_3(aq) \rightarrow \text{FeCl}_2(aq) + \text{SO}_2(g) + \text{H}_2\text{O}(l)
   \]
   
   \[
   2\text{H}^+(aq) + \text{SO}_3^{2-}(aq) \rightarrow \text{SO}_2(g) + \text{H}_2\text{O}(l)
   \]

5. acid + metal sulfide → salt + hydrogen sulfide gas
   
   \[
   2\text{HCl}(aq) + \text{FeS}(s) \rightarrow \text{FeCl}_2(aq) + \text{H}_2\text{S}(g)
   \]
   
   \[
   2\text{H}^+(aq) + \text{FeS}(s) \rightarrow \text{Fe}^{2+}(aq) + \text{H}_2\text{S}(g)
   \]

6. acid + metal oxide (basic oxide) → salt + water
   
   \[
   2\text{HNO}_3(aq) + \text{MgO}(s) \rightarrow \text{Mg(NO}_3)_2(aq) + \text{H}_2\text{O}(l)
   \]
   
   \[
   2\text{H}^+(aq) + \text{MgO}(s) \rightarrow \text{Mg}^{2+}(aq) + \text{H}_2\text{O}(l)
   \]

7. acid + base (metal hydroxide) → salt + water
   
   \[
   \text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + 2\text{H}_2\text{O}(l)
   \]
   
   \[
   \text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)
   \]

Revision questions

1. Write both full and ionic equations for the reactions that result when zinc and HCl are mixed.
2. Write both full and ionic equations for the reactions that result when the following substances are mixed. Assume all reactions go to completion.
   (a) magnesium and sulfuric acid
   (b) aluminum and sulfuric acid
   (c) copper(II) carbonate and nitric acid
   (d) potassium hydrogen carbonate and sulfuric acid
   (e) sodium sulfite and nitric acid
   (f) lead(II) sulfide and phosphoric acid, H$_3$PO$_4$
   (g) copper(II) oxide and hydrochloric acid
   (h) potassium hydroxide and sulfuric acid

The properties of bases

Bases have many common properties. They:
- usually taste bitter
- feel slippery (they react with the natural oils in the skin and produce soap, which gives the characteristic ‘slippery feel’)
- turn litmus from red to blue
- are electrolytes (substances that conduct electricity)
The Brønsted–Lowry theory of acids and bases

An acid–base theory suggested independently by Lowry and Brønsted in 1923 is still used today. This theory is summarised by the following two statements:

- An acid is a proton donor.
- A base is a proton acceptor.

This means that a reaction between an acid and a base involves a proton transfer from the acid to the base. This occurs only when both an acid and a base are present.

The proton described by Brønsted and Lowry is simply a hydrogen ion, \( \text{H}^+ \). The hydrogen ion is transferred from one substance to another. The substance that loses an \( \text{H}^+ \) ion is the acid and the one that accepts it is the base.

\[
\text{HNO}_3(\text{l}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{NO}_3^-(\text{aq})
\]

acid  base
loses a  gains a
proton  proton

According to the Brønsted–Lowry theory, the water in this equation is acting as a base since it accepts a proton.

Revision question

3. Which of the following reactions can be classified as proton-transfer reactions?
   - (a) \( \text{HNO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) \)
   - (b) \( \text{NH}_4(\text{aq}) + \text{HNO}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{NO}_3^-\text{(aq)} \)
   - (c) \( \text{Ag}^+(\text{aq}) + \text{Cl}^-\text{(aq)} \rightarrow \text{AgCl}(\text{s}) \)
Ionisation occurs when a substance reacts to form ions. A common example is when acids and bases react with water to form H₃O⁺ or OH⁻ ions. Reactions such as these may also be termed hydrolysis reactions.

Formation of a hydronium ion. Experimental evidence indicates that each proton is associated with a particular water molecule for about 10–11 seconds at room temperature.

The following equation shows hydrochloric acid acting as an acid and donating a proton to water. The reaction is called ionisation.

\[
\text{HCl}(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]

Ionisation is a reaction in which a substance reacts with water to produce ions. Along with the chloride ion, a hydronium ion is produced. A proton, H⁺, cannot exist by itself as it is attracted to the negative end of the polar water molecule to form a hydronium ion, H₃O⁺.

Hydrolysis

When an ionic substance dissolves in water, the resulting solution is often not neutral. This is because the ions can act as acids or bases when reacting with water and produce solutions that are either acidic or basic; this is known as hydrolysis. A hydrolysis reaction is one in which a substance reacts with water to form OH⁻ or H₃O⁺ ions. Note that hydrolysis reactions can also be called ionisation reactions.

Not all cations and anions hydrolyse. For example, if sodium chloride, NaCl, is dissolved in water, the solution is neutral as neither Na⁺ nor Cl⁻ hydrolys to any significant extent.

However, if ammonium chloride, NH₄Cl, is dissolved in water, ammonium and chloride ions are produced. The ammonium ions react as follows:

\[
\text{NH}_4^+(aq) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{NH}_3(aq)
\]

This solution is acidic due to the formation of H₃O⁺(aq).

If sodium carbonate, Na₂CO₃, is dissolved in water, sodium ions and carbonate ions are produced. The carbonate ions react as follows:

\[
\text{CO}_3^{2-}(aq) + \text{H}_2\text{O}(l) \rightarrow \text{HCO}_3^-(aq) + \text{OH}^-(aq)
\]

This solution is therefore alkaline owing to the formation of OH⁻ ions.

Dissociation of bases

When ionic bases dissolve in water, they dissociate or separate into their constituent ions. They do not ionise since they do not actually react with the water to produce ions as acids do.

\[
\text{NaOH}(s) \xrightarrow{\text{H}_2\text{O}} \text{Na}^+(aq) + \text{OH}^-(aq)
\]

Molecular bases such as NH₃ do not contain ions, but react with water instead. For example, ammonia ionises according to:

\[
\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)
\]

Note: The arrows drawn in this manner indicate that the reaction proceeds partially to the right. This means that, at any given time, there are more reactant species than product species present.
Neutralisation is the process of an acid reacting with a base. The acid and base properties can be cancelled out if the exact amounts of both are used. A neutralisation reaction produces a salt and water as its products. All neutralisation reactions can be represented by the following ionic equation:

\[
\text{H}^+(\text{aq}) + \text{OH}^-\text{(aq)} \rightarrow \text{H}_2\text{O}(l)
\]

or, alternatively,

\[
\text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-\text{(aq)} \rightarrow 2\text{H}_2\text{O}(l)
\]

Neutralisation is the process of an acid reacting with a base. When an acid and base react, they usually produce a salt and water. This is called **neutralisation**. For example:

\[
\text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O}(l)
\]

The ionic equation of this reaction is:

\[
\text{H}^+(\text{aq}) + \text{Cl}^-\text{(aq)} + \text{Na}^+(\text{aq}) + \text{OH}^-\text{(aq)} \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-\text{(aq)} + \text{H}_2\text{O}(l)
\]

Since Na\(^+\) and Cl\(^-\) are **spectator ions** and do not take part in the reaction, neutralisation may be shown simply as:

\[
\text{H}^+(\text{aq}) + \text{OH}^-\text{(aq)} \rightarrow \text{H}_2\text{O}(l)
\]

This equation is called the **ionic equation of neutralisation**. The equation can also be written as follows:

\[
\text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-\text{(aq)} \rightarrow 2\text{H}_2\text{O}(l)
\]

since we have already seen that the **hydrogen ion**, H\(^+\), does not exist by itself in water.

Neutralisation reactions are one way of producing pure samples of salts. For example, if hydrochloric acid and calcium hydroxide are mixed together, an aqueous solution of calcium chloride results.

\[
2\text{HCl(aq)} + \text{Ca(OH)}_2\text{(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + 2\text{H}_2\text{O}(l)
\]

The water can be removed by evaporation. Table 13.3 lists some common salts and some of their uses.

One application of neutralisation reactions is in antacid tablets. Although the hydrochloric acid in the gastric juices in our stomachs helps break down the food we eat, if too much hydrochloric acid is produced, indigestion may result. Antacid tablets contain a base such as magnesium hydroxide, aluminium hydroxide or sodium bicarbonate that neutralises the acid in the stomach.

### Table 13.3 Some common salts and their applications

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Application</th>
</tr>
</thead>
<tbody>
<tr>
<td>ammonium sulfate</td>
<td>(NH(_4))(_2)SO(_4)</td>
<td>fertiliser</td>
</tr>
<tr>
<td>barium sulfate</td>
<td>BaSO(_4)</td>
<td>X-ray of digestive tract</td>
</tr>
<tr>
<td>calcium sulfate dihydrate</td>
<td>CaSO(_4)2H(_2)O</td>
<td>plasterboard</td>
</tr>
<tr>
<td>copper sulfate pentahydrate</td>
<td>CuSO(_4)5H(_2)O</td>
<td>dyeing; fungicides</td>
</tr>
<tr>
<td>sodium chloride</td>
<td>NaCl</td>
<td>table salt</td>
</tr>
<tr>
<td>potassium chloride</td>
<td>KCl</td>
<td>sodium-free salt substitute</td>
</tr>
</tbody>
</table>

**Revision question**

4. Use the following acids and bases to write neutralisation equations to produce the salts in Table 13.3. You may assume that the water in the hydrated salts remains in solution.

**Acids:** H\(_2\)SO\(_4\), HCl

**Bases:** NaOH, KOH, NH\(_4\)OH, Cu(OH)\(_2\), Ca(OH)\(_2\), Ba(OH)\(_2\)
In an acid–base reaction, the substance acting as the acid gives away a proton and forms a conjugate base. The substance acting as a base, after accepting a proton, forms its conjugate acid. An acid–base reaction therefore forms two conjugate pairs, with the formulas of each pair member differing by the $H^+$ ion that was transferred.

Conjugate base = acid – $H^+$
Conjugate acid = base + $H^+$

### Acid–base terms

#### Conjugate acid–base pairs

When an acid and a base react, a conjugate acid and base are formed. In an equation where an acid donates a proton to a base, the conjugate pairs are represented as shown in the following two examples:

\[
\text{HNO}_3(\text{aq}) + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{NO}_3^-(\text{aq})
\]

\[
\text{base}_1 \quad \text{acid}_2 \quad \text{acid}_1 \quad \text{base}_2
\]

In this case, the conjugate pairs are ($\text{HNO}_3/\text{NO}_3^-$) and ($\text{H}_3\text{O}^+/$$\text{H}_2\text{O}$).

\[
\text{NH}_3(\ell) + \text{H}_2\text{O}(\ell) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})
\]

\[
\text{base}_1 \quad \text{acid}_2 \quad \text{acid}_1 \quad \text{base}_2
\]

The conjugate pairs in this reaction are $\text{NH}_4^+$/NH$_3$ and $\text{H}_2\text{O}/\text{OH}^-$.

Conjugate acid–base pairs differ by a proton, $H^+$. To find the conjugate base of an acid, we subtract one $H^+$; to find the conjugate acid of a base we add one $H^+$.

\[
\begin{array}{c}
\text{conjugate acid} \\
\text{conjugate base}
\end{array}
\]

\[
\begin{array}{c}
\text{conjugate acid} \\
\text{conjugate base}
\end{array}
\]

For example:

\[
\text{H}_3\text{PO}_4(\text{aq}) \rightleftharpoons \text{H}_2\text{PO}_4^-(\text{aq})
\]

Note that a conjugate acid and a conjugate base can re-form the original acid and base by once again transferring a proton in a ‘backwards’ reaction according to our definitions.

#### Polyprotic acids

A polyprotic acid can donate more than one proton. Acids can be classified as monoprotic, diprotic, or triprotic depending on the actual number of protons that can be donated.

A polyprotic acid such as sulfuric acid is called a diprotic acid, as it can donate two protons. This occurs in two stages, as shown in the following equations:

\[
\begin{align*}
\text{Stage 1} & \quad \text{H}_2\text{SO}_4(\text{aq}) + \text{H}_2\text{O}(\text{aq}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{HSO}_4^-(\text{aq}) \\
\text{Stage 2} & \quad \text{HSO}_4^-(\text{aq}) + \text{H}_2\text{O}(\text{aq}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})
\end{align*}
\]

The first stage is more extensive than the second.

Hydrochloric acid, HCl, and ethanoic acid, CH$_3$COOH, are examples of monoprotic acids, which can donate only one proton. Phosphoric acid, H$_3$PO$_4$, is an example of a triprotic acid as it has three protons that it can donate.

#### Amphiprotic substances and ampholytes

Some substances can act as acids or bases, according to their chemical environment. This means that they can donate or accept protons. Such substances are described as amphiprotic.

Water can act as an acid or base depending on whether it is reacting with a stronger acid or base. This is illustrated in the examples in the section on
Amphiprotic substances and ampholytes are substances that can act as either acids or bases. The way they react depends on the relative strengths of the acids and bases they are reacting with. Water is a common example of an amphiprotic substance. Ampholytes are electrolytes of ionic substances such as HSO$_4^-$ and HPO$_4^{2-}$.

**Self ionisation of water**

Because water can act as both an acid and a base, it can also react with itself according to the equation:

$$\text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightarrow \text{H}_3\text{O}^+(aq) + \text{OH}^-(aq)$$

Careful measurements have shown that this does indeed occur and that the concentration of these ions in neutral water is $10^{-7}$ M (at 25°C). Such values indicate that this reaction occurs only to an extremely small extent. Furthermore, measurements have shown that, even when dissolved acids and bases are present, the mathematical product of these two ions is always $10^{-14}$. If the concentration of H$_3$O$^+$ ions increases due to the presence of a dissolved acid, the concentration of OH$^-$ ions decreases such that the product is still $10^{-14}$. The reverse applies if a source of OH$^-$ ions (a dissolved base) is present. Therefore, for any acidic, basic or neutral solution, we may write:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$$

where $K_w$ is termed the **self-ionisation constant** for water, and [H$_3$O$^+$][OH$^-$] is called the **ionic product**.

It follows that:
- if [H$_3$O$^+$] > [OH$^-$], the solution is acidic
- if [OH$^-$] > [H$_3$O$^+$], the solution is basic
- if [H$_3$O$^+$] = [OH$^-$], the solution is neutral.

However, in all cases, the product [H$_3$O$^+$][OH$^-$] equals $10^{-14}$.

**Revision questions**

5. Write the formulas of the conjugate bases of the following acids.
   - (a) H$_2$SO$_4$
   - (b) H$_2$S
   - (c) HS$^-$
   - (d) H$_2$O
   - (e) NH$_4^+$

6. Write the formulas of the conjugate acids of the following bases.
   - (a) OH$^-$
   - (b) HCO$_3^-$
   - (c) S$^2$-
   - (d) H$_2$O
   - (e) CN$^-$

7. Consider the following equation and select the substances that are acting as acids in this reaction.

   $$\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$$

   - (a) NH$_3$(aq) and H$_2$O(l)
   - (b) NH$_3$(aq) and OH$^-$(aq)
   - (c) H$_2$O(l) and NH$_4^+(aq)$
   - (d) H$_2$O(l) and OH$^-$(aq)

8. (a) Write equations to illustrate H$_3$PO$_4$ acting as a triprotic acid in water.
    (b) Identify the conjugate acid–base pairs for the equations you wrote in part (a).

9. Write equations to illustrate the amphiprotic nature of the following species.
   - (a) HS$^-$
   - (b) HSO$_4^-$
Strengths of acids and bases

The strength of an acid or base is related to the ease with which it donates or accepts a proton.

A strong acid donates protons readily.
A strong base accepts protons readily.
Weak acids or bases do not donate or accept protons readily.

Other properties of an acid, such as its conductivity, are related to its strength. For example, hydrochloric acid is a strong acid and its reaction in water is nearly complete, producing many ions. The solution therefore is a good electrolyte.

\[
\text{virtually complete reaction} \quad \text{HCl(g)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)
\]  

strong acid

However, ethanoic acid is weak, does not react to completion and its solution is a poor conductor of electricity because not many ions are formed.

\[
\text{weak acid} \quad \text{CH}_3\text{COOH(l)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(aq) + \text{CH}_3\text{OO}^-(aq)
\]  

partial reaction

Acid strength is an indication of how easily an acid donates protons. A strong acid donates protons readily and forms a large number of ions in water.

Strong acids form solutions that are good conductors of electricity. Weak acids do not donate protons readily. As a result, only a few ions are formed by reaction with water and the resulting solutions are poor electrical conductors.

Strangely shaped pillars create an eerie atmosphere in moist limestone caves. They are formed by a long process that begins with rainwater dissolving carbon dioxide in the atmosphere to form carbonic acid. When this acidic rain drips through limestone rocks, an acid–base reaction occurs. The calcium carbonate in the rocks reacts with the weak carbonic acid to form calcium hydrogen carbonate, which trickles down into the caves. As the droplets evaporate slowly, the calcium hydrogen carbonate decomposes to re-form carbon dioxide and solid calcium carbonate, creating stalactites and stalagmites.

The strength of a base also affects its conductivity. A base is said to be strong if it produces many hydroxide ions in solution because the hydroxide ion readily accepts protons. For example, sodium hydroxide is a strong base because it dissociates completely, producing many hydroxide ions. Sodium hydroxide is therefore a good electrolyte.

\[
\text{NaOH(s)} \rightarrow \text{Na}^+(aq) + \text{OH}^-(aq)
\]
However, a weak base such as ammonia does not readily ionise and accept protons. It forms very few ions in solution. Ammonia is therefore a poor conductor of electricity.

\[ \text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq}) \]

As this conductivity probe experiment shows, solutions of hydrochloric acid (a strong acid), sodium hydroxide and sodium chloride are strong electrolytes as they produce many ions in solution. Water is a weak electrolyte as it produces very few ions and is therefore able to conduct only a limited amount of electricity.

### TABLE 13.4 Relative strengths of common acids and their conjugate bases

<table>
<thead>
<tr>
<th>Name of acid</th>
<th>Formula</th>
<th>Conjugate base</th>
<th>Name of base</th>
</tr>
</thead>
<tbody>
<tr>
<td>hydrochloric</td>
<td>HCl</td>
<td>Cl(^-)</td>
<td>chloride ion</td>
</tr>
<tr>
<td>nitric</td>
<td>HNO(_3)</td>
<td>NO(_3^-)</td>
<td>nitrate ion</td>
</tr>
<tr>
<td>sulfuric</td>
<td>H(_2)SO(_4)</td>
<td>HSO(_4^-)</td>
<td>hydrogen sulfate ion</td>
</tr>
<tr>
<td>hydronium ion</td>
<td>H(_3)O(^+)</td>
<td>H(_2)O</td>
<td>water</td>
</tr>
<tr>
<td>phosphoric</td>
<td>H(_3)PO(_4)</td>
<td>H(_3)PO(_4^-)</td>
<td>dihydrogen phosphate ion</td>
</tr>
<tr>
<td>hydrofluoric</td>
<td>HF</td>
<td>F(^-)</td>
<td>fluoride ion</td>
</tr>
<tr>
<td>ethanoic (acetic)</td>
<td>CH(_3)COOH</td>
<td>CH(_3)COO(^-)</td>
<td>ethanoate (acetate) ion</td>
</tr>
<tr>
<td>carbonic</td>
<td>H(_2)CO(_3)</td>
<td>HCO(_3^-)</td>
<td>hydrogen carbonate ion</td>
</tr>
<tr>
<td>hydrogen sulfide</td>
<td>H(_2)S</td>
<td>HS(^-)</td>
<td>hydrogen sulfide ion</td>
</tr>
<tr>
<td>ammonium ion</td>
<td>NH(_4^+)</td>
<td>NH(_3)</td>
<td>ammonia</td>
</tr>
<tr>
<td>hydrogen carbonate ion</td>
<td>HCO(_3^-)</td>
<td>CO(_3^{2-})</td>
<td>carbonate ion</td>
</tr>
<tr>
<td>hydrogen sulfide ion</td>
<td>HS(^-)</td>
<td>S(^{2-})</td>
<td>sulfide ion</td>
</tr>
<tr>
<td>water</td>
<td>H(_2)O</td>
<td>OH(^-)</td>
<td>hydroxide ion</td>
</tr>
<tr>
<td>hydroxide ion</td>
<td>OH(^-)</td>
<td>O(^{2-})</td>
<td>oxide ion</td>
</tr>
<tr>
<td>hydrogen</td>
<td>H(_2)</td>
<td>H(^-)</td>
<td>hydride ion</td>
</tr>
</tbody>
</table>

**Notes**
1. All strong acids react completely with water and their anions have essentially no ability to attract protons.
2. The hydronium ion is listed so that water can appear as a base on the table.
3. Oxide ions do not exist in water as they always revert to hydroxide ions.
‘Strong’ and ‘concentrated’ have different meanings when applied to acids and bases. Likewise, the terms ‘dilute’ and ‘weak’ also have different meanings.

**Revision questions**

10. Using HNO₃ and HS⁻ as examples, write equations, with appropriate arrows, to illustrate the difference between a strong acid and a weak acid. Explain the difference in strength in terms of the conductivity of the resulting solution.

11. Dimethylamine, HN(CH₃)₂, is a weak base that reacts in water like ammonia, NH₃. Write an equation to illustrate the reaction of dimethylamine with water.

12. Acids, such as nitric and hydrochloric, are used by artists to make engravings on metal plates made from copper, zinc or iron.
   (a) Use table 13.4 to find the conjugate pairs of the acids used by the artists for their etchings.
   (b) What safety precautions would you recommend an artist working with strong acids should take?

**Strength versus concentration**

The strength of an acid is different from the concentration of an acid. The strength of a solution is determined by the number of ions present. A strong acid is completely ionised in solution (producing many ions). A weak acid does not ionise to any great extent and so contains a larger number of molecules compared with the number of ions produced in solution. Concentration, however, refers to the amount of an acid or base that is dissolved in a given volume of water. A large amount always produces a concentrated solution whereas a small amount in the same volume of water produces a dilute solution. It is possible to have a weak, concentrated acid or a dilute solution of a strong acid.

![Diagram of concentrated strong and weak acids](image)

**The pH scale**

The pH scale is usually applied over a range from 1 to 14. Using this scale, a neutral solution has a pH of 7. Values lower than 7 indicate an acidic solution. The stronger the acid the lower the pH value. For example, a solution of pH 3 is stronger than a solution of pH 4.
Any pH value greater than 7 indicates a basic solution. In this case, the higher the value, the stronger the base. The pH scale of some common acids and bases is shown below.

<table>
<thead>
<tr>
<th>pH</th>
<th>Acidity/Alkalinity</th>
</tr>
</thead>
<tbody>
<tr>
<td>1.0</td>
<td>Increasing acidity</td>
</tr>
<tr>
<td>2.0</td>
<td>2.5</td>
</tr>
<tr>
<td>8.8</td>
<td>9.5</td>
</tr>
</tbody>
</table>

The pH scale for common acids and bases

Have you ever wondered why the pink hydrangeas in your garden can bloom a different shade from year to year? The colour of the hydrangea bloom depends on the acidity of the soil. If the soil becomes too acidic, pink hydrangea blooms turn blue. To correct this, calcium oxide, CaO, generally known as ‘lime’ or ‘quicklime’ to the gardener, is added. Calcium oxide is a base that dissolves in water to form calcium hydroxide, Ca(OH)$_2$, or ‘slaked lime’. This base changes the colour of the blossoms back to pink.
Definition of pH

pH is defined according to the formula:

\[ \text{pH} = -\log_{10}[H_3O^+] \]

This formula enables pH to be calculated for any known \( H_3O^+ \) or \( OH^- \) concentration by using a calculator with a log function.

Sample problem 13.1

A solution is found to have an \( H_3O^+ \) concentration of 0.0001 M. Calculate its pH.

Solution:

**STEP 1**

\[ [H_3O^+] = 0.0001 = 10^{-4} \text{ M} \] (by inspection)

**STEP 2**

\( \text{pH} = 4 \) (multiply the power above by \(-1\))

Sample problem 13.2

A solution is found to have an \( H_3O^+ \) concentration of \( 3.5 \times 10^{-5} \) M. Calculate its pH.

Solution:

**STEP 1**

\[ [H_3O^+] = 3.5 \times 10^{-5} = 10^{-4.5} \text{ M} \] (using log function on calculator)

**STEP 2**

\( \text{pH} = 4.5 \)

Sample problem 13.3

A solution is found to have an \( OH^- \) concentration of \( 2.8 \times 10^{-4} \) M. Calculate its pH.

Solution:

**STEP 1**

\[ [OH^-] = 2.8 \times 10^{-4} = 10^{-3.6} \text{ M} \] (using calculator)

**STEP 2**

Using the ionic product for water:

\[ [H_3O^+][10^{-3.6}] = 10^{-14} \]

\[ [H_3O^+] = 10^{-10.4} \text{ M} \]

**STEP 3**

\( \text{pH} = 10.4 \)

pH values of strong acids and bases

As we have seen, strong acids undergo virtually complete ionisation with water. For a monoprotic acid such as hydrochloric acid, HCl, this means that the concentration of acid molecules originally dissolved equals the concentration of \( H_3O^+ \) ions produced when these react with water according to:

\[ \text{HCl(g)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq) \]

The pH of such solutions can then be predicted using calculations such as those above.

The pH of strong bases can also be calculated in exactly the same way, using the ionic product of water as an extra step.
Sample problem 13.4
The label on a bottle of nitric acid states that its concentration is 0.005 M. Calculate its pH.

Solution:  
STEP 1  
As nitric acid is a strong acid, the concentration of $\text{H}_3\text{O}^+$ ions is also 0.005 M, due to complete ionisation.

STEP 2  
$[\text{H}_3\text{O}^+] = 0.005 = 10^{-2.3} \text{ M (using calculator)}$

STEP 3  
$pH = 2.3$

Sample problem 13.5
The label on a bottle of potassium hydroxide states that its concentration is 0.02 M. Calculate its pH.

Solution:  
STEP 1  
As potassium hydroxide is a strong base, the concentration of $\text{OH}^-$ ions is also 0.02 M.

STEP 2  
$[\text{OH}^-] = 0.02 = 10^{-1.7} \text{ M (using calculator)}$

STEP 3  
Using the ionic product for water, $[\text{H}_3\text{O}^+] = 10^{-12.3} \text{ M}$

STEP 4  
$pH = 12.3$

Revision questions
13. What is pH and how can it be calculated?
14. What is the difference between a dilute and a concentrated solution of an acid or base?
15. Determine the pH of the following given that $[\text{H}_3\text{O}^+]$ is:
   (a) $10^{-5} \text{ M}$  
   (b) $0.1 \text{ M}$.
16. Determine the pH of the strong acid $\text{HClO}_4$ whose concentration is 0.001 M.
17. Determine the pH of the following given that $[\text{OH}^-]$ is:
   (a) $10^{-4} \text{ M}$  
   (b) $0.1 \text{ M}$.
18. 90 mL of water is added to 10 mL of $\text{HNO}_3$ with pH 2. What is the new pH?
19. Determine the pH of a solution where:
   (a) $[\text{H}_3\text{O}^+] = 9.2 \times 10^{-4} \text{ M}$  
   (b) $[\text{OH}^-] = 1.8 \times 10^{-5} \text{ M}$.
20. Using the formula for pH, explain why a neutral solution has a pH of 7 (at 25 °C).

Dilution and pH
Dilution is the process of adding water to a solution. For solutions that are acidic or basic, this affects the concentration of $\text{H}_3\text{O}^+$ ions that are present and, hence, the pH. Because pH is on a log scale, the factor by which $\text{H}_3\text{O}^+$ changes is not the same as that by which pH changes. For example, if a solution whose $\text{H}_3\text{O}^+$ concentration is 0.01 ($10^{-2}$) M is diluted by a factor of 100 to 0.0001 ($10^{-4}$) M, the pH changes from 2 to 4.
A solution of hydrochloric acid of concentration 0.050 M is diluted by adding an equal volume of water. This halves the concentration to 0.025 M. Calculate the change in pH.

**Solution:**

**STEP 1**

Hydrochloric acid is a strong acid, so all acid concentrations are the same as the $\text{H}_3\text{O}^+$ concentrations, due to complete ionisation.

Therefore, we can write initial $[\text{H}_3\text{O}^+] = 0.050 = 10^{-1.3}$ M and initial pH = 1.3.

**STEP 2**

After dilution, $[\text{H}_3\text{O}^+] = 0.025 = 10^{-1.6}$ M.

Therefore, pH after dilution = 1.6.

Therefore, pH changes from 1.3 to 1.6.

Note the common-sense check here. We would expect the solution to be less acidic after dilution and therefore have a higher pH.

**Relative strengths of acids and bases and pH**

The strength of an acid or base is measured by its tendency to donate (in the case of acids) or accept (in the case of bases) protons. The strength of an acid has to be taken into consideration when calculating the pH, as it determines the acid’s ability to ionise.

It has been found experimentally that not all acids of 0.1 M concentration have the same pH, and do not conduct electricity to the same extent. Remember that the conductivity of a solution is directly related to the number of ions in solution. Consider two acids, HCl and CH$_3$COOH, with the same concentration of 0.1 M at 25 °C. Experimental results confirm that hydrochloric acid has a pH of 1; however, ethanoic acid has a pH of 3. This is explained by the differing abilities of the two acids to ionise. Almost every HCl molecule donates protons to water. This is consistent with the fact that HCl is a strong acid and ionises virtually completely to produce a good electrolyte. CH$_3$COOH, on the other hand, is a weak acid. Only one in every hundred CH$_3$COOH molecules loses a proton, and therefore a solution of ethanoic acid is a poor electrolyte and hence a poor conductor of electricity.

The pH of a basic solution is also affected by the ability of the base to dissolve or dissociate in water. For example, a 0.1 M solution of NaOH, a strong base, has a higher pH than a 0.1 M solution of NH$_3$, a weak base. The sodium hydroxide solution is therefore a stronger base than the ammonia solution.

This dependence of pH on acid strength can be used to compare the relative strengths of a range of acids. Provided that the concentrations of the acids are all the same, the lower the pH the stronger the acid is. This is illustrated in table 13.5. Note that nitrous acid is the strongest acid of those listed (lowest pH) while hypochlorous acid (highest pH) is the weakest. It does not matter which concentration is used to make this comparison, as long as it is the same for all the acids that you are comparing.

**TABLE 13.5 pH values for solutions of a range of acids**

<table>
<thead>
<tr>
<th>Acid</th>
<th>Formula</th>
<th>pH (0.01 M solution)</th>
<th>pH (0.001 M solution)</th>
</tr>
</thead>
<tbody>
<tr>
<td>benzoic acid</td>
<td>C$_6$H$_5$COOH</td>
<td>3.1</td>
<td>3.6</td>
</tr>
<tr>
<td>ethanoic acid</td>
<td>CH$_3$COOH</td>
<td>3.4</td>
<td>3.9</td>
</tr>
<tr>
<td>hypochlorous acid</td>
<td>HCl</td>
<td>4.8</td>
<td>5.3</td>
</tr>
<tr>
<td>lactic acid</td>
<td>HC$_3$H$_5$O$_3$</td>
<td>2.9</td>
<td>3.4</td>
</tr>
<tr>
<td>methanoic acid</td>
<td>HCOOH</td>
<td>2.9</td>
<td>3.4</td>
</tr>
<tr>
<td>nitrous acid</td>
<td>HNO$_2$</td>
<td>2.6</td>
<td>3.1</td>
</tr>
</tbody>
</table>
The growth of algae and bacteria is controlled in swimming pools through the use of chlorine compounds. Sodium hypochlorite, NaOCl, and calcium hypochlorite, Ca(OCl)$_2$, are used, undergoing hydrolysis to produce the weak acid hypochlorous acid, HOCl, according to the equation:

$$\text{OCl}^- (aq) + \text{H}_2\text{O}(l) \rightarrow \text{HOCl}(aq) + \text{OH}^- (aq)$$

However, if the pH is too high, the concentration and effectiveness of HOCl are reduced. To lower the pH, an acid solution (such as hydrochloric acid or solid sodium hydrogen sulfate) can be used to neutralise the excess hydroxide ions. If the pH is too low, the acid content of the water can cause eye and skin irritation and damage to pool fittings. The pH can be raised by neutralising with sodium carbonate or sodium hydrogen carbonate.

**Measuring pH**

**Indicators** are dyes that usually are themselves weak acids or bases. They have the important characteristic that they show a different colour in an acidic solution from that in a basic solution. They can be used to determine whether a substance is acidic or basic and also how acidic or basic it is. Litmus, phenolphthalein, methyl orange and universal indicator are the most commonly used indicators. **Universal indicator** is a mixture of several indicators. It changes colour gradually, from red to green to violet as the solution changes from acidic to neutral to basic.

Indicators are useful tools, although they have certain limitations. The indicator colours found in charts are quoted for 25 °C; at other temperatures an indicator may change colour at a different pH. If a solution being tested has a colour of its own,
The pH of a solution can be measured by adding universal indicator and matching the colour shade obtained against a special chart. More accurate readings can be obtained through the use of specially designed pH meters.

the colour of the indicator may be distorted. As the naked eye is used to make the comparison between the indicator colour and the pH chart, this can be a problem, especially for people who are colourblind.

The pH meter

Most laboratories in industry have a pH meter that is used to make rapid, accurate measurements of pH. It can be connected to a computer to monitor pH changes continuously.

The voltage of the electrode changes with the $\left[ H_3O^+ \right]$ in the solution into which it is dipped. Values of pH obtained by a pH meter are accurate to within 0.01 units of the true pH and are not affected by the colour and cloudiness of the unknown solution. Hospitals use pH meters to find small but significant changes of pH in blood and other body fluids. Soils, ponds, rivers, sewage and industrial effluents are easily monitored with a pH meter.

Amino acids polymerise to make hair, a natural polymer whose chains are held together by hydrogen bonding; salt bridges between the acid and base groups; and disulfide bonds. Hair is normally acidic (pH 4 to 5), and for this reason alkaline shampoos are followed by an 'acidic rinse' or conditioner. In strongly acidic solutions (pH 1 to 2), the H–bonding and salt bridges of the hair are broken. Slightly alkaline (pH 8.5) solutions break some of the disulfide bonds, causing the outer surface of hair strands to become ruffled and appear dull, as light is unevenly reflected from their surfaces.

Pure hydrogen chloride, HCl, is a gas; it can be dissolved in water to produce a concentrated solution. Both concentrated HCl and ammonium hydroxide, NH₄OH, produce fumes. These fumes react to produce a fine white solid, ammonium chloride, NH₄Cl. This is an example of an acid–base reaction that does not involve water.
When acids cause trouble

The presence of pollutant gases in the atmosphere, together with increasing levels of carbon dioxide has, and is, leading to some environmental issues that may have serious implications, on both local and global scales.

Acid rain

Normal rain is slightly acidic due to dissolution of carbon dioxide in water according to the equation:

\[
\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{CO}_3(\text{aq})
\]

The product is the weak acid carbonic acid, which reacts slightly with water to produce a few hydronium ions as per the equation:

\[
\text{H}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) = \text{H}_3\text{O}^+(\text{aq}) + \text{HCO}_3^-(\text{aq})
\]

Many industrial processes involve the burning of coal, oil or some other fossil fuel. Many of these fuels contain sulfur in varying amounts. When sulfur is burned in air, it forms sulfur dioxide, \(\text{SO}_2\). This gas is often released into the air in vast quantities.

It must be noted that sulfur dioxide is released in many natural processes as well. In particular, active volcanoes release a large amount of sulfur dioxide into the air. The atmosphere can cope with large quantities of sulfur dioxide if it is given time to disperse the gas and break it down. When a large number of industries are all producing the gas over a small area, it cannot disperse in the air fast enough and becomes too concentrated to be safe.

Sulfur dioxide (\(\text{SO}_2\)) can react with water to produce sulfurous acid (a weak acid). It can also react with oxygen in the air to produce sulfur trioxide, \(\text{SO}_3\), which then reacts with water to produce the strong acid sulfuric acid. The equations for these three processes are:

\[
\begin{align*}
\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) & \rightarrow \text{H}_2\text{SO}_3(\text{aq}) \\
2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) & \rightarrow 2\text{SO}_3(\text{g}) \\
\text{SO}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) & \rightarrow \text{H}_2\text{SO}_4(\text{aq})
\end{align*}
\]

Other gases contribute to the acid rain problem and many of these are produced by industry and by every car that uses the internal combustion engine.

In a car engine, the temperature produced by the spark plug is sufficient to cause the nitrogen and oxygen in the air to combine and form nitrogen monoxide, NO. Nitrogen monoxide combines spontaneously with oxygen in the air to form nitrogen dioxide, \(\text{NO}_2\). This nitrogen dioxide reacts with rainwater to form nitric acid, another contributor to the problem known as acid rain.

The equation for this reaction is:

\[
3\text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow 2\text{HNO}_3(\text{aq}) + \text{NO}(\text{g})
\]

It must be noted, however, that carbonic acid and the oxides of sulfur are largely responsible for acid rain.

Impacts and control

Acid rain has been responsible for the defoliation of significant amounts of forest in both Europe and North America. In Australia it is one of the principal causes of the ‘lunar’ landscape that occurred around Queenstown in Tasmania where large amounts of sulfur dioxide were produced as a result of copper ore smelting during the last century.

Where acid rain has run off into lakes, there have been instances where the pH of such lakes has dropped to the point where they have been unable to sustain life.
A range of responses has now evolved to reduce the input of gases into the atmosphere that cause acid rain. These include:
- catalytic converters in car exhausts to change nitrogen oxides into nitrogen and oxygen
- the use of natural gas as a fuel (natural gas contains lower levels of sulfur impurity) in place of other fuels such as oil and coal
- switching to coal with a lower sulfur content
- taking measures to assist the dispersal of SO₂, such as by using taller chimney stacks
- scrubbing. This refers to treating exhaust gases to remove undesirable emissions. An increasing number of processes are now removing SO₂ from their emissions and concentrating it so that it can be used as a feedstock to produce sulfuric acid. This has an extra advantage because sulfuric acid is a valuable industrial acid and can be sold to offset the costs of the original process.

Ocean acidification
As we have seen previously, carbon dioxide dissolves in water to produce carbonic acid. The oceans are slightly alkaline, but research has shown that, since the start of the Industrial Revolution, their pH has dropped from 8.2 to 8.1. While there are many localised variations in these figures, such a change is significant. This is directly attributable to the increased level of carbon dioxide in the atmosphere. Although the effect of this is the subject of ongoing research, problems that might arise as a result of this acidification include:
- coral bleaching
- interference with marine organisms, particularly those at the lower end of the food chain
- reproductive disorders in certain marine organisms
- interference with shell building in shellfish and other similar organisms. The shells of many marine species have calcium carbonate as an important constituent. Others, including corals, form skeletons using calcium carbonate. Although calcium carbonate is virtually insoluble, there are Ca²⁺ and CO₃²⁻ ions in sea water at concentrations that are low but nevertheless important. However, carbonate ions can be removed by a reaction between themselves, water and carbon dioxide according to:
  \[
  \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{CO}_3^{2-}(aq) \rightarrow 2\text{HCO}_3^-(aq)
  \]
If the carbonate ion concentration is lowered too much, it is possible that the calcium carbonate in the shells of marine organisms may dissolve in an attempt to replace the removed carbonate ions.
  \[
  \text{CaCO}_3(s) \rightarrow \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq)
  \]

Introducing redox reactions
Oxidation-reduction or redox reactions are one of the largest groups of chemical reactions. Redox chemistry is an important aspect of everyday life.
Our bodies work by redox reactions — the food we eat is oxidised to enable us to obtain the energy we need to live. If we break a bone, electric currents generated around the damaged area are involved in the healing process. Redox chemistry is involved in the conduction of impulses in nerves. The batteries we use in radios, clocks, cars and calculators rely on spontaneous redox reactions. Many metals are extracted by reduction reactions. Municipal water supplies and swimming pools are treated with oxidants that act as bactericides and algicides. Bushfires are uncontrolled redox reactions on a large scale, and explosions are very fast redox reactions. Redox reactions occur all around us.

Redox reactions are very common. The term ‘oxidation’ was used to describe combining with oxygen, but this definition has been expanded.
An explosion is a very fast redox reaction.

What is oxidation–reduction?

Early chemists knew that a gas called oxygen was essential for the survival of living things. This prompted them to study the reactions of oxygen with other substances in an effort to find out more about its behaviour. They found that many substances combined with oxygen, so the term oxidation was used to describe these reactions. The combustion reactions of hydrocarbons such as the burning of propane:

\[ \text{C}_3\text{H}_8(g) + 5\text{O}_2(g) \rightarrow 3\text{CO}_2(g) + 4\text{H}_2\text{O}(g) \]

and the burning of iron in air:

\[ 4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s) \]

were described as oxidation reactions because a reactant, or some part of it, combined with oxygen.

Conversely, reactions that involved the decomposition of a compound, with the loss of oxygen, were called reduction reactions because the compound seemed to be ‘reduced’ to something simpler. For example, copper(II) oxide may be reduced to copper by hydrogen:

\[ \text{CuO}(s) + \text{H}_2(g) \rightarrow \text{Cu}(s) + \text{H}_2\text{O}(g) \]

Similarly, iron(III) oxide is reduced to iron by carbon monoxide in a blast furnace:

\[ \text{Fe}_2\text{O}_3(s) + 3\text{CO}(g) \rightarrow 2\text{Fe}(s) + 3\text{CO}_2(g) \]

As one reactant is reduced, the other reactant is oxidised. This can be seen when magnesium burns in steam to form magnesium oxide and hydrogen:

\[ \text{Mg}(s) + \text{H}_2\text{O}(g) \rightarrow \text{MgO}(s) + \text{H}_2(g) \]
Oxidants and reductants

In any oxidation–reduction reaction:
- The **oxidant** is the species that *causes* oxidation of another substance and is itself reduced.
- The **reductant** is the species that *causes* reduction of another substance and is itself oxidised.

The definitions of oxidation and reduction on the one hand, and oxidants and reductants on the other, can seem confusing at first. It is useful to remember that oxidation and reduction are processes. Oxidants and reductants, however, are substances that these processes happen to.

**An electron transfer view**

The modern definition of oxidation and reduction processes describes redox reactions in terms of electron transfer.

For example, the combustion of magnesium in oxygen to form magnesium oxide involves the oxidation of magnesium.

\[ 2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s) \]

A similar reaction occurs when magnesium chloride is formed by the combustion of magnesium in chlorine, but in this case no oxygen is involved.

\[ \text{Mg}(s) + \text{Cl}_2(g) \rightarrow \text{MgCl}_2(s) \]

Analysis of these two reactions in terms of a transfer of electrons reveals that they are essentially the same process. Both products are ionic substances and so contain Mg\(^{2+}\) ions. In the process of oxidation, each atom of magnesium has lost two electrons.

\[ \text{Mg} \rightarrow \text{Mg}^{2+} + 2e^- \]

In general, a substance that is oxidised is one that loses electrons and is therefore an electron donor.

When magnesium reacts in this way with either oxygen or chlorine, these atoms accept these electrons and become O\(^{2-}\) and Cl\(^{-}\) anions in the ionic lattice of the products.

\[ \text{O}_2 + 4e^- \rightarrow 2\text{O}^{2-} \]
\[ \text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^- \]

In general, a substance that is reduced is one that gains electrons and is therefore an electron acceptor.

Oxidation can be defined as a loss of electrons to another substance. Reduction can be defined as an acceptance of electrons from another substance. Redox reactions, therefore, are those in which electrons are transferred from one reactant to another.
Oxidation must be accompanied by reduction; that is, oxidation cannot occur unless reduction occurs simultaneously. When writing equations for redox reactions, we do not show electrons as all electrons given off during oxidation are taken in during reduction. However, if we are considering oxidation or reduction reactions separately, it is appropriate (and necessary) to write equations that do show the electrons, as has been done in the examples on the previous page. Such equations are called *partial equations* or *half-equations*. It should be noted that, if you multiply the two half-equations so that there are the same number of electrons on each side, and then add them together, you do, in fact, get the overall ionic equation.

The reactant that is losing electrons (undergoing oxidation) is called the reductant or reducing agent, because it transfers electrons to another substance and causes it to be reduced.

The reactant that is gaining electrons (undergoing reduction) is called the oxidant or oxidising agent, as it accepts the electrons from the reductant and causes it to be oxidised. This can be illustrated by the following diagram.

![Diagram showing oxidation and reduction](image)

When steel wool burns in chlorine, iron in the wool is oxidised to form iron(III) ions, Fe\(^{3+}\). At the same time, chlorine is reduced to form chloride ions, Cl\(^{-}\).

We can summarise these ideas in another way.

<table>
<thead>
<tr>
<th>Reactant</th>
<th>Reactant undergoes</th>
<th>Reactant is called</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mg</td>
<td>oxidation (loses e(^-))</td>
<td>reductant</td>
</tr>
<tr>
<td>O(_2)</td>
<td>reduction (gains e(^-))</td>
<td>oxidant</td>
</tr>
</tbody>
</table>

Consider the reaction of steel wool burning in chlorine gas:

\[
2\text{Fe}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{FeCl}_3(s) (\text{Fe}^{3+}, 3\text{Cl}^-)
\]

<table>
<thead>
<tr>
<th>Fe</th>
<th>Cl(_2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>donates electrons</td>
<td>accepts electrons</td>
</tr>
<tr>
<td>is oxidised</td>
<td>is reduced</td>
</tr>
<tr>
<td>is the reductant</td>
<td>is the oxidant</td>
</tr>
</tbody>
</table>

**Revision question**

21. Most of the following equations represent redox reactions. Use a diagram similar to that above to demonstrate which reactant is being oxidised and which reactant is being reduced. Identify the reactions that are not redox, and explain why.

(a) \(\text{Mg}(s) + \text{S}(s) \rightarrow \text{MgS}(s)\)
(b) \(\text{Fe}(s) + 2\text{Ag}^+(aq) \rightarrow 2\text{Ag}(s) + \text{Fe}^{2+}(aq)\)
(c) \(\text{Cl}_2(aq) + 2\text{Br}^-(aq) \rightarrow \text{Br}_2(aq) + 2\text{Cl}^-(aq)\)
(d) \(\text{Fe}^{3+}(aq) + 3\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_3(s)\)
The filaments of some light bulbs are often made from tungsten metal, W. If hot tungsten is exposed to air it oxidises to form tungsten oxide. This is why inert argon gas is used to fill the inside of light globes.

Half-equations

Half-equations are a useful way of understanding the processes involved in a redox reaction. The following example describes how half-equations are written.

When an iron nail is placed in a blue copper sulfate solution, the nail becomes coated with metallic copper and the blue colour of the solution fades. A redox reaction has taken place as electrons have been transferred from the iron nail to the copper ions in the solution, allowing solid copper to form. The full equation for the reaction is as follows:

\[
\text{Fe(s)} + \text{CuSO}_4(aq) \rightarrow \text{FeSO}_4(aq) + \text{Cu(s)}
\]

Since copper sulfate and iron sulfate are in solution, we can write an ionic equation for this reaction, eliminating the sulfate spectator ions.

\[
\text{Fe(s)} + \text{Cu}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{SO}_4^{2-}(aq) + \text{Cu(s)}
\]

The ionic equation becomes:

\[
\text{Fe(s)} + \text{Cu}^{2+}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{Cu(s)}
\]

Although the oxidation and reduction reactions occur simultaneously, it is possible to consider the two reactions separately. To do this we separate the conjugate pair of oxidant and reductant. A conjugate redox pair is made up of two species that differ by a certain number of electrons. Each has its own half-equation.

So for the oxidation conjugate pair we can write:

\[
\text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq)
\]

and for the reduction conjugate pair we can write:

\[
\text{Cu}^{2+}(aq) \rightarrow \text{Cu(s)}
\]
Another way to remember oxidation and reduction is to look at the half-equations. Electrons on the left mean that it is reduction (red is on the left of the term ‘redox’). Electrons on the right mean that it is oxidation (ox is on the right of the term ‘redox’).

In order to balance the conjugate pair to produce proper half-equations, electrons need to be shown. So, an iron atom must have lost two electrons during the reaction to go from iron metal to iron ions. So the half-equation for the oxidation part of the reaction becomes:

\[ \text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq) + 2e^- \]

Similarly, each copper ion must have gained two electrons in order to become a copper atom with no charge. So the half-equation for the reduction part of the reaction becomes:

\[ \text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu(s)} \]

These half-equations are balanced with respect to both atoms and charge. Combining these two half-equations will yield the ionic equation, shown on the previous page, for the reaction as a whole.

**Revision questions**

22. The balanced ionic equation for the displacement of silver from an aqueous silver nitrate solution by metallic lead is:

\[ 2\text{Ag}^+(aq) + \text{Pb}(s) \rightarrow 2\text{Ag}(s) + \text{Pb}^{2+}(aq) \]

(a) Write balanced oxidation and reduction half-equations for the reaction.
(b) Which reactant accepts electrons?
(c) Which reactant is oxidised? What is it oxidised to?
(d) Which reactant is the reductant?
(e) From which reactant are electrons taken?

23. The reaction of aluminium with hydrogen ions in a dilute solution of hydrochloric acid can be represented by the following half-equations:

\[ \text{Al}(s) \rightarrow \text{Al}^{3+}(aq) + 3e^- \]
\[ 2\text{H}^+(aq) + 2e^- \rightarrow \text{H}_2(g) \]

(a) Which reactant accepts electrons?
(b) Which reactant is reduced? What is it reduced to?
(c) Which reactant is the reductant?
(d) From which reactant are electrons taken?
(e) Which reactant is oxidised? What is it oxidised to?
(f) Which reactant is the oxidant?

**Writing balanced half-equations for ions in aqueous solution**

At first glance, the statement that \( \text{MnO}_4^- \) is reduced to \( \text{Mn}^{2+} \) seems incorrect because, according to the charges on the ions, there appears to be a loss, rather than gain, of electrons:

\[ \text{MnO}_4^-(aq) \rightarrow \text{Mn}^{2+}(aq) \]

It is only when the entire half-equation for the change is written that its true nature as reduction becomes obvious, with five electrons being accepted by each permanganate ion:

\[ \text{MnO}_4^- (aq) + 8\text{H}^+(aq) + 5e^- \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \]

The steps involved in obtaining such equations, using as an example the oxidation of chloride ions by permanganate ions in acidified solution are given below. The products are molecular chlorine, \( \text{Cl}_2(aq) \), and manganese(II) ions, \( \text{Mn}^{2+}(aq) \). Do not attempt to balance a single equation.
To balance half-equations: balance the elements, add water molecules, balance by adding hydrogen ions, and balance the difference in charge with electrons.

**STEP 1** Write the half-equations for oxidation and reduction showing conjugate pairs. If you are having trouble identifying the conjugate pairs, use oxidation numbers to identify which substance has been oxidised and which has been reduced.

- reduction \( \text{MnO}_4^-(aq) \rightarrow \text{Mn}^{2+}(aq) \)
- oxidation \( \text{Cl}^- (aq) \rightarrow \text{Cl}_2(aq) \)

**STEP 2** Balance all elements except hydrogen and oxygen, which will be balanced later.

\[
\text{MnO}_4^-(aq) \rightarrow \text{Mn}^{2+}(aq) \\
2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(aq)
\]

**STEP 3** Balance oxygen atoms, where needed, by adding water.

\[
\text{MnO}_4^-(aq) \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \\
2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(aq)
\]

**STEP 4** Balance hydrogen atoms, where needed, by adding \( \text{H}^+ \).

\[
\text{MnO}_4^-(aq) + 8\text{H}^+(aq) \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \\
2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(aq)
\]

**STEP 5** Balance the charge of each half-equation by adding electrons.

- \( \text{MnO}_4^-(aq) + 8\text{H}^+(aq) \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \)
  - charges as the equation stands: \(-1\ +8\ +2\ 0\)
  - net charge: \(+7\ +2\)

  Since the net charge of the left-hand side is greater than that of the right-hand side, it is necessary to add five electrons to the left-hand side to make the net charge on both sides equal at +2. The reduction half-equation becomes:

\[
\text{MnO}_4^-(aq) + 8\text{H}^+(aq) + 5\text{e}^- \rightarrow \text{Mn}^{2+}(aq) + 4\text{H}_2\text{O}(l) \\
2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(aq)
\]

- charges as the equation stands: \(2 \times -1\ 0\)
- net charge: \(-2\ 0\)

  Since the net charge of the left-hand side is less than that of the right-hand side, it is necessary to add two electrons to the right-hand side to make the net charge on both sides equal at -2. The oxidation half-equation becomes:

\[
2\text{Cl}^-(aq) \rightarrow \text{Cl}_2(aq) + 2\text{e}^-
\]

Each half-equation is now balanced. The net charge on each side of the reduction equation is +2 and the net charge on each side of the oxidation equation is -2.

**STEP 6** The previous step has resulted in an uneven number of electrons, which will not cancel if the half-equations are added. To overcome this, multiply each half-equation by factors that will lead to the same number of electrons in each half-equation. In the case of the reduction half-equation, multiply by 2, and for the oxidation half-equation multiply by 5. The reduction half-equation becomes:

\[
2\text{MnO}_4^-(aq) + 16\text{H}^+(aq) + 10\text{e}^- \rightarrow 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l)
\]
The oxidation half-equation becomes:
\[
10\text{Cl}^- (aq) \rightarrow 5\text{Cl}_2 (aq) + 10e^- 
\]

**STEP 7** Add the half-equations.
\[
2\text{MnO}_4^- (aq) + 10\text{Cl}^- (aq) + 16\text{H}^+ (aq) + 10e^- \rightarrow 2\text{Mn}^{2+} (aq) + 5\text{Cl}_2 (aq) + 8\text{H}_2\text{O}(l) + 10e^- 
\]

Then cancel the electrons.
\[
2\text{MnO}_4^- (aq) + 10\text{Cl}^- (aq) + 16\text{H}^+ (aq) \rightarrow 2\text{Mn}^{2+} (aq) + 5\text{Cl}_2 (aq) + 8\text{H}_2\text{O}(l) 
\]

This method of balancing equations is called the half-reaction or half-equation method.

### Summary of steps for balancing half-equations

1. Write the half-equations for oxidation and reduction showing conjugate pairs.
2. Balance all elements except hydrogen and oxygen.
3. Balance oxygen atoms, where needed, by adding water.
4. Balance hydrogen atoms, where needed, by adding H\(^+\).
5. Balance the charge by adding electrons.
6. Multiply each half-equation by factors that will lead to the same number of electrons in each half-equation.
7. Add the half-equations and omit the electrons.

### Sample problem 13.7

When copper metal is placed in concentrated nitric acid, the brown gas nitrogen dioxide is formed and the solution turns green. Nitrogen dioxide is often seen as a pollutant in industrial areas. Write a balanced equation for this redox reaction.

**Solution:**

**STEP 1**

Write the half-equations for oxidation and reduction showing conjugate pairs.

- Reduction: \( \text{NO}_3^- (aq) \rightarrow \text{NO}_2 (g) \)
- Oxidation: \( \text{Cu}(s) \rightarrow \text{Cu}^{2+} (aq) \)

**STEP 2**

It is not necessary in this example to balance elements other than hydrogen and oxygen as the main elements are already balanced.

**STEP 3**

Balance the oxygen atoms (where needed) by adding water.

\( \text{NO}_3^- (aq) \rightarrow \text{NO}_2 (g) + \text{H}_2\text{O}(l) \)

**STEP 4**

Balance hydrogen by adding H\(^+\) where needed.

\( \text{NO}_3^- (aq) + 2\text{H}^+ (aq) \rightarrow \text{NO}_2 (g) + \text{H}_2\text{O}(l) \)

**STEP 5**

Balance the charges by adding electrons so that the net charge on each side is equal to zero.

\[
\text{NO}_3^- (aq) + 2\text{H}^+ (aq) + e^- \rightarrow \text{NO}_2 (g) + \text{H}_2\text{O}(l) \\
\text{Cu}(s) \rightarrow \text{Cu}^{2+} (aq) + 2e^- 
\]
STEP 6
Balance the electrons by multiplying the first equation by 2. The two half-equations are

\[ 2NO_3^- (aq) + 4H^+ (aq) + 2e^- \rightarrow 2NO_2(g) + 2H_2O(l) \]
\[ Cu(s) \rightarrow Cu^{2+}(aq) + 2e^- \]

STEP 7
Add the two half-equations:

\[ Cu(s) + 2NO_3^- (aq) + 4H^+ (aq) + 2e^- \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l) + 2e^- \]

Cancel the electrons

\[ Cu(s) + 2NO_3^- (aq) + 4H^+ (aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l) \]

The reactivity of metals

When zinc is placed in copper sulfate solution a reaction occurs, causing copper metal to form on the zinc, as shown below. This is a displacement reaction: zinc removes copper from the solution and, as a result, the deep blue colour of the solution pales. If you place a copper strip, however, in a solution of zinc sulfate, no reaction occurs. Therefore, zinc is more reactive than copper.

Different metals have varying abilities to react with other substances. Potassium is so reactive that it is found naturally as a compound rather than a pure metal. When extracted, it must be stored in oil to prevent exposure to moisture and oxygen. Gold, on the other hand, is very unreactive. It occurs in nature as a pure element and can be found by mining or panning. Ornamental gold is valuable because it remains untarnished for centuries.

Potassium and gold are at opposite ends of the reactivity series of metals, shown in table 13.6 (see page 265). This series, sometimes called the activity series, lists the metals in order of how readily they react with oxygen, water, steam, dilute acids and salt solutions. Each of these reactions is a redox reaction, in which the metal is always the reductant. The reactivity series can be established experimentally. The most direct method to determine the relative reactivity of metals is to use displacement reactions.

To understand how displacement reactions may be used, consider our earlier example, where zinc displaced copper ions from solution to form copper metal and zinc ions according to the equation:

\[ Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s) \]

Copper metal does not displace zinc from solution. This demonstrated difference in reactivity forms the basis of the displacement method of determining relative reactivity. Samples of each metal are immersed in separate solutions of metal cations. These cations accept electrons from any other more reactive solid metal atoms that are introduced, forming solid metal atoms themselves. The more reactive metal therefore loses electrons and forms ions in solution.
When a strip of zinc is placed in a blue copper sulfate solution, the zinc becomes coated with metallic copper and the blue colour of the solution fades.

**TABLE 13.6 Reactivity series of metals**

<table>
<thead>
<tr>
<th>Element</th>
<th>Appearance of metal</th>
<th>Reaction with oxygen</th>
<th>Reaction with water or steam</th>
<th>Displacement of metals</th>
<th>Reaction with dilute acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>K</td>
<td>dull; stored under oil</td>
<td>oxides in air at room temperature to give oxides</td>
<td>hydrogen formed from cold water</td>
<td>Any metal in this series displaces any other that occurs lower in the series (i.e. change the metal ion back to the metal). The further apart the metals are, the easier the displacement.</td>
<td>hydrogen formed violently</td>
</tr>
<tr>
<td>Na</td>
<td>generally dull</td>
<td>oxidises when heated in air or oxygen to give oxides</td>
<td>hydrogen formed with steam</td>
<td></td>
<td>hydrogen formed with cold acid</td>
</tr>
<tr>
<td>Ca</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mg</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Al</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Zn</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fe</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sn</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Pb</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cu</td>
<td>generally shiny</td>
<td>no reaction</td>
<td>no reaction</td>
<td>no visible reaction</td>
<td>no reaction</td>
</tr>
<tr>
<td>Hg</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ag</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Au</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*Note: Aluminium metal often appears to be less reactive than the reactivity series indicates because it has a coating of aluminium oxide and this protects it from reacting further with oxygen. This coating can be thickened by a process called anodising.*

**Sample problem 13.8**

Use the reactivity series to describe what happens when a piece of lead is dipped into a solution of silver nitrate. Write an equation for the reaction and determine the oxidant and reductant.

**Solution:** Since lead is above silver in the reactivity series, a displacement reaction occurs between lead metal and silver ions. Lead displaces silver from solution, producing silver metal according to the equation:

\[
Pb(s) + 2Ag^{+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)
\]
Corrosion

When a metal corrodes it suffers damage as a result of chemical change. Corrosion is the oxidation of metals by materials in their environment. There are two types of corrosion:

1. **Dry corrosion** occurs when a metal reacts directly with oxygen. Some metals are very reactive and must be stored under oil. Potassium and sodium combine readily with oxygen, so they undergo dry corrosion very quickly. Iron undergoes dry corrosion, but at normal temperatures its rate of dry corrosion is very slow.

2. **Wet corrosion.** This occurs when a metal reacts directly with oxygen in the presence of water. Water can speed up some corrosion reactions, such as that of iron, which corrodes much more quickly in a moist environment.

Corrosion returns metals to their natural states — the original ores. Except for gold and platinum, which are virtually unreactive, most metals are found in the ground as oxidised metals or ores. Since corroded metal often loses its structural purpose and attractiveness, this oxidation reaction has a great economic impact. In Australia, billions of dollars are spent each year preventing and treating corrosion.

**The wet corrosion of iron — an electrochemical process**

The wet corrosion of iron is commonly known as rusting. Iron rusts when a water drop containing dissolved impurities lands on it. Oxidation usually takes place at a ‘stress site’ on the iron, such as a dent in a car or a scratch on the surface. Iron is oxidised, releasing electrons.

\[
\text{Fe(s)} \rightarrow \text{Fe}^{2+}(aq) + 2e^- 
\]

From this equation we can deduce that the two half-equations are:

\[
Pb(s) \rightarrow Pb^{2+}(aq) + 2e^- 
\]

\[
Ag^+(aq) + e^- \rightarrow Ag(s) 
\]

Lead is, therefore, oxidised and is the reductant. Silver ions are reduced and are, therefore, the oxidant.

**Revision questions**

25. Use the reactivity series to determine if a reaction occurs when the following are mixed. Write an equation for any reaction that occurs and identify the oxidant and the reductant.

(a) Mg(s) + Cu^{2+}(aq)  
(b) Cu(s) + Fe^{2+}(aq)  
(c) Al(s) + Pb^{2+}(aq)  
(d) Ag(s) + Zn^{2+}(aq)

26. Use the reactivity series to predict what reaction, if any, occurs when the following reagents are mixed. Write half-equations and overall equations for any reaction that you predict.

(a) Pb^{2+}(aq) and Fe(s)  
(b) Al^{3+}(aq) and Fe(s)  
(c) Sn^{2+}(aq) and Zn(s)  
(d) Zn^{2+}(aq) and Al(s)

27. A student used the reactivity series to predict that a strip of aluminium metal would react with copper(II) sulfate. When he tried the experiment he observed no reaction.

(a) Explain his observations.

(b) What steps would he need to take to prove that his prediction was correct?
These electrons travel through the metal and are accepted by oxygen, usually at the point where the edge of a water drop is in contact with the air. Oxygen can also come from dissolved oxygen in the water drop. Oxygen accepts electrons and is reduced.

$$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^- (\text{aq})$$

Fe$^{2+}$ and OH$^-$ ions produced migrate toward each other and react to produce an insoluble green iron(II) hydroxide precipitate. The overall equation is:

$$2\text{Fe}(s) + \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{Fe(OH)}_2(s)$$

These nails are in brown agar gel in which an indicator has been dissolved. Corrosion forms on nails at weak points or tiny pits, shown here in blue. When rust forms, Fe$^{2+}$ ions are formed, shown as blue spots. The pink spots are where phenolphthalein shows that hydroxide ions form and no rusting occurs.

(a) From the left, iron nails are placed in water and air, a drying agent, distilled water topped with oil, and salty water (far right).

(b) After several weeks, rust (hydrated iron oxide) is seen in the test tube with water and air. The process has been accelerated in the test tube with salt water and air. The nails exposed to air but no water and to water but no oxygen have not rusted.
Iron(II) hydroxide is very unstable in air and quickly reacts with oxygen to produce brown iron(III) hydroxide.

$$4\text{Fe(OH)}_2(s) + \text{O}_2(g) + 2\text{H}_2\text{O}(l) \rightarrow 4\text{Fe(OH)}_3(s)$$

This brown precipitate then partially dehydrates (loses water) to produce iron(III) oxide, $\text{Fe}_2\text{O}_3\cdot x\text{H}_2\text{O}$, where $x$ is less than 3. The degree of hydration (the value of $x$) affects the colour of the rust.

This process is responsible for the corrosion of cars, tools, bridges, buildings, machinery and virtually anything made of iron and steel that is not protected in some way. Because it is soft and flaky, rust is easily dislodged, which accelerates further rusting. Water is vital in the process; without a ‘salt bridge’, iron does not rust. Cars last much longer in dry areas; however, the dissolved salts in the air in seaside areas accelerate rusting. The dissolved ions increase the conductivity of the moisture and therefore encourage corrosion. Other factors increase corrosion, including the presence of acids and pollutants such as sulfur dioxide and nitrogen dioxide, and imperfections in the metal surface.

**Adverse effects of corrosion**

- Metal oxides, or rust, have less tensile strength and less elasticity than uncorroded metal. Corroded buildings, bridges and machinery may develop weaknesses that result in malfunction or breakage.
- A severely corroded metal cannot conduct electricity. Products of corrosion formed in a car battery, between the terminals and the leads, may cause the electrical system to malfunction.

- Corroded copper pipes and hot water tanks leak as the copper compounds dissolve in the water flowing through them.
- Products of corrosion are brittle and flake off, resulting in holes in car mufflers or iron roofing. In car radiators, iron oxide flakes may cause blockages.
- A corroded metal occupies a larger volume than the original metal. Corroded nuts and bolts may jam machinery.

**Corrosion protection**

It is important to protect iron and steel from corrosion. There are three main methods: surface protection (as described in chapter 4), *alloying* and *electrochemical protection*.

**Surface protection**

If a metal surface can be prevented from coming into contact with oxygen or water then it will not corrode. An obvious way of achieving this is to coat the surface in some way. Commonly used surface coatings are plastic, paint, grease or oil, noble metals and sacrificial metals. Specific details of each of these methods may be found in chapter 4.

**Alloying**

Iron can be alloyed with small quantities of metals, such as chromium, nickel, manganese and molybdenum, to produce stainless steel. A wide variety of grades of stainless steel is produced for specific purposes. The most widely used stainless steel contains approximately 18% chromium and 8% nickel. This
is called 18-8 stainless steel and is used for kitchen sinks and cutlery. The formation of a stable film of chromium(III) oxide, Cr₂O₃, provides a very strong surface protection for this alloy, although the metal may still be attacked by chloride ions. It is important, therefore, to choose an appropriate alloy for any specific use. For example, stainless steel containing molybdenum is more resistant to chloride ion attack and is therefore recommended for use in sea water; it is used in valves on scuba diving equipment.

Electrochemical protection

Electrochemical protection involves placing a more reactive metal in electrical contact with the metal that needs protection. It involves ‘sacrificing’ one metal to save another.

If a block of a metal higher in the reactivity series is connected to iron, then that metal is corroded while iron remains intact. Underground pipes may be protected by attaching bags of magnesium scraps at intervals and replacing them when they have been corroded away. The hulls of ships can be protected by attaching blocks of zinc, which are sacrificed to protect the iron. This is sometimes referred to as sacrificial protection or cathodic protection. It involves the selection of an appropriate sacrificial anode. In the diagram above left, magnesium wrapped around an iron nail protects the nail from corrosion.

Another method of cathodic protection is to make iron the cathode by connecting it to the negative terminal of a battery while a conductor such as graphite is connected to the positive terminal. The formation of Fe³⁺ ions is inhibited by the negative impressed current on the iron.

Electrochemical protection involves placing a more reactive metal in electrical contact with the metal that needs protection.

**Impressed-current cathodic protection**

Impressed-current cathodic protection supplies electrons to the protected metal by connecting it to the negative terminal. The current is monitored by a reference electrode that keeps the potential difference between the cathode and the anode high enough to ensure protection.

A demonstration of cathodic protection. Identical nails are introduced to two beakers of agar, each containing phenolphthalein indicator and potassium hexacyanoferrate, K₃Fe(CN)₆. Nail (a) is unprotected. Nail (b) has magnesium ribbon wrapped around its centre. The pink colouring indicates rusting.
**Oxidation numbers**

So far you have learned to identify oxidation and reduction in terms of production and consumption of electrons, respectively. From an overall ionic equation, it has been possible to identify conjugate redox pairs and then write the half-equations. From the positions of the electrons in these half-equations, it has then been possible to identify the oxidation and reduction processes. This procedure is excellent for consolidating an understanding of redox reactions but can be time consuming where reactions are complicated.

In order to simplify the identification of oxidation and reduction processes, chemists have developed *oxidation numbers*. These are assigned to each of the atoms in an equation using another easily learned set of rules.

Oxidation numbers will be introduced in Unit 3.

---

**Revision questions**

28. *Corrosion collage*

   In a small group, collect samples of rusted metals to produce a display. Design and perform experiments that investigate the corrosion of the metal samples your group has collected. Label each sample with:
   - (a) the name of the metal
   - (b) common uses of the metal
   - (c) the chemical reaction that describes the corrosion of the metal
   - (d) methods of corrosion control for the metal
   - (e) experimental investigation results relating to the metal’s corrosion rate
   - (f) a description of the reactivity of the metal with respect to its position in the periodic table
   - (g) the cost to society of the corrosion of the metal.

29. Prepare a multimedia presentation that shows examples of metal corrosion in your home and wider environment. Your presentation should include the following:
   - chemical reactions for the examples of metal corrosion featured
   - a discussion of the effects of corrosion on the properties and thus the uses of your featured metals
   - the environmental factors that contribute to the corrosion
   - the cost of such corrosion to society.

   You should also describe the practices currently employed to control corrosion of the examples shown, and suggest other reasonable methods of corrosion minimisation. A digital camera could be used to document your examples.

30. Find out from a local industry or small business how they deal with either corrosion prevention or replacement due to corrosion.
Reactions in water

Summary

- Common acids include HCl, H$_2$SO$_4$, HNO$_3$, H$_3$PO$_4$, CH$_3$COOH and H$_2$CO$_3$. These have similar properties. They taste sour, are corrosive and molecular in structure and turn litmus red.
- Chemical reactions of acids include:
  - acid + metal (not Cu, Hg or Ag) → salt + hydrogen gas
  - acid + metal carbonate → salt + carbon dioxide gas + water
  - acid + metal hydrogen carbonate → salt + carbon dioxide gas + water
  - acid + metal sulfite → salt + sulfur dioxide gas + water
  - acid + metal sulfide → salt + hydrogen sulfide gas
  - acid + metal oxide (basic oxide) → salt + water
  - acid + base (metal hydroxide) → salt + water.
- Common bases include metal hydroxides (e.g. LiOH, NaOH, KOH) and molecular ammonia, NH$_3$. These have similar properties. They taste bitter, feel slippery, are corrosive and turn litmus blue.
- Alkalis are bases that dissolve in water (e.g. group 1 hydroxides).
- The Bronsted–Lowry theory defines acids and bases as follows:
  - An acid is a proton donor.
  - A base is a proton acceptor.
- Acid–base reactions thus involve proton transfer, in which one hydrogen ion, H$^+$, is transferred from the acid to the base.
- Acids and bases may behave in water in the following ways:
  - Ionisation occurs when an acid donates a proton to water. The water acts as a base and forms the hydronium ion, H$_3$O$^+$.
  - Hydrolysis occurs when an anion reacts with water to produce OH$^-$, or a cation reacts with water to produce H$_3$O$^+$.
  - When an ionic base dissolves in water, it dissociates or separates into its constituent ions.
- When an acid and base react, the process is called neutralisation and a salt and water are usually produced. The ionic equation of neutralisation is:

$$H^+ (aq) + OH^- (aq) \rightarrow H_2O(l)$$

(where H$^+$ is written in place of H$_3$O$^+$).
- A conjugate acid–base pair is formed when an acid reacts with a base.
  - These pairs differ by a proton, H$^+$
  - (e.g. HCl/Cl$^-$, H$_3$O$^+$/H$_2$O).
  - To find a conjugate pair, subtract one H$^+$ from the acid or add it to the base.
- Acids can be classified according to their ability to donate protons:
  - Monoprotic acids can donate one proton (e.g. HCl).
  - Polyprotic acids can donate more than one proton, diprotic acids can donate two protons (e.g. H$_2$SO$_4$), and triprotic acids can donate three protons (e.g. H$_3$PO$_4$).
  - Amphiprotic substances can act as acids or bases depending on their chemical environment (e.g. H$_2$O, HS$^-$, HSO$_4^-$).
- The strengths of acids (or bases) relate to their ability to readily donate (or accept) protons. By contrast, concentration of acids (or bases) relates to the amount of the substance that is dissolved in a given volume of water. A large amount results in a concentrated solution while a small amount gives a dilute solution.
  - Strong acids include nitric, sulfuric and hydrochloric acids. Ethanoic acid is a weak acid.
  - Strong bases include sodium hydroxide and potassium hydroxide. Ammonia is a weak base.
- The pH scale is a convenient way of measuring the level of acidity or basicity in a solution.
  - In neutral solutions, pH = 7.
  - In acidic solutions, pH < 7.
  - In basic solutions, pH > 7.
- pH depends on H$_3$O$^+$ concentration and is defined by the formula:

$$pH = -\log_{10}[H_3O^+]$$

- In any aqueous solution at 25 °C:

$$[H_3O^+][OH^-] = 10^{-14}$$

- pH measurements can be used to compare the relative strengths of acids (and bases) providing the concentrations of the acids (or bases) are the same.
- Two significant environmental issues associated with acid–base chemistry are acid rain and ocean acidification.
- An indicator is a dye that is a weak acid or base. It has a different colour in acidic and basic solutions. Universal indicator is a commonly used indicator.
- Oxidation–reduction (or redox) reactions are those in which electrons are transferred from the reductant to the oxidant.
- Oxidation and reduction are complementary processes; one cannot occur unless the other occurs simultaneously.
- Oxidation can be defined as the loss of electrons.
Reduction can be defined as the gain of electrons.

An oxidant can be defined as a substance:
- that accepts electrons
- that undergoes reduction.

A reductant can be defined as a substance:
- that donates electrons
- that undergoes oxidation.

Oxidation–reduction reactions can be represented as the sum of two half-equations — one for oxidation and one for reduction. A conjugate redox pair is made up of two species that differ by a certain number of electrons, and each has its own half-equation.

These must be balanced to show the relative amounts of oxidant and reductant such that the number of electrons taken from the reductant is the same as the number of electrons accepted by the oxidant.

Metals can be ranked in an approximate order of reactivity, called the reactivity (or activity) series of metals, according to how readily each reacts with oxygen, water, steam, dilute acids and salt solutions. Displacement reactions, where a more reactive solid metal displaces the ions of another metal from solution, can be used to determine this ranking.

- The order of some selected metals in the reactivity series is, from most reactive to least reactive: K, Na, Ca, Mg, Al, Zn, Fe, Sn, Pb, Cu, Hg, Ag, Au.
- More reactive metals displace the cations of less reactive metals from solution.

Corrosion is the oxidation of metals by materials in their environment. It has two types, wet and dry corrosion.

Rusting is the corrosion of iron by oxygen in a moist environment.
- Rust can be represented by the formula Fe₂O₃·xH₂O where x is a number from 1 to 3.
- Rusting requires the presence of oxygen, as well as water and dissolved salts, to act as the electrolyte.
- Corrosion of iron occurs where the metal lattice has been deformed.

Corrosion affects the properties of metals, which can be protected from corrosion by one of three methods:
- surface protection, which may be achieved using plastic, paint, grease or oil, or through the application of metal coatings, such as noble coating and sacrificial coating
- alloying
- electrochemical protection, also known as sacrificial protection or cathodic protection, where a sacrificial anode of another more reactive metal (such as zinc or magnesium) is introduced.

Common metal coatings discussed in this chapter include:
- Galvanised iron has a sacrificial coating of zinc, which is more reactive than iron. At a break in the sacrificial coating, protection of the steel is maintained.
- Stainless steel is produced by alloying and is corrosion resistant because of a protective layer of chromium(III) oxide.
- Aluminium is corrosion resistant because of a tightly bound layer of aluminium oxide.

Multiple choice questions
1. Which of the following is the best definition of an acid?
   A An acid is an electron acceptor.
   B An acid is a substance that tastes sour.
   C An acid is a substance that donates H⁺.
   D An acid is a substance that accepts H⁺.

2. Consider the following equation and state the species that are acting as bases in this reaction.
   HCN(aq) + H₂O(l) → H₃O⁺(aq) + CN⁻(aq)
   A HCN(aq) and H₂O(l)
   B HCN(aq) and H₂O⁺(aq)
   C HCN(aq) and CN⁻(aq)
   D H₂O(l) and CN⁻(aq)

3. The sourness of food is a reasonable guide to its acidity, since sourness increases with increasing acidity. If food A is more sour than food B, then it is likely that:
   A A has a higher pH than B
   B A has more OH⁻(aq) ions than B
   C B is a stronger acid than A
   D A has a higher concentration of H₃O⁺(aq) ions than B.

4. Hydrogen sulfide, H₂S, is a gas at room temperature and dissolves in water. When this solution is tested, it conducts electricity and turns litmus solution red. The solution is acidic because:
   A the H₂S has gained a proton
   B the H₂O has gained a proton
   C the H₂S and H₂O have gained protons
   D the H₂S and H₂O have lost protons.

5. The conjugate base of the species H₂PO₃⁻(aq) is:
   A H₃PO₃(aq)
   B HPO₄²⁻(aq)
   C H₂PO₄(aq)
   D HPO₄³⁻(aq).

6. Consider the following equations and select those that can represent acid–base reactions.
   (i) Zn(s) + 2H₂O⁺(aq) → Zn²⁺(aq) + H₂(g) + 2H₂O(l)
   (ii) Cl⁻(aq) + H₂O⁺(aq) → HCl(aq) + H₂O(l)
   (iii) H₂O⁺(aq) + OH⁻(aq) → 2H₂O(l)
   A (i) and (ii)
   B (i) and (iii)
   C (ii) and (iii)
   D (i), (ii) and (iii)
1. Ammonia solution is alkaline due to the presence of ammonia ions.

2. When hydrogen chloride gas is dissolved in water, the resulting solution can conduct an electric current. The best explanation is that the hydrogen chloride gas is released when dissolved in water. 

3. The strength of an acid is determined by its degree of ionisation in aqueous solution. The relative strengths of these acids, in order from strongest to weakest, is: A, B, C, D. The values obtained were 4.2, 2.3, 5.3, and 4.1 respectively. 

4. Which of the following species acts as an amphiprotic species in aqueous solution?
   A. SO₃²⁻  B. HCO₃⁻  C. PO₄³⁻  D. Cl⁻ 

5. The pH of a solution having an OH⁻ concentration of 0.00001 M is:
   A. 5  B. 7  C. 9  D. 14

6. Which of the following materials is the most acidic?
   A. Grapefruit juice of pH 3.0  B. Detergent of pH 10.5  C. Gastric juice of pH 1.0  D. Sea water of pH 8.5

7. When dissolved in rain, which of the following gases does not cause it to be acidic?
   A. O₂  B. CO₂  C. NO₂  D. SO₃

8. Oxidation is best described as a process that
   A. the gain of protons  B. the loss of oxygen  C. the gain of electrons  D. the loss of electrons.
23. The half-equation for the reduction of \( \text{NO}_3^- \) to \( \text{N}_2\text{O} \) can be represented as follows:

\[
a\text{NO}_3^-(aq) + b\text{H}^+(aq) + 8e^- \rightarrow c\text{N}_2\text{O}(g) + d\text{H}_2\text{O}(l)
\]

where the values of \( a \) and \( b \) respectively are:
- A 2 and 5
- B 2 and 10
- C 1 and 4
- D 1 and 5.

24. The more reactive a metal:
- A the less easily it is oxidised
- B the more easily it gains electrons
- C the more easily it is reduced
- D the greater is its tendency to form positive ions.

25. A small piece of silver was placed in a solution containing both magnesium nitrate and copper(II) nitrate.
Which one of the following will occur?
- A No reaction occurs.
- B The silver dissolves and only copper is precipitated.
- C The silver dissolves and only magnesium is precipitated.
- D A mixture of copper and magnesium forms on the silver.

26. If a piece of shiny aluminium metal is placed in a CuSO\(_4\) solution:
- A sulfur dioxide gas is liberated
- B no reaction occurs
- C copper oxide is formed on the aluminium metal
- D copper metal is deposited on the aluminium metal.

27. What would happen if you tried to store FeSO\(_4\) solution in a container made of Cu metal?
- A The FeSO\(_4\) solution would be stored quite safely.
- B The copper of the container would dissolve and Fe metal would form.
- C The copper container would dissolve and Fe\(^{2+}\) ions would form.
- D The copper container would dissolve and O\(_2\) gas would form.

28. Which of the following metals could be used to make a container to store an aqueous copper sulfate solution?
- A Sn
- B Pb
- C Ag
- D Fe

29. Aluminium, although a fairly reactive metal, does not react with oxygen gas in air because:
- A it is protected by small amounts of its own impurities
- B its oxide forms a hard protective layer on the metal
- C nitrogen and carbon dioxide gas in the air prevent oxidation
- D it forms a hard protective layer by a reaction with sulfur in the air.

30. The underground pipe used to deliver gas to some homes passes through damp marshland. Iron pipes are particularly susceptible to corrosion in the environment. The overall corrosion reaction involves:
- A the reduction of iron
- B the oxidation of iron to Fe\(^{3+}\) ions
- C the oxidation of oxygen to OH\(^-\) ions
- D the oxidation of iron with acid from dissolved carbon dioxide.

31. The diagrams below represent four tests in an experiment designed to find the conditions necessary for a steel nail to rust.

Which diagram represents the most favourable conditions for rust to appear on the nail?
- A A
- B B
- C C
- D D

32. Iron coated with a complete layer of tin does not rust because:
- A the tin acts as a sacrificial metal, becoming oxidised to prevent the iron rusting
- B the tin prevents the iron coming into contact with moisture and air
- C the tin combines with the iron to form a new compound that does not rust
- D particles of rust are unable to stick to the tin surface.

33. A large piece of magnesium buried in the ground and connected to an iron pipe prevents the corrosion of the iron pipe.
The best explanation for this fact is that:
- A the Mg forms a protective coating of Mg(OH)\(_2\) on the iron
- B the Mg is sacrificially oxidised, and the electrons released flow to the Fe and prevent its corrosion
C a protective coating of Fe\(^{2+}\) forms on the iron as electrons flow from Fe to Mg
D Mg atoms fill the vacancies in the iron crystals left by the Fe atoms lost during corrosion.

34. Iron rubbish bins coated with a complete layer of zinc do not rust because:
   A the zinc acts as a sacrificial metal, becoming oxidised to prevent the iron rusting
   B iron is a more reactive metal than zinc
   C the zinc combines with the iron to form a new compound that does not rust
   D particles of rust are unable to stick to the zinc surface.

Review questions

Properties of acids and bases

1. (a) What are the characteristic properties of acids?
   (b) Illustrate two of these properties with an appropriate equation.

2. (a) List three properties of bases.
   (b) Give two uses of bases that illustrate these properties.

3. (a) Define a base and give two examples.
   (b) Define an alkali and give two examples.

4. Classify the following list of properties and uses as:
   acidic; basic; both acidic and basic; or neutral.
   (a) A sour taste
   (b) The bubbles in champagne
   (c) Corrosive
   (d) Dissolves oxide coating on metals
   (e) Dissolves marble statues
   (f) Feels slippery
   (g) Foods that taste bitter
   (h) Breaks down vegetable matter
   (i) Can be used to clean drains
   (j) Produced in stomach
   (k) Used for upset stomachs

Reactions of acids

5. Write both full and ionic equations for the reactions that result when the following substances are mixed. Assume all reactions go to completion.
   (a) lead(II) sulfide and phosphoric acid (H\(_3\)PO\(_4\))
   (b) copper(II) oxide and hydrochloric acid
   (c) potassium hydroxide and sulfuric acid
   (d) aluminium and hydrochloric acid
   (e) calcium and carbonic acid
   (f) lithium carbonate and nitric acid

The Brønsted–Lowry theory of acids and bases

6. In terms of the Brønsted–Lowry theory, define an acid, a base and a proton-transfer reaction. You may use an equation in your explanations.

7. Distinguish between the terms ‘hydrolysis,’ ‘ionisation’ and ‘dissociation.’

8. Explain, using equations, why a solution of hydrochloric acid conducts electricity but pure hydrogen chloride does not.

9. Give two common uses for sodium carbonate, Na\(_2\)CO\(_3\).

10. Antacid tablets are used to neutralise the hydrochloric acid in the stomach. Use the three bases mentioned in the discussion of antacids on page 243 to write neutralisation equations for the reactions that may happen in our bodies.

11. A white cloud is formed when \textit{ammonia} is neutralised by HCl. Write an equation for the formation of the white cloud in this reaction.

12. (a) What are conjugate acid–base pairs?
   (b) State the formulas of the conjugate acid of the following bases:
      (i) Cl\(^-\)
      (ii) CO\(_3\)^{2-}
      (iii) HSO\(_4\)\(^-\)
   (c) State the formulas of the conjugate base of the following acids:
      (i) HP
      (ii) HCO\(_3\)\(^-\)
      (iii) HSO\(_4\)^{2-}

13. Identify the conjugate pairs in the following equations.
   (a) HS\(^-\) (aq) + NH\(_3\) (aq) \rightarrow NH\(_4\)^+ (aq) + S\(^2-\) (aq)
   (b) NH\(_4\)^+ (aq) + CH\(_3\)COO\(^-\) (aq) \rightarrow NH\(_3\) (aq) + CH\(_3\)COO\(^-\) (aq)
   (c) HSO\(_4\)^{2-} (aq) + H\(_2\)O(l) \rightarrow SO\(_4\)^{2-} (aq) + H\(_3\)O\(^+\) (aq)
   (d) HPO\(_4\)^{3-} (aq) + H\(_2\)O(l) \rightarrow OH\(^-\) (aq) + H\(_3\)PO\(_4\) (aq)
   (e) HNO\(_3\) (aq) + OH\(^-\) (aq) \rightarrow H\(_2\)O(l) + NO\(_3\)\(^-\) (aq)

14. Write two equations to illustrate the ampholyte HCO\(_3\)^{2-}.

Strengths of acids and bases

15. Sulfuric acid is a strong, diprotic acid. Explain, using equations, what this term means.

16. (a) Explain why an aqueous solution of an acid conducts electricity.
   (b) What is the term given for such a solution?
   (c) Explain why a strong acid conducts electricity to a greater extent than a weak acid.

17. Andrew took the lid off a bottle of glacial ethanoic acid, CH\(_3\)COOH, and immediately noticed the very potent smell of vinegar. ‘This must be a very strong acid,’ he exclaimed. ‘No,’ said his sister Thalia, ‘it is merely a concentrated acid.’ Using appropriate definitions, explain who is right.

18. Artists often use acids to etch into metals.
   (a) What property of hydrochloric and nitric acids enables artists to etch into metals?
   (b) Why would HCl ‘bite’ into an artist’s plate more deeply than HNO\(_3\)?
The pH scale

19. (a) Identify each of the following solutions as acidic, basic or neutral at 25 °C.
   (i) pH = 6.3
   (ii) pH = 7
   (iii) pH = 8.4
   (iv) pH = 1.2
   (b) Which solution has the highest \([\text{H}_3\text{O}^+]\)? Give a reason for your answer.

20. A person who ingests too much aspirin (acetyl salicylic acid) may die. Explain what happens to the pH and hydrogen ion concentration.

21. In a pH test on a swimming pool, the water is found to have a pH of 4.5.
   (a) What would you add to the water to raise the pH?
   (b) The pH should be about 7.5 for best conditions. Is this acidic or alkaline?

22. Soil contains a number of different elements. Plants need to take up these elements in order to grow and flower. If the pH of soil is too high or too low, plants cannot take up these elements.

   The figure below shows how the pH of the soil influences the amount of an element that can be taken up by a plant. The narrower the bar, the harder it is for the plant to take up the element.

   ![Diagram of pH scale with bars for different elements]

   (a) Which elements can easily be absorbed at a pH below 4.5?
   (b) Which elements can be absorbed at a pH of 8?
   (c) Which elements cannot be taken up easily if the pH is 6?
   (d) Which elements cannot be taken up easily if the soil pH is 8?
   (e) Fuchsias require the soil to be acidic. Which elements would be most important to these plants?
   (f) Azaleas require a pH of around 5.5. Which elements are not important to these plants?
   (g) A gardener comes to you with a problem. He says that he cannot grow lettuce successfully any longer. It is not an insect problem or a water problem. Design an experiment you could do to find out the problem with his soil. (Note: Lettuces need a lot of iron.)

pH calculations

23. Calculate the pH of solutions that have the following \(\text{H}_3\text{O}^+\) concentrations.
   (a) 0.0010 M
   (b) 0.00050 M
   (c) \(4.2 \times 10^{-4}\) M
   (d) \(6.7 \times 10^{-9}\) M

24. Explain why, when an acid is diluted by a factor of 2, its pH does not change by the same factor.

25. Calculate the pH of solutions that have the following \(\text{OH}^-\) ion concentrations.
   (a) 0.0010 M
   (b) 0.00050 M
   (c) \(4.2 \times 10^{-4}\) M
   (d) \(6.7 \times 10^{-9}\) M

Environmental issues

26. Choosing either acid rain or ocean acidification, produce a poster or a PowerPoint presentation.

   Your poster or presentation must include a description of the problem and its scope, the chemistry involved, including relevant equations, the effect on living things and possible preventative action that could be taken to reduce the problem.

Oxidation and reduction

27. Define oxidation and reduction reactions in terms of the transfer of electrons.

28. What is an oxidant?

29. Use the electron transfer approach to identify the oxidant, reductant, direction of electron transfer and the direction of oxidation and reduction in the following equations. (The diagram on page 259 may be used as a guide.)

   (a) \(\text{Sn}^{2+}(aq) + 2\text{Ce}^{4+}(aq) \rightarrow \text{Sn}^{4+}(aq) + 2\text{Ce}^{3+}(aq)\)
   (b) \(\text{Pb}^{2+}(aq) + \text{Cd}(s) \rightarrow \text{Pb}(s) + \text{Cd}^{2+}(aq)\)
   (c) \(2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{Al}^{3+}(aq) + 6\text{Cl}^{-}(aq)\)
   (d) \(2\text{H}^+(aq) + \text{Mg}(s) \rightarrow \text{H}_2(g) + \text{Mg}^{2+}(aq)\)
   (e) \(\text{Na}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Na}^+(aq) + 2\text{OH}^-(aq) + \text{H}_2(g)\)

Redox equations

30. Identify the following half-equations as being either oxidation or reduction.

   (a) \(\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^-\)
   (b) \(\text{Al}^{3+}(aq) + 3\text{e}^- \rightarrow \text{Al}(s)\)
   (c) \(\text{O}_2(aq) + 2\text{H}_2\text{O}(l) + 4\text{e}^- \rightarrow 4\text{OH}^-(aq)\)

31. Show, using oxidation half-equations and reduction half-equations, that the following reactions are redox reactions.

   (a) \(4\text{Ag} + \text{O}_2 \rightarrow 2\text{Ag}_2\text{O}\)
   (b) \(\text{Fe} + \text{S} \rightarrow \text{FeS}\)
   (c) \(\text{Pb} + \text{Cl}_2 \rightarrow \text{PbCl}_2\)
   (d) \(2\text{Hg} + \text{O}_2 \rightarrow 2\text{HgO}\)

   Identify the oxidants and reductants.
32. Balance the following equations using half-equations.
   (a) $\text{Br}^-(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{SO}_2(g) + \text{Br}_2(l)$
   (b) $\text{Al}(s) + \text{Cl}_2(g) \rightarrow \text{AlCl}_3(s)$
   (c) $\text{I}_2(s) + \text{H}_2\text{S}(g) \rightarrow \Gamma(aq) + \text{S}(s)$
   (d) $\text{Cu}(s) + \text{HNO}_3(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{NO}(g)$
   (e) $\text{Cu}(s) + \text{HNO}_3(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{NO}_2(g)$
   (f) $\text{CuO}(s) + \text{NH}_3(g) \rightarrow \text{N}_2(g) + \text{Cu}(s)$
   (g) $\text{PbS}(s) + \text{H}_2\text{O}_2(l) \rightarrow \text{PbSO}_4(s) + \text{H}_2\text{O}(l)$
   (h) $\text{Cr}_2\text{O}_7^{2-}(aq) + \text{CH}_3\text{CH}_2\text{OH}(g) \rightarrow \text{CH}_3\text{COOH}(aq) + \text{Cr}^{3+}(aq)$

**Reactivity series of metals**

33. Explain why silver and gold are used for ornamental purposes more often than zinc and iron.

34. What is a displacement reaction? Explain using an example.

35. How does the reactivity of a metal relate to the reducing strength of that metal?

36. Where are the most reactive metals found in the periodic table?

37. Which of the following reactions is more likely to happen if the reactants are mixed?
   
   $\text{Zn}(s) + \text{Pb}^{2+}(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{Pb}(s)$
   
   $\text{Pb}(s) + \text{Zn}^{2+}(aq) \rightarrow \text{Pb}^{2+}(aq) + \text{Zn}(s)$

38. Develop overall redox equations by first writing the oxidation and reduction equations for the following reactions. Label the oxidant and reductant in each case.
   (a) Zinc solid is placed in a copper sulfate solution.
   (b) An iron nail is placed in a solution of tin(II) nitrate.
   (c) A small piece of lead is placed in silver nitrate solution.
   (d) A small piece of aluminum is reacted with zinc chloride solution.

39. Copper sulfate solutions cannot be stored in tin containers. Why is this so?

40. An iron nail is placed in the following aqueous solutions: $\text{AgNO}_3$, $\text{Mg(NO}_3)_2$, $\text{CuSO}_4$, $\text{Pb(NO}_3)_2$, $\text{Zn(NO}_3)_2$.
   (a) In which of the solutions would you expect a coating of another metal on the iron nail?
   (b) Write a redox equation for any reaction that may occur.

41. A metal $X$ is placed in (a) a solution of NaCl or (b) a solution of CuSO$_4$. There is no reaction with sodium chloride, but $X$ is covered with a thin film of copper in the second solution. What can you say about the relative reactivity of $X$, Na and Cu?

42. Tin displaces copper metal from CuSO$_4$ solution and cadmium displaces tin metal from SnCl$_2$ solution. The three metals and their cations are as follows: Sn/Sn$^{2+}$, Cu/Cu$^{2+}$, Cd/Cd$^{2+}$.
   (a) Ignoring spectator ions, write an ionic equation for the two displacement reactions described above.
   (b) Use the information given to deduce the relative order of metals from most reactive to least reactive.
   (c) State whether you would expect the following reactions to occur.
      (i) $\text{Sn}^{2+}(aq) + \text{Cu}(s) \rightarrow \text{Sn}(s) + \text{Cu}^{2+}(aq)$
      (ii) $\text{Cd}^{2+}(aq) + \text{Cu}(s) \rightarrow \text{Cd}(s) + \text{Cu}^{2+}(aq)$
      (iii) $\text{Sn}(s) + \text{Cd}^{2+}(aq) \rightarrow \text{Sn}^{2+}(aq) + \text{Cd}(s)$
      (iv) $\text{Cd}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Cd}^{2+}(aq) + \text{Cu}(s)$

43. Four hypothetical metallic elements, $A$, $B$, $C$ and $D$, form soluble nitrates with the formulas $A$ : NO$_3$, $B$ : NO$_3$, $C$ : NO$_3$ and $D$ : NO$_3$.
   Strips of each of the four metals were immersed in solutions of the other metal nitrates and the following observations were recorded.
   - Metal $B$ underwent reaction in all solutions.
   - Metal $A$ reacted only with CNO$_3$.
   (a) State the formulas of the metal cations in the soluble nitrates.
   (b) Deduce the position of $D$ in the series by listing the three metals $A$, $B$ and $C$ in decreasing order, with respect to their relative reactivities.
   (c) List the four metals in order from most reactive to least reactive.
   (d) Which metal cation is the weakest oxidant?

44. Six different strips of metal, $A$ to $F$, were placed in solutions of their metal nitrates to see if there was any reaction. The results are shown in the table below.

<table>
<thead>
<tr>
<th>Metal</th>
<th>$A$ nitrate</th>
<th>$B$ nitrate</th>
<th>$C$ nitrate</th>
<th>$D$ nitrate</th>
<th>$E$ nitrate</th>
<th>$F$ nitrate</th>
</tr>
</thead>
<tbody>
<tr>
<td>$A$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$x$</td>
<td>$-$</td>
</tr>
<tr>
<td>$B$</td>
<td>$x$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
</tr>
<tr>
<td>$C$</td>
<td>$x$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
</tr>
<tr>
<td>$D$</td>
<td>$x$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
</tr>
<tr>
<td>$E$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
</tr>
<tr>
<td>$F$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
<td>$-$</td>
</tr>
</tbody>
</table>

$\checkmark$ = metal displaced  $\times$ = no reaction

(a) Deduce the order of reactivity of the metals $A$ to $F$ and list them in order, beginning with the most reactive.
(b) Identify the six metals.

**Corrosion**

45. (a) What is corrosion?
   (b) Why is corrosion sometimes referred as a ‘cancer’?

46. Explain, using equations, why the corrosion of iron to form rust is a redox process.
47. The properties of metals are changed during corrosion. Explain, using three examples.

48. Salt is often spread on roads in cold countries to melt the ice. Explain why cars rust quickly in these countries.

**Corrosion protection**

49. List the methods of corrosion protection of steel that depend on:
   (a) prevention of water and air from coming into contact with the metal
   (b) providing electrons to the steel.

50. Explain what would happen if the following coatings of a steel object are damaged.
   (a) Plastic  (b) Paint  (c) Tin  (d) Zinc

51. (a) What does the term ‘noble coating’ mean?
    (b) Why is it undesirable to coat screws with tin as a protection from corrosion?

52. Food containers are tin-plated. Explain why:
   (a) galvanising steel tins is undesirable
   (b) it is dangerous to eat the food from dented tin-lined cans.

53. Which alloy is used for sinks and cutlery? Explain how it protects the steel.

54. It has been estimated that over 20% of the iron produced in the world is used to replace that lost by corrosion.
   (a) List the substances needed for the rusting of iron to occur.
   (b) Name three other substances that speed up the rusting of iron.
   (c) Explain why iron objects are less likely to rust in the following environments:
      (i) in a desert
      (ii) in outer space
      (iii) on the ocean floor.

55. Which of the following methods would be successful in preventing iron from rusting?
   (a) Plating the iron with zinc
   (b) Plating the iron with copper
   (c) Connecting the iron with a wire to a block of zinc
   (d) Connecting the iron with a wire to a block of copper
   (e) Alloying the iron to produce stainless steel

56. An iron sculpture was produced to commemorate the anniversary of the founding of a small town. To stop it rusting, the mayor attached it by a wire to a block of zinc that was stored underground out of sight.
   (a) Explain how the mayor’s action would prevent corrosion.
   (b) What would happen if a block of lead were to be attached to the iron rather than the block of zinc?
   (c) List two other ways in which corrosion of the statue could be prevented.
Exam practice questions

In a chemistry examination you will be required to write a number of short and extended response questions.

Extended response questions

1. Ethanoic (acetic) acid is a weak monoprotic acid.
   (a) What is meant by the term ‘monoprotic’? 1 mark
   (b) Write the equation for the ionisation of ethanoic acid in water. 1 mark
   (c) Write an equation for the neutralisation of ethanoic acid by sodium hydroxide. 1 mark
   (d) What is meant by the term ‘weak’? Use your equation from part (b) to illustrate your answer. 2 marks
   (e) Draw the structure of an ethanoic acid molecule. Given that it is monoprotic, circle the hydrogen atom that is donated when it reacts with a base. 2 marks

2. The hydrogen carbonate ion is amphiprotic.
   (a) What is meant by the term ‘amphiprotic?’ 1 mark
   (b) Write two different equations for the reaction of the hydrogen carbonate ion with water that demonstrate its amphiprotic nature. 2 marks
   (c) Solutions of the hydrogen carbonate ion in water are slightly basic. Comment on this with respect to the two equations that you have written for part (b). 1 mark

3. A clean piece of tin metal is placed into a solution of lead nitrate and left for a period of time. It is then observed that the tin has become coated with a layer of dull lead metal.
   (a) Write the half-equation for the oxidation reaction that has occurred. 1 mark
   (b) Write the half-equation for the reduction reaction that has occurred. 1 mark
   (c) State the oxidant for this reaction. 1 mark
   (d) State the reductant for this reaction. 1 mark
   (e) If the tin is removed, cleaned of lead, and then placed into a solution of magnesium chloride, would a further redox reaction occur? Explain. 2 marks