OXFORD

# CHEMISTRY

## FOR QUEENSLAND

UNITS 3 & 4

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# CONTENTS

CHAPTER 6

# Redox reactions

Redox reactions involve the transfer of electrons, and therefore the movement of electrons, from one atom to another. The movement of electrons generates an electrical current, making redox reactions one of the most important types of chemical reaction. Redox reactions convert chemical potential energy into electrical energy, so are one of the principle sources of energy in the world.

Redox reactions can also be highly destructive. Some are responsible for killing bacteria and make excellent household cleaners. They can also sterilise unsanitary environments or hospitals and destroy certain metal structures and biomolecules. For example, the rusting of metals (especially cars) and the browning (or oxidation) of fruit are the result of redox reactions.

## OBJECTIVES

→ Recognise that a range of reactions, including displacement reactions of metals, combustion, corrosion and electrochemical processes, can be modelled as redox reactions involving oxidation of one substance and reduction of another substance.

→ Understand that the ability of an atom to gain or lose electrons can be predicted from the atom's position in the periodic table, and explained with reference to valence electrons, consideration of energy and the overall stability of the atom.

→ Identify the species oxidised and reduced, and the oxidising agent and reducing agent, in redox reactions.

→ Understand that oxidation can be modelled as the loss of electrons from a chemical species, and reduction can be modelled as the gain of electrons by a chemical species; these processes can be represented using balanced half-equations and redox equations (acidic conditions only).

→ Deduce the oxidation state of an atom in an ion or compound and name transition metal compounds from a given formula by applying oxidation numbers represented as roman numerals.

→ Use appropriate representations, including half-equations and oxidation numbers, to communicate conceptual understanding, solve problems and make predictions.



## PRACTICALS

| | |
|---|---|
| MANDATORY PRACTICAL | **6.1** Performing single displacement reactions |



**FIGURE 1** The rust on the wreck of the SS *Maheno* at Fraser Island is caused by oxidation of metal exposed to the atmosphere and ocean.

# 6.1 Reduction and oxidation reactions

**KEY IDEAS**

In this section, you will learn about:

+ how reduction is a loss of electrons while oxidation is a gain of electrons
+ valence electrons and their involvement in the transfer of electrons and energy
+ terminology such as 'reduction', 'oxidation', 'reducing agent' and 'oxidising agent'.

Negatively charged electrons move between chemical species. When they move, energy is exchanged. This exchange of electrons between chemicals is like a financial transaction. When a substance loses one or more electrons, another substance accepts them.

**Redox** is an abbreviation for a pair of reactions in which electrons are exchanged between reactants and that occur simultaneously. These reactions are called **oxidation** and **reduction** reactions, and one cannot occur without the other.

**redox**
a chemical reaction involving the transfer of electrons from one reactant to another

**oxidation**
a loss of electrons from one atom to another atom

**reduction**
a gain of electrons from one atom to another atom

## A transfer of electrons

Reduction and oxidation reactions are characterised by the transfer of electrons. As one reactant loses one or more electrons, the second reactant gains them to form new products.

- Oxidation occurs when a chemical species loses one or more electrons.
- Reduction occurs when a chemical species gains one or more electrons.

### Valence electrons

The reaction between sodium and oxygen to form sodium oxide is an example of a redox reaction:

$$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$$

The Bohr model of electron configuration of the sodium and oxygen atoms (Figure 1) demonstrates that sodium has one **valence electron**, which it must lose in order to become stable. Oxygen has six valence electrons and must gain two to complete its octet, have a complete valence shell and be stable. This means that two sodium atoms are required for every oxygen atom because sodium loses one electron and oxygen gains two.

This example demonstrates the transfer of electrons between reactants. Sodium loses an electron and oxygen gains two electrons. The transfer of electrons results in Na becoming the ion $Na^+$ and O becoming the ion $O^{2-}$. Sodium loses an electron and develops a positive charge. Oxygen gains electrons and develops a negative charge. Therefore, sodium has undergone oxidation and oxygen has undergone reduction. Sodium has been **oxidised** and oxygen has been **reduced**. The general process is summarised in Figure 2.

**Study tip**

A useful method of remembering oxidation and reduction is OIL RIG. This stands for 'Oxidation is Loss' and 'Reduction is Gain'.

**valence electron**
an electron in the outermost shell of an atom, according to the Bohr model of electron configuration

**oxidise**
to gain electrons

**reduce**
to lose electrons

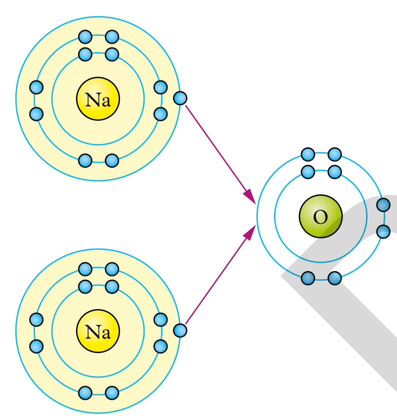

FIGURE 1 Sodium has one valence electron. Oxygen has six valence electrons. Two sodium atoms each transfer an electron to an oxygen atom to form sodium oxide.

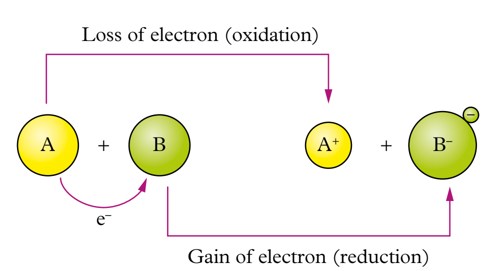

FIGURE 2 The transfer of an electron from A to B results in a positive cation ($A^+$) and negative anion ($B^-$), forming an ionic compound.







### Terminology

In the example between sodium and oxygen to form sodium oxide, sodium is oxidised, and oxygen is reduced. However, sodium will not lose this electron unless oxygen is able to accept it. For this reason, the oxygen causes the sodium to lose an electron and, similarly, the sodium causes the oxygen to gain an electron. The chemical species responsible for causing oxidation and reduction are called **oxidising agents** and **reducing agents** (Figure 3).

**oxidising agent**
the substance that causes oxidation and is itself reduced

**reducing agent**
the substance that causes reduction and is itself oxidised

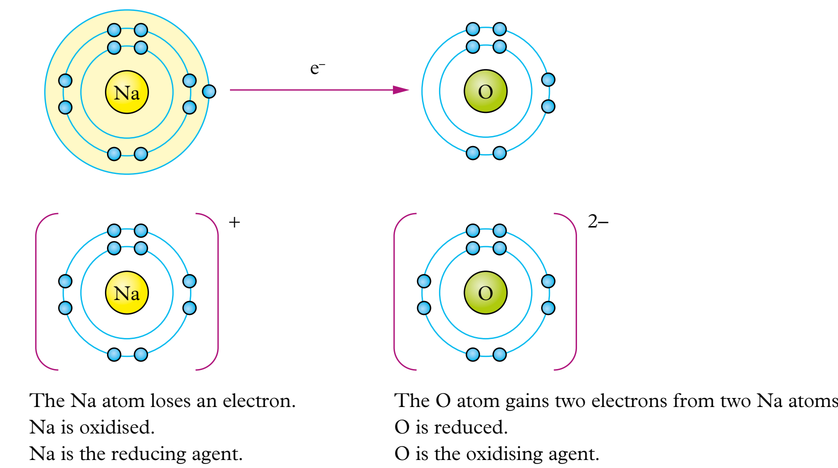

FIGURE 3 A summary of the redox process and the terminology used to describe this process.

### Predicting redox reactions from the periodic table

Elements in groups 1 and 2 on the periodic table readily lose their valence electrons to form more stable positive cations. By losing these electrons, they are oxidised and act as reducing agents. Metals that donate their valence electrons more readily are stronger reducing agents.

Figure 4 lists the first ionisation energy of each element on the periodic table. The first ionisation energy is the energy required to remove one electron from an element. Elements in the bottom left of the table have the lowest **ionisation energies**, and lose their valence electrons more readily.

**ionisation energy**
the energy (in $kJ\ mol^{-1}$) required by a gaseous atom to remove an electron from its valence shell

| Period \ Group | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 | H 1310 | | | | | | | | | | | | | | | | | He 2370 |
| 2 | Li 520 | Be 900 | | | | | | | | | | | B 800 | C 1090 | N 1400 | O 1310 | F 1680 | Ne 2080 |
| 3 | Na 495 | Mg 735 | | | | | | | | | | | Al 580 | Si 780 | P 1060 | S 1005 | Cl 1255 | Ar 1527 |
| 4 | K 420 | Ca 590 | Sc 630 | Ti 660 | V 650 | Cr 660 | Mn 710 | Fe 760 | Co 760 | Ni 730 | Cu 740 | Zn 910 | Ga 580 | Ge 780 | As 960 | Se 950 | Br 1140 | Kr 1350 |
| 5 | Rb 400 | Sr 550 | Y 620 | Zr 660 | Nb 670 | Mo 680 | Tc 700 | Ru 710 | Rh 720 | Pd 800 | Ag 730 | Cd 870 | In 560 | Sn 700 | Sb 830 | Te 870 | I 1010 | Xe 1170 |
| 6 | Cs 380 | Ba 500 | La 540 | Hf 700 | Ta 760 | W 770 | Re 760 | Os 840 | Ir 890 | Pt 870 | Au 890 | Hg 1000 | Tl 590 | Pb 710 | Bi 800 | Po 810 | At ... | Rn 1030 |
| 7 | Fr ... | Ra 510 | | | | | | | | | | | | | | | | |

FIGURE 4 The first ionisation energies ($kJ\ mol^{-1}$) of elements on the periodic table.

Group 17 elements readily gain valence electrons to form more stable negative anions. In gaining electrons, they are reduced and act as oxidising agents. Non-metals that accept electrons more readily are stronger oxidising agents.

**electronegativity** the attraction between a positively charged nucleus and the negatively charged electrons of a neighbouring atom

Figure 5 lists the **electronegativities** of elements. You can see that the elements in the top right-hand side of the periodic table (except the noble gases) have the strongest electronegativities – they more readily accept electrons.

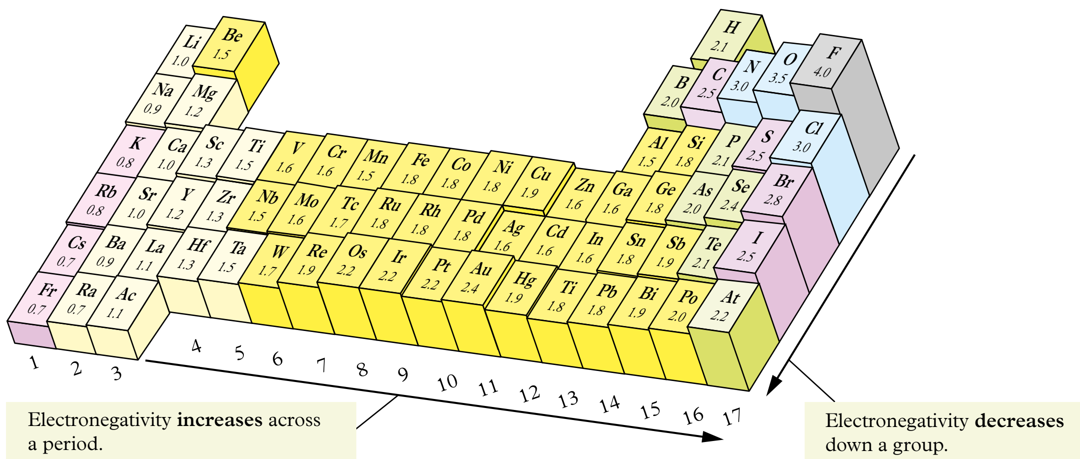


**FIGURE 5** The electronegativities of elements on the periodic table.

**WORKED EXAMPLE 6.1**

Consider the chemical reaction between fluorine and lithium to form lithium fluoride:

$$F_2(g) + 2Li(s) \rightarrow 2LiF(s)$$

1 Draw the Bohr electron configurations of lithium and fluorine and show the electron transfer.
2 Identify the chemicals that gain and lose electrons.
3 Identify the chemicals oxidised and reduced.
4 Identify the oxidising agent and the reducing agent.

**SOLUTION**

1

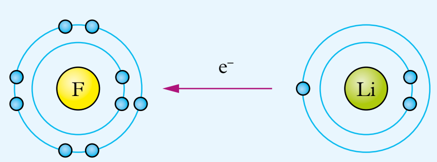


**FIGURE 6** The reaction of fluorine with lithium to form lithium fluoride.

2 Fluorine gains an electron. Lithium loses an electron.
3 Lithium is oxidised. Fluorine is reduced.
4 Fluorine is the oxidising agent. Lithium is the reducing agent.



MANDATORY PRACTICAL
6.1 Single displacement reaction
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# Redox can be modelled using a range of reactions

Several different types of chemical reactions can be regarded as redox reactions. Because redox reactions involve a transfer of electrons, they can often be identified by reactions involving elements, ions and metals, but not always. Reactions that involve a transfer of electrons include:

- displacement reactions
- **combustion** reactions
- **corrosion** reactions.

**combustion** a chemical reaction with oxygen to form a metal oxide, a covalent compound or carbon dioxide and water

**corrosion** the degradation of a metal to form a more stable metal oxide when exposed to gases and liquids

## Displacement reactions of metals

Displacement reactions of metals are often called **single displacement** reactions. This type of reaction is covered in detail in Chapter 9 of *Chemistry for Queensland Units 1 & 2*. Single displacement reactions occur when a stronger reducing agent replaces a weaker reducing agent. Figure 7 shows the reactivity series of the metals.

**single displacement** a chemical reaction in a more reactive metal ion replaces a less reactive metal ion in a molecule

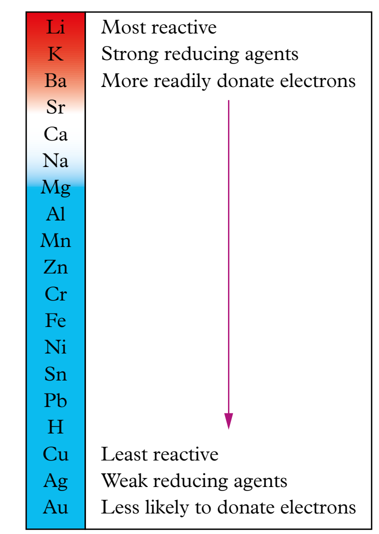


**FIGURE 7** A metal reactivity chart.

Consider the reaction of magnesium with zinc nitrate:

$$Zn(NO_3)_2(aq) + Mg(s) \rightarrow Mg(NO_3)_2(aq) + Zn(s)$$

Magnesium is a stronger reducing agent – it gives away its electrons more readily than zinc does. Therefore, magnesium displaces the zinc atom (knocks it off the nitrate ion) to form magnesium nitrate and zinc metal.

Zinc is the cation in an ionic compound. It gains electrons to form zinc metal. Therefore, zinc undergoes reduction.

Magnesium metal loses electrons to form a positive metal cation in the ionic compound. Therefore, magnesium undergoes oxidation.

## Combustion

A combustion reaction is a reaction with oxygen. Metals and non-metals react in very different ways to form an ionic or a covalent compound.

A metal combusts (reacts with oxygen) to form a metal oxide, according to the following equation:

$$\text{metal} + \text{oxygen} \rightarrow \text{metal oxide}$$

Consider the reaction between copper and oxygen:

$$2Cu(s) + O_2(g) \rightarrow 2CuO(s)$$

Copper metal loses electrons to form a positive metal cation. Therefore, copper undergoes oxidation. Oxygen gas gains electrons to form a negative non-metal anion. Therefore, oxygen undergoes reduction.

## Corrosion

Corrosion is a process in which metals react with chemicals in the atmosphere (including rain and atmospheric gases such as oxygen and carbon dioxide) to form a more stable chemical compound. For the reaction to occur, the metal must be more stable as an ion than as a solid metal. Thus, the metal readily donates an electron according to the following general equation (where M represents the metal):

$$M(s) \rightarrow M^+(aq) + e^-$$

For example, iron is a highly corrosive metal. When water interacts with the surface of the iron metal (Figure 8), iron loses two electrons to form aqueous iron ions and is therefore oxidised according to the following chemical equation:

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

Oxidation cannot occur without reduction, so the reduction occurs at the surface of the water droplet where water reacts with oxygen to form hydroxide ions:

$$O_2(g) + H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

The iron ions and hydroxide ions react with each other to form iron(II) hydroxide:

$$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(aq)$$

The iron hydroxide, which is now a part of the water droplet solution, also reacts with oxygen at the surface to form iron(III) hydroxide:

$$4Fe(OH)_2(aq) + O_2(g) + 2H_2O(l) \rightarrow 4Fe(OH)_3(aq)$$

The iron(III) hydroxide decomposes because of the presence of oxygen to form a brown hydrated iron oxide complex:

$$4Fe(OH)_3(aq) \rightarrow Fe_2O_3{\cdot}H_2O(aq)$$

Iron oxide is the chemical that gives rust its characteristic brown colour (Figure 8).

**FIGURE 8** The corrosion of iron metal forms rust, seen here as red-brown patches.





### CASE STUDY 6.1

#### Why is it called 'reduction' if it is a gain in electrons?

The earliest chemists realised that they could extract pure metals in metal ore from the ground by melting (or smelting) the ore. The metal obtained at the end of the process had less mass than the original ore, so these chemists called the process 'reduction'.

The French chemist Antoine Lavoisier (1743–1794) was the first to determine that this loss of mass was due to the ore losing oxygen. At the time, scientists did not know what electrons were and they were only developing an understanding of what atoms were. Therefore, it is understandable that the process was named without any true understanding of the electrons that are transferred from one atom to another.

Today, students struggle with the idea of a gain in electrons being called reduction. It is only when a deeper level of understanding is obtained, and oxidation numbers are investigated, that a link between the term 'reduction' and a decrease in oxidation number can be made (see section 6.2).

**FIGURE 9** Smelting of copper ore to make copper plates.

## CHALLENGE 6.1

### Ionisation energy, electronegativity and redox

How does the ionisation energies of atoms and their electronegativities relate to the ability of an atom to gain or lose electrons and therefore their strength as an oxidant or reductant? Use francium and fluorine as examples to support your answer.

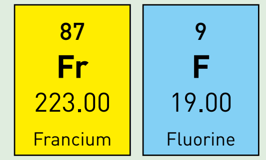


**FIGURE 10** Fracium and fluorine are both chemical elements.

## CHECK YOUR LEARNING 6.1

### Describe and explain

1 **Explain** the processes of reduction and oxidation. How are they related?

2 **Describe** what happens to the valence shell electrons during redox reactions.

### Apply, analyse and interpret

3 **Use** your knowledge of valence shell diagrams to demonstrate the transfer of electrons between the following pairs of atoms.

- **a** Lithium metal and sulfur (S)
- **b** Magnesium metal and fluorine ($F_2$)
- **c** Aluminium metal and oxygen ($O_2$)

**Determine:**

- which atom is oxidised and which atom is reduced
- the reducing and oxidising agent
- the product of the reaction.

4 **Determine** the oxidised and reduced species in the following reactions.

- **a** $H_2(g) + Cl_2(g) \rightarrow 2HCl(aq)$
- **b** $Cu(NO_3)_2(aq) + Fe(s) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$
- **c** $2AgNO_3(aq) + Pb \rightarrow Pb(NO_3)_2(aq) + 2Ag(s)$

### Investigate, evaluate and communicate

5 Oxidation is a name given to many chemical reactions. It can also be defined as a gain in oxygen or a loss of hydrogen. **Evaluate** these definitions and provide examples of these types of reactions.

Using this information, **devise** a more accurate definition for oxidation and reduction.

6 Your teacher claimed that there are circumstances when the corrosion of a metal can be beneficial, but does not explain why this is so. Your friend was absent from this chemistry lesson and asks you to explain it to them. **Investigate** a circumstance where this is true and **devise** the answer that you would give to your friend.

You can find the following resources for this section on your obook assess:

- » Student book questions
  Check your learning 6.1
- » Mandatory practical
  6.1 Single displacement reactions
- » Challenge
  6.1 Ionisation energy, electronegativity and redox
- » Weblink
  Redox





# 6.2 Oxidation numbers

**KEY IDEAS**

In this section, you will learn about:

- the rules for assigning oxidation numbers
- the use of oxidation numbers to determine which chemical species is oxidised and which is reduced.

**overall equation** a reaction that combines the two half-equations after electrons have been balanced and cancelled out

**oxidation state** oxidation number

Often, redox reactions are complex, and the **overall equation** involves more than two chemical species. For this reason, it can be difficult to identify which species have gained or lost electrons without identifying their **oxidation states** (or oxidation numbers) before and after reacting.

## Assigning oxidation numbers

The oxidation number (often called oxidation state) of an element can help you to determine whether electrons have been gained or lost. Oxidation numbers do not always represent the charge of individual chemical elements but are used to keep track of how many electrons an atom has. Oxidation numbers can be calculated for elements, ions or covalent molecules. Table 1 outlines the rules for assigning oxidation numbers to atoms.

**TABLE 1** Rules for assigning oxidation numbers

| Rule | Examples |
|---|---|
| Elements have an oxidation number of 0. | $O_2$, $F_2$, He, Fe, Zn, Li |
| Certain elements when present in compounds have common oxidation number. | Group 1 metals are always +1 ($Li^+$, $Na^+$, $K^+$).<br>Group 2 metals are always +2 ($Mg^{2+}$, $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$).<br>Hydrogen is +1 (except in metal hydrides, where it is –1, e.g. LiH).<br>Oxygen is –2 (except in peroxide ($H_2O_2$), where it is –1). |
| For monatomic ions, the oxidation number is given by the charge on the ion. | $Cu^{2+}$ has an oxidation number of +2.<br>$Na^+$ has an oxidation number of +1. |
| In polyatomic ions, the sum of oxidation numbers is equal to the charge of the ion. | In $PO_4^{3-}$, the sum of the oxidation numbers is –3. P + (4 × O) = –3.<br>Oxygen is –2 and 4 × –2 = –8.<br>Therefore, P + –8 = –3, and P has an oxidation number of +5. |
| In a neutral compound, the sum of oxidation numbers is equal to 0. | HCl is a neutral compound and has an oxidation number of 0. As hydrogen is +1, chlorine must be –1. |
| The most electronegative element has a negative oxidation number. | $NO_2$ is a neutral compound. Oxygen is more electronegative and therefore has a negative oxidation number. Oxygen is –2 and 2 × –2 = –2. The sum of the oxidation states is 0, so nitrogen has an oxidation state of +4. |

### Oxidation and reduction using oxidation numbers

**oxidation number** the number of electrons gained or lost by an atom

When an atom gains one or more electrons, it gains some negative charge and its oxidation state decreases. Therefore, reduction, a gain in electrons, causes a decrease in the **oxidation number**. In reduction, the oxidation number reduces.

When an atom loses one or more electrons, it loses some negative charge and consequently becomes more positive. Therefore, oxidation, a loss of electrons, causes an increase in oxidation number. In oxidation, the oxidation number increases.

Worked examples 6.2A, 6.2B and 6.2C show you how to work out oxidation states.

**WORKED EXAMPLE 6.2A**

Determine the oxidation state of a magnesium ion.

**SOLUTION**

1 On the periodic table, magnesium is in group 2.
2 When Mg becomes an ion, it loses 2 valence electrons, resulting in a charge of 2+ ($Mg^{2+}$).
3 Therefore, the oxidation state of a magnesium ion is +2.

**WORKED EXAMPLE 6.2B**

Determine the oxidation state of chlorine gas ($Cl_2$).

**SOLUTION**

1 $Cl_2(g)$ is an uncharged molecule, as it is in its elemental form.
2 Therefore, the oxidation state of a chlorine molecule is 0.

**WORKED EXAMPLE 6.2C**

Determine the oxidation state of all elements in a water molecule ($H_2O$).

**SOLUTION**

1 $H_2O$ is an uncharged compound; therefore, the sum of all oxidation numbers must be 0.

$$H_2O = 0$$
$$(2 \times H) + (1 \times O) = 0$$

2 The oxidation state of oxygen is always –2 (unless in hydrogen peroxide) and the oxidation state of hydrogen is +1 (unless in a metal hydride). Because there are 2 hydrogens, their oxidation states must add to +2:

$$(2 \times +1) + (1 \times -2) = 0$$

3 Therefore, hydrogen has an oxidation state of +1 and oxygen has an oxidation state of –2.



## Oxidation states of transition metals

Transition metals have multiple oxidation states. This is true of several non-transition metals as well. When determining oxidation numbers, if there is no rule for an atom then its oxidation number must be determined using other atoms in the chemical species.

For example, permanganate exists as $MnO_4^-$. To determine the oxidation state of the manganese atom, you must determine all other pieces of information.

- The molecular ion has a charge of 1–. Therefore, its oxidation state is –1.
- The oxidation state of oxygen is –2, unless in peroxide.

$$MnO_4 = -1$$
$$Mn + (4 \times -2) = -1$$
$$Mn + (-8) = -1$$
$$Mn + (-8) + 8 = -1 + 8$$
$$Mn = +7$$

Therefore, the oxidation state of manganese is +7.

**Study tip**

The terms 'oxidation number' and 'oxidation state' are interchangeable. They mean the same thing.



Oxidation states of transition metals are indicated by roman numerals in brackets after the atom. For example, manganese +7 is written manganese(VII) (or $Mn^{VII}$). Roman numerals are also used in the name of a molecule to identify the oxidation state of the metal.

For example:

- copper has an oxidation state of +1 or +2 (rarely +3). Therefore, it is $Cu^{I}$ or $Cu^{II}$. When naming copper in copper sulfate, it is either copper(I) sulfate or copper(II) sulfate
- iron has an oxidation state of +2 or +3. Therefore, it is $Fe^{II}$ or $Fe^{III}$. When naming iron in iron nitrate, it is either iron(II) nitrate or iron(III) nitrate.

Worked example 6.2D shows you how to determine the oxidation state of a transition metal.

The corroded layer formed a protective coat on the surface of the statue, not unlike a coat of paint, protecting it from further damage.

It took 30 years to form the outer patina (green coating) and it is now a part of the iconic appearance of the Statue of Liberty.

**Study tip**

Molecules with a charge of 1– or 1+ are just written as – or + (e.g. $Cl^-$ or $K^+$).

**WORKED EXAMPLE 6.2D**

Determine the oxidation state of all elements in dichromate ($Cr_2O_7^{2-}$).

**SOLUTION**

1 $Cr_2O_7^{2-}$ is a charged compound; therefore, the sum of all oxidation numbers must be equal to its charge (–2).

$$Cr_2O_7^{2-} = -2$$
$$(2 \times Cr) + (7 \times O) = -2$$

2 The oxidation state of oxygen is always –2 (unless in hydrogen peroxide); chromium is a transition metal and has multiple oxidation states. Therefore, it must be determined.

$$(2 \times Cr) + (7 \times -2) = -2$$
$$(2 \times Cr) + (-14) = -2$$
$$(2 \times Cr) + (-14) + 14 = -2 + 14$$
$$2 \times Cr = +12$$

3 Therefore, chromium has an oxidation state of +6.

Because chromium is a transition metal, it can be identified as Cr(VI).

**CHALLENGE 6.2**

### Oxidation of alcohol

The alcohol ethanol undergoes oxidation by the permanganate ion in an acidic environment according to the equation:

$$3CH_3CH_2OH(aq) + 2MnO_4^-(aq) + 4H^+(aq)$$
$$\rightarrow 3CH_3COOH(aq) + 2Mn^{2+}(aq) + 5H_2O(l)$$

Identify the species that is oxidised and the species that is reduced as well as the oxidising and reducing agent.

## CASE STUDY 6.2

### What happened to the Statue of Liberty?

The Statue of Liberty was gifted to the American people by the French people and was erected in New York Harbour on 19 June 1885. The statue has an outer coating of copper, about the thickness of two Australian 20-cent coins placed together. The internal metals are cast iron and stainless steel. When it was first erected, the statue was brown due to the external copper coating.

Over time, the copper reacted with oxygen in the air, corroding to form copper oxide. Copper loses electrons to oxygen, forming $Cu^{+}$ and $O^{2-}$. Therefore, copper is oxidised and oxygen is reduced, forming $Cu_2O$ a red solid.

$$4Cu(s) + O_2(g) \rightarrow 2Cu_2O(s)$$

The $Cu^{+}$ in the $Cu_2O$ is further oxidised, forming $Cu^{2+}$ in CuO – a black solid.

$$2Cu_2O(s) + O_2(g) \rightarrow 4CuO(s)$$

In the late 1800s, the large amount of coal that was burnt released sulfur dioxide and carbon dioxide into the atmosphere. This caused further reactions such as the precipitation of the copper-based minerals malachite (green), azurite (blue) and brochantite (green).

$$2CuO(s) + CO_2(g) + H_2O(l) \rightarrow Cu_2CO_3(OH)_2(s)$$

malachite

$$3CuO(s) + CO_2(g) + H_2O(l) \rightarrow Cu_3(CO_3)_2(OH)_2(s)$$

azurite

$$4CuO(s) + SO_3(g) + 3H_2O(l) \rightarrow Cu_4SO_4(OH)_6(s)$$

brochantite

**FIGURE 1** Malachite is a green mineral with formula $Cu_2CO_3(OH)_2$.

**FIGURE 2** Azurite is a blue mineral with formula $Cu_3(CO_3)_2(OH)_2$.

**FIGURE 3** Brochantite is a green mineral with formula $Cu_4SO_4(OH)_6$.





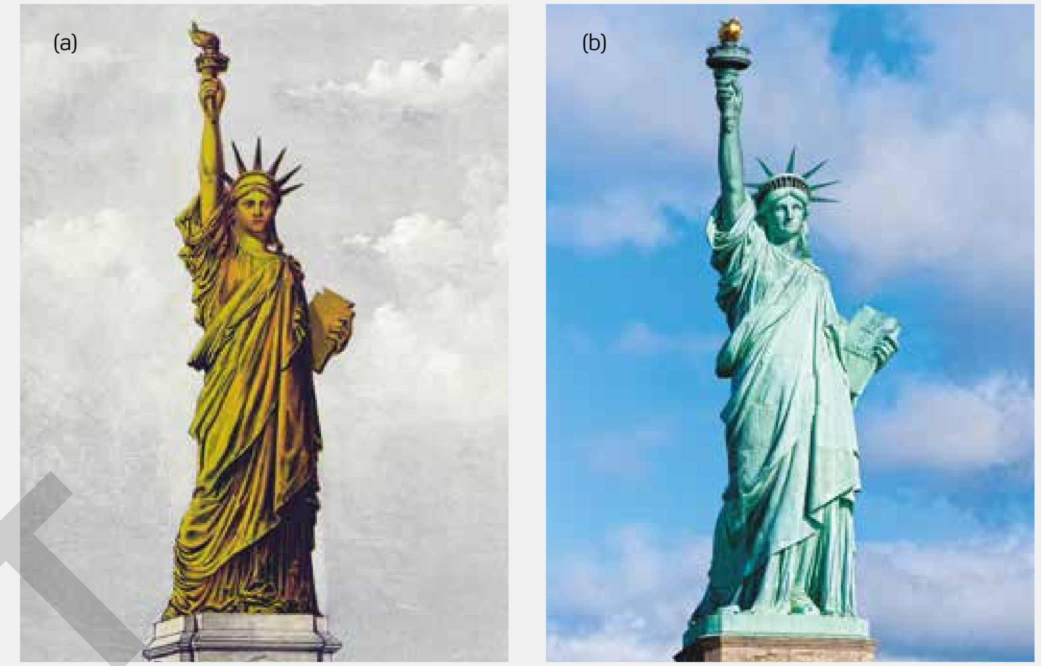


**FIGURE 4** The Statue of Liberty (a) as it was originally erected in 1886 and (b) as it is today.

## CHECK YOUR LEARNING 6.2

### Describe and explain

1 **Define** 'oxidation state'.

2 **Explain** what happens to oxidation numbers in reduction reactions.

### Apply, analyse and interpret

3 **Determine** the oxidation numbers of the atoms in the following chemical substances.

- **a** $O_2$
- **b** $NO_2$
- **c** $SO_4^{2-}$
- **d** $CH_3COO^-$
- **e** NaOH
- **f** $H_2O_2$
- **g** NaH
- **h** $PO_4^{3-}$

4 **Determine** the oxidised and reduced atoms in the following equations. Use their oxidation numbers to **justify** your answers.

- **a** $2Fe(OH)_3(aq) + 3OCl^-(aq) \rightarrow 2FeO_4^{2-}(aq) + 3Cl^-(aq) + H_2O(l) + 4H^+(aq)$
- **b** $Fe^{2+}(aq) + 6H^+(aq) + VO_4^{3-} \rightarrow Fe^{3+}(aq) + VO^{2+}(aq) + 3H_2O(l)$
- **c** $2Cr_2O_7^{2-}(aq) + 16H^+(aq) + C_2H_5OH(l) \rightarrow 4Cr^{3+}(aq) + 2CO_2(g) + 11H_2O(l)$

### Investigate, evaluate and communicate

5 **Investigate** steel and stainless steel.

- **a** What are both materials made of?
- **b** What are they used for?
- **c** Do they corrode? **Justify** your answer with a chemical equation or an explanation.
- **d** What are the advantages and disadvantages of using both materials?

6 In an experiment, copper metal was placed in a solution of silver nitrate. A silver metal and a blue solution of copper(II) nitrate formed. **Evaluate** these results and determine the oxidation and reduction equations as well as a balanced overall equation.

**You can find the following resources for this section on your obook assess:**

- » Student book questions: Check your learning 6.2
- » Challenge: 6.2 Oxidation of alcohol
- » Weblink: Corrosion of steel
- » Weblink: Corrosion of stainless steel

# 6.3 Half-equations and overall redox equations

**KEY IDEAS**

In this section, you will learn about:

- oxidation and reduction half-equations
- combining oxidation and reduction half-equations to develop an overall redox equation.

**half-equation**
an equation that represents either an oxidation or a reduction half of a chemical equation; it includes electrons to demonstrate electron transfer

Redox chemical equations can become quite complicated because of the number of chemical species involved. For this reason, redox equations are balanced in two **half-equations** – the oxidation half-equation and the reduction half-equation. These equations are then combined to form an overall chemical redox equation.

## Identifying and writing half-equations

The oxidation half-equation demonstrates an atom losing electrons, while the reduction half-equation demonstrates an atom gaining electrons.

The following reaction between copper(II) sulfate and zinc metal shows you how to identify the chemical species being oxidised and reduced and how to write the oxidation and reduction half-equations from an overall equation.

$$CuSO_4(aq) + Zn(s) \rightarrow Cu(s) + ZnSO_4(aq)$$

Assign oxidation states to identify chemicals being oxidised and reduced. In this case, it is easier to deal with sulfate because its oxidation state is equal to its charge of –2.

$$\overset{+2\ -2}{CuSO_4}(aq) + \overset{0}{Zn}(s) \rightarrow \overset{0}{Cu}(s) + \overset{+2\ -2}{ZnSO_4}(aq)$$

**spectator ion**
an ion that has no change in oxidation state from the left to the right side of a redox reaction

Both copper and zinc are participating in the chemical equation because they are changing oxidation states. Sulfate is a **spectator ion** and so is excluded from the half-equation.

The oxidation state of copper decreases from +2 to 0. Therefore, copper undergoes reduction. To represent its half-equation, rewrite the copper species:

$$\overset{+2}{Cu^{2+}}(aq) \rightarrow \overset{0}{Cu}(s)$$

Although copper is balanced in the equation, the charge is not. To balance charge, add electrons to one side of the equation. Electrons are negative, so they will decrease the charge on whichever side they are added. In the case of the copper half-equation, the reactant has a charge of +2 and the product has a charge of 0. Therefore, add a charge of –2 to the reactant side to balance the equation.

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$$

The oxidation state of zinc increases from 0 to +2. Therefore, it undergoes oxidation.

$$\overset{0}{Zn}(s) \rightarrow \overset{+2}{Zn^{2+}}(aq)$$

Balance the +2 charge on the product side by adding two electrons.

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





## Overall redox equations

When you have identified the oxidation and reduction half-equations, combined them to form an overall redox equation. To do this, both half-equations must have the same number of electrons so that they can cancel out in the final equation.

The zinc and copper half-equations combine easily because they both have two electrons. When you combine the half-equations, the electrons can be cancelled out to give the following balanced equation:

$$Cu^{2+}(aq) + \cancel{2e^-} + Zn(s) \rightarrow Cu(s) + Zn^{2+}(aq) + \cancel{2e^-}$$

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

An additional example using aluminium ions and magnesium metals can be found on your obook assess.

**WORKED EXAMPLE 6.3**

Identify the oxidation and reduction half-equations and the overall redox equation in the following reaction:

$$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$$

**SOLUTION**

1 Assign oxidation states to identify the atoms that have been oxidised and reduced.

$$\overset{0}{Fe}(s) + \overset{+1\ -1}{2HCl}(aq) \rightarrow \overset{+2\ -1}{FeCl_2}(aq) + \overset{0}{H_2}(g)$$

Chlorine has no change in oxidation state so it is a spectator in the reaction.

2 The oxidation state of iron increases from 0 to +2, so iron is oxidised. The oxidation half-equation is:

$$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$$

3 The oxidation state of hydrogen decreases from +1 to 0, so hydrogen is reduced. $H_2$ is formed, so balance the hydrogens before adding electrons. The reduction half-equation is:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

4 Combine the reduction and oxidation half-equations to get the overall equation (you do not need to multiply these equations because they have the same number of electrons):

$$\text{Overall: } Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$$

**Study tip**

When combining half-equations, they must have the same number of electrons. Multiply the coefficients of the electrons in both equations if the charges are not balanced.

**CHALLENGE 6.3A**

### Oxidation of butanol

Butanol is oxidised to butanoic acid by a solution of acidified sodium dichromate according to the following half-equations:

$$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$$

$$CH_3CH_2CH_2CH_2OH(aq) + H_2O(l) \rightarrow CH_3CH_2CH_2COOH(aq) + 4H^+(aq) + 4e^-$$

Combine the half-equations and write the balanced overall equation for the reaction.

## CASE STUDY 6.3

### Photography

Today, taking a photo is as simple as using your smartphone or digital camera. However, 100 years ago, photography relied on a camera containing a piece of photographic film. When exposed to light, a chemical reaction would occur on the film, resulting in an image. This film was refined and improved until the invention of digital cameras, which became commercially available in the late 1980s.

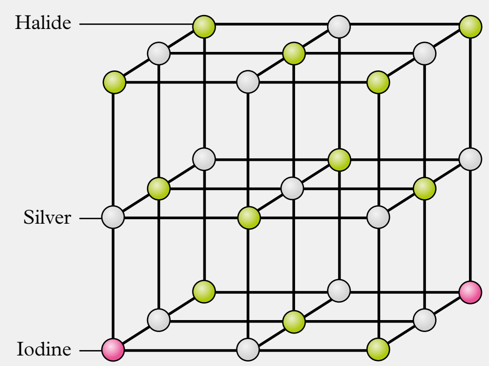


**FIGURE 1** The structure of silver halide crystals.

Photographic film is covered with an emulsion that contains light-sensitive silver halide crystals. The crystal is formed from a 3D structure containing silver, bromide, chloride and iodide ions (Figure 1).

Photographic film works when photons of light strike the silver bromide on the film and initiate a redox reaction. The photon causes an electron to dissociate from the bromide ion. The silver ions migrate towards the free electron (where the light has struck the film) forming solid silver (Ag). Where the solid silver forms, a latent image is located. An increase in light in an area of the film means that more silver will precipitate in that location.

The overall redox equation for this process is:

$$AgBr(s) \rightarrow Ag(s) + Br_2(g)$$

The film is kept in a darkened casing until it is developed. If it is exposed to light before it is developed, the film will be completely black when it is developed because all of the silver ions will have precipitated as silver metal.

Once the film has been developed (Figure 2) and the chemicals have been fixed, no more reaction occurs, and the film can be viewed in the light. Photographers sometimes manipulate the images on the film by scratching the darkened silver metal from the film for artistic effect.

**FIGURE 2** A developed film roll: the darkened areas represent where the silver has precipitated.



## CHALLENGE 6.3B

### Oxidation of olivine

Olivine, a heavy metal silicate, has two different forms, fayalite ($Fe_2SiO_4$) and fosterite ($Mg_2SiO_3$). Construct the two different redox reactions for the oxidation of olivine.

**FIGURE 3** Olivine crystals show two different colours depending on whether they are oxidised or not. Orange crystals are oxidised and green are not.

## CHECK YOUR LEARNING 6.3

### Describe and explain

1 **Describe** what all oxidation half-equations have in common.

2 **Describe** what all reduction half-equations have in common.

3 **Explain** why overall redox equations must have the same number of electrons on the reactant and product sides.

4 **Construct** the half-equations and overall equations for the following reactions. Identify the half-equations as reduction or oxidation.

   a $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$

   b $S(s) + F_2(g) \rightarrow SF_6(g)$

   c $N_2(g) + Cl_2(g) \rightarrow N_2Cl_4(g)$

   d $NO(g) + O_2(g) \rightarrow NO_2(g)$

5 **Identify** the oxidants and reductants in the reaction equations developed in question 4.

### Investigate, evaluate and communicate

6 Copper has been found in ancient burial sites and has been dated as early as 9000 BCE. **Investigate** how copper was smelted in ancient times compared to today. **Discuss** the key differences in the processes.

7 **Investigate** tarnished silverware and **create** a procedure for cleaning the silverware by using redox chemistry. Include chemical equations in your method.

**You can find the following resources for this section on your obook assess:**

- Student book questions
  Check your learning 6.3
- Challenge
  6.3A Oxidation of butanol
- Challenge
  6.3B Oxidation of olivine
- Increase your knowledge
  Extra overall redox equation

SCIENCE AS A HUMAN ENDEAVOUR

# 6.4 Redox reactions in breathalysers

**KEY IDEAS**

In this section, you will learn about:

+ the incidence of drink driving in Australia and Queensland
+ the chemistry involved with breathalysers.

**BAC**
blood alcohol concentration

Drink driving has always been a problem on Australian roads. In 2015, Australia was ranked fourth for the percentage of road incident deaths involving alcohol, with 30% (Figure 1) of road fatalities connected to alcohol consumption and blood alcohol concentration (**BAC**).

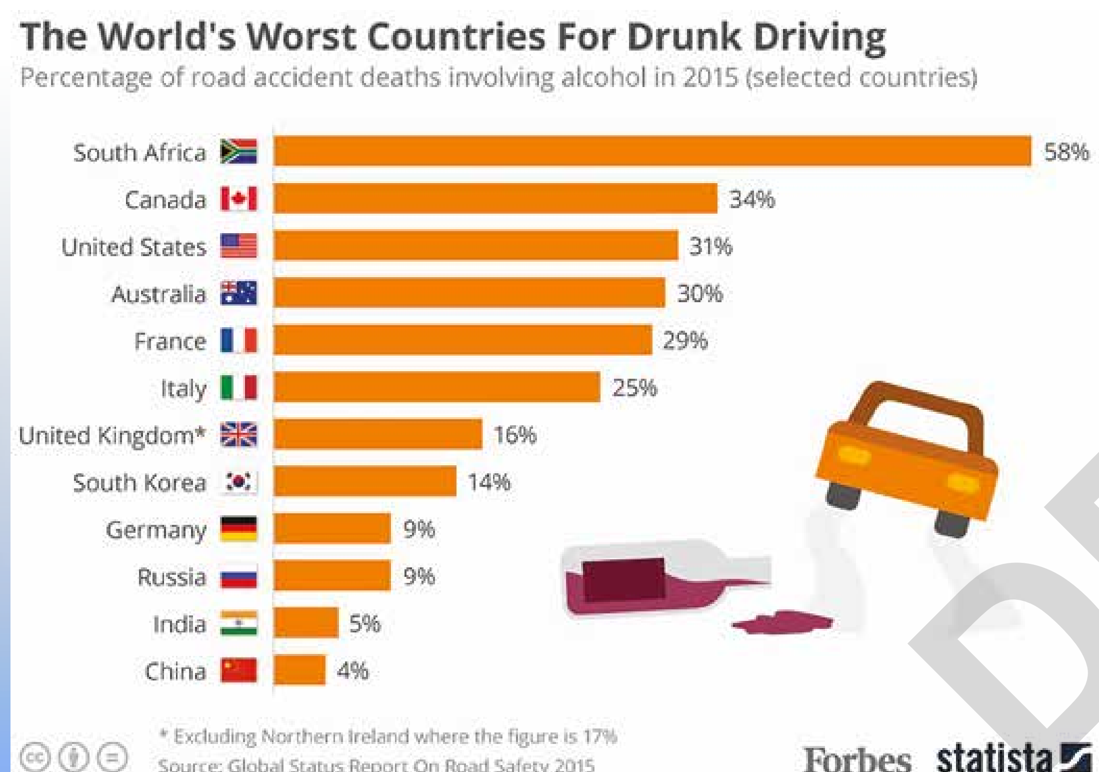


**FIGURE 1** The percentage of road deaths caused by drink driving around the world.

## Incidence of drink driving in Queensland

**breathalyser**
a device for determining the BAC of a motorist

In Queensland, the incidence of drink driving has decreased since the introduction of **breathalysers** in 1982 and their formal commencement in 1989. The 2016–2017 Annual Statistical Review, published by the State of Queensland (Queensland Police Service), reported statistics for drink-driving offences since 1980 (Figure 2).



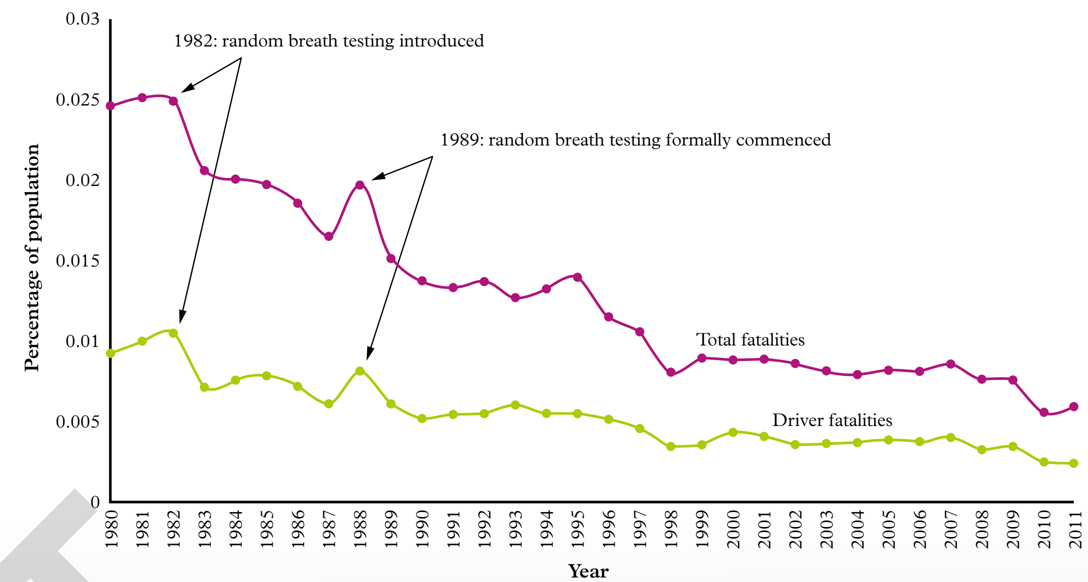


**FIGURE 2** Incidence of fatalities due to drink driving in Queensland between 1980 and 2011.



## Blood alcohol concentration and breathalysers

Henry's law states that the amount of dissolved gas is proportional to its partial pressure in the gas phase. In terms of BAC, this means that the amount of alcohol that is present in the blood is proportional to the amount of alcohol in a motorist's breath. Therefore, if you measure the alcohol concentration on the breath, you can determine the blood alcohol concentration. This will tell you whether it is safe to operate heavy machinery, such as cars, motorbikes and trucks. Instruments that measure the amount of alcohol in a motorist's breath are called breathalysers.

### Drunkometer and Intoximeter – using permanganate

The first breathalysers, called the Drunkometer and Intoximeter, were developed in the US in the 1930s by scientists Rolla Harger and Glen Forrester. The motorist blew into a balloon located within the device, which was then pumped through a solution of acidified potassium permanganate. Potassium permanganate is a dark purple colour but becomes colourless when the permanganate ion ($MnO_4^-$) is reduced to the manganese ion ($Mn^{2+}$). The half-equation for this is:

$$\underset{\text{purple}}{MnO_4^-(aq)} + 8H^+(aq) + 5e^- \rightarrow \underset{\text{colourless}}{Mn^{2+}(aq)} + 4H_2O(l)$$

Any alcohol (ethanol) present in the breath sample is oxidised to ethanoic acid. This reaction uses some of the permanganate and the solution loses colour and becomes paler:

$$\underset{\text{ethanol}}{CH_3CH_2OH(aq)} + H_2O(l) \rightarrow \underset{\text{ethanoic acid}}{CH_3COOH(aq)} + 4H^+(aq) + 4e^-$$

The greater the colour change from purple to pale purple and eventually to clear/colourless, the more alcohol is present in the breath sample. The overall redox equation for this process is:

$$4MnO_4^-(aq) + 5CH_3CH_2OH(aq) + 6H^+(aq) \rightarrow 5CH_3COOH(aq) + 4Mn^{2+}(aq) + 11H_2O(l)$$

## Alcometer – using starch

The Alcometer was an improvement on the permanganate method. In this instrument, a solution of iodine was reduced to iodide. In the presence of a starch indicator, the iodine solution is clear or white, but becomes blue in the presence of iodide:

$$\underset{\text{clear–white}}{I_2(s)} + 2e^- \rightarrow \underset{\text{blue}}{2I^-(aq)}$$

The ethanol is oxidised according to the same half-equation, and the overall redox reaction for the process is:

$$CH_3CH_2OH(aq) + H_2O(l) + 2I_2(s) \rightarrow CH_3COOH(aq) + 4H^+(aq) + 4I^-(aq)$$

This method to determine BAC proved less stable and less reliable than the original instruments. So it was quickly discarded in favour of better alternatives.

## Photocells – using dichromate

After the 1940s, photocells were used to detect the colour change more accurately. In this instrument, acidified dichromate was used instead of permanganate. The dichromate ion ($Cr_2O7_2^-$) appears yellow–orange and undergoes reduction in an acidified solution to form the chromium ion ($Cr^{3+}$), which is green:

$$\underset{\text{orange–yellow}}{Cr_2O_7^{2-}(aq)} + 14H^+(aq) + 6e^- \rightarrow \underset{\text{green}}{2Cr^{3+}(aq)} + 7H_2O(l)$$

**FIGURE 3** A photocell is a sensor that allows you to detect light.

The ethanol is again oxidised to form ethanoic acid:

$$\underset{\text{ethanol}}{CH_3CH_2OH(aq)} + H_2O(l) \rightarrow \underset{\text{ethanoic acid}}{CH_3COOH(aq)} + 4H^+(aq) + 4e^-$$

The overall chemical reaction for the process is:

$$3CH_3CH_2OH(g) + 2Cr_2O_7^{2-}(aq) + 16H^+(aq) \rightarrow 3CH_3COOH(aq) + 2Cr^{3+}(aq) + 11H_2O(l)$$

Once the dichromate has reacted with the ethanol in the breath sample, the resultant solution is compared to a separate sample of unreacted dichromate. The colours are compared by using a photocell. The photocell creates an electric current that is proportional to the amount of colour change.

However, organic substances similar to ethanol that are consumed as a part of a balanced diet often gave false readings in permanganate and dichromate tests. All alcohols and aldehydes are capable of being oxidised by the same chemical methods.



## Electrochemical cells – measuring the flow of electrons

To avoid using colour-based indicators, a fuel cell was developed as an alternative. The half-equations involve the transfer of electrons, so an electric current is produced. This current can be measured and used to determine the concentration of ethanol in the breath and therefore the blood.

Atmospheric oxygen is oxidised in acidic conditions to form water:

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

The ethanol again undergoes reduction to produce the following overall redox reaction:

$$CH_3CH_2OH(l) + O_2(g) \rightarrow CH_3COOH(l) + H_2O(l)$$

**CHALLENGE 6.4**

### False positive BAC readings

What products that are available in supermarkets may interfere with a BAC reading? How do they effect the BAC and which ingredients are responsible for this effect?



**CHECK YOUR LEARNING 6.4**

**Describe and explain**

1 **Explain** what a breathalyser measures.

2 **Define** the limit for blood alcohol concentration in Queensland.

**Apply, analyse and interpret**

3 **Apply** your understanding of spectroscopy to **determine** why a permanganate and dichromate breathalyser must be frequently calibrated.

4 There have been many myths about how to cheat a breathalyser. **Determine** if any of these be accurate from your knowledge of BAC.

**Investigate, evaluate and communicate**

5 **Research** the factors that influence an individual's BAC. **Describe** what these factors are and **determine** how they affect an individual's BAC.

6 In Australia, you receive a fine if you refuse to take a breathalyser test. However, in the US, a motorist can request a blood test rather than blowing into a breathalyser. **Investigate** both types of testing and communicate the differences of each technique. **Describe** the potential advantages and disadvantages for the motorist.

**You can find the following resources for this section on your obook assess:**

- » Student book questions
  Check your learning 6.4
- » Challenge
  6.4 False positive BAC readings
- » Weblink
  Blood alcohol levels

CHAPTER 6

# Review

## Chapter summary

6.1
- Redox reactions occur when electrons are transferred from one reactant to another. A reactant undergoes oxidation if it loses electrons (OIL) or undergoes reduction if it gains electrons (RIG).
- Electrons move between the valence shells of the reactants. The reactant that loses one or more electrons causes reduction in the other reactant and so is called the reductant. The other reactant takes one or more electrons, causing oxidation in the other reactant and so is called the oxidant.
- Displacement, oxidation and corrosion are examples of redox reactions.

6.2
- Oxidation numbers are used to determine if oxidation or reduction has occurred. There are rules to follow to determine the oxidation number of an atom. If the oxidation number decreases or reduces, then reduction has occurred. If the oxidation number increases, then oxidation has occurred.
- The oxidation states of transition metals are represented by roman numerals after the element's name or symbol, e.g. chromium(III), Cr(III).

6.3
- A half-equation represents one half of a redox equation, either the oxidation or the reduction reaction. Half-equations show electrons being gained or lost. Half-equations do not include spectator ions and can be combined to form overall redox reactions.
- When combining the reactions, the number of electrons must balance. The whole half-equation must be multiplied to ensure that both half-equations have the same number of electrons before combining them.

6.4
- Redox reactions are used in breathalysers to determine blood alcohol content, which is proportional to the amount of gaseous alcohol on the breath.

## Key terms

- BAC
- breathalyser
- combustion
- corrosion
- electronegativity
- half-equation
- ionisation energy
- overall equation
- oxidation
- oxidation number
- oxidation state
- oxidise
- oxidising agent
- redox
- reduce
- reducing agent
- reduction
- single displacement
- spectator ion

## Key formulas

| | |
|---|---|
| Metal combusts to form a metal oxide | metal + oxygen → metal oxide |



## Revision questions

The relative difficulty of these questions is indicated by the number of stars beside each question number: ★ = low; ★★ = medium; ★★★ = high.

### Multiple choice

The following equation relates to questions 1 and 2.

$$2Au^{3+}(l) + 6Br^{-}(l) \rightarrow 2Au(s) + 3Br_2(l)$$

**1** Identify the atoms that undergo oxidation and reduction.

**A** Au is oxidised; $Br_2$ is reduced.
**B** $Br_2$ is oxidised; Au is reduced.
**C** $Au^{3+}$ is oxidised; $Br^-$ is reduced.
**D** $Br^-$ is oxidised; $Au^{3+}$ is reduced.

**2** Identify the oxidant and reductant.

**A** Au is the oxidant; $Br_2$ is the reductant.
**B** $Br_2$ is the oxidant; Au is the reductant.
**C** $Au^{3+}$ is the oxidant; $Br^-$ is the reductant.
**D** $Br^-$ is the oxidant; $Au^{3+}$ is the reductant.

**FIGURE 1** Gold is an atom

The following equation relates to questions 3 and 4.

$$H_2Te(g) + Se(s) \rightarrow Te(s) + H_2S(g)$$

**3** Identify the atoms that undergo oxidation and reduction.

**A** $H_2Te$ is oxidised; Se is reduced.
**B** Se is oxidised; $H_2Te$ is reduced.
**C** Te is oxidised; $H_2S$ is reduced.
**D** $H_2S$ is oxidised; Te is reduced.

**4** Identify the oxidant and reductant.

**A** $H_2Te$ is the oxidant; Se is the reductant.
**B** Se is the oxidant; $H_2Te$ is the reductant.
**C** Te is the oxidant; $H_2S$ is the reductant.
**D** $H_2S$ is the oxidant; Te is the reductant.

**5** Identify the oxidation state of copper in $Cu_3(CO_3)_2(OH)_2$.

**A** +3
**B** +2
**C** +1
**D** –1

**6** Identify the oxidation state of phosphorus in $P_2O_7^{2-}$.

**A** +10
**B** +7
**C** +5
**D** +3.5

**7** Identify the oxidation half-equation.

**A** $NO_3^-(aq) + 4H^+(aq) + 3e^- \rightarrow NO(g) + 2H_2O(l)$
**B** $O_2(g) + 4H^+(aq) + 2e^- \rightarrow 2H_2O(l)$
**C** $AuCl_4^-(aq) + 3e^- \rightarrow Au(s) + 4Cl^-(aq)$
**D** $2\,Ag(s) + S^{2-}(aq) \rightarrow Ag_2S(s) + 2e^-$

**8** Identify the reduction half-equation.

**A** $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
**B** $Mn^{2+}(aq) + 4H_2O(l) \rightarrow MnO_4^-(aq) + 8H^+(aq) + 5e^-$
**C** $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
**D** $H_2(g) \rightarrow 2H^+(aq) + 2e^-$

**9** Reduction means there is:

**A** a loss of electrons.
**B** a gain of electrons.
**C** an electron in the outermost shell of an atom.
**D** a negative charge.

**10** For monatomic ions, the oxidation number is:

**A** the same as the charge on the ion.
**B** always +1.
**C** always neutral.
**D** always negative.

## Short answer

### Describe and explain

★ **11** **Describe** why a reducing agent is different to reduction.

★ **12** **Describe** the difference between charge of an ion and an oxidation state.

★ **13** **Explain** why corrosion is an example of oxidation.

★ **14** **Construct** the Bohr electron configuration of $Ca(s) + Cl_2(g) \rightarrow CaCl_2(s)$ and show the electron configuration.

★ **15** **Identify** the missing word in the following statement 'the attraction between a positively charged nucleus and the negatively charged electrons of a neighbouring atom is called ______________'.

★ **16** **Explain** why the term reduction is a gain of electrons.

★ **17** **Define** the term ionisation energy.

★ **18** **Describe** how the location of atoms on the periodic table affects their ability to gain or lose electrons.

★ **19** **Explain** the terms 'oxidising agent' and 'reducing agent'. How would you use them to **describe** redox reactions?

★ **20** **Describe** what is common to breathalyser redox reactions. Why does this commonality exist?


**FIGURE 2** A breathalyser can be used to estimate BAC from a breath sample.

★★ **21** A student placed a piece of iron metal into a solution of lead(II) nitrate. **Construct** two half-equations to demonstrate the oxidation and reduction processes. **Construct** a balanced overall equation to represent the chemical reaction.

★★ **22** A piece of solid nickel was placed in a solution of blue copper(II) sulfate. Over time, the solution became paler in colour and an orange coloured precipitate formed. **Construct** two half-equations to demonstrate the oxidation and reduction processes. **Construct** a balanced overall equation to represent the chemical reaction.

### Apply, analyse and interpret

★ **23** **Determine** the oxidation number of:

- **a** nitrogen in $NH_3$
- **b** silicon in $CaSiO_3$
- **c** chromium in $BaCrO_4$
- **d** vanadium in $VO_2^+$.

★ **24** **Determine** the oxidation number of:

- **a** cadmium in $Br_2CdO_6$
- **b** thorium in $ThO_2$
- **c** manganese in $MnCO_3$
- **d** zinc in $Zn(OH)_2$.

★ **25** **Determine** the oxidation number of iron in the following compounds:

- **a** $Fe_2O_3$
- **b** FeO
- **c** $Fe_2SiO_4$
- **d** Fe(s).

★★ **26** **Deduce** why iron can have many different oxidation numbers.

★★★ **27** **Consider** the following unbalanced chemical equation:

$$CO_2(g) + H_2 \rightarrow CO(g) + H_2O$$

- **a** **Identify** oxidation numbers of all atoms.
- **b** **Determine** which atom has been oxidised and which has been reduced.
- **c** **Determine** which atoms are the oxidising and reducing agents.
- **d** **Construct** the oxidation and reduction half-equations.
- **e** **Construct** the balanced overall redox reaction.



★★★ **28** **Consider** the following unbalanced chemical equation.

$$Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$$

- **a** **Identify** oxidation numbers for all atoms.
- **b** **Determine** which atom has been oxidised and which has been reduced.
- **c** **Determine** which atoms are the oxidising and reducing agents.
- **d** **Construct** the oxidation and reduction half-equations.
- **e** **Construct** the balanced overall redox reaction.

★★★ **29** **Consider** the following unbalanced chemical equation:

$$Ag(s) + H_2S \rightarrow Ag_2S(g) + H_2(g)$$

- **a** **Identify** oxidation numbers of all atoms.
- **b** **Determine** which atom has been oxidised and which has been reduced.
- **c** **Determine** which atoms are the oxidising and reducing agents.
- **d** **Construct** the oxidation and reduction half-equations.
- **e** **Construct** the balanced overall redox reaction.

### Investigate, evaluate and communicate

★★ **30** A student proposed that the best method to produce sodium hydroxide is the redox reactions between sodium metal and water to form sodium hydroxide and hydrogen gas. The half-equation for the water reaction is:

$$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$$

- **a** **Construct** the two half-equations to demonstrate the oxidation and reduction processes.
- **b** **Construct** a balanced overall equation to represent the chemical reaction.
- **c** **Evaluate** the student's claim and **discuss** the strengths and weaknesses of this reaction.


**FIGURE 3** Hydrogen gas

★★★ **31** A claim was made that the oxidation numbers of all atoms in an ytterbium complex ($YBa_2Cu_3O_7$) cannot be determined. **Evaluate** this claim and **discuss** your opinion, using the theory learnt in this chapter.

★★★ **32** A student placed a piece of silver metal in a blue solution of copper(II) sulfate. After 10 minutes, the student observed that the solution had not changed colour. **Evaluate** the student's results and **discuss** a reason for this observation, using the terminology learnt in this chapter.

**FIGURE 4** A piece of silver metal and a copper (II) sulfate solution.

**You can find the following resources for this section on your obook assess:**

- » Student book questions: Chapter 6 Revision questions
- » Revision notes: Chapter 6
- » assess quiz: Auto-correcting multiple-choice quiz
- » Flashcard glossary: Chapter 6