21 Redox Reactions

Introduction

The concept of redox is needed to understand the chemistry involved in many kinds of chemical reactions, including the rusting, displacement reactions, and reactions of metals with air, water and acids. As students have already studied these kinds of reactions, the chapter is well suited for a built-in revision of material from earlier chapters. The chapter is also useful to further developing students' appreciation of how scientific knowledge changes and accumulates over time.

Chapter Opener (page 324)

1. Begin the chapter by discussing the following questions. Precise answers are not needed at this stage.

   When magnesium burns in air, oxidation occurs. Suggest what oxidation is and how it occurs in this reaction.

   Answer: Oxidation is the addition of oxygen to a substance. In this reaction, magnesium is oxidised as oxygen is added to it to form magnesium oxide.

   What is an oxidising agent? Explain using your answer to the first question?

   Answer: An oxidising agent is a substance that brings about oxidation. In the reaction of magnesium with oxygen, oxygen is the oxidising agent.

   What is the contribution of Lavoisier to our knowledge of oxidation and reduction?

   Answer: Refer to 'Chemistry in Society' on page 329 of the Textbook.

2. Carry out an 'Inquiry Preview.'

Learning Outcomes

After completing this chapter, you should be able to:

- explain the term redox reaction
- list examples of redox reactions
- define oxidation and reduction (redox) in terms of oxygen/hydrogen gain/loss, electron transfer and changes in oxidation state
- identify various kinds of redox reactions in terms of gain/loss of oxygen/hydrogen, electron transfer and change in oxidation state
- describe the use of aqueous potassium iodide and acidified potassium manganate(VII) in testing for oxidising and reducing agents from the resulting colour changes

Teaching pointers

21.1 What are Oxidation and Reduction? (page 325)

Discuss, and possibly demonstrate, examples of redox reactions, such as some of those shown in Figure 21.1 on page 325 of the Textbook. Use the questions in Skills Practice to get students to think about the concepts of oxidation and reduction based on their existing knowledge.

Skills Practice (page 325)

1. For example, the word 'oxidation' makes one think of the addition of oxygen to a substance while 'reduction' reminds one of reducing or removing a substance such as the removal of oxygen from a substance.
2. $2Mg(s) + O2(g) \longrightarrow 2MgO(s)$

   Oxidation is the addition of oxygen to a substance, in this case, magnesium.
3. (a) $Zn(s) + CuO(s) \longrightarrow ZnO(s) + Cu(s)$
   (b) Oxygen is added to the zinc.
   (c) Copper(II) oxide is reduced as oxygen is removed from it.

(page 325)

Mystery Clue

The iron/steel of the bridge reacts with oxygen and water (moisture) in the atmosphere to form rust. The reaction is oxidation as oxygen is added to the iron.

Notes for Teachers

Applications of Redox Reactions
Most of the examples of redox reactions are so common in our daily lives that we sometimes take them for granted. Some of these examples are:

- Bleaching agents
- Photosynthesis
- Metabolism
- Nitrogen fixation
- Electrochemistry
- Combustion of fuels
- Refining metal ores

Teaching pointers

21.2 Oxidation and Reduction as Gain or Loss of Oxygen (page 326)

1. The definitions of oxidation and reduction in this chapter follow the historical order in the development of redox. As the understanding of atomic structure and bonding increased, the definitions of oxidation and reduction were modified to include reactions that, using earlier definitions, would not have been classified as redox reactions. Thus, while definitions of redox in terms of oxygen and hydrogen go back to the 18^{th} Century, definitions in terms of electron transfer and oxidation state date from the early 20^{th} Century.
2. Link ideas in this section with those from previous chapters, namely, reduction of metal oxides in Chapter 13 and the extraction of metals in Chapter 14. All these reactions are redox reactions, including electrolytic reactions.

21.3 Oxidation and Reduction as Gain or Loss of Hydrogen (page 328)

1. Oxidation and reduction in terms of gain and loss of hydrogen is included in this course but is more of historical interest nowadays.
2. Chemistry in Society on page 329 of the Textbook and Exercise 21.1 on the Theory Workbook look at Priestley and Lavoisier and the history of oxidation and reduction in terms of oxygen. Link point 4 on Lavoisier and acids with his definition of an acid in Exercise 21.2 of the Theory Workbook.

Skills Practice (page 328)

1.

Reaction	Substance oxidised (gain oxygen)	Substance reduced (lose oxygen)	Oxidising agent	Reducing agent
(a)	Zn	PbO	PbO	Zn
(b)	Mg	H_2O	H_2O	Mg
(c)	CO	Fe_3O_4	Fe_3O_4	CO
(d)	PbS	H_2O_2	H_2O_2	PbS

2. Metals are reducing agents as they reduce the oxides of other metals, i.e. remove oxygen from these oxides.
3. (a) H_2S is oxidised because it loses hydrogen. Br_2 is reduced because it gains hydrogen.
   (b) CH_4 is oxidised as it loses hydrogen. Cl_2 is reduced as it gains hydrogen.
4. In the reaction between lead(II) oxide and hydrogen, oxygen is lost from the oxide and gained by the hydrogen. In the reaction between hydrogen sulfide and chloride, hydrogen is lost from the hydrogen sulfide and gained by the chlorine.

Chemistry in Society (page 329)

A Little History — Oxygen and Oxidation

Exercise

1. Priestley was interested in science, which was a way to explain how things work and happen, despite being a minister in religion. He is also willing to share his discovery with other scientists and help in progressing scientific knowledge.

2. He improved the way to make gunpowder and used science to help the poor.

3. (a) Various possible responses, e.g., "Yes" as gunpowder is used for good as well as bad purposes.
   (b) Points to think about: A scientist's primary task is to learn about nature even if the knowledge obtained will be used by others for bad purposes. Scientists cannot remain neutral if they know their knowledge is going to be abused (hence Einstein's later opposition to atomic warfare); there may be fewer abuses if scientists speak out.

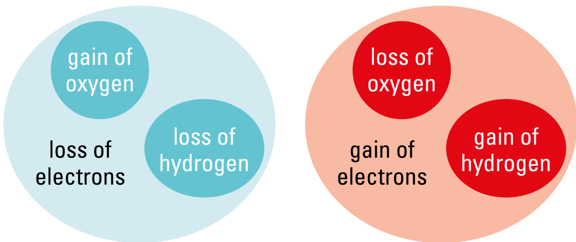
   Note: This question 3 helps to link Chemistry to history and also to further show that scientific knowledge can be used for good and bad purposes. Link this question with earlier discussions in the Additional Exercise 3 in Chapter 6 on Pauling and Additional Exercise 1 in Chapter17 on Haber and Faraday.

Teaching pointers

21.4 Oxidation and Reduction as Gain or Loss of Electrons (page 330)

1. Link the definitions of oxidation and reduction in terms of electron transfer to the formation of ions in Section 5.4 on pages 76–78 of the Textbook and the transfer of electrons in the formation of ionic bonds in Section 6.2 on pages 85–87.

2. Emphasise that many redox reactions involving oxygen and hydrogen can also be explained in terms of electron transfer and so the definitions in terms of oxygen and hydrogen are subsets of the definitions involving electron transfer. The Venn diagrams in Figure 21.4 on page 330 in the Textbook show this but could be extended to include hydrogen. That is:



oxidation reduction

3. In the reaction of metals with dilute acids, the acid is both an acid (because it forms hydrogen gas) and an oxidising agent (because the H^+ ions in the acid are reduced).

4. The displacement reactions of halogens show an important result, that is, halogens are *oxidising* agents.

Skills Practice (page 332)

1. **(a)** Oxidation as Al loses electrons to become Al^{3+}.
   (b) Reduction as Ag^+ gains electrons to become Ag.
   (c) Oxidation as Fe^{2+} loses electrons to become Fe^{3+}.
   (d) Reduction as H^+ gains electrons to become H_2.
   (e) Oxidation as Br^- gains electrons to become Br_2.

2.

Reaction	Substance oxidised (loses electrons)	Substance reduced (gains electrons)	Oxidising agent	Reducing agent
(a)	Cu	Cl_2	Cl_2	Cu
(b)	I^-	Br_2	Br_2	I^-
(c)	Mg	Ag^+	Ag^+	Mg
(d)	Fe	H^+	H^+	Fe
(e)	Fe^{2+}	Cl_2	Cl_2	Fe^{2+}

Teaching pointers

21.5 What is Oxidation State (page 333)

1. Emphasise that oxidation state (number) is a number that can be given to any atom in an element or compound (for both ionic and covalent substances).
2. Students may have the misconception that an oxidation state indicates that an element in a compound must have a charge. Sulfur and its compounds could be used to show the application of oxidation number to both ionic and covalent substances. For example, the oxidation number of sulfur in sulfide ion, S^{2-} is –2, the oxidation number of sulfur in sulfur dioxide, SO_2 is +4 and the oxidation number of sulfur in sulfate ion, SO_4^{2-} is +6.
3. In covalent compounds, we imagine that the substances are ionic in order to give the atoms oxidation states. Thus, if water were ionic, it would consist of H^+ and O^{2-} ions; so, hydrogen and oxygen in water have oxidation states of +1 and –2 respectively. See also the example of dichromate(VI) ions in Worked Example 1(c) on page 334 of the Textbook.
4. Note that all oxidation states have a + or – sign. Roman numerals such as (II) and (VI), used to represent oxidation states, are assumed to be positive.
5. Most of the metals with variable oxidation number are transition elements. Refer again to the position of transition metals in the Periodic Table.
6. When discussing Table 21.1, you might get the class to name the compounds in the table with the inclusion of Roman numerals in the names where appropriate. Example are: manganese(II) chloride, manganese(IV) oxide, potassium manganate(VII), chromium(II) chloride, chromium(III) chloride, potassium dichromate(VI), iron(II) chloride, iron(III) chloride, iron(II) sulfide.

Skills Practice (page 335)

1. (a) +4
   (b) −2
   (c) +4
   (d) +6
   (e) +7
2. The oxidation state of nitrogen in ammonium is −3; the oxidation state of nitrogen in nitrate is +5.

Skills Practice (page 335)

(a) $PbCl_2$
(b) $Cu(OH)_2$
(c) FeO
(d) Fe_2O_3
(e) $Cr_2(SO_4)_3$

Teaching pointers

21.6 Oxidation and Reduction as Increase or Decrease in Oxidation State (page 336)

The only new idea in this section is the application of oxidation state. As the reactions in this section have been discussed in previous chapters, this lightens the cognitive load for students.

Skills Practice (page 337)

1. Oxidation — increase in oxidation state. Reduction — decrease in oxidation state. Therefore:

   (a) $\overset{0}{4Al}(s) + \overset{0}{3O_2}(g) \longrightarrow \overset{+3\ -2}{2Al_2O_3}(s)$

   Aluminium is being oxidised; oxygen is being reduced.

   (b) $\overset{0}{Mg}(s) + \overset{+1}{H_2}O(g) \longrightarrow \overset{+2}{Mg}O(s) + \overset{0}{H_2}(g)$

   Magnesium is being oxidised; hydrogen in steam is being reduced.

   (c) $\overset{0}{Fe}(s) + 2\overset{+1}{H}Cl(aq) \longrightarrow \overset{+2}{Fe}Cl_2(aq) + \overset{0}{H_2}(g)$

   Iron is being oxidised; hydrogen (ions) in acid is being reduced.

   (d) $2\overset{+2}{Cu}O(s) + \overset{0}{C}(s) \longrightarrow 2\overset{0}{Cu}(s) + \overset{+4}{C}O_2(g)$

   Carbon is being oxidised; copper in copper(II) oxide is being reduced.

   (e) $\overset{0}{Br_2}(aq) + 2K\overset{-1}{I}(aq) \longrightarrow 2K\overset{-1}{Br}(aq) + \overset{0}{I_2}(g)$

   Iodide ions (in KI) are being oxidised; bromine is being reduced.

2. (a) (i) +3
   (ii) +3

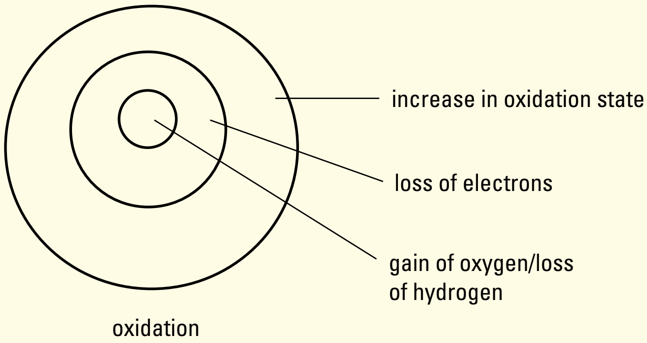
   (b) (i) Aluminium is being oxidised because its oxidation state increases from 0 in Al to +3 in Al_2O_3.
   (ii) Chromium in Cr_2O_3 is being reduced because its oxidation state decreases from +3 in Cr_2O_3 to 0 in Cr.

3. The oxidation state of iron increases from +2 in iron(II) sulfate to +3 in iron(III) sulfate, so iron is oxidised.
   The oxidation state of chromium decreases from +6 in potassium chromate(VI) to +3 in chromium(III) sulfate, so chromium is reduced.

Notes for Teachers

Venn diagrams for oxidation and reduction

The Venn diagram below shows the relationships between the four definitions of oxidation. A similar diagram can be constructed for reduction. As the circles in the diagram expand, the definitions become broader and encompass more chemical reactions.



Teaching pointers

21.7 How Can We Test for Oxidising and Reducing Agents? (page 338)

Note: Experiments with oxidising agents

Because of safety concerns relating to the use of potassium dichromate(VI), this reagent has been replaced by potassium manganate(VII) in Experiments 21.1 and 21.2.

1. In Experiment 21.1 of the Practical Workbook, students get to carry out tests for oxidising and reducing agents. If this experiment is not carried out, you should demonstrate the tests to students (as described in the experiment).
2. Acidified potassium manganate(VII) can also be used as a test for reducing agent. You may demonstrate this test to students. See 'Notes for Teachers' on the next page.
3. Potassium manganate(VII) solution is used to test for reducing agents; this could be demonstrated to the class. Refer to 'Notes for Teachers' below.
4. Experiment 21.2 of the Practical Workbook combines tests for oxidising and reducing agents in a qualitative analysis exercise to identify such agents.

Notes for Teachers

Tests for reducing agents using potassium manganate(VII) solution

The test shown below shows the desired observations of colour change from purple to colourless.

Procedure: Fill a test-tube with sodium sulfite solution. Add a few drops of dilute sulfuric acid followed by a few drops of potassium manganate(VII) solution.

Observation: The purple $KMnO_4$ solution is decolourised.

Conclusion: Sodium sulfite is a reducing agent.

Changes: Sulfite ions are oxidised to sulfate ions:

$$\overset{+4}{S}O_3^{2-}(aq) \longrightarrow \overset{+6}{S}O_4^{2-}(aq)$$

colourless → colourless

Manganate(VII) ions are reduced to manganese(II) ions:

$$\overset{+7}{Mn}O_4^{-}(aq) \longrightarrow \overset{+2}{Mn}^{2+}(aq)$$

purple → colourless

Teaching pointers

21.8 Are All Reactions Redox Reactions? (page 339)

1. Several types of reactions are not redox reactions.
2. Students often have the misconception that a substance is *either* an oxidising agent *or* a reducing agent.

   Use the example of sulfur dioxide on page 340 of the Textbook to point out that is not the case; some substances can be *both* oxidising agent and a reducing agent. This occurs when the elements in the compounds have variable oxidation states and the oxidation state can either increase or decrease.

   Another example is iron(II) compounds; the oxidation state of +2 can increase to +3 in some reactions or it can decrease to 0 in others.

   For example:

$$\overset{+2}{FeSO_4}(aq) \xrightarrow{Mg} \overset{0}{Fe}(s)$$

   Fe(II) in $FeSO_4$ is being reduced; $FeSO_4$ acts as an oxidising agent.

$$\overset{+2}{FeSO_4}(aq) \xrightarrow{KI} \overset{+3}{Fe_2}(SO_4)_3(aq)$$

   Fe(II) in $FeSO_4$ is being oxidised; $FeSO_4$ acts as a reducing agent.

   Whether a substance acts as an oxidising or reducing agent depends relatively on the strength of the other compound. If the other compound is a stronger oxidising agent, then the first compound will behave like a reducing agent.

Skills Practice (page 340)

1. **(a)** Oxidation state of Zn is 0.
   Oxidation state of Zn in $ZnCl_2$ is +2.
   Therefore, Zn is being oxidised.

   Oxidation state of H in HCl is +1.
   Oxidation state of H in H_2 is 0.
   Therefore, H^+ in HCl is being reduced.

   This is a redox reaction.

   (b) Oxidation state of Mg is 0.
   Oxidation state of Mg in $MgSO_4$ is +2.
   Therefore, Mg is being oxidised.

   Oxidation state of Fe in $FeSO_4$ is +2.
   Oxidation state of Fe is 0.
   Therefore, Fe^{2+} in $FeSO_4$ is being reduced.

   This is a redox reaction.

   (c) Oxidation state of H in HNO_3 is +1.
   Oxidation state of H in H_2O is +1.
   There is no change of oxidation state. H is neither oxidised nor reduced.

   (d) Oxidation state of Cu^{2+} is +2.
   Oxidation state of Cu in $Cu(OH)_2$ is +2.
   There is no change of oxidation state. Cu is neither oxidised nor reduced.

2. **(a)** **(i)** Oxidation state of Na is 0.
   Oxidation state of Cl in Cl_2 is 0.
   Oxidation state of I^- is –1.
   Oxidation state of Fe^{2+} is +2.
   Oxidation state of Fe^{3+} is +3.

   (ii) Only Fe^{2+}(aq)

   (b) **(i)** Cl_2(g) and Fe^{3+}(aq) are oxidising agents.
   (ii) Na(s) and I^-(aq) are reducing agents.
   (iii) Fe^{2+} is both an oxidising and a reducing agent.

21 Chapter Review

Self-Management

Misconception Analysis (page 341)

1. **False** This was the original meaning of oxidation. But now oxidation also refers to the loss of electrons by a substance or an increase in oxidation state. By extending the definition of oxidation, more reactions are classified as oxidation. Even if oxygen is not involved, the same basic chemical changes occur.

2. **True** Metals in ionic compounds exist as positive ions. These ions are reduced to metals in redox reactions.

3. **True** By convention, oxidation states of elements are zero.

4. **False** In some cases this is true. For example, sodium and potassium have an oxidation state of +1 in all their compounds. Other elements have variable oxidation states. For example, iron occurs in compounds with oxidation states of +2 and +3.

5. **False** For example, when potassium dichromate(VI) $K_2Cr_2O_7$ reacts, only the oxidation state of chromium changes. The oxidation states of potassium and oxygen remains the same.

6. **False** For example, neutralisation reactions are not redox reactions.

Practice

Structured Questions (pages 342)

1. Magnesium gains oxygen, loses electrons to form Mg^{2+} ion and its oxidation state increases from 0 to +2.

2. (a) $2Al(s) + Fe_2O_3(s) \longrightarrow Al_2O_3(s) + 2Fe(l)$
   (b) (i) Aluminium is oxidised. (It gains oxygen. / There is an increase in oxidation state from 0 in Al to +3 in Al_2O_3).
   (ii) Iron(III) oxide is reduced. (It loses oxygen. / There is a decrease in oxidation state from +3 in Fe_2O_3 to 0 in Fe).
   (c) Number of moles of iron
   = 2 × Number of moles of iron(III) oxide
   = 2 × 1000 ÷ [2(56) + 3(16)]
   = 12.5 moles

   Mass of iron = 12.5 × 56
   = 700 g
   (d) The molten iron produced in the thermite reaction can be used to weld railway lines together.

3. (a) Hydrogen is being oxidised as it gains oxygen, or its oxidation state increases from 0 in H_2 to +1 in HCl. Chlorine is being reduced as it gains hydrogen, or its oxidation state decreases from 0 in Cl_2 to –1 in HCl.
   (b) Iron is being oxidised as it loses electrons, or its oxidation state increases from 0 in Fe to +3 in $FeCl_3$. Chlorine is being reduced as it gains electrons, or its oxidation state decreases from 0 in Cl_2 to –1 in $FeCl_3$.
   (c) Hydrogen is being oxidised as its oxidation state increases from 0 in H_2 to +1 in NH_3. Nitrogen is being reduced as its oxidation state decreases from 0 in N_2 to –3 in NH_3.
   (d) Ammonia is being oxidised as it loses hydrogen to become nitrogen, or the oxidation state of nitrogen increases from –3 in NH_3 to 0 in N_2. Chlorine is being reduced as it gains hydrogen, or its oxidation state decreases from 0 in Cl_2 to –1 in HCl.
   (e) Lead(II) sulfide is being oxidised as it gains oxygen, or the oxidation state of sulfur increases from –2 in PbS to +6 in $PbSO_4$. Hydrogen peroxide is being reduced as it loses oxygen, or the oxidation state of oxygen decreases from –1 in H_2O_2 to –2 in H_2O.

4. (a) and (e)

5. (a) Copper metal
   (b) $Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$
   (c) Mg is being oxidised as it loses electrons to form Mg^{2+} ions. The Cu^{2+} ion is being reduced as it gains electrons to form Cu.
   (d) The copper(II) ion, Cu^{2+} is the reducing agent.
   (e) If magnesium is not in excess, the colour of the solution will be pale blue. This is due to the unreacted blue copper(II) sulfate which remains in solution.

6. (a) $(NH_4)_2Cr_2O_7(s) \longrightarrow 4H_2O(g) + Cr_2O_3(s) + N_2(g)$
   (b) Chromium is reduced because its oxidation state decreases from +6 in $(NH_4)_2Cr_2O_7$ to +3 in Cr_2O_3.

7. (a) Oxidation state of Fe in iron(II) sulfate is +2. Oxidation state of Fe in iron(III) sulfate is +3.
   (b) (i) A (dirty) green precipitate forms, which slowly turns brown in air.
   (ii) $Fe^{2+}(aq) + 2OH^-(aq) \longrightarrow Fe(OH)_2(s)$
   (iii) It is *not* a redox reaction as there is no change in the oxidation state of iron. The oxidation state of iron is +2 in both Fe^{2+} and $Fe(OH)_2$.
   (c) (i) This change is oxidation as the oxidation state of iron increases from +2 in iron(II) sulfate to +3 in iron(III) sulfate.
   (ii) Reagents such as potassium dichromate(VI) solution or potassium manganate(VII) solution can be used. They are oxidising agents and oxidise iron(II) sulfate to iron(III) sulfate while themselves being reduced. (The oxidation state of Cr decreases from +6 in $K_2Cr_2O_7$ to +3 in Cr^{3+}; the oxidation state of Mn decreases from +7 from $KMnO_4$ to +2 in Mn^{2+}).

8. (a) Dilute sulfuric acid
   (b) (i) $K_2Cr_2O_7$
   (ii) +6
   (c) (i) Orange to green
   (ii) The oxidation state of chromium decreases from +6 in $K_2Cr_2O_7$ to +3 in Cr^{3+} ions.
   (d) (i) For example, iron(II) sulfate and sodium sulfite. (Refer to Tests 1 and 2 in Experiment 19.1 of the Practical Workbook.)
   (ii) Refer to Questions 1–3 in Experiment 19.1 of the Practical Workbook.

Free Response Questions (page 343)

1. Responses to this question may include the following points: (Any three examples of redox reactions.)
   (a) Rusting: Iron changes to (hydrated) iron(III) oxide. Iron is being oxidised as it gains oxygen, or the oxidation state of iron increases from 0 in iron to +3 in Fe_2O_3.
   (b) Extraction of iron: Haematite, iron(III) oxide reacts with carbon monoxide to form iron.

   $$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(l) + 3CO_2(g)$$

   Iron(III) oxide is being reduced as it loses oxygen to form iron. Carbon monoxide is being oxidised as it gains oxygen to form carbon dioxide.
   (c) Metal displacement reactions: For example, the reaction between zinc metal and a solution of copper(II) ions to form copper metal and a solution of zinc ions.

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

   The oxidation state of zinc increases from 0 in Zn to +2 in Zn^{2+} (oxidation). The oxidation state of magnesium decreases from +2 in Mg^{2+} to 0 in Mg (reduction).

2. Responses to this question may include the following points:
   - Ionic equation: $Cl_2(aq) + 2I^-(aq) \longrightarrow 2Cl^-(aq) + I_2(aq)$
   - Oxidation is a loss of electrons or an increase in oxidation state.
     Reduction is a gain of electrons or a decrease in oxidation state.
   - In the above reaction, chlorine atoms gain electrons to form chloride ions, Cl^-. The oxidation state decreases from 0 in Cl_2 to –1 in Cl^-. This is reduction.
   - At the same time, the iodide ions lose electrons to become iodine. The oxidation state increases from –1 in I^- to 0 in I_2. This is oxidation.
   - As both oxidation and reduction occur together, this reaction is a redox reaction.

3. Responses to this question may include the following points:
   - Early knowledge: Oxygen was discovered. The addition of oxygen in reactions gives oxidation; the removal of oxygen gives reduction.
   - More knowledge and changes: Knowledge of electrons and their involvement in ionic bonding/reactions give new definitions to oxidation and reduction in terms of transfer of electrons.
   - Later knowledge: The concept of oxidation number helps to explain redox reactions that cannot be explained in terms of gain or loss of oxygen or electron transfer.

Extension (page 343)

The VRB

Batteries are devices that store chemical energy and generate electricity by a redox reaction. The VRB consists of two halves separated by a membrane. In one half, an oxidation reaction occurs (i.e. loss of electrons) which produces electrons. These electrons flow through an external circuit (e.g. a light bulb) to the other half where a reduction reaction (gain of electrons) occurs using these electrons. The VRB works because vanadium compounds exist in different oxidation states.

(a) The oxidation half of the battery contains a solution of vanadium(II) sulfate VSO_4 (oxidation state of the vanadium is +2). This loses electrons and is oxidised to vanadium(III) sulfate $V_2(SO_4)_3$ (oxidation state of the vanadium is +3). The ionic equation for the reaction is:

$$V^{2+}(aq) \longrightarrow V^{3+}(aq) + e- \quad \text{[oxidation]}$$

The reduction half contains a compound called pervanadyl sulfate $(VO_2)_2SO_4$ in which the vanadium has an oxidation state of +5. This gains electrons and is reduced to a compound called vanadyl(IV) sulfate $(VO)SO_4$ in which the vanadium has an oxidation state of +4. The ionic equation for the reaction is:

$$VO_2^+(aq) + 2H^+(aq) + e^- \longrightarrow VO^{2+}(aq) + H_2O(l) \quad \text{[reduction]}$$

(b) The main disadvantages with the VRB is that it is much more complex in comparison with other batteries and the energy output is low.

(c) The electrolyte often consists of vanadium compounds dissolved in sulfuric acid. In these regions where the temperature can be very low, sulfuric acid could crystallise. Thus the team needs to ensure that there is sufficient insulation around the system such that the electrolytes remain in liquid form.

Additional Teaching Material

Additional Exercise 1: Use of Mapping Techniques

Objective

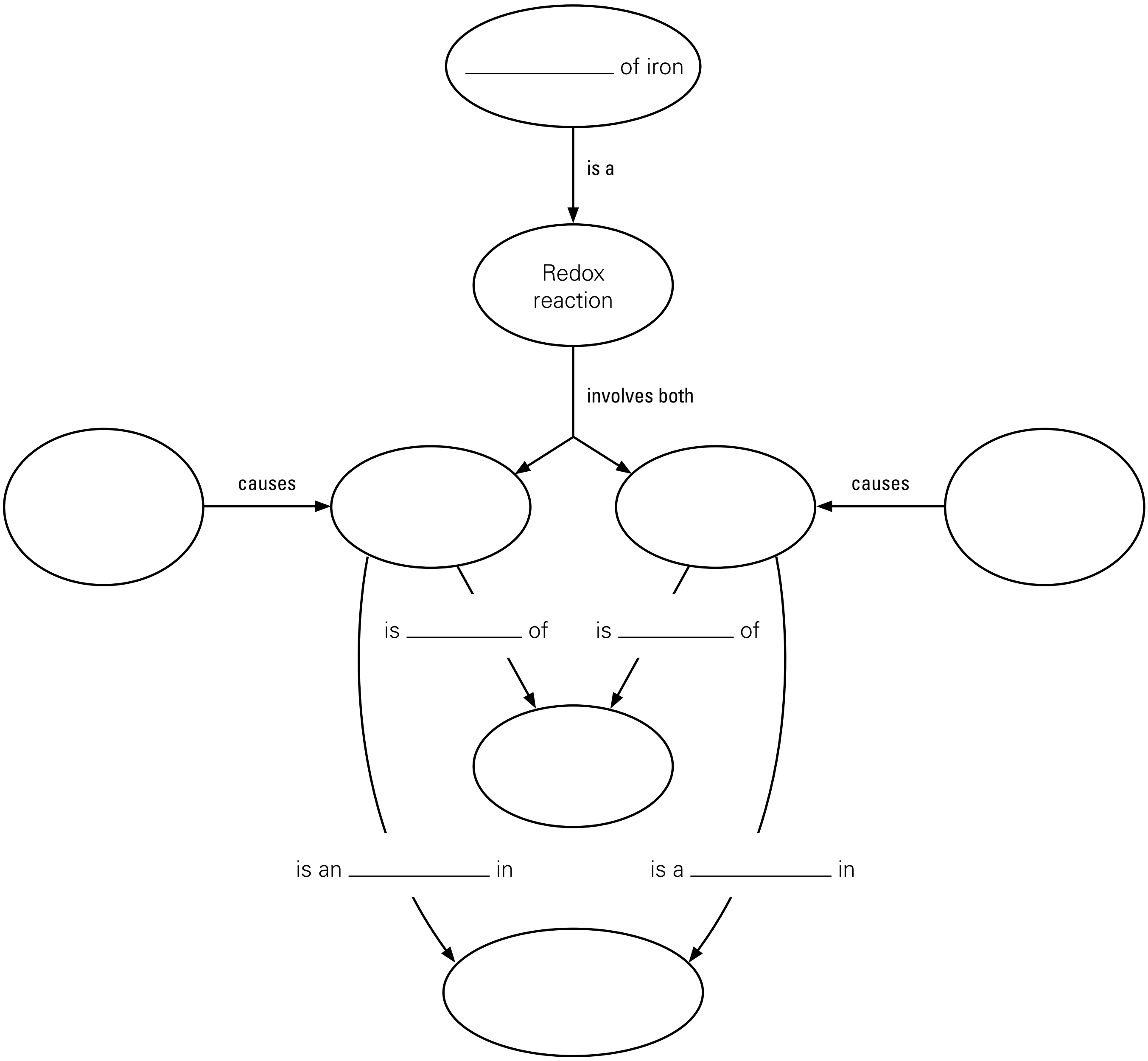
- To complete a concept map and construct a mind map for the concept of a redox reaction

Key Competencies

CIT: sound reasoning [*comparing*]
ICS: communicating effectively [*mapping, peer review, feedback*]

A Concept map

1. Fill in suitable concept words and link words in the following concept map.



2. **(a)** Add the following additional concept words and link words to your concept map.

   Concept words: *cell reaction, displacement reaction, carbon, potassium dichromate(VI)*
   Link words: is a (3 times), is an

   (b) Add any other terms you can think of which are connected to oxidation and reduction.

B Mind map

On a large sheet of paper, or by means of appropriate mapping software, create a mind map for the topic of redox reactions that includes the main ideas you learnt in this chapter. Get other classmates to review your mind map and suggest ways to improve it. Then revise your mind map.

Additional Teaching Material

Additional Exercise 2: What is the Answer?

Objective

- To come up with questions for which their answer is the given statement

Key Competency

CIT: sound reasoning [*comparing*]

(a) The answer to a Chemistry question is:

The hydrogen ions are reduced.

Think of as many questions as possible to which this is the answer.

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(b) The answer to a Chemistry question is:

The zinc is the reducing agent.

Think of as many questions as possible to which this is the answer.

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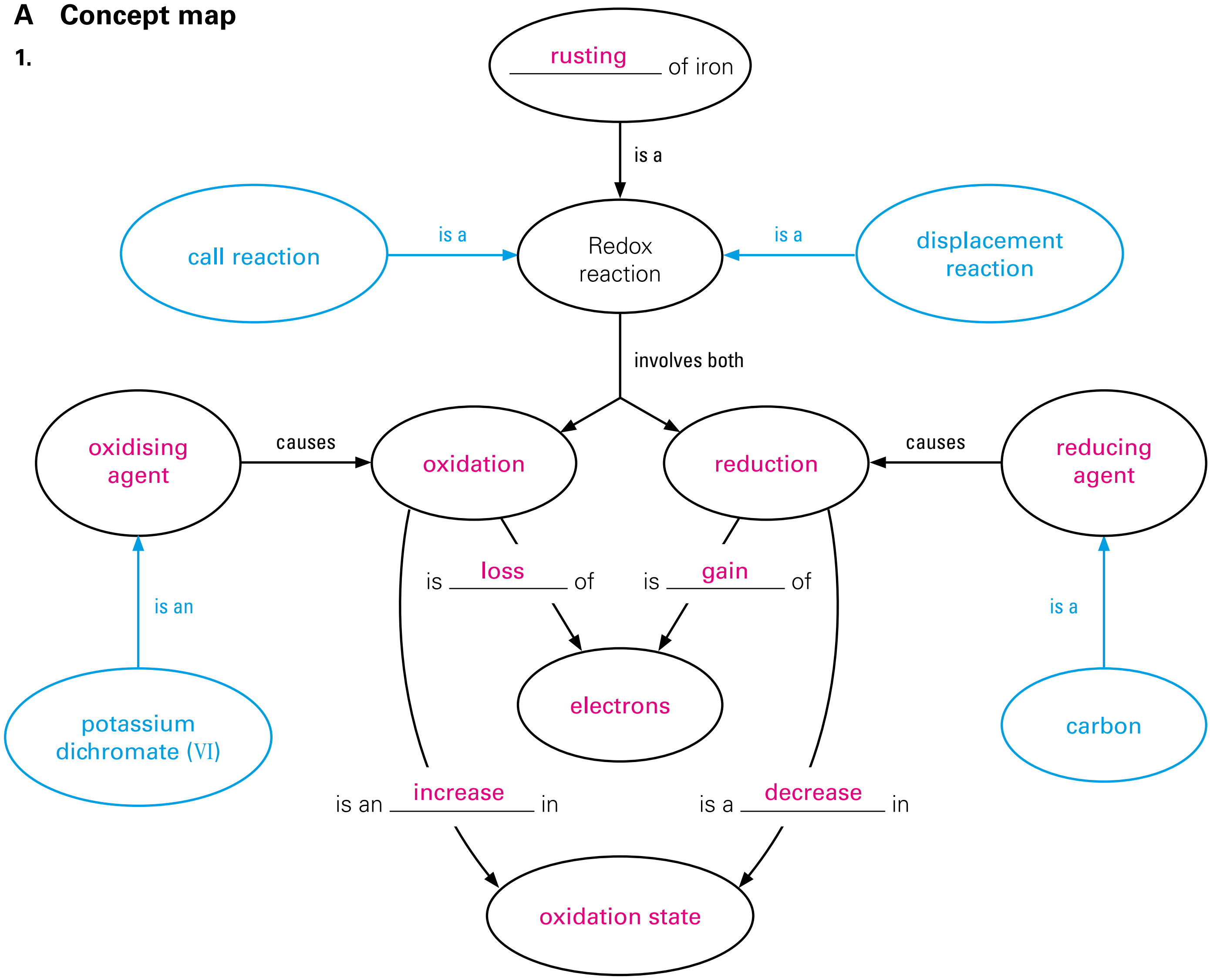
Answers

Additional Exercise 1:

A Concept map

1.



Additional Exercise 2:

(a) Possible questions:

- In terms of oxidation and reduction, what change happens to the hydrogen ions in the following reaction?

$$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$$

Note: Iron and magnesium can be used instead of zinc in the above reaction.

- In terms of oxidation and reduction, what change happens to the hydrogen ions in the following reaction?

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

(b) Possible questions:

- What is the reducing agent in the following reactions?

(i) $Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$

(ii) $CuO(s) + Zn(s) \longrightarrow ZnO(s) + Cu(s)$

(iii) $Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$

(iv) $Zn(s) + 2Ag^+(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s)$