Topic 11: **Answers**
Reduction and Oxidation

SYLLABUS RELEVANCE & TEXTBOOK CHAPTERS		
O-LEVEL PURE (5072)	✓	Chapter 13
O-LEVEL SCIENCE (5116)	✓	Chapter 12
N-LEVEL SCIENCE (5155)	✗	

Lesson Package & Accompanying Slides Designed by Alex Lee (2009)
Last Modified by Alex Lee (2011)

1. Reduction and Oxidation – An Introduction

Individual atoms, when combined with other elements to form compounds, may be oxidised or reduced in the process. For example, when carbon (C) burns to become carbon dioxide (CO_2), we can say that the carbon atom has become oxidised.

Let us look at the three definitions of reduction and oxidation.

Reduction	Oxidation
Gain of Hydrogen	Loss of Hydrogen
Gain of Electrons	Loss of Electrons
Loss of Oxygen	Gain of Oxygen

Identify if the underlined elements have been oxidised or reduced, and why.

Chemical Reaction	Oxidised or Reduced?	Why?
$\underline{N_2} + 3\,H_2 \longrightarrow 2\,\underline{N}H_3$	Reduced	Gained Hydrogen
$\underline{Si}H_4 \longrightarrow \underline{Si} + 2\,H_2$	Oxidised	Lost Hydrogen
$2\,\underline{S}O_2 + O_2 \longrightarrow 2\,\underline{S}O_3$	Oxidised	Gained Oxygen
$2\,\underline{Cl_2}O \longrightarrow 2\,\underline{Cl_2} + O_2$	Reduced	Lost Oxygen
$2\,\underline{H_2} + O_2 \longrightarrow 2\,\underline{H_2}O$	Oxidised	Gained Oxygen
$2\,H_2 + \underline{O_2} \longrightarrow 2\,H_2\underline{O}$	Reduced	Gained Hydrogen
$Mg \longrightarrow Mg^{2+}$	Oxidised	Lost Electrons
$Cl_2 \longrightarrow 2\,Cl^-$	Reduced	Gained Electrons
$\underline{Mg} + F_2 \longrightarrow \underline{Mg}F_2$	Oxidised	Lost Electrons (Mg^{2+})
$2\,K + \underline{Cl_2} \longrightarrow 2\,K\underline{Cl}$	Reduced	Gained Electrons (Cl^-)

However, in many situations there may be a concurrent gain or loss in hydrogen, oxygen and electrons on the same atom. For example, when nitrogen (N_2) is converted into a nitrate ion (NO_3^-), there is a concurrent gain in oxygen (i.e. oxidation) and gain in electrons (i.e. reduction).

Hence, to find out whether an atom has been oxidised or reduced, it is **most accurate to calculate its oxidation state**.

2. **Calculating Oxidation State**

Pure elements (e.g. O_2, H_2, N_2) have an oxidation state of **zero**.

When two elements combine to form a compound, generally one will become oxidised and one will become reduced. For example, carbon monoxide:

$$CO$$

oxidation state: +2 **oxidation state: –2**

This gives rise to two methods to calculate oxidation state.

Method #1: Algebraic Approach

1. Let the oxidation state of the unknown element be x.

2. Allocate the oxidation states for all other elements in the compound.
   i. The oxidation state of monoatomic ions is equivalent to their ionic charge (e.g. the oxidation state of Mg^{2+} is +2, and the oxidation state of N^{3-} is -3).
   ii. Each oxygen atom, when bonded, has an oxidation state of **-2** (with the exception of hydrogen peroxide, H_2O_2, and in element form, O_2).
   iii. Each hydrogen atom, when bonded, has an oxidation state of **+1**.

3. The sum of the oxidation states of all the atoms making up the compound is equivalent to its total ionic charge of the compound. Often, this value is zero.

Illustration 1 – Sulfur in Potassium Sulfate (K_2SO_4)

$$K_2\underline{S}O_4$$

oxidation state: +1 *(ionic charge of K^+)*

oxidation state: x

oxidation state: –2

$$2\,(K) + S + 4\,(O) = 0$$
$$2\,(+1) + x + 4\,(-2) = 0$$
$$+2 + x - 8 = 0$$
$$x = +6$$

Illustration 2 – Nitrogen in Ammonium Ion (NH_4^+)

$$\underline{N}H_4^{\,+}$$

oxidation state: x **oxidation state: +1**

$$N + 4\,(H) = +1$$
$$x + 4\,(+1) = +1$$
$$x + 4 = +1$$
$$x = -3$$

Method #2: Short-Cut Approach

1. If the substance is an ionic compound, split it into the various ions. Observe only the ion containing the unknown element, and ignore the other ions.

2. Apply the following rules:
   i. For **each oxygen attached** to the atom, **add 2**.
   ii. For **each hydrogen attached** to the atom, **subtract 1**.
   iii. Add or subtract the ionic charge accordingly.

3. If there is more than one atom of the same element, **divide** the oxidation state accordingly.

Illustration 3 – Carbon in Calcium Carbonate ($CaCO_3$)

$$Ca\underline{C}O_3$$

split into ions

Ca^{2+} — ignore this one

CO_3^{2-} — calculate this one

Attached to carbon:
a) 3 oxygen atoms
b) no hydrogen atoms
c) 2- ionic charge

Oxidation State = 3 (+2) + (–2)
= + 4

Illustration 4 – Sodium in Sodium Chloride (NaCl)

$$\underline{Na}Cl$$

split into ions

Na^{+} — calculate this one

Cl^{-} — ignore this one

Attached to sodium:
a) no oxygen atoms
b) no hydrogen atoms
c) 1+ ionic charge

Oxidation State = + 1

Calculate the oxidation states for the underlined elements in the following compounds:

(a) $\underline{N}H_3$

O.S. = -3

(b) $\underline{N}O_2$

O.S. = +4

(c) $\underline{N}O_3^{-}$

O.S. = +5

(d) $\underline{N}O_2^{-}$

O.S. = +3

(e) $Cu\underline{Cl}_2$

O.S. = -1

(f) $\underline{Cl}_2O$

O.S. = +2

(g) $H_2\underline{O}$

O.S. = -2

(h) $NH_4\underline{N}O_3$

O.S. = +5

(i) $\underline{N}H_4NO_3$

O.S. = -3

(j) $K\underline{Mn}O_4$

O.S. = +7

(k) $\underline{S}_8$

O.S. = 0

(l) $H_2\underline{S}$

O.S. = -2

(m) $\underline{S}O_2$

O.S. = +4

(n) $\underline{S}O_3$

O.S. = +6

(o) $\underline{S}O_3^{2-}$

O.S. = +4

3. **Redox Reactions**

The most effective way of determining whether a substance gets reduced or oxidised is by observing its oxidation state. A **decrease in oxidation state** implies that a substance has been **reduced**; an **increase in oxidation state** implies that a substance has been **oxidised**.

$$2\ \underline{Fe}SO_4\ (aq) + \underline{Cl}_2\ (g) + H_2SO_4\ (aq) \rightarrow \underline{Fe}_2(SO_4)_3\ (aq) + 2\ H\underline{Cl}\ (g)$$

Oxidation State:	Oxidation State:	Oxidation State:	Oxidation State:
+2	0	+3	-1

Element oxidised: Fe Element reduced: Cl

$$\underline{N}H_3\ (g) + \underline{Cu}O\ (s) \rightarrow \underline{N}_2\ (g) + 3\ \underline{Cu}\ (s) + 3\ H_2O\ (l)$$

Oxidation State:	Oxidation State:	Oxidation State:	Oxidation State:
-3	+2	0	0

Element oxidised: N Element reduced: Cu

$$\underline{Fe}\ (s) + 3\ \underline{Cl}_2\ (g) \rightarrow 2\ \underline{FeCl}_3\ (aq)$$

Oxidation State:	Oxidation State:	Oxidation State:	Oxidation State:
0	0	+3	-1

Element oxidised: Fe Element reduced: Cl

$$\underline{Cl}_2\ (g) + 2\ \underline{I}^-\ (aq) \rightarrow 2\ \underline{Cl}^-\ (aq) + \underline{I}_2\ (s)$$

Oxidation State:	Oxidation State:	Oxidation State:	Oxidation State:
0	-1	-1	0

Element oxidised: I Element reduced: Cl

4. **Review Questions**

(a) Indicate '**Redox**' or '**Not Redox**' next to the chemical equations below.

$Ag_2O\ (s) + H_2O_2\ (l) \longrightarrow 2\ Ag\ (s) + H_2O\ (l) + O_2\ (g)$	Redox
$AgNO_3\ (aq) + KCl\ (aq) \longrightarrow AgCl\ (s) + KNO_3\ (aq)$	Not Redox
$CuO\ (s) + H_2SO_4\ (aq) \longrightarrow CuSO_4\ (aq) + H_2O\ (l)$	Not Redox
$2\ H^+\ (aq) + CO_3^{2-}\ (aq) \longrightarrow H_2O\ (l) + CO_2\ (g)$	Not Redox
$2\ K\ (s) + 2\ H_2O\ (l) \longrightarrow 2\ KOH\ (aq) + H_2\ (g)$	Redox
$NH_3\ (g) + HCl\ (g) \longrightarrow NH_4Cl\ (s)$	Not Redox
$2\ NH_3\ (g) + 3\ CuO\ (s) \longrightarrow N_2\ (g) + 3\ Cu\ (s) + 3\ H_2O\ (l)$	Redox
$SO_3\ (g) + H_2O\ (l) \longrightarrow 2\ H^+\ (aq) + SO_4^{2-}\ (aq)$	Not Redox

(b) Define oxidation

(i) in terms of oxygen,

Oxidation is the gain of oxygen.

(ii) in terms of electrons,

Oxidation is the loss of electrons.

(c) For each of the following reactions, name the substance being oxidised and explain how you arrived at your answer.

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

Zinc metal; oxidation state increases from 0 to +2.

(ii) $Cl_2\ (g) + 2\ KI\ (aq) \longrightarrow 2\ KCl\ (aq) + I_2\ (s)$

Iodide ions / potassium iodide; oxidation state increases from -1 to 0.

(d) What is the oxidation state of chromium in each of the following substances?

(i) Cr_2O_3	(ii) $CrCl_2$	(iii) $K_2Cr_2O_7$
+3	+2	+6

(e) What is the oxidation state of oxygen in each of the following substances?

(i) H_2O	(ii) O_2	(iii) H_2O_2
-2	0	-1

5. <u>Oxidising & Reducing Agents</u>

Oxidising Agents	**Reducing Agents**
A chemical which causes **other** substances to be ......oxidised....... In the process, itself gets ......reduced.......	A chemical which causes **other** substances to be ......reduced....... In the process, itself gets ......oxidised.......
Examples of good oxidising agents: 1. potassium manganate(VII), $KMnO_4$ 2. potassium dichromate(VI), $K_2Cr_2O_7$ 3. reactive halogens, e.g. F_2, Cl_2 4. conc. sulfuric acid, H_2SO_4 5. oxygen, O_2 6* hydrogen peroxide, H_2O_2	*Examples of good reducing agents:* 1. potassium iodide, KI 2. reactive metals, e.g. K, Na, Li, Ca 3. carbon, C & carbon monoxide, CO 4. sulfur dioxide, SO_2 5. hydrogen, H_2 6* hydrogen peroxide, H_2O_2

(a) (i) Briefly explain why alkali metals are good reducing agents.

Alkali metals readily give away electrons, reducing other substances.

(ii) Which is a stronger reducing agent – francium (Fr) or sodium (Na)? Be sure to give a reason for your answer.

Francium. It is more reactive than sodium and therefore gives away an electron more readily than sodium, hence reducing another substance more easily.

(b) (i) Briefly explain why halogens are good oxidising agents.

Halogens readily take in electrons, oxidising other substances.

(ii) Which is a stronger oxidising agent – fluorine (F_2) or chlorine (Cl_2)? Be sure to give a reason for your answer.

Fluorine. It is more reactive than chlorine and therefore takes in an electron more readily than chlorine, hence oxidising another substance more easily.

(c) In the reaction below, identify the oxidising agent and the reducing agent.

$$2\ NO_2\ (g) + 7\ H_2\ (g) \longrightarrow 2\ NH_3\ (g) + 4\ H_2O(l)$$

Oxidising Agent: Nitrogen Dioxide Reducing Agent: Hydrogen

Item Oxidised: Hydrogen Item Reduced: Nitrogen Dioxide

6. **Redox in Qualitative Analysis**

In testing for **an oxidising agent**, we are testing whether a substance can be reduced.
We thus use a reducing agent to do this.

In testing for **a reducing agent**, we are testing whether a substance can be oxidised.
We thus use a an oxidising agent agent to do this.

We can **test for an oxidising agent** by using either of these two aqueous reducing agents:

Potassium iodide (KI) is a powerful reducing agent which is colourless in appearance. After it is oxidised, we observe that its colour changes to become a brown solution. This is due to the displacement of the iodide ions.

$$\underset{\textit{colourless}}{\mathbf{2\,I^- (aq)}} \longrightarrow \underset{\textit{brown}}{\mathbf{I_2 (aq)}} \mathbf{+ 2\,e^-}$$

Aqueous iron(II) solutions (e.g. $Fe(NO_3)_2$, $FeCl_2$) are easily oxidised and hence may act as reducing agents. After it is oxidised, we observe that its colour changes from green to become brown.

$$\underset{\textit{green}}{\mathbf{Fe^{2+} (aq)}} \longrightarrow \underset{\textit{yellow}}{\mathbf{Fe^{3+} (aq)}} \mathbf{+ 2\,e^-}$$

We can **test for a reducing agent** by using these either of these two aqueous oxidising agents:

Acidified potassium permanganate (also known as potassium manganate(VII)), $KMnO_4$, is a powerful oxidising agent which is purple in colour. After it is reduced, we observe that the solution becomes decolourised to form a colourless solution.

$$\underset{\textit{purple}}{\mathbf{MnO_4^- (aq)}} \mathbf{+ 8\,H^+ (aq) + 5\,e^-} \longrightarrow \underset{\textit{colourless}}{\mathbf{Mn^{2+} (aq)}} \mathbf{+ 4\,H_2O (l)}$$

Acidified potassium dichromate(VI), $K_2Cr_2O_7$, is a powerful oxidising agent which is orange in colour. After it is reduced, we observe that its colour changes to become a green solution.

$$\underset{\textit{orange}}{\mathbf{Cr_2O_7^{2-} (aq)}} \mathbf{+ 14\,H^+ (aq) + 6\,e^-} \longrightarrow \underset{\textit{green}}{\mathbf{2\,Cr^{3+} (aq)}} \mathbf{+ 7\,H_2O (l)}$$

7. Review Questions

(a) When orange crystals of ammonium dichromate(VI) are heated, the products are green chromium(III) oxide, nitrogen and water vapour.

(i) Construct the equation, including state symbols, for this reaction.

$(NH_4)_2Cr_2O_7\ (s) \longrightarrow Cr_2O_3\ (s) + N_2\ (g) + 4\ H_2O\ (g)$

(ii) By naming the reagents and giving the observations, describe how to confirm that

① the orange crystals contain the ammonium ion,

Warm the crystals with aqueous sodium hydroxide. A colourless, pungent gas should be evolved which turns moist red litmus paper blue.

② the reaction products contain water.

Test the products of the reaction with anhydrous cobalt(II) chloride paper. The paper should turn from blue to pink.

(iii) The action of heat on the crystals starts a redox reaction. Has the chromium been oxidised or reduced? Explain your answer.

Reduced. The oxidation state of chromium decreases from +6 in $K_2Cr_2O_7$ to +3 in Cr_2O_3.

(b) In the following redox reactions, write down the **formulae** of the oxidising agent and the reducing agent. Also state your predicted observations, in terms of colour changes.

(i) A sample of chlorine gas is bubbled through a solution of potassium iodide.

Oxidising Agent: Cl_2 Reducing Agent: I^-

Observations: The colourless potassium iodide solution turns brown.

(ii) A sample of sulfur dioxide is bubbled through aqueous potassium permanganate.

Oxidising Agent: MnO_4^- Reducing Agent: SO_2

Observations: The purple potassium permanganate solution turns colourless.

(iii) Silver chloride, exposed to light, decomposes to become silver and chlorine gas.

Oxidising Agent: Ag^+ Reducing Agent: Cl^-

Observations: A metallic-grey deposit forms; green-yellow gas evolved.

Self-Designed Summary

Supplementary Questions

1. (a) Oxidation can be described as an increase in oxidation number. State three other definitions of oxidation.
   (b) Describe a reaction to show that aqueous chlorine is an oxidizing agent.

2. Three reactions are given below.
   (i) $Zn\ (s) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2\ (g)$
   (ii) $Cl_2\ (g) + 2\ KI\ (aq) \rightarrow 2\ KCl\ (aq) + I_2\ (g)$
   (iii) $Fe^{2+}\ (aq) + NO_3^-\ (aq) + 2\ H^+\ (aq) \rightarrow Fe^{3+}\ (aq) + NO_2\ (g) + H_2O\ (l)$

   For each of the above reactions, identify
   (a) the substance being oxidized,
   (b) the oxidizing agent,
   (c) the substance being reduced,
   (d) the reducing agent.

3. Explain what is meant by the "(II)" and "(VI)" in copper(II) sulfate(VI), and the "(VII)" in potassium manganate(VII). Also suggest why we do not need these roman numerals in naming silver fluoride.

4. Ammonia is oxidized in the manufacture of nitric acid. Hydrazine (N_2H_4) is a more powerful reducing agent than ammonia. Suggest what products, other than nitrogen, would be formed when hydrazine is added to
   (a) chlorine,
   (b) a solution containing silver ions, Ag^+ (aq).

5. Anhydrous iron(III) chloride can be prepared by the reaction between hot iron and dry chlorine, freed from hydrogen chloride.
   (a) In this reaction the chlorine oxidizes the iron. Justify the use of the word 'oxidizes'.
   (b) Name a reagent which will reduce aqueous iron(III) chloride to aqueous iron(II) chloride and write an equation for the reduction.

Supplementary Questions (Answers)

Question 1

(a) Oxidation is the increase in oxygen, decrease in hydrogen and decrease in electrons.
(b) To a solution of aqueous potassium iodide, add a few drops of aqueous chlorine. The colourless potassium iodide solution should turn brown, proving that aqueous chlorine is an oxidizing agent (as it was able to react with a reducing agent).

Question 2

(i) (a) Oxidized: Zn
    (b) Oxidizing Agent: H_2SO_4 (or H^+ ions)
    (c) Reduced: H_2SO_4 (or H^+ ions)
    (d) Reducing Agent: Zn
(ii) (a) Oxidized: KI (or I^- ions)
    (b) Oxidizing Agent: Cl_2
    (c) Reduced: Cl_2
    (d) Reducing Agent: KI (or I^- ions)
(iii) (a) Oxidized: Fe^{2+}
    (b) Oxidizing Agent: NO_3^-
    (c) Reduced: NO_3^-
    (d) Reducing Agent: Fe^{2+}

Question 3

The roman numerals refer to the oxidation state of that particular ion/atom. In copper(II) sulfate(VI), the copper ion has an oxidation state of 2 and the sulfate ion has an oxidation state of 6. In potassium manganate(VII), the manganate ion has an oxidation state of 7.
However, silver have only one possible oxidation state (+1), and fluoride ions have only one possible oxidation state (-1). Hence it is unnecessary to indicate their oxidation states in their chemical name.

Question 4

(a) hydrogen chloride / hydrochloric acid
(b) silver metal

Question 5

(a) Iron changes from Fe to Fe^{3+}, losing three electrons in the process. Its oxidation state also increases from zero to +3. Hence it has been oxidized.
(b) Potassium iodide; $2\ FeCl_3 + 2\ KI \longrightarrow 2\ FeCl_2 + 2\ KCl + I_2$ *(other possible answers)*

Lecture Slides

Reduction and Oxidation

Anglo-Chinese School (Barker Road)

CHEMISTRY
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Chapter Overview

In This Chapter, We Will Learn …

1. Definitions – oxygen, hydrogen, electrons
2. Oxidation State
3. Identifying Redox Reactions
4. Reducing & Oxidising Agents
5. Qualitative Analysis

Three Definitions

Reduction	Oxidation
loss of oxygen	gain of oxygen
gain of hydrogen	loss of hydrogen
gain of electrons *(negative charge)*	loss of electrons *(positive charge)*

- Reduction and oxidation occur simultaneously. Why?
- We hence call this a **'redox' reaction**.

Calculating Oxidation State

- Sometimes, there may be a concurrent oxidation and reduction:

$$N_2 \longrightarrow NO_3^-$$

- The nitrogen atom has gained oxygen (oxidised) and gained electrons (reduced) at the same time!
- Hence a more accurate way to determine oxidation or reduction is in calculating the degree of oxidation or reduction, i.e. **calculating oxidation state**.

Calculating Oxidation State

- The oxidation state of elements is zero, e.g. O_2, H_2, P_4, S_8.
- The sum of the oxidation states is equal to the ionic charge of the molecule/compound.
- **For example, carbon monoxide (zero charge):**

CO

C: oxidation state: +2
O: oxidation state: –2

sum of the oxidation states is zero

Calculating Oxidation State

Method #1: Algebraic Approach

Construct an algebraic expression for the sum of the oxidation states of the compound.

Oxidation States of Common Elements:

Oxygen **– 2**
Hydrogen **+ 1**
Metallic Ions **same as ionic charge**

For unknown element, define it as **x**.

Calculating Oxidation State

Method #1: Algebraic Approach

Equate the sum of the oxidation states to the ionic charge of the substance (usually zero).

For example, $K_2\underline{S}O_4$

No net charge.
O.S. of K is +1.
O.S. of S is x.
O.S. of O is –2.

Oxidation States of 2 K + S + 4 O = 0

Hence, 2 (+1) + x + 4 (–2) = 0

x – 6 = 0

x = +6

Must have either plus or minus sign!

Calculating Oxidation State

Method #1: Algebraic Approach

Equate the sum of the oxidation states to the ionic charge of the substance (usually zero).

For example, $\underline{Cr}_2O_7^{2-}$

Net charge of –2.
O.S. of Cr is x.
O.S. of O is –2.

Oxidation States of 2 Cr + 7 O = –2

Hence, 2x + 7 (–2) = –2

2x = 12

x = +6

Must have either plus or minus sign!

Calculating Oxidation State

Method #2: Short-Cut Approach

Add up the 'oxidising effect' of oxygen, hydrogen and electrons on the unknown atom.

Apply These Simple Rules:

Every Oxygen Attached + 2

Every Hydrogen Attached – 1

Every Lost Electron (+ve Charge) + 1

Every Extra Electron (– ve Charge) – 1

Calculating Oxidation State

Method #2: Short-Cut Approach

Apply These Rules:

Every Oxygen Attached	+2
Every Hydrogen Attached	–1
Every Positive Charge	+1
Every Negative Charge	–1

$K_2\underline{S}O_4$ → $2K^+$ + $\underline{S}O_4^{2-}$

Ignore this one, as it does not contain sulfur.

4 Oxygen Atoms
2 Extra Electrons

O.S. = 4(+2) + 2(-1)
= +6

Must have either plus or minus sign!

Calculating Oxidation State

Method #2: Short-Cut Approach

Apply These Rules:

Every Oxygen Attached	+2
Every Hydrogen Attached	–1
Every Positive Charge	+1
Every Negative Charge	–1

$\underline{C}_2H_4O$

Covalent compound – **may not be split.**

1 Oxygen Atom
4 Hydrogen Atoms
shared by 2 Carbon Atoms

O.S. = [1(+2) + 4(-1)] ÷ 2
= -1

Must have either plus or minus sign!

Calculating Oxidation State

Method #2: Short-Cut Approach

Apply These Rules:

Every Oxygen Attached	+2
Every Hydrogen Attached	–1
Every Positive Charge	+1
Every Negative Charge	–1

$\underline{Cl}O_3^-$

Molecular ion – **has already been split.**

3 Oxygen Atoms
1 Extra Electron

O.S. = 3(+2) + (-1)
= +5

Must have either plus or minus sign!

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Redox Reactions

- An **oxidation state** helps us to **quantify** the degree of oxidation or reduction that an atom has undergone.
- Hence, **comparing oxidation states** is the most accurate manner in which to determine whether a substance has been oxidised or reduced.
- If the oxidation state **increases**, the substance has been **oxidised**. If the oxidation state **decreases**, the substance has been **reduced**.

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Redox Reactions

- Determine which elements are oxidised or reduced.

$2\ \underline{Fe}SO_4\ (aq) + \underline{Cl}_2\ (g) + H_2SO_4\ (aq) \rightarrow \underline{Fe}_2(SO_4)_3\ (aq) + 2\ H\underline{Cl}\ (g)$

$\underline{Fe}^{2+}$ SO_4^{2-}	element	$2\ \underline{Fe}^{3+}$ $3\ SO_4^{2-}$	H^+ $\underline{Cl}^-$
O.S. = <u>+2</u>	O.S. = <u>0</u>	O.S. = <u>+3</u>	O.S. = <u>-1</u>

oxidation state increased - OXIDATION

oxidation state decreased - REDUCTION

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Redox Reactions

- Determine which elements are oxidised or reduced.

$\underline{N}H_3\ (g) + \underline{Cu}O\ (s) \rightarrow \underline{N}_2\ (g) + 3\ \underline{Cu}\ (s) + 3\ H_2O\ (l)$

covalent	Cu^{2+} O^{2-}	element	element
O.S. = <u>-3</u>	O.S. = <u>+2</u>	O.S. = <u>0</u>	O.S. = <u>0</u>

oxidation state increased - OXIDATION

oxidation state decreased - REDUCTION

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Redox Reactions

- Determine which elements are oxidised or reduced.

$\underline{Fe}\ (s) + 3\ \underline{Cl}_2\ (g) \rightarrow 2\ \underline{FeCl}_3\ (aq)$

element	element	Fe^{3+} $3\ Cl^-$	Fe^{3+} $3\ Cl^-$
O.S. = <u>0</u>	O.S. = <u>0</u>	O.S. = <u>+3</u>	O.S. = <u>-1</u>

oxidation state increased - OXIDATION

oxidation state decreased - REDUCTION

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Redox Reactions

- Determine which elements are oxidised or reduced.

$\underline{Cl}_2\ (g) + 2\ \underline{I}^-\ (aq) \rightarrow 2\ \underline{Cl}^-\ (aq) + \underline{I}_2\ (s)$

element	ion	ion	element
O.S. = <u>0</u>	O.S. = <u>-1</u>	O.S. = <u>-1</u>	O.S. = <u>0</u>

oxidation state decreased - REDUCTION

oxidation state increased - OXIDATION

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Redox Reactions

We must always calculate the oxidation states <u>before</u> and <u>after</u> each reaction.

A substance is <u>oxidised</u> if there is an <u>increase</u> in oxidation state.

A substance is <u>reduced</u> if there is a <u>decrease</u> in oxidation state.

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Oxidising & Reducing Agents

What is an agent?

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Oxidising & Reducing Agents

Oxidising Agent

oxidises others

gets reduced itself

examples

- acidified potassium manganate(VII), $KMnO_4/H^+$
- acidified potassium dichromate(VI), $K_2Cr_2O_7/H^+$
- reactive halogens, e.g. F_2, Cl_2
- oxygen, O_2
- hydrogen peroxide, H_2O_2

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Oxidising & Reducing Agents

Reducing Agent

reduces others

gets oxidised itself

examples

- aqueous potassium iodide, KI
- aqueous iron(II) chloride, $FeCl_2$
- reactive metals, e.g. Na, K, Ca
- gasesous reducing agents, e.g. H_2, SO_2, CO, NH_3
- hydrogen peroxide, H_2O_2

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chemistry | reduction and oxidation

Oxidising & Reducing Agents

Why do alkali metals make good reducing agents?
Alkali metals lose its one valence electron readily, hence becoming oxidised, and reducing another substance.

Which metal is a better reducing agent – francium (Fr) or sodium (Na)? Why?
Francium, as it is more reactive. It loses its valence electron more easily, hence becoming oxidised more easily. In doing so, it reduces another substance more readily.

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chemistry | reduction and oxidation

Oxidising & Reducing Agents

Why do halogens make good oxidising agents?
Halogens gain a valence electron readily, hence becoming reduced, and oxidising another substance.

Which halogen is a better oxidising agent – chlorine (Cl_2) or fluorine (F_2)? Why?
Fluorine, as it is more reactive. It gains an extra electron more easily, hence becoming reduced more easily. In doing so, it oxidises another substance more readily.

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Qualitative Analysis

- In testing for an **oxidising agent**, we are actually testing whether a substance **can be reduced**.
  - As such, we use a **reducing agent** to do this.
- In testing for a **reducing agent**, we are actually testing whether a substance **can be oxidised**.
  - As such, we use an oxidising agent to do this.

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Test for Reducing & Oxidising Agents

To **test for an oxidising agent**, we add a **reducing agent**.

Powerful Reducing Agent

Aqueous Potassium Iodide, KI

Colourless solution turns into a brown solution.

$$2\ I^- (aq) \longrightarrow I_2 (aq) + 2\ e^-$$

colourless → brown

O.S. = –1 → O.S. = 0

iodide ions get displaced!

Reducing agents get OXIDISED.

Test for Reducing & Oxidising Agents

To **test for an oxidising agent**, we add a **reducing agent**.

easily oxidised, hence acting as a reducing agent

Aqueous Iron(II) Chloride, $FeCl_2$ (or Nitrate, $Fe(NO_3)_2$)

Green solution turns into a yellow solution.

$$Fe^{2+} (aq) \longrightarrow Fe^{3+} (aq) + e^-$$

green → yellow

O.S. = +2 → O.S. = +3

Reducing agents get OXIDISED.

Test for Reducing Agents

To **test for a reducing agent**, we add an **oxidising agent**.

Powerful Oxidising Agent

Acidified Potassium Manganate(VII), $H^+/KMnO_4$

Purple solution decolourises to become colourless.

$$MnO_4^- (aq) + 8\ H^+ (aq) + 5\ e^- \longrightarrow Mn^{2+} (aq) + 4\ H_2O\ (l)$$

purple → colourless

O.S. = +7 → O.S. = +2

Oxidising agents get REDUCED.

Test for Reducing Agents

To **test for a reducing agent**, we add an **oxidising agent**.

Powerful Oxidising Agent

Acidified Potassium Dichromate(VI), $H^+/K_2Cr_2O_7$

Orange solution turns into a green solution.

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

orange → green

O.S. = +6 → O.S. = +3

Oxidising agents get REDUCED.