Redox Reactions

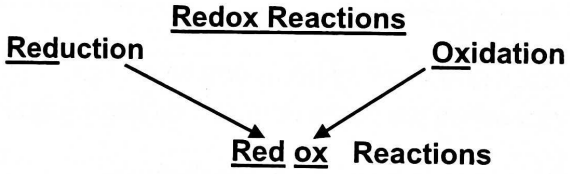
Learning Objectives

- define oxidation and reduction (redox) in terms of oxygen/hydrogen gain/loss
- define redox in terms of electron transfer and changes in oxidation state
- calculate the oxidation states of element in a substance
- identify redox reactions in terms of oxygen/hydrogen gain/loss, electron gain/loss and changes in oxidation state
- describe the use of reagents e.g. aqueous potassium iodide and acidified potassium manganate (VII) in testing for oxidising and reducing agents from the resulting colour changes
- describe and explain the redox reaction involving displacement reactions between a halogen and a compound of another halogen (revision)

A. Introduction

We encounter many chemical reactions in our daily lives.
Some chemical reactions can be described as *redox reactions*.

A redox reaction is a reaction where **oxidation** and **reduction** occur at the same time.



Redox reactions that occur in daily life include:

- Respiration
- Photosynthesis
- Corrosion of metals
- Combustion of fuels
- Decay of food
- Colouring of hair
- Bleaching action
- Nitrogen fixation
- Electrolytic reactions
- Chemical reactions in batteries

B. Oxidation and Reduction

Question:
How do we identify that a reaction involves oxidation and reduction (i.e. is a redox reaction)?

Oxidation is the:	**Reduction** is the:
• Gain of oxygen by a substance	• Loss of oxygen by a substance
• Loss of hydrogen by a substance	• Gain of hydrogen by a substance
• Loss of electrons by a substance	• Gain of electron by a substance
• Increase in oxidation of an element in a substance	• Decrease in Oxidation of an element in a substance

1. Involving gain or loss of OXYGEN

Oxidation : ..gain..... of oxygen by a substance (with no other changes to the substance)

Reduction: ..loss..... of oxygen by a substance (with no other changes to the substance)

E.g. 1

$$CuO(s) + C(s) \rightarrow Cu(s) + CO(g)$$

- C is ..oxidised..... because C .....gains.. oxygen to form CO.
- CuO is ..reduced..... because CuO ..loses..... oxygen to form Cu.

- CuO is the ..oxidising.... agent because CuO ..loses..... **oxygen** to C and **oxidises** C to CO.
- C is the ..reducing..agent because C ..gains..... **oxygen** from CuO and **reduces** CuO to Cu.

An <u>oxidising agent</u> is a substance that ..oxidises..... another substance while itself is being ..reduced.....

A <u>reducing agent</u> is a substance that ..reduces..... another substance while itself is being ..oxidised.....

E.g. 2

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

- CO is ..oxidised..... because CO ..gains..... oxygen to form CO_2.
- H_2O is ..reduced..... because H_2O ..loses..... oxygen to form H_2.

- CO is the ..reducing.... agent because CO ...gains..... **oxygen** from H_2O and ..reduces..... H_2O to H_2.
- H_2O is the ..oxidising..... agent because H_2O..loses..... **oxygen** to CO and ..oxidises.. CO to CO_2.

2. Involving gain or loss of HYDROGEN

Oxidation: loss **of hydrogen by a substance (with no other changes to the substance)**

Reduction: gain **of hydrogen by a substance (with no other changes to the substance)**

E.g. 1

$$Br_2(g) + N_2H_2(g) \rightarrow N_2(g) + 2HBr(g)$$

- Br_2 is reduced because Br_2 gains hydrogen to form HBr
- N_2H_2 is oxidised because N_2H_2 loses hydrogen to form N_2

- Br_2 is the oxidising agent because Br_2 gains hydrogen from N_2H_2 and oxidises N_2H_2 to N_2
- N_2H_2 is the reducing agent because N_2H_2 loses hydrogen to Br_2 and reduces Br_2 to HBr

E.g. 2

$$Cl_2(g) + H_2S(g) \rightarrow 2HCl(g) + S(s)$$

- Cl_2 is reduced because Cl_2 gains hydrogen to form HCl
- H_2S is oxidised because H_2S loses hydrogen to form S

- Cl_2 is the oxidising agent because Cl_2 gains hydrogen from H_2S and oxidises H_2S to S
- H_2S is the reducing agent because H_2S loses hydrogen to Cl_2 and reduces Cl_2 to HCl

Quick Check Exercise 1

1 For each of the following reactions, state and explain which substance is oxidised and which is reduced.

(a) $Zn + CuO \rightarrow ZnO + Cu$

- Zn is **oxidised** because Zn gains oxygen to form ZnO
- CuO is **reduced** because CuO loses oxygen to form Cu

(b) $2HI + Br_2 \rightarrow 2HBr + I_2$

- HI is **oxidised** because HI loses hydrogen to form I_2
- Br_2 is **reduced** because Br_2 gains hydrogen to form HBr

2 For each of the following reactions, state and explain which substance is the oxidising agent and which is the reducing agent.

(a) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

- Fe_2O_3 is the **oxidising agent** because Fe_2O_3 loses oxygen to CO and oxidises CO to CO_2
- CO is the **reducing agent** because CO gains oxygen from Fe_2O_3 and reduces Fe_2O_3 to Fe

(b) $2NH_3 + 3Cl_2 \rightarrow N_2 + 6HCl$

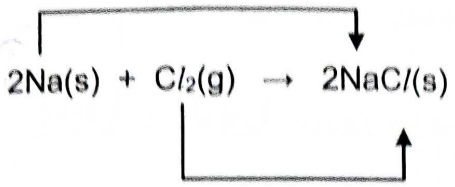
- Cl_2 is the **oxidising agent** because Cl_2 gains hydrogen from NH_3 and oxidises Cl_2 to HCl.
- NH_3 is the **reducing agent** because NH_3 loses hydrogen to Cl_2 and reduces NH_3 to N_2.

3. INVOLVING GAIN OR LOSS OF ELECTRONS

Oxidation: ...loss..... **of electrons by a substance**

Reduction: ...Gain..... **of electrons by a substance**

E.g. 1



- Na is ...oxidised..... because Na ...loses..... electron to form ...Na+.....

  Half equation: ...Na → Na+ + e-.....

- Cl_2 is ...reduced..... because Cl_2 ...gains..... electrons to form ...Cl-.....

  Half equation: ...Cl2 + 2e- → 2Cl-.....

- Na is the ...reducing.... agent because Na ...loses..... electrons to Cl_2 and ...reduces..... Cl_2 to ...Cl-.....

- Cl_2 is the ...oxidising....agent because Cl_2 ...gains..... electrons from Na and ...oxidises.. Na to ...Na+.....

** The separate equations showing which substance gains electron(s) and which substance loses electron(s) are known as ...half equations.....

Tip to Remember (Gain / Loss of electron):

OIL : **O**xidation **I**s **L**oss of electrons
RIG : **R**eduction **I**s **G**ain of electrons

3rd of of

OR

LEO : **L**oss of **E**lectrons is **O**xidation.
GER : **G**ain of **E**lectrons is **R**eduction
LEO the lion says **GER**

QUICK CHECK EXERCISE 2a

Complete the following half equations and identify if the process is oxidation or reduction.

Half equation	Oxidation / Reduction?
$Ca \rightarrow Ca^{2+} + 2e^-$	oxidation
$Al \rightarrow Al^{3+} + 3e^-$	oxidation
$Fe^{2+} + 2e^- \rightarrow Fe$	Reduction
$Cr^{3+} + 3e^- \rightarrow Cr$	Reduction
$2H^+ + 2e^- \rightarrow H_2$	Reduction
$2I^- \rightarrow I_2 + 2e^-$	Reduction oxidation
$2O^{2-} \rightarrow O_2 + 4e^-$	Reduction oxidation
$Br_2 + 2e^- \rightarrow 2Br^-$	Oxidation. reduction.
$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$	Reduction
$Fe^{2+} \rightarrow Fe^{3+} + e^-$	Oxidation.

E.g. 2

$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

$Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

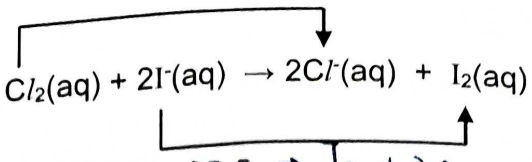
Ionic equation is $Mg(s) + SO_4^{2-}(aq) \rightarrow MgSO_4(aq)$

Half equation for oxidation: $Mg \rightarrow Mg^{2+} + 2e^-$

Half equation for reduction: $2H^+ + 2e^- \rightarrow H_2$

- Mg is oxidised because Mg loses electrons to form $MgSO_4$
- H_2SO_4 is reduced because SO_4 gains electrons to form $MgSO_4$ H_2
- H_2SO_4 is the oxidising agent because H^+ gains electrons from Mg and oxidises Mg to Mg^{2+}
- Mg is the reducing agent because Mg loses electrons to H^+ and reduces H^+ to H_2

E.g. 3



Half equation for oxidation: $2I^- \rightarrow I_2 + 2e^-$

Half equation for reduction: $Cl_2 + 2e^- \rightarrow 2Cl^-$

- I is oxidised because I loses electrons to form I_2
- Cl is reduced because Cl gains electrons to form Cl^-
- Cl_2 is the oxidising agent because Cl_2 gains electrons from I^- and oxidises I^- to I_2
- 2I is the reducing agent because 2I loses electrons to Cl_2 and reduces Cl_2 to Cl^-

QUICK CHECK EXERCISE 2b

1 Do the following processes involve oxidation, reduction, both oxidation and reduction or none of these?

(a) $Cu^+ \rightarrow Cu^{2+} + e$ oxidation

(b) $Ag^+ + Cl^- \rightarrow AgCl$ none both ✗

(c) $Br_2 + 2e \rightarrow 2Br^-$ reduction

(d) $2Cu^+ \rightarrow Cu^{2+} + Cu$

2 For the following reactions, write the **half equations** for each reaction and explain why they are redox reactions in terms of **electrons**.

(a) $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

Half equation for Oxidation: $Al \rightarrow Al^{3+} + 3e^-$

Half equation for Reduction: $O_2 + 4e^- \rightarrow 2O_2^-$

Explanation:

- Al is oxidised because Al loses electrons to form Al^{3+}
- O_2 is reduced because O_2 gains electrons to form O_2^-

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

Half equation for Oxidation: $Cu - 2e^- \rightarrow Cu^{2+}$

Half equation for Reduction: $2Ag^+ + 2e^- \rightarrow 2Ag$

Cu is oxidised because Cu loses electrons to form Cu^{2+}

Ag^+ is reduced because Ag gains electrons to form Ag

(c) $2FeCl_2(aq) + Cl_2(g) \rightarrow 2FeCl_3(aq)$ $2Fe^{2+}(aq) + Cl_2(g) \rightarrow 2Fe^{3+} + 2Cl^-(aq)$

Ionic Equation: $Fe^{2+}(aq) + Cl_2(g) \rightarrow Fe^{3+}(aq)$ ✗

Half equation for Oxidation: $Fe^{2+} - e^- \rightarrow Fe^{3+}$

Half equation for Reduction: $Cl_2 + e^- \rightarrow 3Cl^-$

Fe^{2+} is oxidised because Fe^{2+} loses electrons to form Fe^{3+}

Cl_2 is reduced because Cl gains electrons to form Cl^-

3 $2KI(aq) + Br_2(aq) \rightarrow 2KBr(aq) + I_2(aq)$

Ionic Equation: $2I^-(aq) + Br_2(aq) \rightarrow 2Br^-(aq) + I_2(aq)$

For the above reaction,
(a) write the half equations,
(b) state and explain which substance is the oxidising agent and which is the reducing agent.

Half equation for Oxidation: $2I^- \rightarrow I_2 + 2e$

Half equation for Reduction: $Br_2 + 2e \rightarrow 2Br$

Oxidising agent: Br_2

Explanation:

Br_2 gains electrons from I^- and oxidises I^- to form I_2.

Reducing agent: KI

Explanation:

4. OXIDATION STATE / OXIDATION NUMBER

Chemists have worked out an alternative method to tell if reactions are redox reactions. This method uses the idea of oxidation state/oxidation number.

The **oxidation state / oxidation number** is the **charge** an atom would have if it existed as an **ion**.

An atom in a **covalent**. substance can be assigned oxidation state/oxidation number by assuming that the substance is **ionic**.

This number represents the number of electrons that an atom can gain, lose, or share when chemically bonded with an atom of another element.

.....

RULES GOVERNING OXIDATION STATE/OXIDATION NUMBER

To work out or assign oxidation states/oxidation numbers to elements in different substances, we follow several rules:

(1) Oxidation state of an element = 0.

e.g. Oxidation state of **sodium** in **Na** = **0**

e.g. Oxidation state of **carbon** in **C** = **0**

e.g. Oxidation state of **chlorine** in **Cl_2** = **0**

e.g. Oxidation state of **phosphorus** in **P_4** = **0**

(2) Oxidation state of an element in a simple ion = charge on the ion.

e.g. Oxidation state of **lithium** in **Li^+** = **+1**

e.g. Oxidation state of **aluminium** in **Al^{3+}** = **+3**

e.g. Oxidation state of **fluorine** in **F^-** = **-1**

e.g. Oxidation state of **oxygen** in **O^{2-}** = **-2**

(3) Sum of oxidation states of all elements in a compound = 0.

e.g. Sum of oxidation states of all elements in **NaCl** = **0**

- Oxidation state of **sodium** + Oxidation state of **chlorine** = $(+1)+(-1)=0$

e.g. Sum of oxidation states of all atoms in **Al_2O_3** = **0**

- 2(Oxidation state of **aluminium**) + 3(Oxidation state of **oxygen**) = $2(+3)+3(-2)=0$

e.g. Sum of oxidation states of all elements in **H_2O** = **0**

- 2(Oxidation state of **hydrogen**) + Oxidation state of **oxygen** = 0

(4) Sum of oxidation states of all elements in a polyatomic ion = charge on the ion.

e.g. Sum of oxidation states of all elements in OH^- = **-1**

- Oxidation state of **hydrogen** + Oxidation state of **oxygen** = ... $-1 + (-2) = -1$ .....

e.g. Sum of oxidation states of all elements in CO_3^{2-} = **-2** $(+4) + 3(-2) = -2$

- Oxidation state of **carbon** + 3(Oxidation state of **oxygen**) = ...

(5) Oxidation state of some elements in their compounds is fixed.

e.g.	Oxidation state of all **Group 1 metals** in their compounds = ..+1.. - sodium in NaCl - rubidium in Rb_2SO_4
e.g.	Oxidation state of all **Group 2 metals** in their compounds = ..+2.. - magnesium in MgO - barium in $BaCO_3$
e.g.	Oxidation state of **hydrogen** in most of its compounds (covalent) = ..+1.. - hydrogen in H_2O, HCl, NH_3, CH_4 ***Exception:** Oxidation state of **hydrogen** in metal hydrides (ionic) = ..-1.. - hydrogen in NaH, MgH_2
e.g.	Oxidation state of **oxygen** in most of its compounds = ..-2.. - oxygen in MgO, $Ca(OH)_2$, H_2O, CO_2 ***Exception:** Oxidation state of **oxygen** in peroxides = ..-1.. - oxygen in H_2O_2, Na_2O_2, BaO_2

(6) Oxidation state of some elements can vary in different compounds.

e.g. Oxidation state of **many transition metals** such as Fe, Cu, Mn, Cr, etc. can ..vary.... in their different compounds.

e.g. Oxidation state of many **non-metals in Groups 14 to 17** such as C, N, S, Cl, Br, I, etc. can ..vary... in their different compounds.

OXIDATION STATE (O.S.) CHART

O.S.	zinc	iron	copper	manganese	chromium	carbon	nitrogen	sulfur	chlorine
+7				$KMnO_4$					
+6				K_2MnO_4	$K_2Cr_2O_7$ K_2CrO_4			H_2SO_4 SO_3	
+5							KNO_3		$NaClO_3$
+4				MnO_2		$CaCO_3$ CO_2	NO_2	SO_2	
+3		$FeCl_3$			$CrCl_3$ Cr_2O_3		HNO_2		
+2	$ZnCl_2$ $ZnSO_4$	$FeCl_2$	$CuSO_4$ CuO	$MnSO_4$	$CrCl_2$	CO	NO	$Na_2S_2O_3$	
+1			$CuCl$ Cu_2O				N_2O		$NaClO$
0	Zn	Fe	Cu	Mn	Cr	C	N_2	S	Cl_2
-1						C_2H_2			$NaCl$
-2						C_2H_4		FeS	
-3						C_2H_6	NH_3		
-4						CH_4			

OXIDATION STATE & ELECTRONEGATIVITY

Question: What is the oxidation state of **nitrogen** in (i) NH_3 (ii) NCl_3?

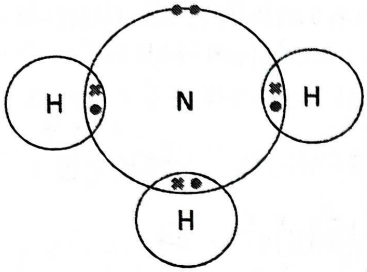
(i) Sum of oxidation states of all atoms in NH_3 = ...
(Oxidation state of **nitrogen**) + 3(Oxidation state of **hydrogen**) =
(Oxidation state of **nitrogen**) + 3( ) =
Oxidation state of **nitrogen** =

(ii) Sum of oxidation states of all atoms in NCl_3 = ...
(Oxidation state of **nitrogen**) + 3(Oxidation state of **chlorine**) =
(Oxidation state of **nitrogen**) + 3( ) =
Oxidation state of **nitrogen** =

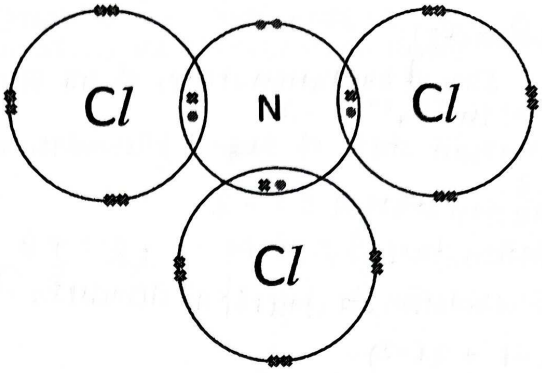
How can we explain the difference in oxidation state of nitrogen in the compounds above?

Explanation

In **covalent compounds,** the **negative oxidation state** is assigned to the **most electronegative atom.**



In NH_3, **N** is ..more.. electronegative than H. Hence oxidation state of N in NH_3 = ..-3..



In NCl_3, **N** is ..less.. electronegative than Cl. Hence oxidation state of N in NCl_3 = ..+3..

What is Electronegativity?

- **Electronegativity** refers to the ability of an atom to **attract** the **shared pair of electrons** towards itself in a covalent bond:
  - **Electronegativity increases across a period**, e.g. **F** is **more electronegative** than **N**.
  - **Electronegativity decreases down a group**, e.g. **Br** is **less electronegative** than **F**.
  - Electronegativity values: **F > O > Cl > N > Br > I > C > H**

Extra reading:
- http://www.chemguide.co.uk/atoms/bonding/electroneg.html
- http://en.wikipedia.org/wiki/Electronegativity

E.g.
- In water, H_2O, oxidation state of oxygen is ..-2.. and oxidation state of hydrogen is ..+1..
- In dichlorine monoxide, Cl_2O, oxidation state of chlorine is ..+1.. and oxidation state of oxygen is ..-2..
- In iodine monochloride, ICl, oxidation state of iodine is ..+1.. and oxidation state of chlorine is ..-1..
- In tetraiodomethane, CI_4, oxidation state of carbon is ..+4.. and oxidation state of iodine is ..-1..
- In sodium chlorate(V), $NaClO_3$, the oxidation state of sodium is ..+1.. and oxidation state of oxygen is ..-2.. and the oxidation state of chlorine is ..+5..

QUICK CHECK EXERCISE 3

1 Determine the **oxidation state** of the **named element** in each given substance:

(a) sulfur in SO_4^{2-} (b) sulfur in Na_2SO_3

(c) nitrogen in NO_2^- (d) nitrogen in $Mg(NO_3)_2$

(e) manganese in $KMnO_4$ (f) chromium in K_2CrO_4

sulfate/sulfate(VI)

Sum of oxidation states of all elements

(a) in SO_4^{2-} = -2

oxidation state of ~~S to~~ S + 4(oxidation state of O) = -2.

Oxidation state of S = -2

Oxidation state of sulfur: -2 + 8 = +6

(b) oxidation states of all elements = 0

2(+1) + oxidation state of S + 3(-2) = 0

oxidation state of S = +6 - 2

= +4

SO_3^{2-} sulfite/sulfate (IV)

(c) oxidation ~~sta~~ states of all elements = -1

-1 - 2(-2)

= -1 + 4

= +3

(d) Oxidation State of ~~all~~ all = 0

Oxidation state of Mg + oxidation state of N

= +2 + 2(state of N) + 6(-2)

state of N = $\frac{10}{2}$ = +5

(e)

K = +1

+1 + Mn + 4(-2) = 0

Mn = +8 - 1

= +7 //

(f)

2(+1) + Cr + 4(-2) = 0

Cr = +8 - 2

= +6 //

2 Determine the oxidation state/oxidation number of the named element in each of the substances below:

hydrogen in PH_3	+1	hydrogen in NaH	-1	hydrogen in HCO_3^-	+1
oxygen in H_2O_2	-1	oxygen in O_3 (ozone)	0	carbon in CO (reduced)	+2
carbon in K_2CO_3 (+4, -6)	+4	carbon in C_2H_6 (-3, 6-)	-3	fluorine in F_2	0
lead in PbO_2	+4	sulfur in $MgSO_4$	+6	sulfur in H_2S	-2
sulfur in $CaSO_3$	+4	sulfur in H_2SO_3	+4	sulfur in SO_3	+6
sulfur in S or S_8	0	sulfur in $Na_2S_2O_3$	+2	nitrogen in NH_3	-3
nitrogen in NO_2	+4	nitrogen in $Cu(NO_3)_2$	+5	nitrogen in NH_4^+	-3
nitrogen in $Al(NO_2)_3$	+3	chlorine in ClO^- (-1, -1-2 = -3)	+1	chlorine in $NaClO_3$	+5
phosphorus in K_2HPO_4	+5	phosphorus in H_3PO_4	+5	phosphorus in P_4O_{10}	+5
zinc in $Zn(OH)_2$	+2	silver in AgBr	+1	copper in Cu_2O	+1
iron in $FeBr_2$	+2	iron in Fe_2O_3	+3	iron in $FeCl_3$	+3
chromium in $K_2Cr_2O_7$	+6	chromium in Na_2CrO_4	+6	chromium in $Cr_2(SO_4)_3$	+3
vanadium in $(VO_2)_2SO_4$	+5	manganese in MnO_2	+4	iodine in PbI_2	-1

$(VO_2)_2SO_4$: $2+$ and SO_4^{2-}

Chemistry/ Redox Reaction

VO_2^+

OXIDATION & REDUCTION AS CHANGES IN OXIDATION STATE

Oxidation is the increase in **oxidation state** of an **element** in a substance during a reaction.

Reduction is the decrease in **oxidation state** of an **element** in a substance during a reaction.

E.g. 1

$$Mg + CuO \rightarrow \overset{+2\ -2}{MgO} + Cu$$

- Mg is oxidised because the oxidation state of magnesium increases from 0 in Mg to +2 in MgO
- CuO is reduced because the oxidation state of copper decreases from +2 in CuO to 0 in Cu
- Reducing agent is Mg (itself oxidised)
- Oxidising agent is CuO (itself reduced)

E.g. 2

$$Fe + 2HCl \rightarrow \overset{2+}{FeCl_2} + H_2$$

- Fe is oxidised because the oxidation state of Iron increases from 0 in Fe to +2 in $FeCl_2$
- HCl is reduced because the oxidation state of hydrogen decrease from +1 in HCl to 0 in H_2
- Reducing agent is Fe
- Oxidising agent is HCl

E.g. 3

$$Br_2(aq) + SO_2(g) + H_2O(l) \rightarrow 2HBr(aq) + SO_3(g)$$

- SO_2 is oxidised because the oxidation state of sulfur increases from +4 in SO_2 to +6 in SO_3

- Br_2 is reduced because the oxidation state of Bromine decreases from 0 in Br_2 to -1 in HBr

- Reducing agent is SO_2
- Oxidising agent is Br_2

E.g. 4 **Spot (by circling) and correct the mistakes** in the following example to explain why the following reaction is redox.

$$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(aq) + 2H_2O(l)$$

This is a redox reaction because

- Cu is oxidised because the oxidation state of copper changes from 0 in Cu to +2 in $Cu(NO_3)_2$.
- N is reduced because the oxidation state of nitrogen reduces from to +5 in $4HNO_3$ to +2 in $2NO_2$.

E.g. 5

$$2H_2O_2 \rightarrow 2H_2O + O_2$$

- H_2O_2 is oxidised because the oxidation state of oxygen H_2 decreases from -1 in H_2O_2 to 0 in O_2

- H_2O_2 is reduced because the oxidation state of oxygen decreases from -1 in H_2O_2 to -2 in H_2O

- Reducing agent is H_2O_2
- Oxidising agent is H_2O_2

Note:

H_2O_2 is both oxidised and reduced.

This is an example of a disproportionation reaction where a **single substance** is both ***oxidised*** *and* **reduced** in the reaction.

Examples of Chemical Reactions that are Not Redox Reactions

- Note that oxidation states of all elements in the substances remain unchanged during the reaction.

E.g.1 Neutralisation

(a) Neutralisation (between soluble base and acid)

- $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

(b) Neutralisation (between insoluble base and acid)

- $H_2SO_4(aq) + CuO(s) \rightarrow CuSO_4(aq) + H_2O(l)$

E.g.2 Precipitation reactions

(a) Precipitation (formation of insoluble base)

- $CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$

(b) Precipitation (formation of insoluble salt)

- $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

E.g.3 Reaction of acids with carbonates

- $2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

E.g.4 Thermal Decomposition of carbonates

- $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$

QUICK CHECK EXERCISE 4

Identify which of the following are redox reactions. Explain your answer in terms of oxidation state.

(a) $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

(b) oxidised reduced
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

(c) $3HNO_2(aq) \rightarrow HNO_3(aq) + 2NO(g) + H_2O(l)$

(d) reduced oxidised
$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$

(e) $Cr_2O_7^{2-} + 3SO_2 + 2H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + H_2O$
-14

(a) Not redox reaction

(b) Redox reaction
CH_4 is oxidised: Oxidation state of carbon increases from -4 in CH_4 to +4 in CO_2.
O_2 is reduced: oxidation state of oxygen decreases from 0 in O_2 to -2 in H_2O and CO_2.

(c) Redox reaction (disproportionation)
HNO_3 is oxidised: Oxidation state of nitrogen increases from +3 in HNO_2 to +5 in NO_3^-.
HNO_2 is reduced: Oxidation state of nitrogen decreases from +3 in HNO_2 to +2 in NO.

(d) ~~Redox reaction~~ Not a redox reaction!
NH_3 is ~~reduced~~:
~~HCl is oxidised~~:

(e) Redox reaction.
SO_2 is oxidised: Oxidation state of sulfur increases from +4 in SO_2 to +6 in SO_4^{2-}
$Cr_2O_7^{-2}$ is reduced because oxidation state of chromium decreases from +6 in $Cr_2O_7^{2-}$ to +3 in Cr^{3+}.

5. TESTS FOR REDUCING AND OXIDISING AGENTS

We can test for the presence of reducing agents and oxidising agents (in unknowns) by using suitable known oxidising agents or reducing agents and **observing any colour changes** that occur.

Some of the common oxidising agents and reducing agents used in the laboratory for such purpose include:

Oxidising Agent used for Testing Reducing Agent	**Reducing Agent used for Testing Oxidising Agent**
acidified aqueous potassium manganate(VII)	aqueous potassium iodide
acidified aqueous potassium dichromate(VI)	aqueous iron(II) sulfate
aqueous iron(III) chloride	
aqueous iodine	

TESTS FOR REDUCING AGENTS

The following oxidising agents are used to test for the presence of a reducing agent:

Oxidising agent used	**Observation**	**Explanation**
acidified aqueous potassium manganate(VII), $KMnO_4$ (acidified with dilute H_2SO_4)	Purple acidified aqueous potassium manganate(VII) turns colourless.	$MnO_4^- + 8H^+ + 5e \rightarrow Mn^{2+} + 4H_2O$ purple colourless $KMnO_4$ is reduced because the oxidation state of manganese decreases from _____ in MnO_4^- to _____ in Mn^{2+}.
acidified aqueous potassium dichromate(VI), $K_2Cr_2O_7$ (acidified with dilute H_2SO_4)	orange acidified aqueous potassium dichromate(VI) turns green.	$Cr_2O_7^{2-} + 14H^+ + 6e \rightarrow 2Cr^{3+} + 7H_2O$ orange green $K_2Cr_2O_7$ is reduced because the oxidation state of chromium decreases from _____ in $Cr_2O_7^{2-}$ to _____ in Cr^{3+}.
aqueous iron(III) chloride, $FeCl_3$	yellow aqueous iron(III) chloride solution turns pale green. *(How can we test for the presence Fe^{2+} in the reacting mixture?)*	$Fe^{3+} + e \rightarrow Fe^{2+}$ yellow pale green $FeCl_3$ is reduced because the oxidation state of iron decreases from _____ in Fe^{3+} to _____ in Fe^{2+}.

$0 = 1 + N - 3$

$0 = 2 + N$

aqueous iodine, I_2	Brown aqueous iodine turns colourless.	$I_2 + 2e \rightarrow 2I^-$ brown colourless I_2 is reduced because the oxidation state of iodine decreases from _____ in I_2 to _____ in I^-.

TESTS FOR OXIDISING AGENTS

The following reducing agents are used to test for the presence of an oxidising agent:

Reducing agent used	Observation	Explanation
aqueous potassium iodide, KI	colourless aqueous potassium iodide turns brown (black solid may also be seen if excess iodine is formed) *(How can we test for the presence of iodine in the reacting mixture?)*	$2I^- \rightarrow I_2 + 2e$ colourless brown KI is _____ because the oxidation state of iodine increases from _____ in I^- to _____ in I_2
aqueous iron(II) sulfate, $FeSO_4$	Pale green aqueous iron(II) sulfate turns yellow.	$Fe^{2+} \rightarrow Fe^{3+} + e$ pale green yellow $FeSO_4$ is oxidised because the oxidation state of iron increases from _____ in Fe^{2+} to _____ in Fe^{3+}.

OTHER EXAMPLES OF OXIDISING AND REDUCING AGENTS

Oxidising Agent	Reducing Agent
non-metals such as Cl_2, Br_2, I_2	metals such as K, Na, Mg
oxygen	hydrogen gas
nitric acid	carbon
ozone	carbon monoxide
concentrated sulfuric acid	sulfur dioxide