SANGAM SKM COLLEGE NADI

LESSON NOTES

WEEK 1

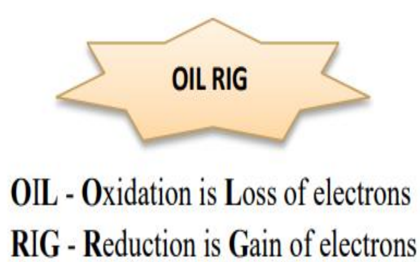
CHEMISTRY

YEAR 12

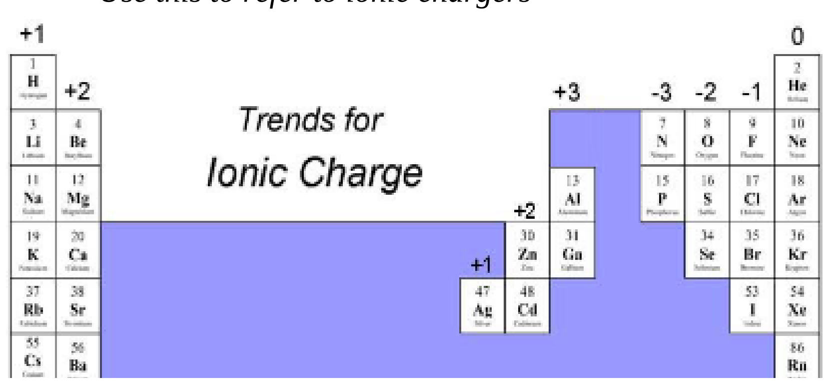
Strand	3. Reactions
Sub-Strand	3.2 Oxidation and Reduction
Content Learning Outcome	Investigate redox reactions and its application in the production of some useful metals

Oxidation Reduction Terminology

Term	Transfer of atoms	Transfer of electrons	Change in oxidation number
Oxidation	Gain of oxygen E.g. $Mg + O \rightarrow MgO$ Loss of Hydrogen E.g. $NaH \rightarrow Na + H$	Loss of Electron E.g. $Al \rightarrow Al^{3+} + 3e$	Increase in oxidation number
Reduction	Loss of oxygen E.g.$6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6O_2$ Gain of Hydrogen E.g. $H_2 + F_2 \rightarrow 2\ HF$	Gain of Electron E.g. $Cu^{2+} + 2e \rightarrow Cu$	Decrease in oxidation number
Oxidant (Oxidising agent)	Substance that loses oxygen or gains hydrogen	Substance that gains electron or an electron acceptor	Substance whose oxidation number has decreased
Reductant (Reducing agent)	Substance that gains oxygen or loses hydrogen	Substance that loses electron or an electron donor	Substance whose oxidation number has increased

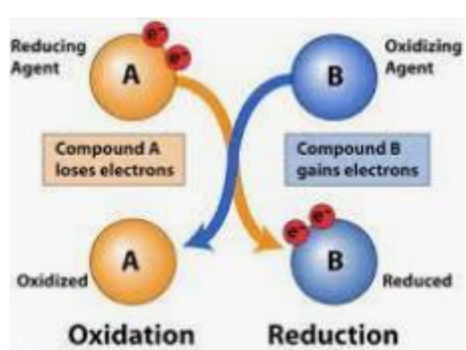


Use this to refer to ionic chargers



Oxidising Agent			
Name	**Appearance**	**Use**	**Reduction Equation**
Oxygen (O_2)	- Colourless gas	- Combustion - Oxidation of metals	$O_2 \rightarrow O^{2-}$
Chlorine (Cl_2)	- Greenish yellow pungent gas	- Bleaching - Disinfectant	$Cl_2 \rightarrow 2Cl^-$
Permanganate ion (MnO_4^-)	- Purple coloured solution	- Breathalyzer test - Water treatment and disinfection - Synthesis of organic compounds - Oxidation of alcohols	$MnO_4^- \rightarrow Mn^{2+}$
Dichromate ion ($Cr_2O_7^{2-}$)	- Orange solution	- Chrome plating to protect metals from corrosion and to improve paint adhesion. - photographic screen printing - Breathalyzer test - Wood treatment - Sulfur dioxide test - Oxidation of alcohols	$Cr_2O_7^{2-} \rightarrow Cr^{3+}$
Hydrogen peroxide (H_2O_2)	- Colourless liquid	- Bleaching - Disinfectant/ Antiseptic	$H_2O_2 \rightarrow H_2O$
Dilute acids (H^+)	- Colourless	- Oxidation of metals	$H^+ \rightarrow H_2$

Reducing Agent			
Name	**Appearance**	**Use**	**Oxidation Equation**
Metal e.g. Zn Mg Fe	- Silvery shiny metals	- Formation of metal oxides.	$Zn \rightarrow Zn^{2+}$ $Mg \rightarrow Mg^{2+}$ $Fe \rightarrow Fe^{2+}$
Carbon	- Black solid- Charcoal	-Used in the smelting process in the production of metals.	$C \rightarrow CO_2$
Sulphur Dioxide	- Colourless gas - Irritating smell - Gives colourless sulphite ion in solution	- Used as a preservative as it delays the oxidation of food by bacteria, as a fumigant and bleaching agent.	$SO_2 \rightarrow SO_4^{2-}$ $SO_3^{2-} \rightarrow SO_4^{2-}$
Ferrous ion	- Pale green solution	- Formation of iron oxides and hydroxides.	$Fe^{2+} \rightarrow Fe^{3+}$
Carbon monoxide (CO)	- Is a colourless and odorless	- Important industrial gas, which is widely used as a fuel.	$CO \rightarrow CO_2$



Exercise

1. What are two different definitions of oxidation?
2. What are two different definitions of reduction?
3. Which substance loses electrons and which substance gains electrons in this reaction?

    i. $2Mg(s) + O_2(g) \rightarrow 2MgO$
    ii. $2Li(s) + O_2(g) \rightarrow Li_2O_2(s)$
    iii. $2Fe(s) + 3I_2(s) \rightarrow 2FeI_3(s)$

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LESSON NOTES

WEEK 2

CHEMISTRY

YEAR 12

Strand	3. Reactions
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LESSON NOTES: 3.2.2 Oxidation number (Oxidation state)

An oxidation number is a number that is assigned to an element in a chemical reaction to show the total number of electrons which have been

-----removed from an element (a positive oxidation state) or

---- added to an element (a negative oxidation state) to get to its present state.

<u>Rules for assigning oxidation number</u>

1. The oxidation number of an atom is zero in a neutral substance that contains only one type of element. Example: For O_2 and Mg, the oxidation number is 0.

   $O_{2(g)} = 0$

   $Mg_{(s)} = 0$ (***note the state)

2. The oxidation number of each oxygen atom in a compound is **-2**, except in peroxides (e.g. H_2O_2) where the oxidation number is **-1.**

Eg. CO_2	Eg: H_2O_2
C = +4	H=+1
O = -2	O = -1

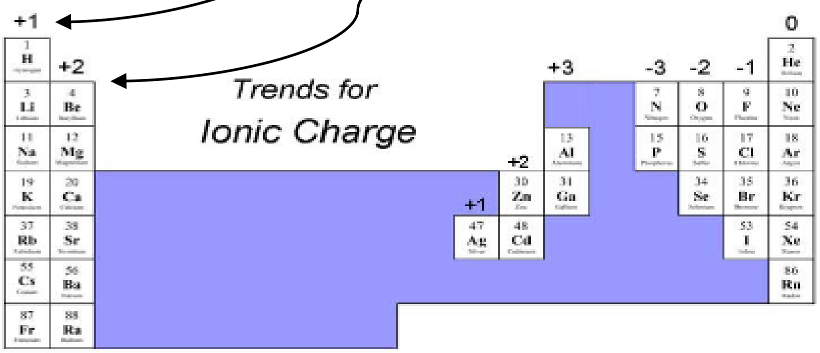
3. The oxidation number of each hydrogen atom in a compound is **+1** except in **metallic hydrides** (example in LiH, NaH, CaH_2, and $LiAlH_4$) where it is **-1**.

Li = +1

H = -1

4. The oxidation number of **Group I metals is +1**.

5. The oxidation number of **Group II metals is** +2.



6. The sum of the oxidation number in a neutral molecule is equal to zero (0). Example: For $C_6H_{12}O_6$, the oxidation number is 0.

7. The oxidation number of an atom in a monoatomic ion is equal to the charge on the ion. Example for: $Na+_{(aq)}$ = +1 ,$Cl-_{(aq)}$ = -1 (***note the state/molten/aqueous state)

8. The sum of the oxidation numbers in a polyatomic ion is equal to the charge on the ion. Example for H_3O^+ , the oxidation number is +1.

Common Polyatomic Ions

Name	Formula	Charge
ammonium	NH_4^+	1+
hydroxide	OH^-	1–
acetate	$C_2H_3O_2^-$	1–
permanganate	MnO_4^-	1–
nitrate	NO_3^-	1–
chlorate	ClO_3^-	1–
bicarbonate	HCO_3^-	1–
carbonate	CO_3^{2-}	2–
sulfate	SO_4^{2-}	2–
phosphate	PO_4^{3-}	3–

all end in *–ide* Monatomic	*typically* end in *–ate* Polyatomic	(*most* have O's)
nitride N^{3-}	nitrate NO_3^-	ammonium NH_4^+
chloride Cl^-	chlorate ClO_3^-	hydroxide OH^-
sulfide S^{2-}	sulfate SO_4^{2-}	
phosphide P^{3-}	phosphate PO_4^{3-}	

Names and Formulas for Compounds with Polyatomic Ions

Same rules as with all compounds.

Example: calcium nitrate

1. symbols of ions — Ca NO_3
2. charges (as superscripts) — Ca^{2+} NO_3^-
3. crisscross (to get subscripts) — Ca^{2+} NO_3^-
4. erase charges — $CaNO_{32}$?
5. use ( ) for multiple polyatomic ions — $Ca(NO_3)_2$

Exercise: CHEMISTRY FOR YEAR 12 Page 73 Questions 1,3, 4

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LESSON NOTES

WEEK 3

CHEMISTRY

YEAR 12

Strand	3. Reactions
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LESSON NOTES: 3.2.3 Balancing REDOX equations

Before actually beginning to balance redox reactions, for each reaction you should be able to:

i. Identify the species which is reacting but should not be involved in the redox equation.

ii. Decide what the products are for each reacting species.

iii. Work out which part is oxidation and which is reduction.

iv. Formulate two half-equations.

Rules for balancing redox equations

1. Obtain the half-equations.
2. For each half-equation:
   i. Balance all atoms except oxygen and hydrogen.
   ii. Balance oxygen by adding water molecules (H_2O).
   iii. Balance hydrogen by adding hydrogen ions, H^+ .
   iv. Balance charges by adding electrons.
3. Add the balanced half-equations so that the electrons cancel out.
4. Identify and cancel out terms common to both sides of the equation.

Example

Balance the following redox equation.

$SO_2 + NO_3^- \rightarrow SO_4^{2-} + NO$

Solution

Reductant: SO_2

Oxidant: NO_3^-

Reducing half-equation	Oxidising half-equation
$NO_3^- \rightarrow NO$	$SO_2 \rightarrow SO_4^{2-}$
Balancing	Balancing
$NO_3^- \rightarrow NO + 2H_2O$	$SO_2 + 2H_2O \rightarrow SO_4^{2-}$
$NO_3^- + 4H^+ \rightarrow NO + 2H_2O$	$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+$
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	$SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-$

Add both the equations

$[NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O] \times 2$

$[SO_2 + 2H_2O \rightarrow SO_4^{2-} + 4H^+ + 2e^-] \times 3$

Multiply both equations by the smallest possible whole number **to cancel out the electrons.**

$2NO_3^- + 8H^+ + 6e^- \rightarrow 2NO + 4H_2O$

$3SO_2 + 6H_2O \rightarrow 3SO_4^{2-} + 12H^+ + 6e^-$

$2NO_3^- + \cancel{8H^+} + \cancel{6e^-} \rightarrow 2NO + \cancel{4H_2O}$

$3SO_2 + \cancel{6H_2O} \rightarrow SO_4^{2-} + \cancel{12H^+} + \cancel{6e^-}$

$2H_2O \qquad 4H^+$

Cancel out terms common to both sides of the equation and identify the "balance" remaining as shown below.

Combine the two equations:

$2NO_3^- + 3SO_2 + 2H_2O \rightarrow 2NO + 3SO_4^{2-} + 4H^+$

Check that the number /types of atoms and the charge are equal in both sides of the equation.

Exercise: CHEMISTRY FOR YEAR 12 Page 73 Questions 1,2, 6