PROSPERITY ACADEMY

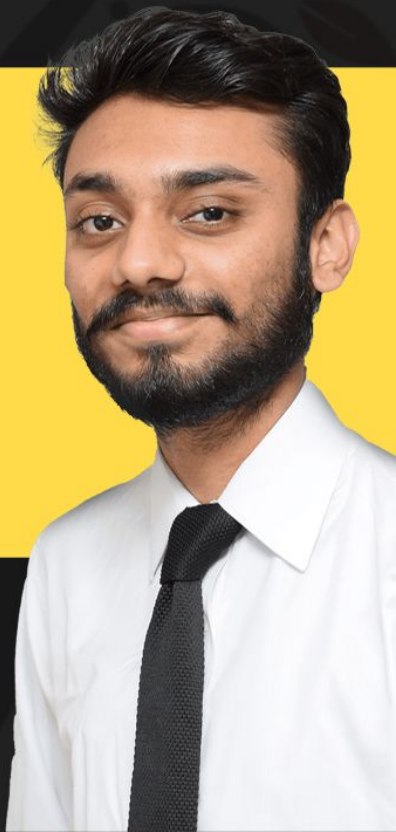
AS CHEMISTRY 9701

Crash Course

RUHAB IQBAL

REDOX REACTIONS

COMPLETE NOTES



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Redox Reactions:-

Oxidation:- loss of electrons or increase in oxidation number (Number of electrons an atom has gained control of or lost control of)

Reduction:- gain of electrons or decrease in oxidation number

Important Keypoints:-

1) Oxidation number of an element is zero(0) e.g. Na, Cl_2, O_2, H_2
2) Oxidation number of an ion is equal to charge on it e.g +1 for Na^+, -1 for Cl^-, -2 for O^{2-}
3) Group 1 elements are always +1
4) Group 2 elements are always +2
5) Group 3 elements are always +3
6) Fluorine is always -1.
7) Oxygen is always -2 unless it is combined with fluorine or is in the form of a peroxide or superoxide
8) Hydrogen is nearly always +1

9) Oxidation numbers of all elements in a neutral compound must add upto zero.

10) Oxidation numbers of all elements in a complex ion must add upto the charge on the ion.

11) Important ions:- NO_3^- (nitrate), CO_3^{2-} (carbonate), SO_4^{2-} (sulphate), OH^- (hydroxide), NH_4^+ (ammonium), HCO_3^- (hydrogen carbonate), PO_4^{3-} (phosphate)

Q. Write the formulae for the following compounds:-

1) Titanium (IV) chloride

Ti^{4+} Cl^-

$TiCl_4$

2) Titanium (IV) Oxide

Ti^{4+} O^{2-}

TiO_2

3) Vanadium (III) Bromide

V^{3+} Br^-

VBr_3

4) Vanadium (V) Oxide

V^{5+} O^{2-}

V_2O_5

5) chromium (III) oxide

Cr^{3+} O^{2-}

Cr_2O_3

6) Chromium (VI) oxide

Cr^{6+} O^{2-}

CrO_3

7) Iron (II) Sulphate

Fe^{2+} SO_4^{2-}

$FeSO_4$

8) Copper (I) oxide

Cu^{1+} O^{2-}

Cu_2O

9) Copper (II) oxide

Cu^{2+} O^{2-}

CuO

10) Iron (III) Sulphate

Fe^{3+} SO_4^{2-}

$Fe_2(SO_4)_3$

11) Sodium hydrogencarbonate

Na^+ HCO_3^-

$NaHCO_3$

12) Magnesium nitrate

Mg^{2+} NO_3^-

$Mg(NO_3)_2$

Q. Calculate oxidation numbers of underlined atoms:-

1) $\underline{H_2}O$

$2x - 2 = 0$

$x = +1$

2) $\underline{Cu_2}O$

$2x - 2 = 0$

$x = +1$

3) $H\underline{N}O_3$

$+1 + x + 3(-2) = 0$

$x = +5$

4) $H_2\underline{S}O_4$

$+2 + x + 4(-2) = 0$

$x = +6$

5) $\underline{Cl_2}O$

$2x - 2 = 0$

$x = +1$

6) $\underline{Cl_2}O_7$

$+2x + 7(-2) = 0$

$x = +7$

7) $Na\underline{Cl}O$

$+1 + x - 2 = 0$

$x = +1$

8) $K\underline{Mn}O_4$

$+1 + x + 4(-2) = 0$

$x = +7$

9) $K_2\underline{Cr_2}O_7$

$2(+1) + 2x + 7(-2) = 0$

$x = +6$

10) $Na_2\underline{S_2}O_3$

$2(+1) + 2x + 3(-2) = 0$

$x = +2$

11) $K\underline{Cl}O_3$

$+1 + x + 3(-2) = 0$

$x = +5$

12) $\underline{Cl_2}$

$x = 0$

13) $\underline{Mn}O_4^{2-}$

$x + 4(-2) = -2$

$x = +6$

14) $\underline{Cr_2}O_7^{2-}$

$2x + 7(-2) = -2$

$x = +6$

15) $\underline{S}O_3^{2-}$

$x + 3(-2) = -2$

$x = +4$

Reducing agent:- A species which reduces other species in the reaction and gets oxidised its ownself

Oxidising agent:- A species which oxidises other species in the reaction and gets reduced its ownself

Another way to identify oxidation ⟶ Addition of oxygen / removal of hydrogen

Another way to identify reduction ⟶ Addition of hydrogen / removal of oxygen

Q. In the following reactions, identify which species was reduced and which species was oxidised as well as the oxidising and reducing agent.

1) $\overset{+4}{M}n\overset{-2}{O}_2 + 4\overset{+1}{H}\overset{-1}{Cl} \longrightarrow \overset{+2}{M}n\overset{-1}{Cl}_2 + \overset{+1}{H}_2\overset{-2}{O} + \overset{0}{Cl}_2$

Oxidised: Cl as its oxidation number increased from -1 to 0.
Reduced: Mn as its oxidation number decreased from +4 to +2.
Oxidising agent: Mn
Reducing agent: Cl

2) $\overset{0}{Cl}_2 + 2Br^- \longrightarrow 2Cl^- + \overset{0}{Br}_2$

Oxidised: Br as oxidation no. $-1 \rightarrow 0$
Reduced: Cl as oxidation no. $0 \rightarrow -1$
Oxidising agent: Cl
Reducing agent: Br

3) $\overset{+1}{H}_2\overset{+6}{S}\overset{-2}{O}_4 + \overset{+1}{N}a\overset{-1}{O}H \longrightarrow \overset{+1}{N}a_2\overset{+6}{S}\overset{-2}{O}_4 + \overset{+1}{H}_2\overset{-2}{O}$

Oxidised: NOT A REDOX REACTION
Reduced:
Oxidising agent:
Reducing agent:

4) $\overset{0}{Br}_2(aq) + \overset{+1}{H}\overset{+2}{C}\overset{-2}{O}_2\overset{+1}{H} \longrightarrow 2\overset{+1}{H}\overset{-1}{Br} + \overset{+4}{C}\overset{-2}{O}_2$

Oxidised: C as $+2 \rightarrow +4$
Reduced: Br as $0 \rightarrow -1$
Oxidising agent: Br
Reducing agent: C

$(x+1+1+2(-2)=0 \quad x=+2)$

Combining half equations:-

1) Oxidation of bromide ions by chlorine molecules

Half equation for oxidation for bromide ions:- $2Br^- \longrightarrow Br_2 + 2e^-$

Half equation for reduction of chlorine molecules:- $Cl_2 + 2e^- \longrightarrow 2Cl^-$

Combined: $2Br^- + Cl_2 + \cancel{2e^-} \longrightarrow Br_2 + \cancel{2e^-} + 2Cl^-$

$2Br^- + Cl_2 \longrightarrow Br_2 + 2Cl^-$

2) Oxidation of sodium atoms by oxygen molecules:-

Half equation for oxidation for sodium atoms :- $(Na \longrightarrow Na^+ + e^-) \times 4$

Half equation for reduction of oxygen molecules:- $O_2 + 4e^- \longrightarrow 2O^{2-}$

Combined: $4Na + O_2 + \cancel{4e^-} \longrightarrow 4Na^+ + 2O^{2-} + \cancel{4e^-}$

$4Na + O_2 \longrightarrow 4Na^{2+} + 2O^{2-}$

3) Oxidation of chloride ions to chlorine by acidified manganate(VII) ions

Half equation for oxidation of chloride ions :- $(2Cl^- \longrightarrow Cl_2 + 2e^-) \times 5$

Half equation for reduction of manganate(VII) ions:- $(MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O) \times 2$

Mn in MnO_4^- => $4(-2) + x = -1$

$x = +7$

Mn goes from +7 to +2 -> we need $5e^-$

Combined: $10Cl^- + 2MnO_4^- + 16H^+ + \cancel{10e^-} \longrightarrow 5Cl_2 + \cancel{10e^-} + Mn^{2+} + 4H_2O$

$10Cl^- + 2MnO_4^- + 16H^+ \longrightarrow 5Cl_2 + Mn^{2+} + 4H_2O$

4) The disproportionation of Cu(I) ions in solution:-

Half equation for oxidation of copper ions :- $Cu^+ \longrightarrow Cu^{2+} + e^-$

Half equation for reduction of copper ions:- $Cu^+ + e^- \longrightarrow Cu$

Combined: $2Cu^+ \longrightarrow Cu^{2+} + Cu$

Disproportionation:- Simultaneous oxidation and reduction of an element.

5) Disproportionation of Chlorate (I) to chloride and chlorate (V)

Half equation for the oxidation of chlorate (I) to chlorate (V): $ClO^- + 2H_2O \longrightarrow ClO_3^- + 4H^+ + 4e^-$

Cl in ClO^- = $x - 2 = -1$, $x = +1$

Cl in ClO_3^- = $x + 3(-2) = -1$, $x = +5$

Cl goes from +1 to +5, we lose $4e^-$

Half equation for the reduction of chlorate (I) to chloride: $(ClO^- + 2H^+ + 2e^- \longrightarrow Cl^- + H_2O) \times 2$

Cl in ClO^- = $x - 2 = -1$, $x = +1$

Cl in Cl^- = $x = -1$

Cl goes from +1 to −1, we need $2e^-$

Combined: $3ClO^- + 4H^+ + 4e^- + 2H_2O \longrightarrow ClO_3^- + 4H^+ + 4e^- + 2Cl^- + 2H_2O$

$3ClO^- \longrightarrow ClO_3^- + 2Cl^-$

(d) Mohr's salt, $(NH_4)_2Fe(SO_4)_2 \cdot xH_2O$, is the hydrated form of ammonium iron(II) sulfate.

x represents the number of moles of water in 1 mole of the salt.

A student wanted to determine the value of x. 0.784 g of the hydrated salt was dissolved in water and this solution was acidified.

All of the solution was titrated with 0.0200 mol dm^{-3} potassium manganate(VII). 20.0 cm^3 of this potassium manganate(VII) solution was required for complete reaction with the Fe^{2+} ions.

(i) Use changes in oxidation numbers to balance the equation for the reaction taking place.

$\overset{+7\ -2}{MnO_4^-}(aq) + \overset{+2}{5}Fe^{2+}(aq) + 8H^+(aq) \rightarrow \overset{+2}{1}Mn^{2+}(aq) + \overset{+3}{5}Fe^{3+}(aq) + 4H_2O(l)$

$x + 4(-2) = -1$ Mn goes from +7 → +2 → we need 5 e^- [1]

(ii) State the role of the Fe^{2+} ions in this reaction.

Explain your answer.

They act as reducing agent as they reduce Mn from +7 to +2 [2]

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

$(Fe^{2+} \longrightarrow Fe^{3+} + e^-) \times 5$

$MnO_4^- + 5e^- + 8H^+ + 5Fe^{2+} \longrightarrow Mn^{2+} + 4H_2O + 5Fe^{3+} + 5e^-$

2 **(a)** The equation shown in **(a)(i)** describes the reaction which occurs when aqueous potassium iodide is added to aqueous copper(II) sulfate. A white precipitate of copper(I) iodide forms in a brown solution of iodine and potassium sulfate.

(i) Balance the equation and include state symbols.

$1\ CuSO_4(aq) + 2\ KI(aq) \rightarrow 1\ CuI(s) + \frac{1}{2}\ I_2(aq) + 1\ K_2SO_4(aq)$ [2]

The table gives the oxidation numbers of iodine in the different species in the equation.

iodine-containing species	oxidation number of iodine
KI	–1
CuI	–1
I_2	0

(ii) Deduce the oxidation number of copper in $CuSO_4$ and CuI.

- oxidation number of copper in $CuSO_4$ +2
- oxidation number of copper in CuI +1

[1]

(iii) Describe the type of reaction shown by the equation in **(a)(i)**. Explain your answer in terms of electron transfer.

Redox reaction. Cu gained an electron and went from +2 to +1. Iodine lost an electron and went from −1 to 0. [2]

(c) **(i)** In the boxes, give the oxidation numbers of nitrogen in the nitrogen-containing species for the reaction in stage 3.

$3\boxed{+4}NO_2 + H_2O \rightarrow 2\boxed{+5}HNO_3 + \boxed{+2}NO$

$x + 2(-2) = 0$, $x = +4$ $\quad +1 + x + 3(-2) = 0$, $x = +5$ $\quad x - 2 = 0$, $x = +2$ [2]

(ii) Explain why the reaction in stage 3 is described as a disproportionation reaction. Include reference to transfer of electrons in your answer.

Because simultaneous reduction and oxidation has occured. Nitrogen in NO_2 was oxidised from +4 to +5 in HNO_3 by losing an electron and also reduced from +4 in NO_2 to +2 in NO by gaining 2 electrons [2]

(b) $Cl_2(g)$ dissolves in cold water and reacts with it.

$$\overset{0}{Cl_2} + H_2O \rightarrow \overset{+1\ -1}{HCl} + \overset{+1\ +1\ -2}{HClO}$$

(i) Identify the oxidation number of chlorine in each of the chlorine-containing species in this reaction.

chlorine-containing species	Cl_2	HCl	$HClO$
oxidation number of chlorine	0	−1	+1

[2]

(ii) Name the type of reaction occurring.

disproportionation [1]

(iii) Explain why chlorine is used in the purification of water.

It lowers the pH of the water and kills harmful micro organisms as it forms HCl and HClO [1]

(c) A mixture of HCl and $HClO$ is added to cold dilute NaOH. One of the products behaves as a bleach.

Suggest the equation for the reaction occurring.

will learn later.

[2]

[Total: 13]

1 When ammonia is converted into nitric acid on a commercial scale, the following reactions can occur.

In which reaction does the greatest change in oxidation number of the nitrogen occur?

reaction	
A	$\overset{-3\ +1}{4NH_3} + 5\overset{0}{O_2} \rightarrow 4\overset{+2\ -2}{NO} + 6H_2O$ → 5
B	$\overset{+4\ -2}{3NO_2} + H_2O \rightarrow 2\overset{+1\ +5\ -2}{HNO_3} + \overset{+2\ -2}{NO}$ → 1, 2
C	$\overset{+2\ -2}{2NO} + O_2 \rightarrow 2\overset{+4\ -2}{NO_2}$ → 2
D	$\overset{-3\ +1}{4NH_3} + 6\overset{+2\ -2}{NO} \rightarrow 5\overset{0}{N_2} + 6H_2O$ → 3, 2

2 Granular urea, CON_2H_4, can be used to remove NO_2 from the flue gases of power stations, converting it into harmless nitrogen.

$$2CON_2H_4 + x\overset{3}{NO_2} \rightarrow 2\overset{4}{CO_2} + y\overset{3.5}{H_2O} + zN_2$$

What are the values of x, y and z in a balanced equation?

	x	y	z
A	1½	2	1¼
B	2	4	3
C	3	4	3½
D	3	4	3

5 The first stage in the manufacture of nitric acid is the oxidation of ammonia by oxygen.

$$w NH_3(g) + \tfrac{5w}{4} O_2(g) \rightarrow w NO(g) + \tfrac{3w}{2} H_2O(g)$$

Which values for **w**, **x**, **y** and **z** are needed to balance the equation?

	w	**x**	**y**	**z**
A	4	5	4	6
B	4	6 ✗	4	5
C	5	6	5	4 ✗
D	6	5	6	4 ✗

$3w = 2z$

$z = \frac{3w}{2}$

$2x = w + \frac{3w}{2} = \frac{5}{4}w$

10 In aqueous solution, the acid HIO disproportionates according to the following equation where **m**, **n**, **p** and **q** are simple whole numbers in their lowest ratios.

$$\mathbf{m}\overset{+1\ +1\ -2}{HIO} \rightarrow \mathbf{n}\overset{0}{I_2} + \mathbf{p}\overset{+1\ +5\ -2}{HIO_3} + \mathbf{q}\overset{+1\ -2}{H_2O}$$

This equation can be balanced using oxidation numbers.

What are the values for **n** and **p**?

	n	**p**
A	1	2
B	2	1
C	4	1
D	4	2

$HIO + O_2 \longrightarrow HIO_3 + 4e^-$

$2(2HIO + 2e^- \longrightarrow I_2 + H_2O + \frac{1}{2}O_2)$

$HIO + 4HIO + O_2 + 4e^- \longrightarrow HIO_3 + 4e^- + 2I_2 + 2H_2O + O_2$

$5HIO \longrightarrow 2I_2 + HIO_3 + 2H_2O$

22 The following half reactions occur when potassium iodate(V), KIO_3, in hydrochloric acid solution oxidises iodine to ICl_2^-.

$$IO_3^- + 2Cl^- + 6H^+ + 4e^- \rightarrow ICl_2^- + 3H_2O$$

$$(I_2 + 4Cl^- \rightarrow 2ICl_2^- + 2e^-)2$$

What is the ratio of IO_3^- to I_2 in the balanced chemical equation for the overall reaction?

A 1:1 **B** 1:2 **C** 1:4 **D** 2:1

36 When copper reacts with a 50% solution of nitric acid, nitrogen monoxide is evolved and a blue solution results.

The balanced equation for this reaction is shown.

$$p\text{Cu} + \tfrac{8}{3}q\text{HNO}_3 \rightarrow p\text{Cu(NO}_3)_2 + \tfrac{4}{3}r\text{H}_2\text{O} + \tfrac{r}{1.5}\text{NO}$$

What are the values of the integers *p*, *q*, *r*, *s* and *t*?

	p	*q*	*r*	*s*	*t*
A	1	4	1	2	2
B	2	6	2	3	2
C	2	8	2	4	4
D	3	8	3	4	2

37 When K_2MnO_4 is dissolved in water, the following reaction occurs.

$$a\text{MnO}_4^{2-}(aq) + b\text{H}_2\text{O}(l) \rightarrow c\text{MnO}_4^-(aq) + d\text{MnO}_2(s) + e\text{OH}^-(aq)$$

What are the values of *a* and *c* in the balanced chemical equation?

	a	*c*
A	2	1
B	3	2
C	4	3
D	5	4

31 Arsenic chloride, $AsCl_3$, reacts with sodium borohydride, $NaBH_4$.

$$p\text{AsCl}_3 + q\text{NaBH}_4 \rightarrow r\text{AsH}_3 + s\text{NaCl} + t\text{BCl}_3$$

What are the numbers **p**, **q**, **r**, **s** and **t** when this equation is balanced correctly?

	p	**q**	**r**	**s**	**t**
A	2	3	2	3	1
B	3	3	3	3	2
C	4	3	4	3	3
D	4	4	4	4	3