Redox Reactions

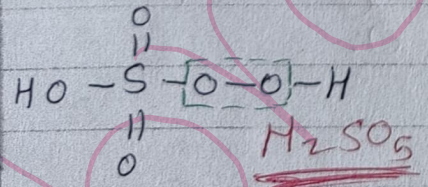
1. Oxidation = Loss of e^-s = ↑ in valency = ↑ in OS
   Reduction = Gain of e^-s = ↓ in valency = ↓ in OS.
2. OA = oxidizes the substrate ⇒ self-reduction.
   RA = reduces the substrate ⇒ self-oxidation.
   OA = Lewis acid
   RA = Lewis base.

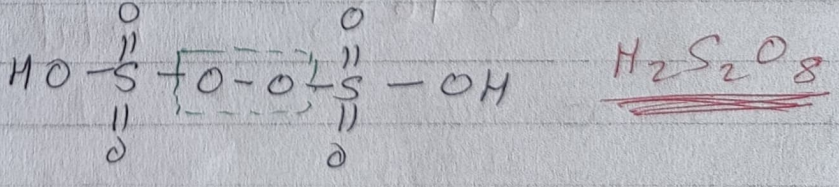
3. Amphoteric oxides = Be, Zn, Al, Sn, Pb, Mn, Ga
   ↳ oxidants and reductants.

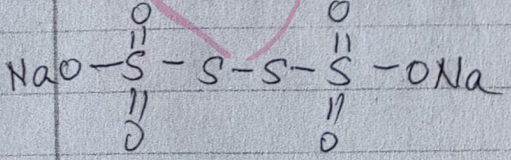
4. OS
   O Group ⇒ OS = O (except Xe)
   OS = +8 (Ru, Os, Xe)
   RuO_4, OsO_4, XeO_4.

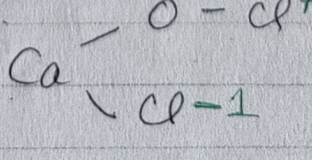
5. In peroxide bond, O = -1
   In super oxides, O = -1/2
   Halogens show +1, +3, +5, +7 also
   except F shows -1 only.
   In elemental form, OS = 0.

6. Ligands -
   → CN^-, OH^-, X^-, NO_2^-, NH_2^- = -1.
   → NH_3, CO, H_2O, C_2H_4 = 0.
   → NO^+ = +1.

7. 
   H_2SO_5

   
   $H_2S_2O_8$

   
   $Na_2S_4O_6$
   Sodium tetrathionate

   
   $CaOCl_2$

8. ↓ OS ↑es ⇒ stability ↑es. (in compounds)

9. Inert Pair Effect (IPE) - For group 13, 14, 15
Lower OS stability ↑es down the group
∴ for last element - Lower OS > Higher OS.
(stability)

$Tl^{+1} > Tl^{+3}$
$Pb^{+2} > Pb^{+4}$ $Bi^{+3} > Bi^{+5}$.

10. In metal amalgums, the OS of metal = 0.
(metal + Hg).

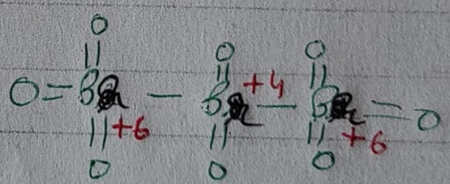
11. Fe_3O_4, Co_3O_4, Mn_3O_4 } Mixed oxides (contain two types of OS).

12. N_3^-
OS = -1/3
azide ion

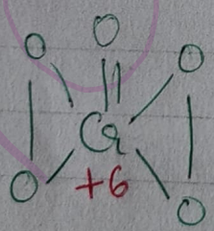
N^{3-}
OS = -3
nitride ion.

13. C_3O_2 (carbon suboxide) $O=\overset{x}{C}=\overset{x}{C}=\overset{x}{C}=O$ (-2) (+2) (0) (+2) (-2) $x = 4/3$

Br_3O_8 (tribromine octoxide)



CrO_5



In metal carbonyl complexes, metal OS = 0.

14. OS ∝ acidic strength ∝ $\frac{1}{\text{oxidizing ability}}$

except $H_3PO_2 > H_3PO_3 > H_3PO_4$ (acidic strength order)
(+1) (+3) (+5)

15. In disproportionation / auto redox change, one element undergoes both oxidation and reduction.

16. In acidic medium

$Cr_2O_7^- \longrightarrow Cr^{+3}$

$S_2O_3^- \longrightarrow S_4O_6^-$

$MnO_4^- \longrightarrow Mn^{+2}$

In basic medium

$MnO_4^- \longrightarrow MnO_4^{2-}$

In neutral medium

$MnO_4^- \longrightarrow MnO_2$.

17. Cl_2 + dil. NaOH $\longrightarrow$ NaCl + NaOCl + H_2O

Cl_2 + conc. NaOH $\longrightarrow$ NaCl + $NaClO_3$ + H_2O.

generally, $C_2O_4^{2-} \longrightarrow CO_2 + H_2O$.
(oxalate ion)

$Fe^{2+} \longrightarrow Fe^{3+}$

$SO_3^{2-} \longrightarrow SO_4^{2-}$

18. ## Balancing Reactions

I Ion - Electron Method -

① In oxidation and reduction half, balance the central atoms.

② Balance the Oxygen atoms by using water molecules which are added to the deficient side.

③ For balancing hydrogen -

In acidic medium	In basic medium
add H^+ ions to the deficient side.	add water molecules to the deficient side and equal no. of OH^- ions to opposite side.

④ Balance e^-s.

II Oxidation Number Method -

Eg - $KMnO_4 + FeSO_4 + H_2SO_4 \longrightarrow MnSO_4 + Fe_2(SO_4)_3 + K_2SO_4 + H_2O$

$MnO_4^- + Fe^{+2} + H^+ \longrightarrow Mn^{+2} + Fe^{+3} + H_2O$

(+7) (+2) (+2) (+3)

1 e^- (oxidation)

5 e^- (reduction)

Multiply the opposite numbers (1 and 5)

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

balance the H and O's.

∴ $4H_2O$

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

∴ 8 ∴ 4

19. In haemoglobin → Fe = +2
In insulin → Zn = +2
In chlorophyll → Mg = +2
In vit. B_{12} → Co = +3.