TURBO

Created by students, for students

Redox Reactions

Class 11 Chemistry • Complete Formula Sheet

Sr.	Concept	Formulas	Other Information
		Basic Definitions	
1	Oxidation	Oxidation = Loss of electrons Example: $Fe^{2+} \rightarrow Fe^{3+} + e^-$	**Memory trick:** OIL RIG Oxidation Is Loss Reduction Is Gain
2	Reduction	Reduction = Gain of electrons Example: $Cu^{2+} + 2e^- \rightarrow Cu$	**Key point:** Always occurs simultaneously with oxidation
3	Oxidizing Agent	Substance that gets reduced Accepts electrons Gets reduced itself	**Examples:** $KMnO_4$, $K_2Cr_2O_7$, O_2, Cl_2
4	Reducing Agent	Substance that gets oxidized Donates electrons Gets oxidized itself	**Examples:** H_2, metals (Na, Zn), H_2S, $SnCl_2$
		Oxidation Number (O.N.) Rules	
5	O.N. in Free State	O.N. of element in free state = 0	**Examples:** O.N. of O in O_2 = 0 O.N. of Na in Na = 0 O.N. of Cl in Cl_2 = 0
6	O.N. of Monoatomic Ions	O.N. = Charge on ion	**Examples:** Na^+: O.N. = +1 Ca^{2+}: O.N. = +2 Cl^-: O.N. = −1 O^{2-}: O.N. = −2
7	O.N. of Oxygen	**General:** O.N. = −2 **Peroxides** (H_2O_2, Na_2O_2): O.N. = −1 **Superoxides** (KO_2, RbO_2): O.N. = $-\frac{1}{2}$ **OF_2** (Oxygen difluoride): O.N. = +2 **O_2F_2** (Dioxygen difluoride): O.N. = +1	**Exception order:** OF_2 ¿ Superoxides ¿ Peroxides ¿ Normal oxides **Remember:** F is more electronegative than O
8	O.N. of Hydrogen	**General:** O.N. = +1 **Metal hydrides** (NaH, CaH_2): O.N. = −1	**Examples:** H in H_2O: +1 H in HCl: +1 H in NaH: −1
9	O.N. of Fluorine	O.N. of F = −1 (always) Most electronegative element	**No exception:** F always has −1 oxidation state
10	O.N. of Group 1 Metals	O.N. = +1 (always) Li, Na, K, Rb, Cs	**In all compounds:** NaCl, Na_2O, NaOH all have Na = +1
11	O.N. of Group 2 Metals	O.N. = +2 (always) Be, Mg, Ca, Sr, Ba	**In all compounds:** $CaCl_2$, MgO, $BaSO_4$

Sr.	Concept	Formulas	Other Information
12	O.N. in Neutral Molecule	Sum of O.N. of all atoms = 0	**Example:** H_2SO_4: $2(+1) + x + 4(-2) = 0$ $x = +6$ (O.N. of S)
13	O.N. in Polyatomic Ion	Sum of O.N. = Charge on ion	**Example:** SO_4^{2-}: $x + 4(-2) = -2$ $x = +6$ (O.N. of S)
		Equivalent Weight Concepts	
14	Equivalent Weight of Salt	Eq. Wt. of salt = $\frac{\text{Molar mass}}{\text{Net positive (or negative) valency}}$	**Example:** $AlCl_3$: Molar mass = 133.5 Valency = 3 Eq. Wt. = 133.5/3 = 44.5 g/eq
15	Equivalent Weight of Acid	Eq. Wt. of acid = $\frac{\text{Molar mass}}{\text{Basicity}}$ Basicity = Number of replaceable H^+ ions	**Examples:** HCl: Basicity = 1, Eq. Wt. = 36.5 H_2SO_4: Basicity = 2, Eq. Wt. = 98/2 = 49 H_3PO_4: Basicity = 3, Eq. Wt. = 98/3
16	Equivalent Weight of Base	Eq. Wt. of base = $\frac{\text{Molar mass}}{\text{Acidity}}$ Acidity = Number of replaceable OH^- ions	**Examples:** NaOH: Acidity = 1, Eq. Wt. = 40 $Ca(OH)_2$: Acidity = 2, Eq. Wt. = 74/2 = 37 $Al(OH)_3$: Acidity = 3, Eq. Wt. = 78/3 = 26
17	Number of Gram Equivalents	No. of gram equivalents (Eq) = $\frac{\text{Weight of compound}}{\text{Equivalent weight}}$	**How to use:** If 49 g H_2SO_4 (Eq. Wt. = 49) Equivalents = 49/49 = 1 eq
		Mole-Equivalent Relationships	
18	Mole and Equivalent Relationship	$n = \frac{w}{M_W}$ (moles) $\text{Eq} = \frac{w}{\text{Eq. Wt.}}$ (equivalents) $\frac{\text{Eq}}{n} = \frac{M_W}{\text{Eq. Wt.}} = \text{n-factor}$	**Key relation:** Equivalents = Moles × n-factor **Example:** 1 mol H_2SO_4 = 2 eq n-factor = 2
19	n-factor for Acids	n-factor = Basicity = Number of H^+ furnished per molecule	**Examples:** HCl: n-factor = 1 H_2SO_4: n-factor = 2 H_3PO_4: n-factor = 3
20	n-factor for Bases	n-factor = Acidity = Number of OH^- furnished per molecule	**Examples:** NaOH: n-factor = 1 $Ca(OH)_2$: n-factor = 2
		Normality and Molarity	
21	Normality Definition	$N = \frac{\text{Eq}}{V}$ N = Normality (eq/L) V = Volume in liters	**Unit:** Normality is in equivalents per liter (N or eq/L)
22	Molarity Definition	$M = \frac{n}{V}$ M = Molarity (mol/L) V = Volume in liters	**Unit:** Molarity is in moles per liter (M or mol/L)
23	Normality-Molarity Relation	$N = M \times \text{n-factor}$ $\frac{N}{M} = \text{n-factor}$	**Example:** 1 M H_2SO_4 = 2 N H_2SO_4 (n-factor = 2)
		Equivalent Weight in Redox	

Sr.	Concept	Formulas	Other Information
24	Equivalent Weight of Oxidizing Agent (O.A.)	Eq. Wt. of O.A. $= \frac{\text{Molar mass}}{\text{Change in O.N. per mole}}$	**How to use:** Find total change in O.N. for entire molecule **Example:** $KMnO_4 \rightarrow Mn^{2+}$ Mn: $+7 \rightarrow +2$ Change = 5 Eq. Wt. = 158/5 = 31.6
25	Equivalent Weight of Reducing Agent (R.A.)	Eq. Wt. of R.A. $= \frac{\text{Molar mass}}{\text{Change in O.N. per mole}}$	**Example:** $FeSO_4 \rightarrow Fe_2(SO_4)_3$ Fe: $+2 \rightarrow +3$ Change = 1 Eq. Wt. = 152/1 = 152
26	n-factor for Oxidizing Agent	n-factor = Total electrons gained per molecule = Change in O.N. per molecule	**Examples:** $KMnO_4$ (acidic): n = 5 $KMnO_4$ (basic): n = 1 $KMnO_4$ (neutral): n = 3
27	n-factor for Reducing Agent	n-factor = Total electrons lost per molecule = Change in O.N. per molecule	**Example:** $H_2C_2O_4$ (oxalic acid): 2C: $+3 \rightarrow +4$ each Total change = 2 n-factor = 2
		Titration Equations	
28	Acid-Base Titration	Gram equivalents of acid = Gram equivalents of base $N_1V_1 = N_2V_2$ where subscript 1 = acid, 2 = base	**How to use:** Use when neutralization occurs **Example:** 25 mL 0.1 N HCl neutralizes how much 0.2 N NaOH? $0.1 \times 25 = 0.2 \times V_2$ $V_2 = 12.5$ mL
29	Redox Titration	Gram equivalents of O.A. = Gram equivalents of R.A. $N_1V_1 = N_2V_2$ where 1 = O.A., 2 = R.A.	**How to use:** In redox reactions **Example:** $KMnO_4$ (0.02 N, 20 mL) oxidizes $FeSO_4$ (V mL, 0.1 N) $0.02 \times 20 = 0.1 \times V$ $V = 4$ mL
		Balancing Redox Reactions	
30	Oxidation Number Method - Steps	**Step 1:** Assign O.N. to all atoms **Step 2:** Identify oxidized and reduced species **Step 3:** Calculate change in O.N. **Step 4:** Equalize total increase and decrease **Step 5:** Balance O and H (add H_2O and H^+/OH^-) **Step 6:** Check atoms and charges	**Example:** $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$ Mn: $+7 \rightarrow +2$ (gain $5e^-$) Fe: $+2 \rightarrow +3$ (lose $1e^-$) Multiply Fe by 5 $MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Sr.	Concept	Formulas	Other Information
31	Half-Reaction Method - Steps	**Step 1:** Write oxidation and reduction half-reactions **Step 2:** Balance atoms except O and H **Step 3:** Balance O by adding H_2O **Step 4:** Balance H by adding H^+ (acidic) or OH^- (basic) **Step 5:** Balance charge by adding e^- **Step 6:** Equalize electrons in both half-reactions **Step 7:** Add half-reactions and simplify	**Example (acidic):** $MnO_4^- \rightarrow Mn^{2+}$ (reduction) Balance O: $MnO_4^- \rightarrow Mn^{2+} + 4H_2O$ Balance H: $MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$ Balance charge: $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
Common Redox Reactions			
32	$KMnO_4$ in Acidic Medium	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ n-factor = 5 Color: Purple → Colorless/Pink	**Use:** Strong oxidizing agent Common in titrations
33	$KMnO_4$ in Basic Medium	$MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$ n-factor = 3 Color: Purple → Brown (MnO_2 ppt)	**Note:** Different n-factor than acidic
34	$KMnO_4$ in Neutral Medium	$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$ n-factor = 3	**Use:** Less common than acidic
35	$K_2Cr_2O_7$ in Acidic Medium	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$ n-factor = 6 Color: Orange → Green	**Note:** Each Cr goes from +6 to +3 2 Cr atoms, so $2 \times 3 = 6$
36	Common Reducing Agents	**Oxalic acid:** $H_2C_2O_4$, n-factor = 2 **Ferrous salts:** $FeSO_4$, n-factor = 1 **Stannous chloride:** $SnCl_2$, n-factor = 2 **Hydrogen peroxide:** H_2O_2, n-factor = 2 (as R.A.)	**H_2O_2 is amphoteric:** Can act as both O.A. and R.A.
37	H_2O_2 as Oxidizing Agent	$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$ O goes from −1 to −2 n-factor = 2	**Example:** Oxidizes KI to I_2
38	H_2O_2 as Reducing Agent	$H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$ O goes from −1 to 0 n-factor = 2	**Example:** Reduces $KMnO_4$
Disproportionation Reactions			
39	Disproportionation	Same element simultaneously oxidized and reduced Element in intermediate O.N.	**Examples:** $Cl_2 + 2OH^- \rightarrow Cl^- + ClO^- + H_2O$ Cl: 0 → −1 and 0 → +1 $3Cl_2 + 6OH^- \rightarrow 5Cl^- + ClO_3^- + 3H_2O$ (hot and conc.)
40	Comproportionation	Opposite of disproportionation Two different O.N. of same element give intermediate O.N.	**Example:** $H_2S + SO_2 \rightarrow S + H_2O$ S: −2 and +4 both → 0