TURBO

Created by students, for students

Some Basic Concepts of Chemistry

Class 11 Chemistry • Complete Formula Sheet

Sr.	Concept	Formulas	How to Use / Other Information
Fundamental Constants & Definitions			
1	Avogadro's Number	$N_A = 6.022 \times 10^{23}$	**Use:** This is the number of particles (atoms/molecules) in 1 mole of any substance. **Example:** 1 mole of H_2O contains 6.022×10^{23} molecules
2	Mole Concept	**Mole** = Collection of Avogadro's number of particles	**Remember:** 1 mole $= 6.022 \times 10^{23}$ particles Whether it's atoms, molecules, ions, or electrons, 1 mole always contains this number
3	Atomic Weight	Atomic weight $= \frac{\text{mass of one atom}}{\frac{1}{12}\times\text{mass of one } ^{12}\text{C atom}}$	**How to use:** Atomic weight is dimensionless (no units) **Example:** Atomic weight of O = 16 means oxygen atom is 16 times heavier than $\frac{1}{12}$ of C-12 atom
4	Molecular Weight	Molecular weight = Sum of atomic masses of all atoms **Example:** $H_2SO_4 = (1 \times 2) + (32 \times 1) + (16 \times 4) = 98$	**Step-by-step:** 1. Identify all atoms in formula 2. Multiply atomic mass by number of atoms 3. Add all values **Common molecular weights:** $H_2O = 18$, $NaCl = 58.5$, $CO_2 = 44$
Mole Calculations			
5	Moles of Atoms	Moles $= \frac{\text{Weight (gm)}}{\text{Gram atomic weight}}$ $= \frac{\text{No. of atoms}}{N_A}$	**Example:** For 1.6 gm of oxygen: Moles $= \frac{1.6}{16} = 0.10$ mole No. of O atoms $= 0.10 \times 6.022 \times 10^{23}$ $= 6.022 \times 10^{22}$ atoms
6	Moles of Molecules	Moles $= \frac{\text{Weight (gm)}}{\text{Gram molecular wt}}$ $= \frac{\text{No. of molecules}}{N_A}$	**Example:** For 3.011×10^{24} molecules of H_3PO_4: Moles $= \frac{3.011\times10^{24}}{6.022\times10^{23}} = 5$ moles Weight $= 5 \times 98 = 490$ gm
7	Avogadro's Hypothesis	At constant T and P: Volume $\propto$ Number of moles Moles $= \frac{\text{Volume at STP (L)}}{22.4\ \text{L}}$	**Key point:** At STP (0°C, 1 atm), 1 mole of any gas occupies 22.4 L **Example:** 2.24 L of NH_3 at STP $= \frac{2.24}{22.4} = 0.10$ mole
8	Loschmidt Number	2.687×10^{19} molecules/mL at STP	**Definition:** Number of molecules in 1 mL of gas at STP
Equivalent Weight & n-factor			
9	Equivalent Weight	Eq. Wt. $= \frac{\text{Molecular weight}}{n\text{-factor}}$	**Important:** n-factor depends on type of substance (acid, base, salt, oxidizing/reducing agent)

Sr.	Concept	Formulas	How to Use / Other Information
10	n-factor for Acids	n-factor = Number of H^+ ions replaced = Basicity of acid	**Examples:** $H_2SO_4 \rightarrow 2NaOH$: n-factor = 2 $H_3PO_4 \rightarrow NaOH$: n-factor = 1 H_3BO_3: n-factor = 1 (accepts OH^-, not donates H^+)
11	n-factor for Bases	n-factor = Number of OH^- lost or H^+ gained = Acidity of base	**Examples:** $NH_3 + HCl$: n-factor = 1 $Mg(OH)_2 + H_2SO_4$: n-factor = 2
12	n-factor for Salts	n-factor = Total cationic or anionic charge	**Example:** $Al_2(SO_4)_3$ Cationic charge = $2 \times 3 = 6$ So n-factor = 6
13	n-factor for Radicals	n-factor = Charge on radical	**Examples:** SO_4^{2-}: n-factor = 2 PO_4^{3-}: n-factor = 3
14	n-factor for Oxidizing Agents	n-factor = Electrons gained per molecule	**Example:** $K_2Cr_2O_7 \rightarrow Cr^{3+}$ Cr: $+6 \rightarrow +3$ ($3e^-$ per Cr) 2 Cr atoms: $2 \times 3 = 6$ n-factor = 6
15	n-factor for Reducing Agents	n-factor = Electrons lost per molecule	**Example:** $C_2O_4^{2-} \rightarrow 2CO_2$ Carbon: $+3 \rightarrow +4$ ($1e^-$ per C) 2 C atoms: $2 \times 1 = 2$ n-factor = 2
16	Gram Equivalent	Gram equivalent = $\frac{\text{Weight (gm)}}{\text{Eq. wt}}$ = Moles × n-factor = Normality × Volume (L)	**Law of Equivalence:** Gram equivalent of all reactants = Gram equivalent of all products
Concentration Terms			
17	Molarity (M)	Molarity = $\frac{\text{Moles of solute}}{\text{Volume of solution (L)}}$ $= \frac{w \times 1000}{M^0 \times V(\text{mL})}$	**Example:** 4g NaOH in 100 mL solution M = $\frac{4 \times 1000}{40 \times 100} = 1$ M **Remember:** Molarity × Volume (L) = Moles
18	Normality (N)	Normality = $\frac{\text{g.eq. of solute}}{\text{Volume (L)}}$ $= \frac{w \times 1000}{E \times V(\text{mL})}$ N = M × n-factor	**Example:** 50 mL of N/20 H_2SO_4 meq = N × V(mL) = $\frac{1}{20} \times 50 = 2.5$ meq **Remember:** N × V(L) = g.eq.
19	Molality (m)	Molality = $\frac{\text{Moles of solute}}{\text{Mass of solvent (kg)}}$ $= \frac{w \times 1000}{M^0 \times w'}$	**Example:** 10g HCl in 250 mL solution (density = 1.2 g/mL) Weight of solution = $250 \times 1.2 = 300$ g Weight of solvent = $300 - 10 = 290$ g m = $\frac{10 \times 1000}{36.5 \times 290} = 0.945$ m
20	Weight Percentage (w/w)	% (w/w) = $\frac{\text{Wt. of solute (g)}}{\text{Wt. of solution (g)}} \times 100$	**Example:** 20g NaCl in 60g water % (w/w) = $\frac{20}{20+60} \times 100 = 25\%$
21	Volume Percentage (v/v)	% (v/v) = $\frac{\text{Vol. of solute (mL)}}{\text{Vol. of solution (mL)}} \times 100$	**Example:** 10 mL ethanol + 120 mL methanol % (v/v) = $\frac{10}{10+120} \times 100 = 7.7\%$
22	Weight/Volume Percentage	% (w/v) = $\frac{\text{Wt. of solute (g)}}{\text{Vol. of solution (mL)}} \times 100$	**Example:** 7.5g KCl in 100 mL % (w/v) = $\frac{7.5}{100} \times 100 = 7.5\%$
23	Relation between Molarity, Density & %	If % (w/w) = x, density = d g/mL Then: M = $\frac{10 \times x \times d}{M^0}$	**Example:** H_2SO_4: 80% by mass, d = 1.71 g/cc M = $\frac{10 \times 1.71 \times 80}{98} = 13.95$ M
24	Molarity of Pure Liquid	M = $\frac{1000 \times d}{M^0}$	**Example:** Pure water (d = 1 g/mL) M = $\frac{1000 \times 1}{18} = 55.55$ M

Sr.	Concept	Formulas	How to Use / Other Information
25	Strength (g/L)	Strength $= \frac{\text{Wt. of solute (g)}}{\text{Vol. of solution (L)}}$ = Molarity × Mol. Mass = Normality × Eq. wt	**Use:** Converts concentration to grams per liter
		Dilution & Formulas	
26	Dilution Formula	$M_1V_1 = M_2V_2$ = Moles (constant) or $N_1V_1 = N_2V_2$ = g.eq. (constant)	**Example:** 100 mL of 0.1 M NaOH diluted to 0.01 M $V_2 = \frac{100\times0.1}{0.01}$ = 1000 mL Water added = 1000 - 100 = 900 mL **Key:** On dilution, moles remain constant
27	Molecular vs Empirical Formula	**Molecular formula** = (Empirical formula) × n where n is whole number	**Empirical:** Simplest ratio of atoms **Molecular:** Actual number of atoms **Example:** Glucose = $C_6H_{12}O_6$ (molecular) = CH_2O (empirical)
		Laws of Chemical Combination	
28	Law of Conservation of Mass	Total mass of reactants = Total mass of products	**Example:** $2Ca + O_2 \rightarrow 2CaO$ 80g + 32g = 112g **Exception:** Nuclear reactions (Einstein's $E = mc^2$)
29	Law of Constant Composition	Same compound always has same elemental composition by mass	**Example:** All samples of CO_2 have C:O = 3:8 by mass All samples of H_2O have H:O = 1:8 by mass
30	Law of Multiple Proportions	When two elements form multiple compounds, masses of one element with fixed mass of other are in simple ratio	**Example:** CO and CO_2 CO: C:O = 12:16 CO_2: C:O = 12:32 Ratio of O masses = 16:32 = 1:2 (simple ratio)
31	Avogadro's Law (Gaseous)	At same T and P, equal volumes of gases contain equal number of molecules	**Example:** $H_2 + Cl_2 \rightarrow 2HCl$ 1 vol + 1 vol → 2 vol Volume ratio = 1:1:2
32	Vapour Density	2 × Vapour Density = Molecular mass	**Use:** To find molecular mass of gas If VD = 22, then Mol. mass = 2 × 22 = 44

TURBO *Your complete JEE/NEET formula companion*