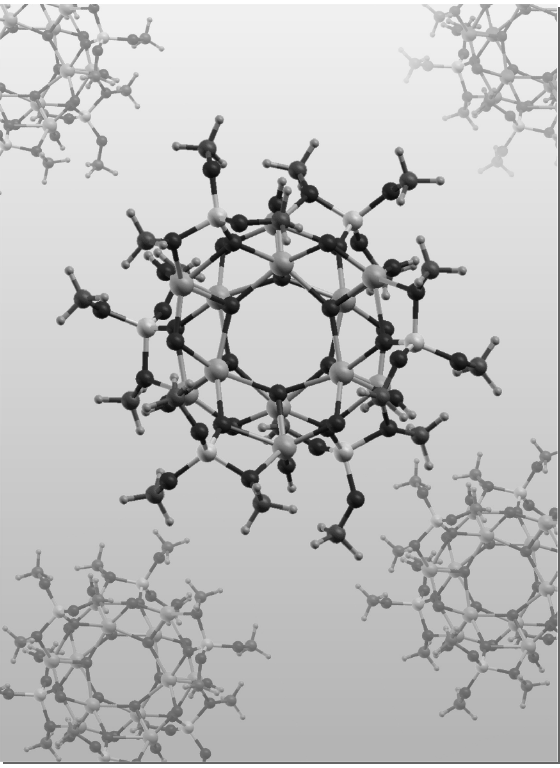


CHEMISTRY

Formula Book for Engineering Entrance Examinations

Best Wishes for

Your Success in Competitive Examinations ahead !!!

SHORT FORMULA (GYAN SUTRA)

CHEMISTRY

PHYSICAL CHEMISTRY

ATOMIC STRUCTURE

Estimation of closest distance of approach (derivation) of α-particle : $R = \frac{4KZe^2}{m_\alpha V_\alpha^2}$

The radius of a nucleus : $R = R_0 (A)^{1/3}$ cm

Planck's Quantum Theory : Energy of one photon = $h\nu = \frac{hc}{\lambda}$

Photoelectric Effect : $h\nu = h\nu_0 + \frac{1}{2} m_e v^2$

Bohr's Model for Hydrogen like atoms :

1. $mvr = n \frac{h}{2\pi}$ (Quantization of angular momentum)

2. $E_n = -\frac{E_1}{n^2} z^2 = 2.178 \times 10^{-18} \frac{z^2}{n^2}$ J/atom $= 13.6 \frac{z^2}{n^2}$ eV ; $E_1 = \frac{-2\pi^2 me^4}{n^2}$

3. $r_n = \frac{n^2}{Z} \times \frac{h^2}{4\pi^2 e^2 m} = \frac{0.529 \times n^2}{Z}$ Å

4. $v = \frac{2\pi ze^2}{nh} = \frac{2.18 \times 10^6 \times z}{n}$ m/s

De–Broglie wavelength :

$\lambda = \frac{h}{mc} = \frac{h}{p}$ (for photon)

Wavelength of emitted photon :

$$\frac{1}{\lambda} = \bar{\nu} = RZ^2\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

No. of photons emitted by a sample of H atom :

$$\frac{\Delta n(\Delta n + 1)}{2}$$

Heisenberg's uncertainty principle :

$\Delta x.\Delta p > \frac{h}{4\pi}$ or $m\,\Delta x.\Delta v \geq \frac{h}{4\pi}$ or $\Delta x.\Delta v \geq \frac{h}{4\pi m}$

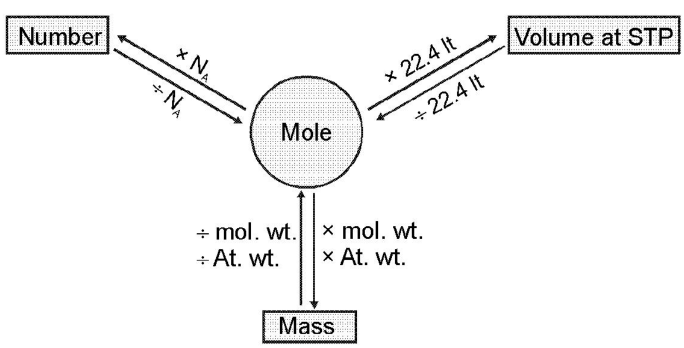
Quantum Numbers :

* Principal quantum number (n) = 1, 2, 3, 4 .... to ∞.
* Orbital angular momentum of electron in any orbit = $\frac{nh}{2\pi}$.
* Azimuthal quantum number (ℓ) = 0, 1, ... to (n – 1).
* Number of orbitals in a subshell = $2\ell + 1$
* Maximum number of electrons in particular subshell = $2 \times (2\ell + 1)$
* Orbital angular momentum $L = \frac{h}{2\pi}\sqrt{\ell(\ell+1)} = \hbar\sqrt{\ell(\ell+1)}$ $\left[\hbar = \frac{h}{2\pi}\right]$

STOICHIOMETRY

☞ Relative atomic mass (R.A.M) = $\frac{\text{Mass of one atom of an element}}{\frac{1}{12}\times\text{mass of one carbon atom}}$ = Total Number of nucleons

☞ <u>Y-map</u>



Density :

Specific gravity = $\frac{\text{density of the substance}}{\text{density of water at } 4^\circ C}$

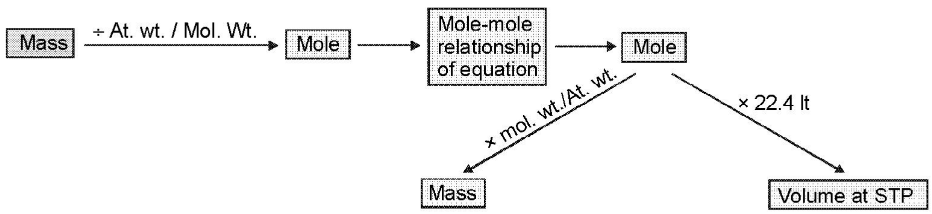
For gases :

Absolute density (mass/volume) = $\frac{\text{Molar mass of the gas}}{\text{Molar volume of the gas}}$ ⇒ $\rho = \frac{PM}{RT}$

Vapour density V.D.= $\frac{d_{gas}}{d_{H_2}} = \frac{PM_{gas/RT}}{PM_{H_2/RT}} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2}$

M_{gas} = 2 V.D.

Mole-mole analysis :



Concentration terms :

Molarity (M) :

$\therefore$ Molarity (M) = $\dfrac{w \times 1000}{(\text{Mol. wt of solute}) \times V_{inml}}$

Molality (m) :

Molality = $\dfrac{\text{number of moles of solute}}{\text{mass of solvent in gram}} \times 1000$ = $1000\, w_1 / M_1 w_2$

Mole fraction (x) :

$\therefore$ Mole fraction of solution (x_1) = $\dfrac{n}{n+N}$ $\therefore$ Mole fraction of solvent (x_2) = $\dfrac{N}{n+N}$

$x_1 + x_2 = 1$

% Calculation :

(i) % w/w = $\dfrac{\text{mass of solute in gm}}{\text{mass of solution in gm}} \times 100$

(ii) % w/v = $\dfrac{\text{mass of solute in gm}}{\text{mass of solution in ml}} \times 100$

(iii) % v/v = $\dfrac{\text{Volume of solution in ml}}{\text{Volume of solution}} \times 100$

Derive the following conversion :

1. Mole fraction of solute into molarity of solution M = $\dfrac{x_2 \rho \times 1000}{x_1 M_1 + M_2 x_2}$

2. Molarity into mole fraction x_2 = $\dfrac{MM_1 \times 1000}{\rho \times 1000 - MM_2}$

3. Mole fraction into molality m = $\dfrac{x_2 \times 1000}{x_1 M_1}$

4. Molality into mole fraction x_2 = $\dfrac{mM_1}{1000 + mM_1}$

5. Molality into molarity M = $\dfrac{m\rho \times 1000}{1000 + mM_2}$

6. Molarity into Molality $m = \frac{M \times 1000}{1000\rho - MM_2}$

M_1 and M_2 are molar masses of solvent and solute. ρ is density of solution (gm/mL)
M = Molarity (mole/lit.), m = Molality (mole/kg), x_1 = Mole fraction of solvent, x_2 = Mole fraction of solute

Average/Mean atomic mass :

$$A_x = \frac{a_1x_1 + a_2x_2 + ... + a_nx_n}{100}$$

Mean molar mass or molecular mass :

$$M_{avg.} = \frac{n_1M_1 + n_2M_2 + ...n_nM_n}{n_1 + n_2 + ...n_n} \quad \text{or} \quad M_{avg.} = \frac{\sum_{j=1}^{j=n} n_jM_j}{\sum_{j=1}^{j=n} n_j}$$

Calculation of individual oxidation number :

Formula : Oxidation Number = number of electrons in the valence shell – number of electrons left after bonding

Concept of Equivalent weight/Mass :

For elements, equivalent weight (E) = $\frac{\text{Atomic weight}}{\text{Valency - factor}}$

For acid/base, $E = \frac{M}{\text{Basicity / Acidity}}$ Where M = Molar mass

For O.A/R.A, $E = \frac{M}{\text{no. of moles of } e^- \text{ gained / lost}}$

Equivalent weight (E) = $\frac{\text{Atomic or moleculear weight}}{\text{v.f.}}$ **(v.f. = valency factor)**

Concept of number of equivalents :

No. of equivalents of solute = $\frac{Wt}{Eq.\ wt.} = \frac{W}{E} = \frac{W}{M/n}$

No. of equivalents of solute = No. of moles of solute × v.f.

Normality (N) :

Normality (N) = $\frac{\text{Number of equivalents of solute}}{\text{Volume of solution (in litres)}}$

Normality = Molarity × v.f.

Calculation of valency Factor :

n-factor of acid = basicity = no. of H^+ ion(s) furnished per molecule of the acid.

n-factor of base = acidity = no. of OH^- ion(s) furnised by the base per molecule.

At equivalence point :

$N_1V_1 = N_2V_2$

$n_1M_1V_1 = n_2M_2V_2$

Volume strength of H_2O_2:

20V H_2O_2 means **one litre** of this sample of H_2O_2 on decomposition gives **20 lt. of O_2** gas at **S.T.P.**

$$\text{Normality of } H_2O_2 \text{ (N)} = \frac{\text{Valume, strength of } H_2O_2}{5.6}$$

$$\textbf{Molarity of } H_2O_2 \textbf{(M)} = \frac{\textbf{Volume strength of } H_2O_2}{11.2}$$

Measurement of Hardness :

$$\textit{Hardness in ppm} = \frac{\text{mass of } CaCO_3}{\text{Total mass of water}} \times 10^6$$

Calculation of available chlorine from a sample of bleaching powder :

$\% \text{ of } Cl_2 = \dfrac{3.55 \times x \times V\text{(mL)}}{W(g)}$ where x = molarity of hypo solution and v = mL. of hypo solution used in titration.

GASEOUS STATE

Temperature Scale :

$\dfrac{C-O}{100-0} = \dfrac{K-273}{373-273} = \dfrac{F-32}{212-32} = \dfrac{R-R(O)}{R(100)-R(O)}$ where R = Temp. on unknown scale.

Boyle's law and measurement of pressure :

At constant temperature, $V \alpha \dfrac{1}{P}$

$P_1V_1 = P_2V_2$

Charles law :

At constant pressure, $V \alpha T$ or $\dfrac{V_1}{T_1} = \dfrac{V_2}{T_2}$

Gay-lussac's law :

At constant volume, $P \alpha T$

$\dfrac{P_1}{T_1} = \dfrac{P_2}{T_2} \rightarrow$ temp on absolute scale

Ideal gas Equation :

$PV = nRT$

$PV = \dfrac{w}{m} RT$ or $P = \dfrac{d}{m} RT$ or $Pm = dRT$

Daltons law of partial pressure :

$P_1 = \dfrac{n_1RT}{v}$, $P_2 = \dfrac{n_2RT}{v}$, $P_3 = \dfrac{n_3RT}{v}$ and so on.

Total pressure = $P_1 + P_2 + P_3 + \ldots\ldots\ldots\ldots$

Partial pressure = mole fraction X Total pressure.

Amagat's law of partial volume :

$$V = V_1 + V_2 + V_3 + ...$$

Average molecular mass of gaseous mixture :

$$M_{mix} = \frac{\text{Total mass of mixture}}{\text{Total no. of moles in mixture}} = \frac{n_1 M_1 + n_2 M_2 + n_3 M_3}{n_1 + n_2 + n_3}$$

Graham's Law :

Rate of diffusion $r \propto \frac{1}{\sqrt{d}}$; d = density of gas

$$\frac{r_1}{r_2} = \frac{\sqrt{d_2}}{\sqrt{d_1}} = \frac{\sqrt{M_2}}{\sqrt{M_1}} = \sqrt{\frac{V.D_2}{V.D_1}}$$

Kinetic Theory of Gases :

$PV = \frac{1}{3} mN \overline{U^2}$ Kinetic equation of gases

Average K.E. for one mole = $N_A \left(\frac{1}{2} m \overline{U^2}\right) = \frac{3}{2} K N_A T = \frac{3}{2} RT$

☞ Root mean suqare speed

$U_{rms} = \sqrt{\frac{3RT}{M}}$ molar mass must be in kg/mole.

☞ Average speed

$$U_{av} = U_1 + U_2 + U_3 + ... U_N$$

$U_{avg.} = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{\pi m}}$ K is Boltzmman constant

☞ Most probable speed

$$U_{MPS} = \sqrt{\frac{2RT}{M}} = \sqrt{\frac{2KT}{m}}$$

Vander wall's equation :

$$\left(P + \frac{an^2}{v^2}\right)(v - nb) = nRT$$

☞ **Critical constants :**

$V_c = 3b$, $P_C = \frac{a}{27b^2}$, $T_C = \frac{8a}{27Rb}$

Vander wall equation in virial form :

$$Z = \left(1+\frac{b}{V_m}+\frac{b^2}{V_m^2}+\frac{b^3}{V_m^3}+.........\right) - \frac{a}{V_m RT} = 1 + \frac{1}{V_m}\left(b-\frac{a}{RT}\right) + \frac{b^2}{V_m^2} + \frac{b^3}{V_m^3} + ...$$

Reduced Equation of state :

$$\left(P_r + \frac{3}{V_r^2}\right)(3V_r - 1) = 8\,T_r$$

THERMODYNAMICS

Thermodynamic processes :

1. **Isothermal process :** T = constant
   dT = 0
   $\Delta T = 0$
2. **Isochoric process :** V = constant
   dV = 0
   $\Delta V = 0$
3. **Isobaric process :** P = constant
   dP = 0
   $\Delta P = 0$
4. **Adiabatic process :** q = 0
   or heat exchange with the surrounding = 0(zero)

IUPAC Sign convention about Heat and Work :

Work done on the system = Positive
Work done by the system = Negative

1st Law of Thermodynamics

$$\Delta U = (U_2 - U_1) = q + w$$

Law of equipartion of energy :

$$U = \frac{f}{2}nRT \qquad \textbf{(only for ideal gas)}$$

$$\Delta E = \frac{f}{2}nR(\Delta T)$$

where f = degrees of freedom for that gas. (Translational + Rotational)

f = 3 for monoatomic
= 5 for diatomic or linear polyatmic
= 6 for non - linear polyatmic

Calculation of heat (q) :
Total heat capacity :

$$C_T = \frac{\Delta q}{\Delta T} = \frac{dq}{dT} = J/^oC$$

Molar heat capacity :

$$C = \frac{\Delta q}{n\Delta T} = \frac{dq}{ndT} = J\ mole^{-1} K^{-1}$$

$$C_P = \frac{\gamma R}{\gamma - 1} \qquad C_V = \frac{R}{\gamma - 1}$$

Specific heat capacity (s) :

$$S = \frac{\Delta q}{m\Delta T} = \frac{dq}{mdT} = Jgm^{-1} K^{-1}$$

WORK DONE (w) :

Isothermal Reversible expansion/compression of an ideal gas :

$$W = -nRT \ln (V_f/V_i)$$

Reversible and irreversible isochoric processes.

Since $dV = 0$

So $dW = -P_{ext} . dV = 0.$

Reversible isobaric process :

$$W = P (V_f - V_i)$$

Adiabatic reversible expansion :

$$\Rightarrow \quad T_2V_2^{\gamma-1} = T_1V_1^{\gamma-1}$$

Reversible Work :

$$W = \frac{P_2V_2 - P_1V_1}{\gamma - 1} = \frac{nR(T_2 - T_1)}{\gamma - 1}$$

Irreversible Work :

$$W = \frac{P_2V_2 - P_1V_1}{\gamma - 1} = \frac{nR(T_2 - T_1)}{\gamma - 1} \quad nC_v (T_2 - T_1) = -P_{ext} (V_2 - V_1) \text{ and use } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Free expansion – Always going to be irrerversible and since $P_{ext} = 0$

so $dW = -P_{ext} . dV = 0$

If no. heat is supplied $q = 0$

then $\Delta E = 0$ so $\Delta T = 0.$

Application of Ist Law :

$\Delta U = \Delta Q + \Delta W \quad \Rightarrow \quad \Delta W = -P\,\Delta V$

$\therefore \quad \Delta U = \Delta Q - P\Delta V$

Constant volume process

Heat given at constant volume = change in internal energy

$\therefore du = (dq)_v$

$du = nC_v dT$

$$C_v = \frac{1}{n} . \frac{du}{dT} = \frac{f}{2} R$$

Constant pressure process :

H ≡ Enthalpy (state function and extensive property)

$$H = U + PV$$

⇒ **$C_p - C_v = R$ (only for ideal gas)**

Second Law Of Thermodynamics :

$\Delta S_{universe} = \Delta S_{system} + \Delta S_{surrounding} > 0$ for a spontaneous process.

Entropy (S) :

$$\Delta S_{system} = \int_A^B \frac{dq_{rev}}{T}$$

Entropy calculation for an ideal gas undergoin a process :

State A $\xrightarrow[\Delta S_{irr}]{irr}$ State B

P_1, V_1, T_1 P_2, V_2, T_2

$$\Delta S_{system} = nc_v \ln \frac{T_2}{T_1} + nR \ln \frac{V_2}{V_1}$$ (only for an ideal gas)

Third Law Of Thermodynamics :

The entropy of perfect crystals of all pure elements & compounds is zero at the absolute zero of temperature.

Gibb's free energy (G) : (State function and an extensive property)

$$G_{system} = H_{system} - TS_{system}$$

Criteria of spontaneity :

(i) If ΔG_{system} is (–ve) < 0 ⇒ process is spontaneous

(ii) If ΔG_{system} is > 0 ⇒ process is non spontaneous

(iii) If $\Delta G_{system} = 0$ ⇒ system is at equilibrium.

Physical interpretation of ΔG :

→ The maximum amount of non-expansional (compression) work which can be performed.

$$\Delta G = dw_{non\text{-}exp} = dH - TdS.$$

Standard Free Energy Change (ΔG°) :

1. $\Delta G^o = -2.303\ RT \log_{10} K$

2. At equilibrium $\Delta G = 0$.

3. The decrease in free energy ($-\Delta G$) is given as :

$$-\Delta G = W_{net} = 2.303\ nRT \log_{10} \frac{V_2}{V_1}$$

4. ΔG_f^o for elemental state = 0

5. $\Delta G_f^o = G^o_{products} - G^o_{Reactants}$

Thermochemistry :

Change in standard enthalpy $\Delta H° = H^0_{m,2} - H^0_{m,1}$

= heat added at constant pressure. = $C_P\Delta T$.

If $H_{products} > H_{reactants}$

→ Reaction should be endothermic as we have to give extra heat to reactants to get these converted into products

and if $H_{products} < H_{reactants}$

→ Reaction will be exothermic as extra heat content of reactants will be released during the reaction.

Enthalpy change of a reaction : $\Delta H_{reaction} = H_{products} - H_{reactants}$

$\Delta H°_{reactions} = H°_{products} - H°_{reactants}$

= positive – endothermic

= negative – exothermic

Temperature Dependence Of ΔH : (Kirchoff's equation) :

For a constant volume reaction

$\Delta H_2° = \Delta H_1° + \Delta C_P (T_2 - T_1)$

where $\Delta C_P = C_P$ (products) – C_P (reactants).

For a constant volume reaction

$$\Delta E_2^0 = \Delta E_1^0 + \int \Delta C_V .dT$$

Enthalpy of Reaction from Enthalpies of Formation :

The enthalpy of reaction can be calculated by

$\Delta H_r° = \Sigma \nu_B \Delta H_f°,_{products} - \Sigma \nu_B \Delta H_f°,_{reactants}$ ν_B is the stoichiometric coefficient.

Estimation of Enthalpy of a reaction from bond Enthalpies :

$$\Delta H = \begin{pmatrix}\text{Enthalpy required to} \\ \text{break reactants into} \\ \text{gasesous atoms}\end{pmatrix} - \begin{pmatrix}\text{Enthalpy released to} \\ \text{form products from the} \\ \text{gasesous atoms}\end{pmatrix}$$

Resonance Energy :

$\Delta H°_{resonance} = \Delta H°_{f,\ experimental} - \Delta H°_{f,\ calclulated}$

$= \Delta H°_{c,\ calclulated} - \Delta H°_{c,\ experimental}$

CHEMICAL EQUILIBRIUM

At equilibrium :

(i) Rate of forward reaction = rate of backward reaction

(ii) Concentration (mole/litre) of reactant and product becomes constant.

(iii) $\Delta G = 0$.