

# TURBO

Created by students, for students

## State of Matter

Class 11 Chemistry • Complete Formula Sheet

Sr.	Concept	Formulas	How to Use / Other Information
<b>Units &amp; Conversions</b>			
1	Pressure Units	$1 \text{ atm} = 760 \text{ mm Hg} = 76 \text{ cm Hg}$ $= 760 \text{ torr} = 101.325 \text{ kPa}$ $= 101,325 \text{ Pa} = 1.01325 \text{ bar}$ $1 \text{ bar} = 10^5 \text{ Pa}$	<b>Most common:</b> <ul style="list-style-type: none"><li>- atm (atmosphere)</li><li>- mm Hg or torr (manometer)</li><li>- Pa (SI unit)</li></ul> <b>Conversion:</b> Divide by 760 to convert mm Hg to atm
2	Volume Units	$1 \text{ L} = 1000 \text{ mL} = 1 \text{ dm}^3$ $= 1000 \text{ cm}^3$ $1 \text{ m}^3 = 1000 \text{ L} = 10^6 \text{ mL}$	<b>Remember:</b> <ul style="list-style-type: none"><li>- <math>1 \text{ mL} = 1 \text{ cm}^3</math></li><li>- <math>1 \text{ L} = 1 \text{ dm}^3</math></li></ul>
3	Temperature Conversion	$K = ^\circ C + 273$ Absolute zero = 0 K = -273°C	<b>Important:</b> <ul style="list-style-type: none"><li>- Always use Kelvin in gas law calculations</li><li>- Room temp <math>\approx 25^\circ C = 298 \text{ K}</math></li><li>- STP: 0°C = 273 K</li></ul>
<b>Gas Laws</b>			
4	Boyle's Law	At constant T and n: $P \propto \frac{1}{V}$ or $PV = \text{constant}$ $P_1 V_1 = P_2 V_2$	<b>How to use:</b> <ul style="list-style-type: none"><li>- Temperature must be constant</li><li>- For same amount of gas</li></ul> <b>Example:</b> Gas at 2 atm, 5 L If P = 5 atm, then V = ? $V_2 = \frac{2 \times 5}{5} = 2 \text{ L}$
5	Charles' Law	At constant P and n: $V \propto T$ or $\frac{V}{T} = \text{constant}$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$	<b>How to use:</b> <ul style="list-style-type: none"><li>- Pressure must be constant</li><li>- T in Kelvin only!</li></ul> <b>Example:</b> Gas at 273 K has 10 L At 546 K: $V_2 = \frac{10 \times 546}{273} = 20 \text{ L}$
6	Gay-Lussac's Law (Pressure-Temp)	At constant V and n: $P \propto T$ or $\frac{P}{T} = \text{constant}$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$	<b>How to use:</b> <ul style="list-style-type: none"><li>- Volume must be constant</li><li>- T in Kelvin</li></ul> <b>Example:</b> Gas at 1 atm, 300 K At 600 K: $P_2 = \frac{1 \times 600}{300} = 2 \text{ atm}$
7	Avogadro's Law	At constant T and P: $V \propto n$ or $\frac{V}{n} = \text{constant}$ $\frac{V_1}{n_1} = \frac{V_2}{n_2}$	<b>Key point:</b> <ul style="list-style-type: none"><li>- Equal volumes of gases at same T, P contain equal moles</li><li>- At STP: 1 mole = 22.4 L</li></ul>
8	Ideal Gas Equation	$PV = nRT$ For 1 mole: $PV = RT$	<b>R values:</b> <ul style="list-style-type: none"><li>- <math>0.0821 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}</math></li><li>- <math>8.314 \text{ J} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}</math></li><li>- <math>1.987 \text{ cal} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}</math></li></ul> <b>How to use:</b> Know which R to use based on pressure units

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9	Density of Gas	$\rho = \frac{PM}{RT}$ where M = molar mass	<b>How to use:</b> - To find molar mass from density - $M = \frac{\rho RT}{P}$ <b>Example:</b> At STP, if $\rho = 1.43$ g/L $M = \frac{1.43 \times 0.0821 \times 273}{1} \approx 32$ g/mol
10	Molar Volume at STP	At STP (0°C, 1 atm): 1 mole of any gas = 22.4 L	<b>Remember:</b> - STP: 273 K, 1 atm - This is constant for ALL gases <b>Use:</b> $n = \frac{V}{22.4}$ at STP
<b>Dalton's Law &amp; Gas Mixtures</b>			
11	Dalton's Law of Partial Pressures	$P_{\text{total}} = P_A + P_B + P_C + \dots$ where $P_i$ = partial pressure of gas i	<b>Key concept:</b> - Total pressure = sum of individual pressures - Each gas behaves independently
12	Partial Pressure	$P_A = X_A \times P_{\text{total}}$ where $X_A = \frac{n_A}{n_{\text{total}}}$ (mole fraction)	<b>How to use:</b> - Find mole fraction first - Multiply by total pressure <b>Example:</b> Mixture: 2 mol O <sub>2</sub> , 8 mol N <sub>2</sub> , P = 10 atm $X_{O_2} = \frac{2}{10} = 0.2$ $P_{O_2} = 0.2 \times 10 = 2$ atm
<b>Graham's Law</b>			
13	Graham's Law of Diffusion/Effusion	$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{d_2}{d_1}}$ Or: $\frac{r_1}{r_2} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}$	<b>How to use:</b> - Rate $\propto \frac{1}{\sqrt{M}}$ - Lighter gas diffuses faster <b>Example:</b> H <sub>2</sub> (M=2) vs O <sub>2</sub> (M=32) $\frac{r_{H_2}}{r_{O_2}} = \sqrt{\frac{32}{2}} = 4$ H <sub>2</sub> diffuses 4 times faster
<b>Kinetic Theory of Gases</b>			
14	Kinetic Gas Equation	$PV = \frac{1}{3}mn c^2$ where m = mass of one molecule n = number of molecules c = RMS speed	<b>Relates:</b> - Pressure to molecular motion - m = mass of 1 molecule = $\frac{M}{N_A}$
15	Kinetic Energy of Gas	For 1 mole: $KE = \frac{3}{2}RT$ For n moles: $KE = \frac{3}{2}nRT$	<b>Key points:</b> - KE depends only on temperature - At same T, all gases have same KE - Independent of nature of gas
16	Molecular Speeds	<b>Most probable:</b> $c_{mp} = \sqrt{\frac{2RT}{M}}$ <b>Average:</b> $\bar{c} = \sqrt{\frac{8RT}{\pi M}}$ <b>RMS:</b> $c_{rms} = \sqrt{\frac{3RT}{M}}$ Ratio = 1 : 1.128 : 1.224	<b>How to use:</b> - M in kg/mol for SI units - R = 8.314 J/(mol·K) <b>Order:</b> $c_{mp} < \bar{c} < c_{rms}$ <b>Example:</b> For N <sub>2</sub> at 300 K: $c_{rms} = \sqrt{\frac{3 \times 8.314 \times 300}{0.028}}$ m/s
<b>Real Gases &amp; van der Waals</b>			
17	Compressibility Factor	$Z = \frac{PV}{nRT}$ Z = 1: Ideal gas Z > 1: Less compressible (positive deviation) Z < 1: More compressible (negative deviation)	<b>Interpretation:</b> - Ideal gas: Z = 1 at all P, T - Real gas: Z varies with P, T - At low P, high T: gas behaves ideally

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18	van der Waals Equation	$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$ For 1 mole: $(P + \frac{a}{V^2})(V - b) = RT$	<b>Corrections:</b> - $\frac{an^2}{V^2}$ : Pressure correction (attractive forces) - nb: Volume correction (molecular volume) <b>How to use:</b> - 'a' large $\rightarrow$ strong intermolecular forces - 'b' large $\rightarrow$ large molecular size
19	van der Waals Constants	<b>Constant 'a':</b> - Units: $\text{atm} \cdot \text{L}^2 \cdot \text{mol}^{-2}$ - Measure of intermolecular forces <b>Constant 'b':</b> - Units: $\text{L} \cdot \text{mol}^{-1}$ - Excluded volume = $4 \times$ actual volume	<b>Remember:</b> - Larger 'a' $\rightarrow$ easier to liquefy - Larger 'b' $\rightarrow$ larger molecules <b>Order of 'a':</b> $\text{NH}_3 > \text{H}_2\text{O} > \text{CO}_2 > \text{O}_2 > \text{H}_2$
<b>Combined Gas Law Problems</b>			
20	Combined Gas Law	$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$ (for constant n)	<b>How to use:</b> - Combines Boyle's, Charles', Gay-Lussac's - Always use Kelvin for T <b>Example:</b> Gas at 2 atm, 3 L, 300 K Find V at 1 atm, 600 K: $V_2 = \frac{2 \times 3 \times 600}{1 \times 300} = 12 \text{ L}$
21	Converting to STP	$V_{\text{STP}} = V \times \frac{273}{T} \times \frac{P}{1}$ where T in K, P in atm	<b>How to use:</b> - To compare gas volumes - STP = 273 K, 1 atm <b>Example:</b> 5 L at 2 atm, 546 K $V_{\text{STP}} = 5 \times \frac{273}{546} \times \frac{2}{1} = 5 \text{ L}$
22	Gas Stoichiometry	Moles = $\frac{PV}{RT} = \frac{V}{22.4}$ (at STP) For reactions: Use mole ratios	<b>Steps:</b> 1. Find moles using $PV = nRT$ 2. Use stoichiometry 3. Convert back to volume <b>Example:</b> $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ If 11.2 L $\text{H}_2$ at STP, $\text{O}_2$ needed? Moles $\text{H}_2 = 0.5$ , so $\text{O}_2 = 0.25 \text{ mol} = 5.6 \text{ L}$
23	Vapor Pressure of Water	When gas collected over water: $P_{\text{total}} = P_{\text{gas}} + P_{\text{water vapor}}$ $P_{\text{gas}} = P_{\text{total}} - P_{\text{H}_2\text{O}}$	<b>How to use:</b> - Subtract vapor pressure of water - Vapor pressure depends on temperature <b>Common values:</b> At 25°C: 23.8 mm Hg At 30°C: 31.8 mm Hg