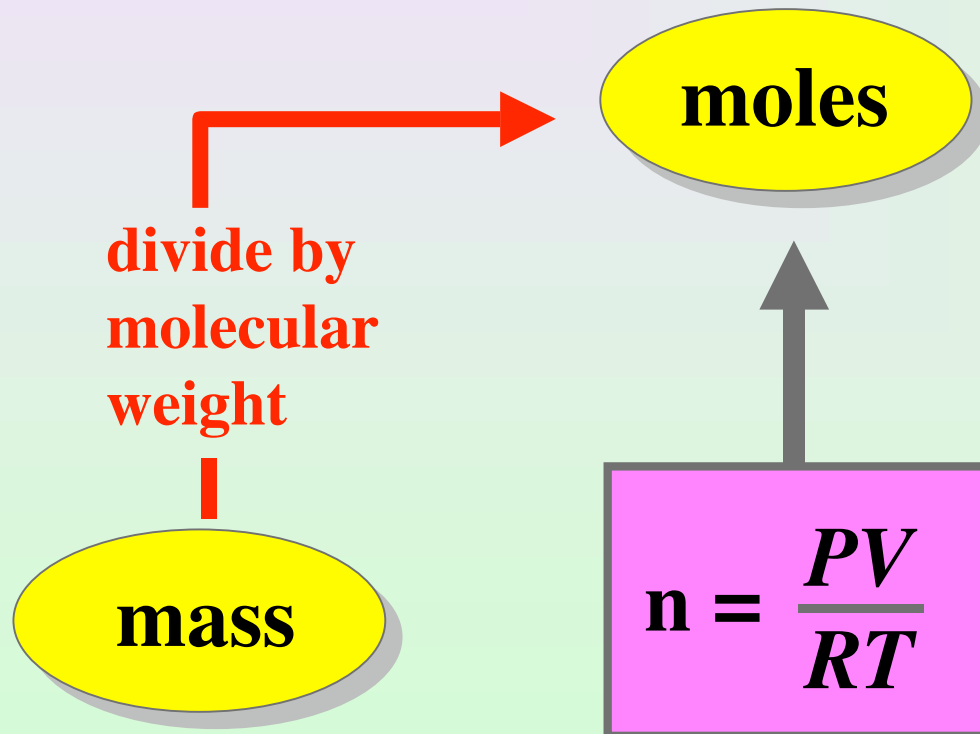


Stoichiometry Involving Gases

The Mole Method



$$V = kn$$

Example

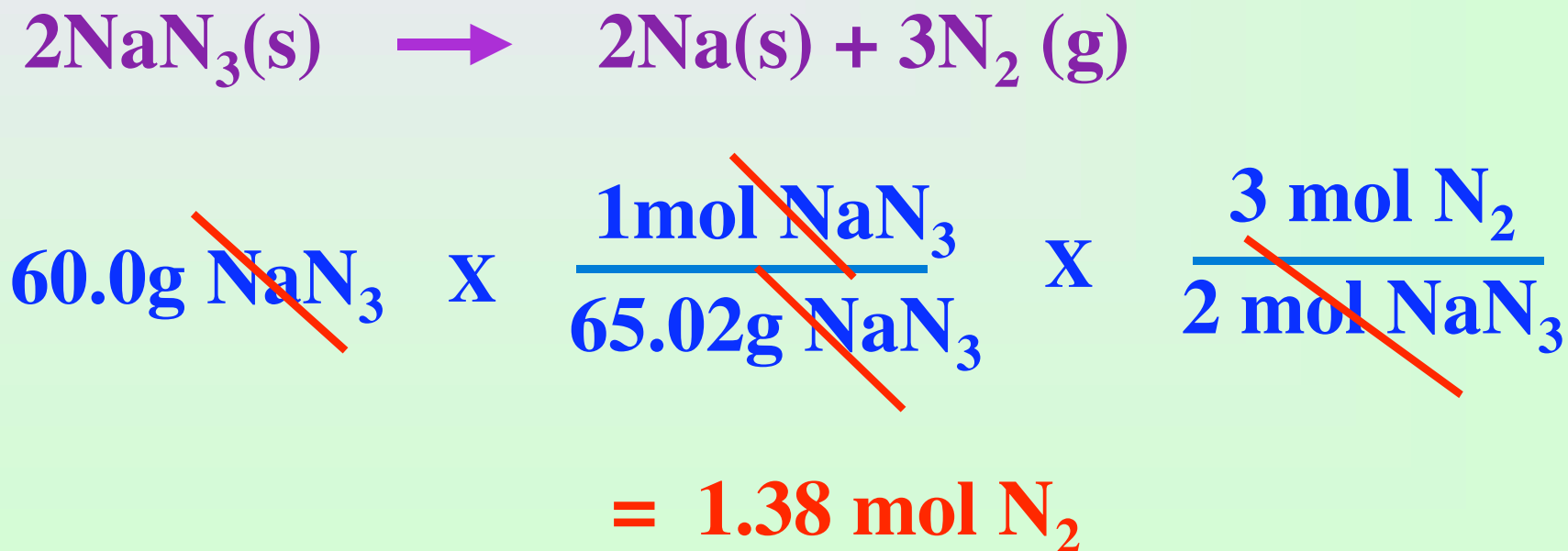
Calculate the volume of O_2 required for the complete combustion of 14.9 L of butane.



$$14.9 \text{ L } C_4H_{10} \times \frac{13 \text{ L } O_2}{2 \text{ L } C_4H_{10}} = 96.9 \text{ L}$$

Example

Calculate the volume of N_2 generated at 21°C and 823 mmHg upon the decomposition of 60.0g of NaN_3



Example Cont.

Calculate the volume of N_2 generated at 21°C and 823 mmHg upon the decomposition of 60.0g of NaN_3

$$n = 1.38 \text{ mol}$$

$$V = \frac{nRT}{P} =$$

$$P = (823/760)\text{atm}$$

$$\frac{(1.38 \text{ mol NaN}_3) (0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(294 \text{ K})}{(823/760)\text{atm}}$$

$$= 30.8 \text{ L}$$

Example

An empty 0.5L sealed container was filled with $F_2(g)$ until the pressure was 1 atm at $25^\circ C$. To this was added a 0.156g sample of $Cu(s)$ which is completely reacted to give $CuF_2(s)$. What is the final pressure of $F_2(g)$ in the container (at $25^\circ C$)?



$$0.156g \text{ Cu} \times \frac{1 \text{ mol Cu}}{63.55g \text{ Cu}} \times \frac{1 \text{ mol } F_2}{1 \text{ mol Cu}} = 0.00245 \text{ mol } F_2$$

Example Cont.

$$P = \frac{nRT}{V} \quad 0.00245 \text{ mol F}_2$$

$$(0.00245 \text{ mol F}_2) (0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})$$

$$0.5 \text{ L}$$

$$= 0.12 \text{ atm F}_2$$

$$1.00 \text{ atm} - 0.12 \text{ atm} = 0.88 \text{ atm}$$

Dalton's Law of Partial Pressures

Dalton's Law of Partial Pressures

The total pressure of a mixture of gases is just the sum of the pressures that each gas would exert if it were present alone.

$$P_{\text{total}} = P_1 + P_2 + P_3 + \dots$$

partial pressure of gas 1
partial pressure of gas 2
partial pressure of gas 3

$$P_{total} = n_{total} \left(\frac{RT}{V} \right)$$

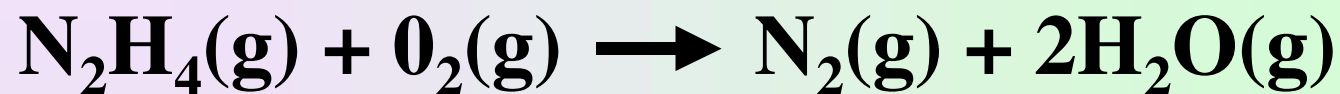
pressure depends only on the number of gas molecules present, not on what gases are present

Example

The reaction shown below was carried out at constant temperature and constant volume. At the start, the reaction vessel contained only N_2H_4 and O_2 , each contributing a partial pressure of 2 atm. The reaction was allowed to proceed until 50% of the reactants were consumed. What was the total pressure in the container at this time?



Example



Start: 2atm 2atm 0 0

Finish: 1atm 1atm 1atm 2atm

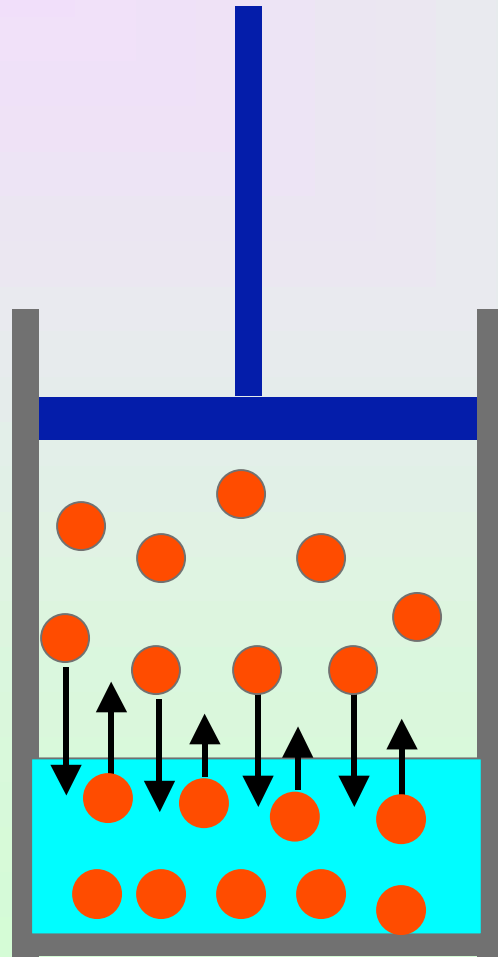
$$\mathbf{P_{\text{total}} = 5 \text{ atm}}$$

Commonly encountered case of partial pressures involves a gas saturated with water.

vapor pressure of water

pressure exerted by $\text{H}_2\text{O}(\text{g})$ in equilibrium with $\text{H}_2\text{O}(\text{l})$ at a particular temperature

Vapor Pressure



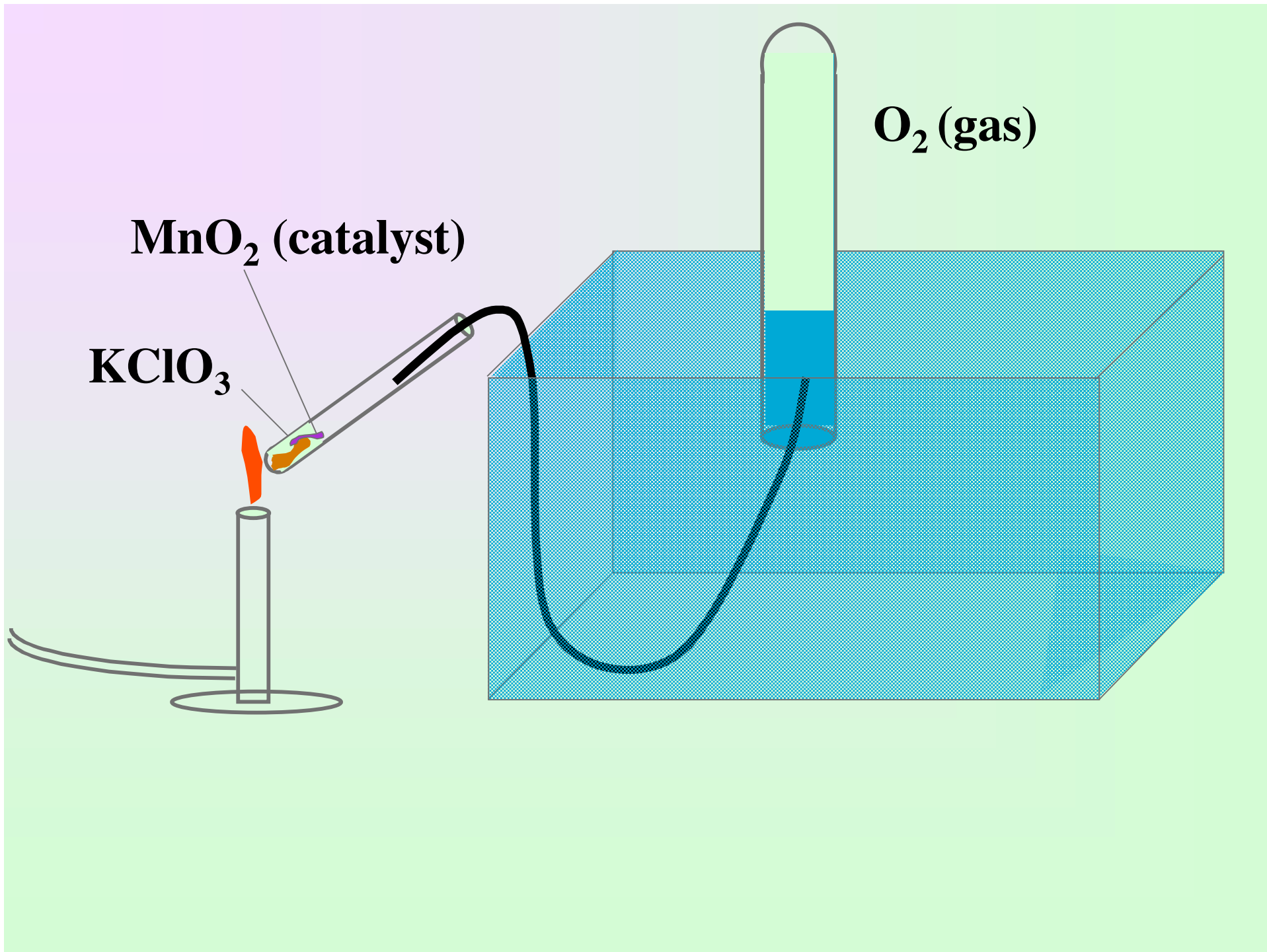
$\text{H}_2\text{O}(g)$

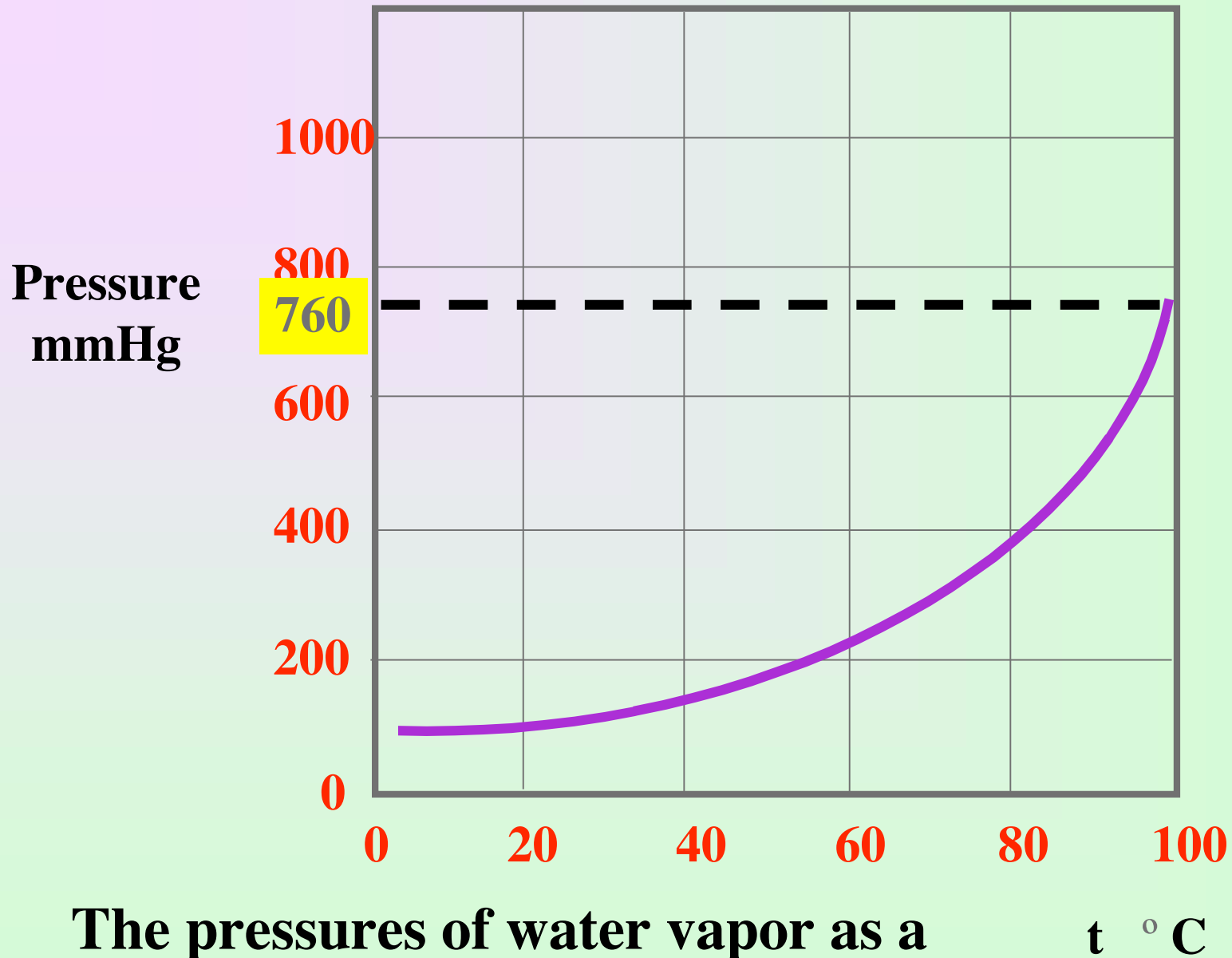
$\text{H}_2\text{O}(l)$

Laboratory preparation of O₂

Oxygen gas can be prepared by heating potassium chlorate:







The pressures of water vapor as a function of temperature.

Temperature (°C)	Water Vapor pressure (mmHg)
20	17.54
25	23.76
30	31.82
35	42.18

Example

Oxygen gas can be prepared by heating potassium chlorate:



On heating KClO_3 , 128 mL of gas was collected over water at 24 °C and 762 mm Hg. (At 24 °C the vapor pressure of water is 22.4 mm Hg.) What is the mass of O_2 collected?

Example



$$P_{(\text{O}_2)} = P_{\text{tot}} - P_{(\text{H}_2\text{O})}$$

$$P_{(\text{O}_2)} = 762 \text{ mm Hg} - 22.4 \text{ mm Hg}$$

$$P_{(\text{O}_2)} = 739.6 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}}$$
$$= 0.973 \text{ atmO}_2$$

Example



$$n = \frac{PV}{RT}$$

$$P = 0.973 \text{ atm O}_2$$

$$n = \frac{(0.973 \text{ atm})(0.128\text{L})}{(0.0821 \text{ L-atm/mol-K})(297\text{K})}$$

$$n = 0.00508 \text{ mol O}_2$$

$$(0.00508 \text{ mol})(32 \text{ g/mol}) = 0.16 \text{ g O}_2$$

Concentration

Mole fraction (X_n)

Moles of component / total moles of all materials

Mole fraction X_A = $\frac{\text{moles of component } n_A}{\text{total moles } n_A + n_B + n_C \dots}$

The partial pressure of a particular component is proportional to the mole fraction of that component times the total pressure

$$P_1 = (X_1) (P_{\text{total}})$$

Example

The mole fraction of nitrogen in the air is **0.7808**. Calculate the partial pressure N_2 in when the atmospheric pressure is **760. torr**.

$$\begin{aligned} P_1 &= (X_1) (P_{\text{total}}) \\ &= 0.7808 \times 760. \text{ torr} \\ &= 593 \text{ torr} \end{aligned}$$

