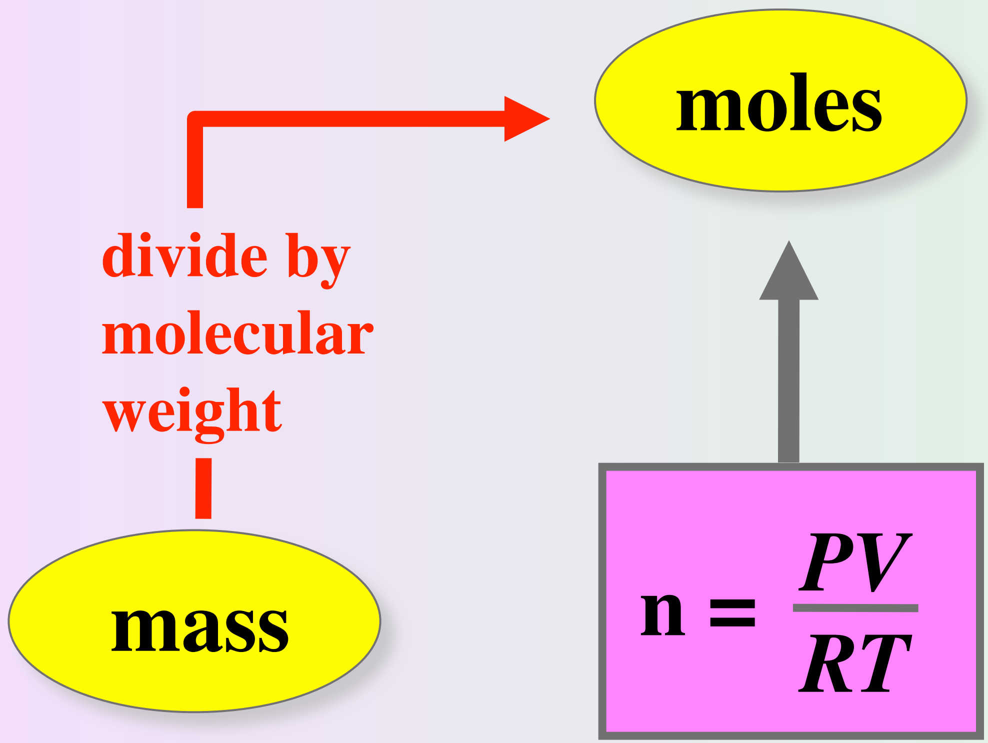
Stoichiometry Involving Gases

The Mole Method

$$A + B \longrightarrow C$$

atoms or molecules



Example

$V = kn$

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

$$2C_4H_{10}(g) + 130_2(g) \longrightarrow 8CO_2(g) + 10H_2O(liq)$$

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

Example

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

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

$$60.0\text{g } \cancel{NaN_3} \times \frac{1\text{mol } \cancel{NaN_3}}{65.02\text{g } \cancel{NaN_3}} \times \frac{3 \text{ mol } N_2}{2 \text{ mol } \cancel{NaN_3}}$$

$$= 1.38 \text{ mol } N_2$$

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 mol

P =(823/760)atm

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

$$\frac{(1.38 \text{ mol } NaN_3)(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(294 \text{ K})}{(823/760)\text{atm}}$$

$$= 30.8 \text{ L}$$

Example

An empty 0.5L sealed container was filled with F_2 (g) until the pressure was 1 atm at 25°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°C)?

$$Cu(s) + F_2(g) \longrightarrow CuF_2(s)$$

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

$$= 0.00245 \text{ mol } F_2$$

Example Cont.

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

0.00245 mol F_2

$$\frac{(0.00245 \text{ mol } F_2)(0.0821 \text{ L} \bullet \text{atm/K} \bullet \text{mol})(298 \text{ K})}{0.5 \text{ L}}$$

$= 0.12$ atm F_2

1.00 atm - 0.12 atm $= 0.88$ 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.

$P_{total} = P_1 + P_2 + P_3 + ...$

partial pressure of gas 3

partial pressure of gas 2

partial pressure of gas 1

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

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?

$$N_2H_4(g) + 0_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

Example

$$N_2H_4(g) + 0_2(g) \longrightarrow N_2(g) + 2H_2O(g)$$

Start:	2atm	2atm	0	0
Finish:	1atm	1atm	1atm	2atm

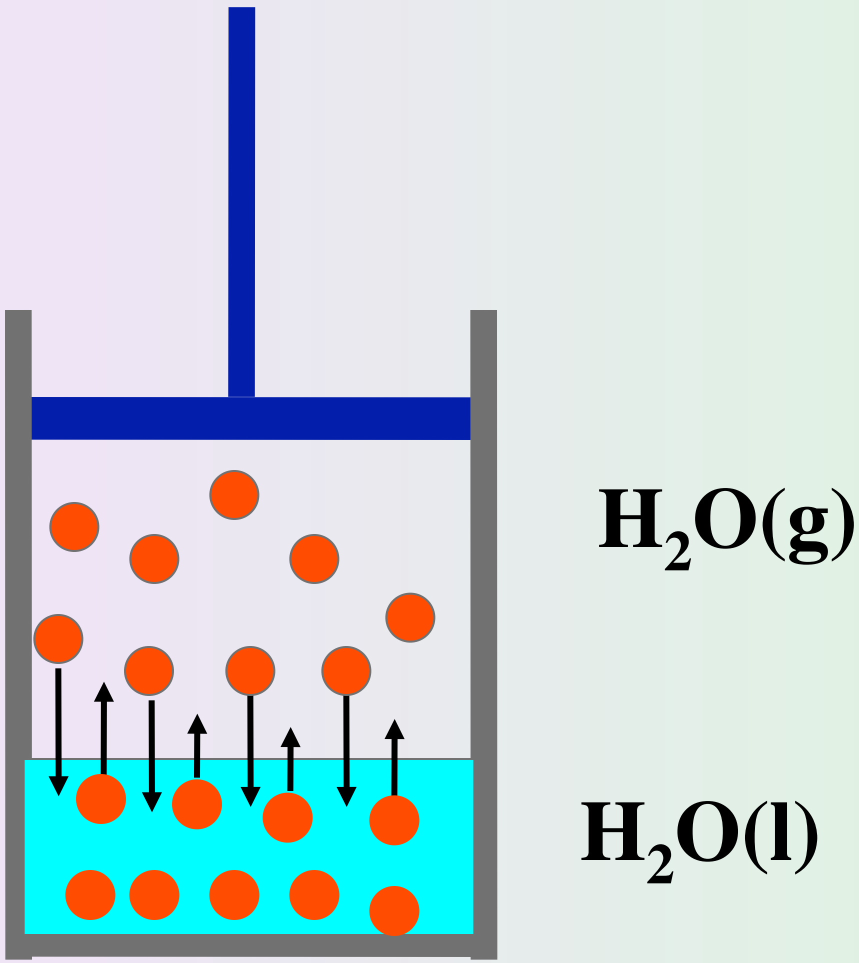
$$P_{total} = 5\ atm$$

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

vapor pressure of water

pressure exerted by $H_2O(g)$ in equilibrium with $H_2O(l)$ at a particular temperature

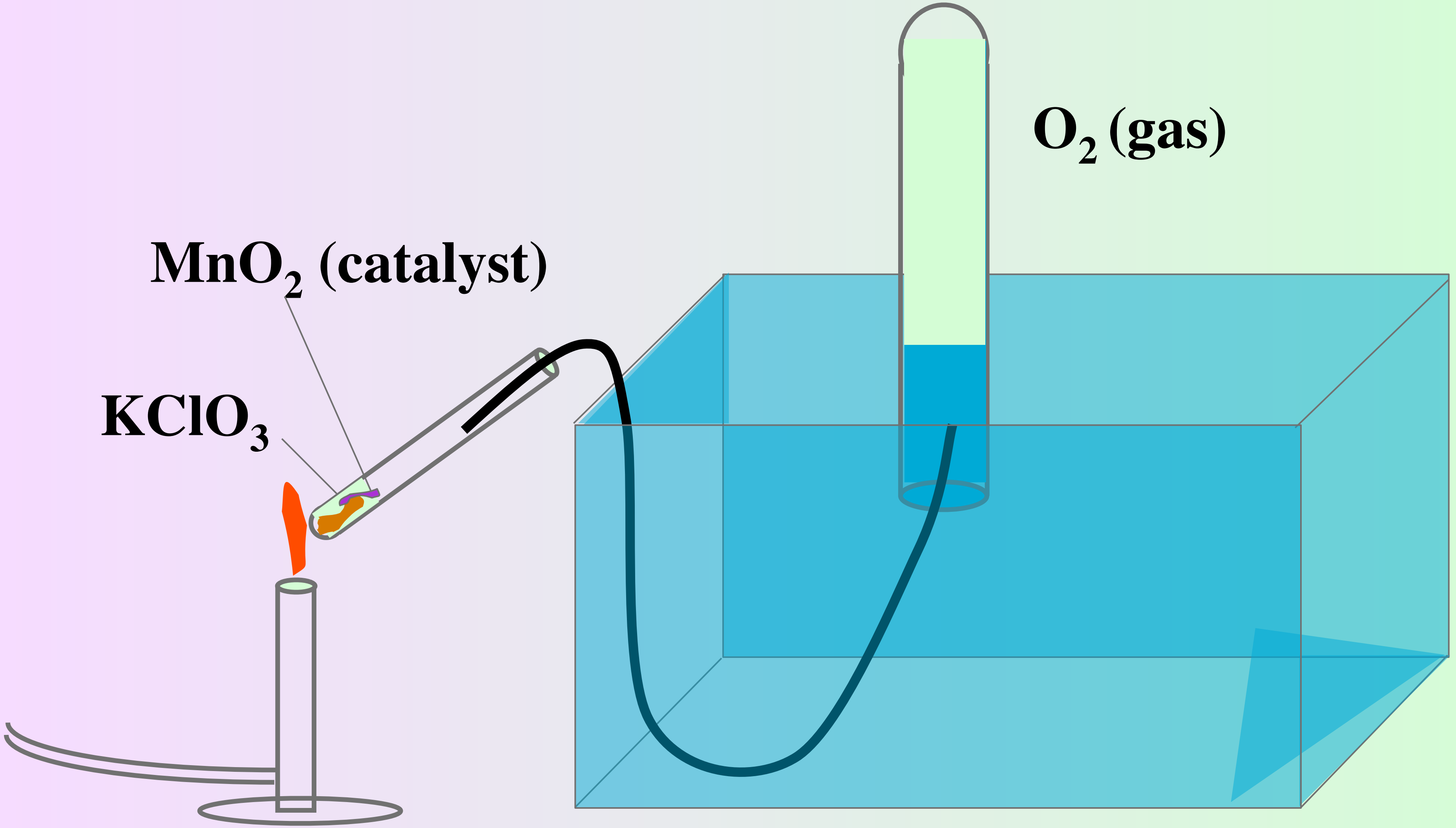
Vapor Pressure

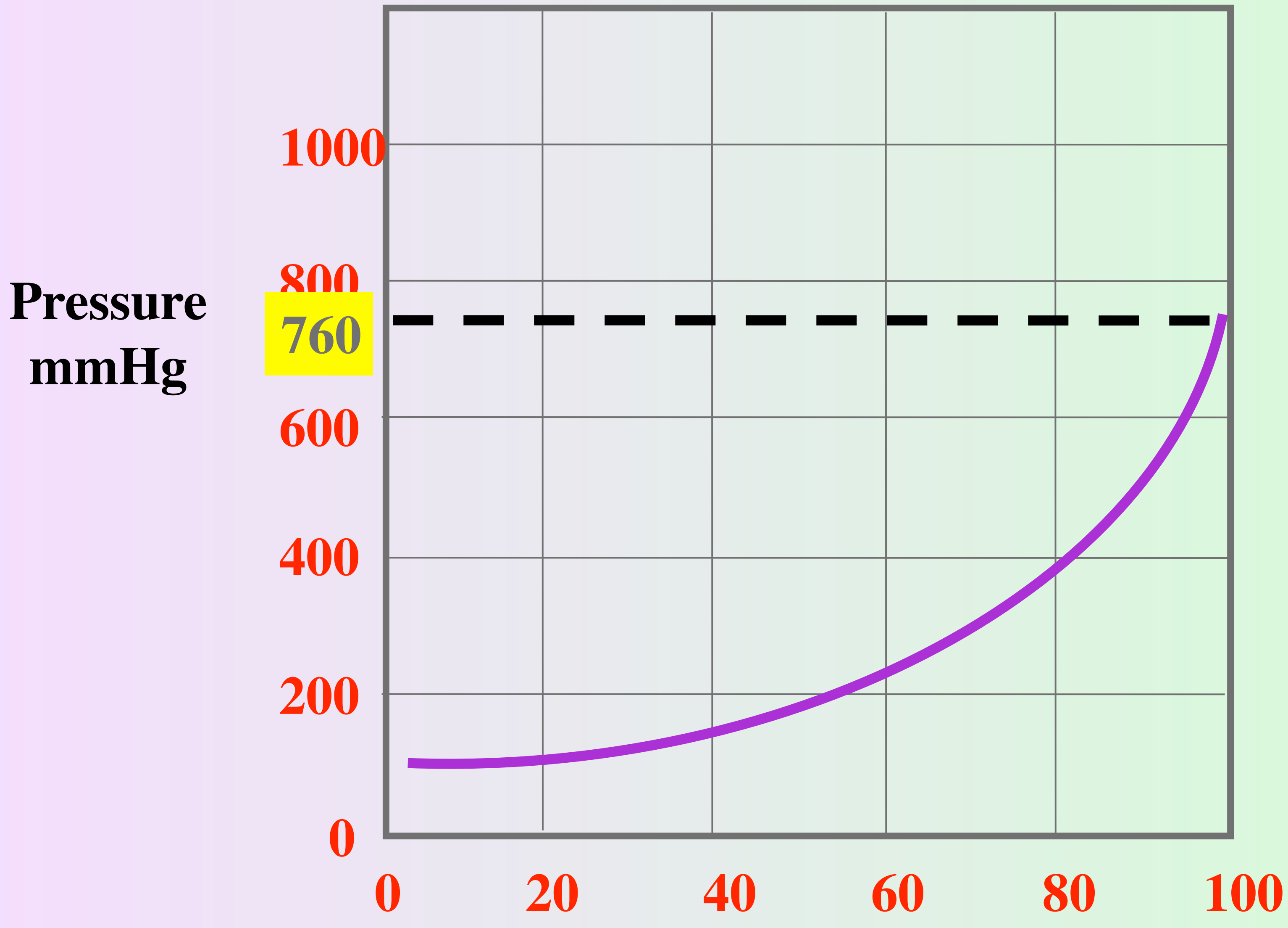


Laboratory preparation of O_2

Oxygen gas can be prepared by heating potassium chlorate:

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$


O_2 (gas)
MnO_2 (catalyst)
$KClO_3$



The pressures of water vapor as a function of temperature.

Temperature (^{o}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:

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

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

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

$$\boxed{P_{(0_2)} = P_{tot} - P_{(H_2O)}}$$

$$P_{(0_2)} = 762 \text{ mm Hg} - 22.4 \text{ mm Hg}$$

$$P_{(0_2)} = 739.6 \text{ mm Hg} \times \frac{1 \text{ atm}}{760 \text{ mm Hg}}$$

$$= 0.973 \text{ atm} O_2$$

Example

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

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

$P = 0.973$ atm O_2

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

$n = 0.00508$ mol O_2

(0.00508 mol)(32 g/mol) $= 0.16$ g O_2

Concentration

Mole fraction (X_n)

Moles of component / total moles of all materials

moles of component

Mole fraction

$$X_A = \frac{n_A}{n_A + n_B + n_C \ldots}$$

total moles

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

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

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.

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

$$= 0.7808 \times 760.\ \text{torr}$$

$$= 593\ \text{torr}$$