Partial Pressures

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GASES

1. Gases take up the volume of the container – has no definite shape or volume
2. Gases mix well – diffusion
3. Gases exert pressure

THINGS WE CARE ABOUT FOR GASES

- Pressure (P)
- Volume (V)
- Temperature (T)
- Moles (n)

We'll come back to these in a moment.

ATMOSPHERIC PRESSURE

Remember that we are always under the pressure of the atmosphere, which is defined as 1 atm.

Any system that is allowed to equilibrate with the pressure of the atmosphere will try to obtain atmospheric pressure.

This is how balloons work because they can change their volume to maintain atmospheric pressure *inside*.

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How to calculate partial pressure of gas A in a mixture:

- Determine moles of the gas — n_A
- Determine the mole ratio — $X_A = n_A/n_{total}$
- Multiply mole ratio by total pressure — $P_A = X_A P_{total}$
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How to calculate partial pressure of gas A in a mixture:
- Determine moles of the gas n_A
- Determine the mole ratio $X_A = n_A/n_{total}$
- Multiply mole ratio by total pressure $P_A = X_AP_{total}$
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- Or apply the ideal gas law on gas A only to find P_A

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$$P_{\mathrm{N_2}} = \frac{n_{\mathrm{N_2}}RT}{V}$$
$$= \frac{(186\ \mathrm{mol})\left(0.08206\ \frac{\mathrm{L\cdot atm}}{\mathrm{mol\cdot K}}\right)(297._{15}\ \mathrm{K})}{50.0\ \mathrm{L}}$$

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For O_2:

$$P_{\mathrm{O_2}} = \frac{n_{\mathrm{O_2}}RT}{V}$$
$$= \frac{(145\ \mathrm{mol})\left(0.08206\ \frac{\mathrm{L\cdot atm}}{\mathrm{mol\cdot K}}\right)(297._{15}\ \mathrm{K})}{50.0\ \mathrm{L}}$$
$$P_{\mathrm{O_2}} = 70.7_1\ \mathrm{atm}$$

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$$P_{\mathrm{O_2}} = 70.7_1\ \mathrm{atm}$$

For total pressure: $P_{\mathrm{total}} = P_{N_2} + P_{O_2} = 90.7_1\ \mathrm{atm} + 70.7_1\ \mathrm{atm} = 161\ \mathrm{atm}$

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Find mole fractions for each gas:

$$X_{\mathrm{N_2}} = \frac{n_{\mathrm{N_2}}}{n_{\mathrm{N_2}} + n_{\mathrm{O_2}}} = \frac{186\ \text{mol}}{331\ \text{mol}} = 0.561_9 \ \text{ and } \ X_{\mathrm{O_2}} = \frac{n_{\mathrm{O_2}}}{n_{\mathrm{N_2}} + n_{\mathrm{O_2}}} = \frac{145\ \text{mol}}{331\ \text{mol}} = 0.438_1$$

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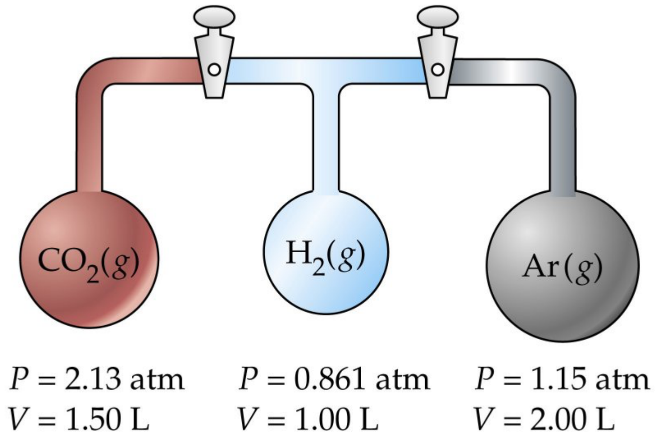
Find partial pressures from mole fractions and P_{total}:

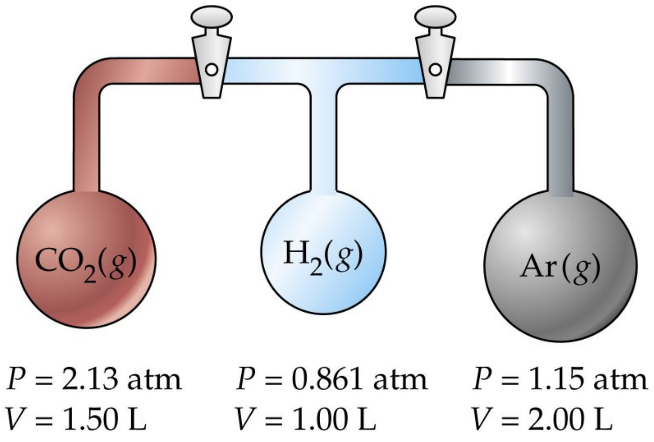
$$P_{\mathrm{N_2}} = X_{\mathrm{N_2}} P_{\mathrm{total}} = (0.561_9)(161._4\ \mathrm{atm}) = 90.7\ \mathrm{atm}$$

$$P_{\mathrm{O_2}} = X_{\mathrm{O_2}} P_{\mathrm{total}} = (0.438_1)(161._4\ \mathrm{atm}) = 70.7\ \mathrm{atm}$$



Imagine that both stopcocks were opened so that the gases mix at 298 K.

What is the partial pressure of each gas after opening?



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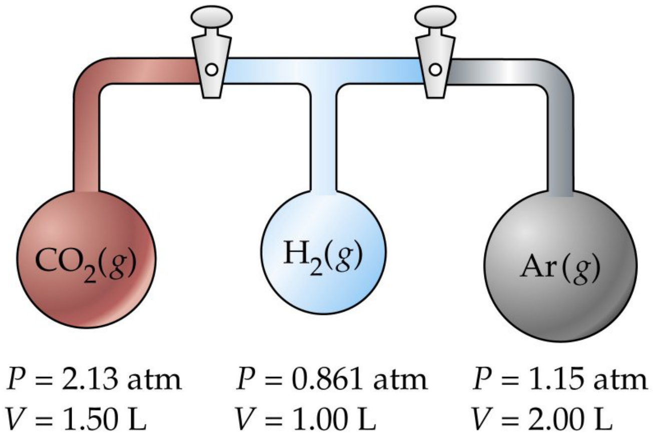

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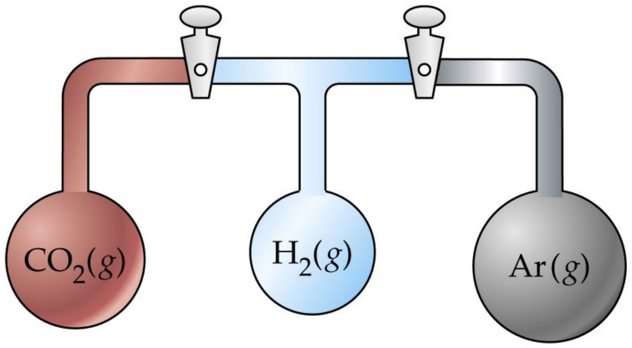
Solve for the number of moles using the pressures of each gas using the ideal gas law:

$$n_{CO_2} = \frac{PV}{RT}$$

$$= \frac{(2.13\ \text{atm})(1.50\ \mathbf{L})}{\left(0.08206\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right)(298\ \text{K})}$$

$$n_{CO_2} = 0.130_7\ \text{mol}$$



$P = 2.13$ atm $P = 0.861$ atm $P = 1.15$ atm
$V = 1.50$ L $V = 1.00$ L $V = 2.00$ L

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$$n_{H_2} = \frac{PV}{RT} = \frac{(0.861\text{ atm})(1.00\text{ L})}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298\text{ K})}$$
$$n_{N_2} = 0.0352_1\text{ mol}$$

$$n_{Ar} = \frac{PV}{RT} = \frac{(1.15\text{ atm})(2.00\text{ L })}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298\text{ K})}$$
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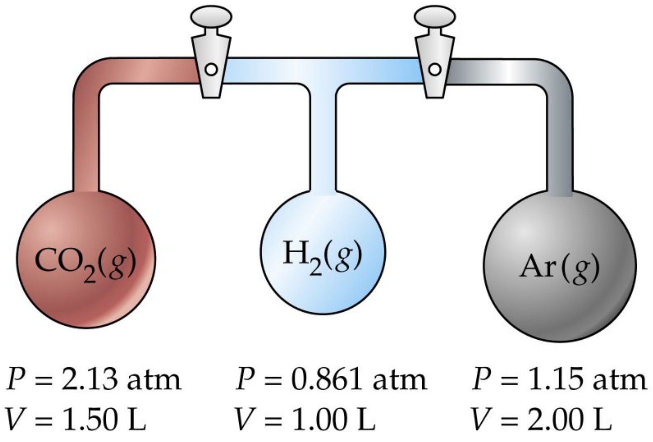

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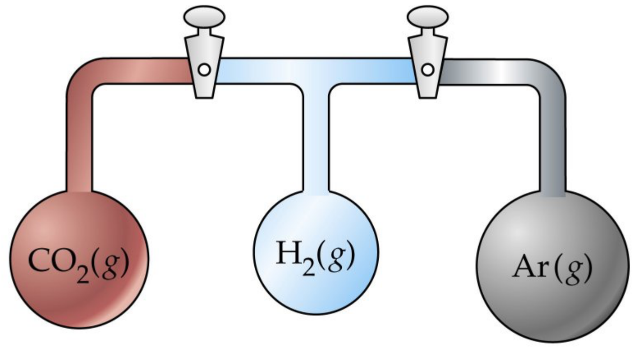
If we open the stopcocks, the total volume changes to 4.50 L. Now solve for new pressures:



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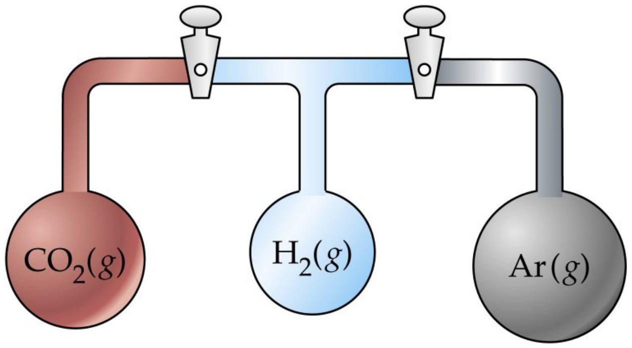
$$P_{CO_2} = 0.710\text{ atm}$$



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$$P_{Ar} = 0.511\text{ atm}$$