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[TP] Consider 1 mol each of two different gases, A and B, each with the same volume (1 L) and temperature (25°C). The molar mass of A is twice the molecular mass of B. What is the relationship between the pressures of A and B?

89% 1. $P_A = 2P_B$
0% 2. $P_A = \frac{P_B}{2}$
0% 3. $P_A = 4P_B$
0% 4. $P_A = \frac{P_B}{4}$
0% 5. $P_A = P_B$
11% 6. Not sure

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Lecture 5 CH131 Summer 1 2021

Wednesday, June 2, 2021

- Ideal gas law
- Gas constant R, STP, and SATP
- Gas law calculational recipe
- Gas density
- Dalton's law of partial pressures
- Gas law calculations in chemical reactions

Next lecture: Kinetic molecular theory; Why the gas law does not depend on mass; Distribution of molecular speeds; Real gases; Gas law for real gases: van der Waals equation; Ch10: Solids, liquids and phase transitions

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Ideal gas law

In the expression $P = c''nT/V$, the remaining constant of proportionality, c'', is known as the gas constant R, $P = RnT/V$.

This expression is written as

$$PV = nRT$$

and it is known as the ideal gas law.

The adjective "ideal" signifies that this relation does not depend of what the gas particles are.

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[TP] Consider 1 mol each of two different gases, A and B, each with the same volume (1 L) and temperature (25°C). The molar mass of A is twice the molecular mass of B. What is the relationship between the pressures of A and B?

8% 1. $P_A = 2P_B$
8% 2. $P_A = \frac{P_B}{2}$
0% 3. $P_A = 4P_B$
0% 4. $P_A = \frac{P_B}{4}$
85% 5. $P_A = P_B$
0% 6. Not sure

$P_A = \frac{n_A RT}{V} = \frac{1\,\text{mol}\, R\, 298\,\text{K}}{1\,\text{L}}$

$P_B = \frac{n_B RT}{V} = \frac{1\,\text{mol}\, R\, 298\,\text{K}}{1\,\text{L}}$

Amazing!!

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Ideal gas law

The ideal gas law

$$PV = nRT$$

is as true for a gas of O_2 molecules as for a gas of Xe atoms.

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Gas constant R

The value of the gas constant has been determined by measuring V for a given P, n, and T.

For example, the volume of 1 mol of gas at 20°C and 1 bar = 100000 Pa = 100000 J/m^3 is

$$V = 0.02437387 \text{ m}^3$$

From this we can evaluate that

$$R = \frac{PV}{nT} = \frac{10^5 \text{J/m}^3 \times 0.02437387 \text{ m}^3}{1 \text{ mol } 293.15 \text{ K}} = 8.314472 \frac{\text{J}}{\text{K mol}}$$

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Gas constant R

Often, pressures are measured in atm (1.01325 bar) and volumes are measured in L (1 L = 10^{-3} m^3).

Since 1 m^3 = 1000 L and 1 Pa = 1 atm/101325,

$$1 \text{ J} = \text{m}^3\text{Pa} = 1000 \text{ L} \times \frac{\text{atm}}{101325} = 0.00986923 \text{ L atm}$$

From this we can evaluate that

$$R = 8.314472 \frac{\text{J}}{\text{K mol}} \times \frac{0.00986923 \text{ L atm}}{\text{J}} = 0.0820575 \frac{\text{L atm}}{\text{K mol}}$$

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Gas constant R

Since

$$1 \text{ J} = 1 \text{ Pa m}^3 = 10^{-5}\text{bar} \times 10^3\text{L} = 10^{-2} \text{ L bar}$$

an alternative expression for the value of the gas constant

$$R = 8.314472 \frac{\text{J}}{\text{K mol}}$$

is

$$R = 0.08314472 \frac{\text{L bar}}{\text{K mol}}$$

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Gas constant R

$R = 8.314472 \frac{\text{J}}{\text{K mol}}$

$R = 0.08314472 \frac{\text{L bar}}{\text{K mol}}$

Finally, pressures are sometimes measured in atm (1.01325 bar) and so a third expression for the value of the gas constant is

$$R = 0.08314472 \frac{\text{L bar}}{\text{K mol}} \times \frac{\text{atm}}{1.01325\ \text{bar}} = 0.0820575 \frac{\text{L atm}}{\text{K mol}}$$

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Gas constant R

$R = 8.314472 \frac{\text{J}}{\text{K mol}}$

$R = 0.08314472 \frac{\text{L bar}}{\text{K mol}}$

$R = 0.0820575 \frac{\text{L atm}}{\text{K mol}}$

In calculations, use the value of R with the most convenient units.

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STP and SATP

Two different "standard" values of temperature and pressure are defined.

STP is Standard Temperature and Pressure,

$T = 273.15\ \text{K} = 0°\text{C}$ and $P = 1\ \text{bar} = 10^5\ \text{Pa}$

SATP is Standard Ambient Temperature and Pressure,

$T = 298.15\ \text{K} = 25°\text{C}$ and $P = 1\ \text{bar} = 10^5\ \text{Pa}$

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[TP] What is the volume (in L) of 1.00 mol of an ideal gas at standard ambient temperature and pressure, SATP? $R = 8.3145 \frac{\text{J}}{\text{mol K}} = 0.083145 \frac{\text{L bar}}{\text{mol K}}$

0% 1. Less than 20 L
0% 2. 20-22 L
0% 3. 22-24 L
100% 4. 24-26 L
0% 5. 26-28 L
0% 6. 28-30 L
0% 7. More than 30 L

$PV = nRT$

$V = \frac{nRT}{P} = \frac{1\ \text{mol}\ R\ 298\ \text{K}}{1\ \text{bar}}$

$T = 25°\text{C} \rightarrow 298\ \text{K}$

$P = 1\ \text{bar}$

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STP and SATP

What is the volume (in L) of 1.00 mol of an ideal gas at standard ambient temperature and pressure, SATP?

$$V = \frac{nRT}{P} = \frac{1\text{ mol } 0.08314\text{ L bar/(K mol) } 298.15\text{ K}}{1.00\text{ bar}} = 24.5\text{ L}$$

STP 22 L
SATP 24.5 L

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Gas law calculational recipe

The recipe is to rearrange $PV = nRT$ so that what does not change is grouped on one side of the equality and what does change is grouped on the other.

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Gas law calculational recipe

A 2-liter cylinder contains a gas at 250 kPa at 18°C. What will be the pressure at 125°C?

Heating the gas makes the particles move faster, there will be more collisions with the container walls in a given time, and so the pressure will increase.

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Gas law calculational recipe

A 2-liter cylinder contains a gas at 250. kPa at 18°C. What will be the pressure at 125°C?

The volume (V) and amount (n) of gas are constant, so rearrangement gives

$$\frac{nR}{V} = \frac{P_1}{T_1} = \frac{P_2}{T_2} \rightarrow P_2 = P_1 \frac{T_2}{T_1}$$

$$P_2 = 250.\text{ kPa} \times \frac{(273.15+125)\text{ K}}{(273.15+18)\text{ K}} = 342\text{ kPa}$$

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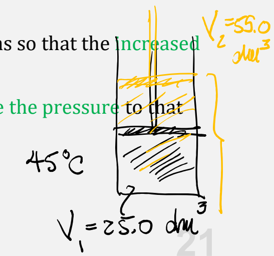
Gas law calculational recipe

A piston containing 25.0 dm^3 of gas at 45°C is expanded to 55.0 dm^3.

The increase in volume means the gas particles must travel farther between collisions with the container walls, there will be fewer collisions with the walls in a given time, and so the pressure will decrease.

The original pressure can be restored by heating the gas so that the increased speed offsets the greater distance.

To what temperature must the gas be heated to restore the pressure to that before the expansion?



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Gas law calculational recipe

A piston containing 25.0 dm^3 of gas at 45°C is expanded to 55.0 dm^3. To what temperature must the gas be heated to maintain the same pressure as before the expansion?

The pressure (P) and amount (n) of gas are constant, so rearrangement gives

$$\frac{nR}{P} = \frac{V_1}{T_1} = \frac{V_2}{T_2} \rightarrow T_2 = T_1 \frac{V_2}{V_1}$$

$$T_2 = (273.15 + 45)\ \text{K} \times \frac{55.0\ \text{dm}^3}{25.0\ \text{dm}^3} = 700.\ \text{K} = 427\text{°C}$$

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Gas law calculations

Calculate the density of water in the gas phase at 100.0°C and 1.00 atm.

Rearrangement of the ideal gas law gives

$$\frac{n}{V} = \frac{P}{RT}$$

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Gas law calculations

Calculate the density of water in the gas phase at 100.0°C and 1.00 atm.

Rearrangement of the ideal gas law gives

$$\frac{n}{V} = \frac{P}{RT} = \frac{1.00\ \text{atm}}{0.0820575 \frac{\text{L atm}}{\text{K mol}} \times 373.25\ \text{K}} = 0.0327 \frac{\text{mol}}{\text{L}}$$

$$\frac{n}{V} = 0.0327 \frac{\text{mol}}{\text{L}} \times \frac{(2 \times 1.008 + 15.999)\ \text{g}}{\text{mol}} \times \frac{\text{L}}{1000\ \text{mL}} = 0.000588 \frac{\text{g}}{\text{mL}}$$

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Gas law calculations

At 100.0°C , the density of water vapor is 0.000588 g/mL and that of liquid water is 0.958 g/mL. How much farther apart of water molecules in the gas phase than those in the liquid phase?

The ratio of the densities,

$$\frac{0.958 \text{ g/mL}}{0.000588 \text{ g/mL}} = 1630.,$$

is how many more liquid water molecules there are in a given volume than there are gaseous water molecules in the same volume.

a lot more $H_2O(l)$ / many fewer $H_2O(g)$

1630 $H_2O(l)$ / 1 $H_2O(g)$

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Gas law calculations

This means the same volume can hold 1630 times as many liquid water molecules as gaseous water molecules.

Since volume is the cube of distance, the cube root of the ratio,

$1630^{1/3} = 11.8$

means the gas particles are about 11.8 times farther apart than liquid particles.

So, about a factor of ten farther apart than in liquids or solids.

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Partial pressures

We have seen that the ideal gas law is the same for all gases, since it depends only on the number (moles) of particles but not what kind of particles they are.

This means each kind of gas contributes to the total pressure only in proportion to its relative number of moles,

$$P = \frac{n_a RT}{V} + \frac{n_b RT}{V} + \cdots = P_a + P_b + \cdots$$

where P_a, P_b, ..., are the partial pressures of gas a, gas b, etc.,

$$P_a = \frac{n_a RT}{V}, \text{ etc.}$$

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Partial pressures

A convenient way to calculate partial pressures is to use mole fractions,

$$x_a = \frac{n_a}{n}, x_b = \frac{n_b}{n}, \ldots$$

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Partial pressures

Since

$$P = \frac{n_a}{n}\frac{nRT}{V} + \frac{n_b}{n}\frac{nRT}{V} + \cdots = x_a \frac{nRT}{V} + x_b \frac{nRT}{V} + \cdots = x_a P + x_b P + \cdots$$

we see that the partial pressure of each kind of gas is

$P_a = x_a P, P_b = x_b P, \ldots$

This is known as Dalton's law of partial pressures.

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[TP] When 1 mol of $H_2(g)$ is added to a cylinder of volume V and at temperature T the pressure is 1 bar. Then an additional 0.5 mol of gas $CO_2(g)$ is added to the same cylinder keeping the temperature the same. What is the new pressure in the cylinder?

0% 1. < 1 bar
0% 2. = 1 bar
0% 3. Between 1 and 1.5 bar
100% 4. 1.5 bar
0% 5. Between 1.5 and 2 bar
0% 6. = 2 bar

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Partial pressures

A mixture consisting of 0.513 g of $H_2(g)$ and 16.1 g of $N_2(g)$ occupies 10.0 L at 20.0°C. Calculate the partial pressures of each gas in atm.

First step, calculate the moles of each gas and the total moles.

$n_{H_2} = 0.254\text{ mol}, n_{N_2} = 0.575\text{ mol}, n = 0.254\text{ mol} + 0.575\text{ mol} = 0.829\text{ mol}$

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Partial pressures

A mixture consisting of 0.513 g of $H_2(g)$ and 16.1 g of $N_2(g)$ occupies 10.0 L at 20.0°C. Calculate the partial pressures of each gas in atm.

$n_{H_2} = 0.254\text{ mol}, n_{N_2} = 0.575\text{ mol}, n = 0.829\text{ mol}$

Second step: Use the total moles to calculate the total pressure.

$$P = \frac{nRT}{V} = 0.829\text{ mol} \times 0.0820575\,\frac{\text{L atm}}{\text{K mol}} \times \frac{293.15\text{ K}}{10.0\text{ L}} = 1.99\text{ atm}$$

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Partial pressures

A mixture consisting of 0.513 g of $H_2(g)$ and 16.1 g of $N_2(g)$ occupies 10.0 L at 20.0°C. Calculate the partial pressures of each gas in atm.

$n_{H_2} = 0.254$ mol, $n_{N_2} = 0.575$ mol, $n = 0.829$ mol, $P = 1.99$ atm

Final step: Calculate the partial pressures.

$P_{H_2} = x_{H_2} P = \frac{0.254}{0.829} \times 1.99 \text{ atm} = 0.612 \text{ atm}$

$P_{N_2} = x_{N_2} P = \frac{0.575}{0.829} \times 1.99 \text{ atm} = 1.38 \text{ atm}$

Check:

$P = P_{H_2} + P_{N_2} = 0.612 \text{ atm} + 1.38 \text{ atm} = 1.99 \text{ atm}$

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Gas law calculations in chemical reactions

$C_3H_8(g)$ and $O_2(g)$ are placed in a sealed, rigid container. After the container has been heated to 125°C, the partial pressures are $P_{C_3H_8} = 1.00$ atm and $P_{O_2} = 1.00$ atm. A spark then initiates the complete combustion of the mixture, forming $CO_2(g)$ and $H_2O(g)$. What is the total pressure in the container at the end of the reaction at 125°C?

Write the balanced chemical equation.

$C_3H_8 + 5\,O_2 \rightarrow 3\,CO_2 + 4\,H_2O$

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Gas law calculations in chemical reactions

$C_3H_8(g)$ and $O_2(g)$ are placed in a sealed, rigid container. After the container has been heated to 125°C, the partial pressures are $P_{C_3H_8} = 1.00$ atm and $P_{O_2} = 1.00$ atm. A spark then initiates the complete combustion of the mixture, forming $CO_2(g)$ and $H_2O(g)$. What is the total pressure in the container at the end of the reaction at 125°C?

$C_3H_8(g) + 5\,O_2(g) \rightarrow 3\,CO_2(g) + 4\,H_2O(g)$

Determine the limiting reagent

1.00 atm $C_3H_8 \rightarrow$ 3.00 atm CO_2

1.00 atm $O_2 \rightarrow$ 0.600 atm CO_2

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Gas law calculations in chemical reactions

$C_3H_8(g) + 5\,O_2(g) \rightarrow 3\,CO_2(g) + 4\,H_2O(g)$

1.00 atm $O_2 \rightarrow$ 0.600 atm CO_2

$1.00 \text{ atm } O_2 \times \frac{4 \text{ atm } H_2O}{5 \text{ atm } O_2} = 0.800$

$1.00 \text{ atm } O_2 \times \frac{1 \text{ atm } C_3H_8}{5 \text{ atm } O_2} = 0.200$

Determine all partial pressures and the total pressure.

	$C_3H_8(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(g)$	P_{total}
Initial	1.00	1.00	0	0	2
Change	−0.200	−1.00	+0.600	+0.800	
Final					

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Gas law calculations in chemical reactions

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

1.00 atm O_2 → 0.600 atm CO_2

Determine all partial pressures and the total pressure.

	$C_3H_8(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(g)$	P_{total}
Initial	1.00	1.00	0	0	2
Change	−0.200	−1.00	+0.600	+0.800	
Final					

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Gas law calculations in chemical reactions

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

1.00 atm O_2 → 0.600 atm CO_2

Determine all partial pressures and the total pressure.

	$C_3H_8(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(g)$	P_{total}
Initial	1.00	1.00	0	0	2
Change	−0.200	−1.00	+0.600	+0.800	
Final	0.80	0	0.600	0.800	2.20

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[TP] $CH_4(g)$ and $O_2(g)$ are placed in a sealed, rigid container. After the container has been heated to 125°C, the partial pressures are P_{CH_4} = 1.00 atm and P_{O_2} = 1.00 atm. A spark then initiates the complete combustion of the mixture, forming $CO_2(g)$ and $H_2O(g)$. What is the total pressure in the container at the end of the reaction at 125°C?

① balanced eqn.
② limiting reagent
③ ICF table

0% 1. 0.50 atm
0% 2. 1.00 atm
0% 3. 1.50 atm
100% 4. 2.00 atm
0% 5. 2.50 atm
0% 6. 3.00 atm

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Gas law calculations in chemical reactions

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

1.00 atm CH_4 → 1.00 atm CO_2

1.00 atm O_2 → 0.500 atm CO_2

Determine all partial pressures and the total pressure.

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Gas law calculations in chemical reactions

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$

1.00 atm O_2 → 0.500 atm CO_2

Determine all partial pressures and the total pressure.

	$CH_4(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(g)$	P_{total}
Initial	1.00	1.00	0	0	2.00
Change					
Final					

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Gas law calculations in chemical reactions

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$

1.00 atm O_2 → 0.500 atm CO_2

Determine all partial pressures and the total pressure.

	$CH_4(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(g)$	P_{total}
Initial	1.00	1.00	0	0	2.00
Change	−0.500	−1.00	+0.500	+1.00	
Final					

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Gas law calculations in chemical reactions

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$ *yield was 37% instead*

1.00 atm O_2 → 0.500 atm CO_2

Determine all partial pressures and the total pressure.

	$CH_4(g)$	$O_2(g)$	$CO_2(g)$	$H_2O(g)$	P_{total}
Initial	1.00	1.00	0	0	2.00
Change	−0.500	−1.00	+0.500	+1.00	
Final	0.50	0	0.500	1.00	2.00

100% means limiting reagent all used up

37%

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