IB DP PHYSICS · THEME B THE PARTICULATE NATURE OF MATTER · B.3 · ~7 MIN READ

B.3 Gas laws

Pressure, the mole and Avogadro's number, the three gas laws and the ideal gas equation in moles and molecules, the kinetic model of an ideal gas, $P = \frac{1}{3}\rho v^2$, the internal energy of a monatomic gas, and when real gases are not ideal.

Pressure · Moles & Avogadro · Boyle, Charles & pressure laws · PV = nRT · Kinetic model · $P = \frac{1}{3}\rho v^2$ · Internal energy · Real gases

RS by Riyaz Sir
mentorriyaz.in

No tag = **SL and HL** · HL ONLY HL students only · DATA BOOKLET given in the exam

M1 A1 how marks are awarded · highlight = comes up again and again · Test yourself at the end

IN 30 SECONDS

- **Pressure** is force per unit area, P = F/A. In a gas it comes from molecules hitting the walls.
- For a fixed amount of ideal gas, **PV/T = constant**. The full ideal gas law is **PV = nRT** (moles) or **PV = Nk_BT** (molecules), with T in kelvin.
- The **kinetic model** links pressure to molecular speeds, $P = \frac{1}{3}\rho\langle v^2\rangle$, and gives the internal energy of an ideal monatomic gas, $U = \frac{3}{2}nRT$.
- Real gases behave like ideal gases at **low pressure, high temperature and low density**.
- **B.3 is the same at SL and HL.**

BY THE END YOU CAN

- Use P = F/A and $n = N/N_A$.
- Apply the gas laws and PV = nRT or $PV = Nk_BT$, with temperatures in kelvin.
- State the assumptions of the kinetic model and explain pressure with it.
- Use $P = \frac{1}{3}\rho\langle v^2\rangle$ and $U = \frac{3}{2}Nk_BT$.
- Explain when and why real gases differ from ideal gases.

1 Pressure and the mole

$$P = \frac{F}{A}$$ DATA BOOKLET

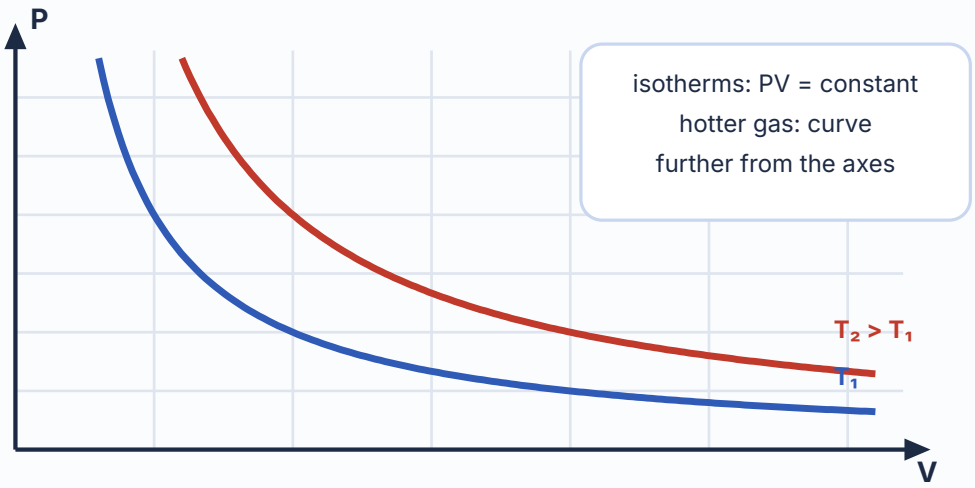
pressure in pascals (Pa = N m^{-2})

$$n = \frac{N}{N_A}$$ DATA BOOKLET

amount of substance in moles; $N_A = 6.02 \times 10^{23}$ mol^{-1}

- **Molar mass** M is the mass of one mole: about 0.028 kg mol^{-1} for N_2, 0.032 for O_2, 0.004 for He. Mass of gas m = nM.

2 The gas laws

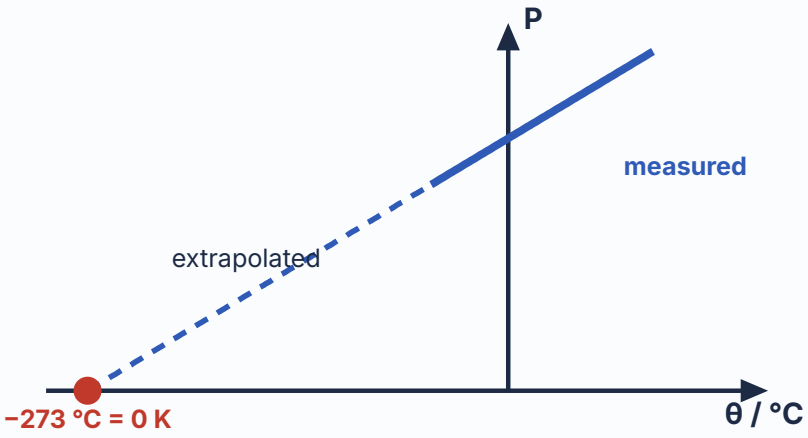


At constant temperature, halving the volume doubles the pressure.

Law	Kept constant	Relationship
Boyle's law	T and n	PV = constant
Charles' law	P and n	V/T = constant
Pressure (Gay-Lussac) law	V and n	P/T = constant

$$\frac{PV}{T} = \text{constant}$$ DATA BOOKLET

for a fixed mass of ideal gas, so $P_1V_1/T_1 = P_2V_2/T_2$



Every gas extrapolates to zero pressure at the same temperature: absolute zero.

Straight lines that all point back to −273 °C are how absolute zero was found.

WORKED EXAMPLE

A tyre is at 2.0 × 10^5 Pa at 20 °C. After a drive its temperature is 40 °C. Find the new pressure (volume constant).

1. T in kelvin: 293 K and 313 K M1
2. $P_2 = 2.0 \times 10^5 \times 313/293$ M1
3. **$= 2.14 \times 10^5$ Pa** A1

WORKED EXAMPLE

A weather balloon holds 2.0 L of gas at 1.0 × 10^5 Pa and 290 K. It rises to 0.50 × 10^5 Pa and 250 K. Find its volume.

1. $V_2 = V_1 \times (P_1/P_2) \times (T_2/T_1)$ M1
2. $= 2.0 \times 2.0 \times (250/290)$ M1
3. **$= 3.4$ L** A1

COMMON MISTAKE

Putting °C into P/T or V/T. 40 °C is not "double" 20 °C: in kelvin the rise is only 313/293 = 1.07.

3 The ideal gas law

$$PV = nRT$$ DATA BOOKLET

$R = 8.31$ J mol^{-1} K^{-1}

$$PV = Nk_BT$$ DATA BOOKLET

$k_B = R/N_A = 1.38 \times 10^{-23}$ J K^{-1}

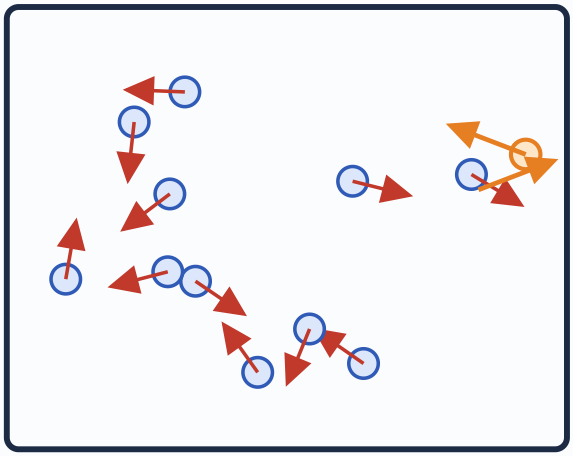
WORKED EXAMPLE

How many moles and molecules of gas are in 0.50 m^3 at 1.0 × 10^5 Pa and 300 K?

1. $n = PV/RT = (1.0 \times 10^5 \times 0.50)/(8.31 \times 300)$ M1
2. $n =$ **20 mol** A1
3. **$N = nN_A = 20 \times 6.02 \times 10^{23} = 1.2 \times 10^{25}$** A1

4 The kinetic model of an ideal gas

- **Assumptions:** a very large number of identical molecules in random motion; the molecules' own volume is negligible compared with the container; collisions (with each other and the walls) are elastic and last a negligible time; no forces between molecules except during collisions; between collisions they move in straight lines at constant speed.



Each collision with a wall reverses the momentum component: $\Delta p = 2mv$. Many collisions per second give a steady force, so a pressure $P = F/A$.

$P = \frac{1}{3}\rho\langle v^2\rangle$

Pressure is the average force per unit area from huge numbers of collisions with the walls.

- **Explaining the laws:** heat the gas at constant volume → faster molecules → harder and more frequent collisions → higher pressure. Halve the volume at constant temperature → twice as many collisions per second on each m^2 → double the pressure.

$$P = \frac{1}{3}\rho\langle v^2\rangle$$ DATA BOOKLET

$\langle v^2\rangle$ is the mean square speed of the molecules

WORKED EXAMPLE

Find the root mean square speed of nitrogen molecules at 300 K.

1. combine $P = \frac{1}{3}\rho\langle v^2\rangle$ with $PV = nRT$, $\rho = nM/V$: $\langle v^2\rangle = 3RT/M$ M1
2. $= 3 \times 8.31 \times 300/0.028 = 2.67 \times 10^5\ m^2\ s^{-2}$ M1
3. **$v_{rms} = 517\ m\ s^{-1}$** A1

$$U = \frac{3}{2}Nk_BT = \frac{3}{2}nRT$$ DATA BOOKLET

internal energy of an IDEAL MONATOMIC gas: all of it is kinetic

WORKED EXAMPLE

Find the internal energy of 2.0 mol of helium at 300 K.

1. $U = 1.5 \times 2.0 \times 8.31 \times 300$ M1
2. **$= 7.5$ kJ** A1

- **Why no potential energy?** In an ideal gas there are no intermolecular forces, so the intermolecular potential energy is zero. The internal energy depends only on temperature.

5 Real gases

- **Real gases deviate** at high pressure and low temperature: the molecules are close together, so their own volume and the attractive forces between them are no longer negligible. Near condensation the ideal gas law fails completely.
- **Good approximation:** low pressure, high temperature, low density (for example air at room conditions).

EXAM TIP

"Explain in terms of molecules" questions want: molecular speed or kinetic energy, the rate of collisions with the walls, and the change of momentum per collision. Name all three.

LINKING QUESTION

The ideal gas is the working substance for heat engines in B.4, and $U = \frac{3}{2}nRT$ is exactly the ΔU in the first law there.

IDEA FOR YOUR IA / EXPLORATION

Test Boyle's law with a syringe and pressure sensor, or the pressure law with a flask in a water bath. Plot P against 1/V (or P against θ) and extrapolate to estimate absolute zero.

Test yourself — cover the answers first

QUICK CHECK

1. What is 1 Pa in base units?
2. State Charles' law.
3. Why must T be in kelvin in PV = nRT?
4. Name two assumptions of the kinetic model.
5. If the absolute temperature of a gas doubles, what happens to v_{rms}?
6. When does a real gas behave least like an ideal gas?

EXAM-STYLE

7. A 0.020 m^3 cylinder holds oxygen (M = 0.032 kg mol^{-1}) at 1.5 × 10^7 Pa and 290 K. Find the mass of oxygen.
8. Air has density 1.2 kg m^{-3} at 1.0 × 10^5 Pa. Estimate the rms speed of its molecules.

Show answers

1. kg m^{-1} s^{-2}
2. at constant pressure, $V \propto T$ (in kelvin) for a fixed mass of gas
3. the equation needs $T \propto$ average kinetic energy, which is zero at 0 K, not at 0 °C
4. elastic collisions; negligible molecular volume (also: no intermolecular forces, random motion)
5. it increases by a factor of √2
6. at high pressure and low temperature
7. $n = PV/RT = 1.5 \times 10^7 \times 0.020/(8.31 \times 290) = 124.5$ mol
   $m = nM = 124.5 \times 0.032$
   = **4.0 kg**
8. $\langle v^2 \rangle = 3P/\rho = 3 \times 1.0 \times 10^5/1.2 = 2.5 \times 10^5$
   v_{rms} = **500 m s^{-1}**

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