LECTURE L2-10

Gas Laws

PV = nRT: Why Your Tires Bulge in Summer and How to Calculate It

Duration: 75 minutes

Advanced lecture script — pre-requisite: Level 1

HOOK (3 minutes)

Teacher holds up (or shows photos of):

A car tire with a pressure gauge

A balloon (inflated)

An aerosol can with "do not incinerate" warning

A soda bottle (fizzing when opened)

Teacher says: "Your tire pressure light comes on in winter — but you didn't lose air. The air just contracted because it's cold. In summer, your tires bulge — the air expanded.

A balloon shrinks in a freezer and expands in warm air. An aerosol can explodes if heated.

- Today's question: How do pressure, volume, and temperature relate — and how do you calculate the changes? *

By the end of this session, you will be able to:

Use the Ideal Gas Law (PV = nRT) to find pressure, volume, moles, or temperature

Apply Combined Gas Law ($P_1V_1/T_1 = P_2V_2/T_2$) for changes in conditions

Use Dalton's Law of Partial Pressures for gas mixtures

Calculate gas density and molar mass from the Ideal Gas Law"

SEGMENT 1: Review from Level 1 and Previous Level 2 Lectures (5 minutes)

Teacher says: "Before we go deeper, let's recall what you already know."

Review from Level 1 (Gas Laws - Basic)

Level 1 Concept	Definition	Household Example
Pressure (P)	Force gas exerts on container walls	PSI in car tires
Volume (V)	Space gas takes up	Balloon size
Temperature (T)	How hot or cold gas is	Hot car vs. freezer
Charles's Law	T ↑ → V ↑ (P constant)	Hot air balloon, chip bag on plane
Boyle's Law	P ↑ → V ↓ (T constant)	Syringe, scuba diving
Gay-Lussac's Law	T ↑ → P ↑ (V constant)	Aerosol can in hot car

Review from Level 2-5 (Moles)

Level 2-5 Concept	Formula	Example
Moles (n)	Mass ÷ Molar mass	10 g H_2O ÷ 18 g/mol = 0.556 mol

Review from Level 2-8 (Temperature)

Level 2-8 Concept	Formula	Example
Kelvin (K)	K = °C + 273.15	25°C = 298 K

Quick check (show of hands / chat): "What happens to tire pressure in winter?" (Decreases — cold air contracts) "What happens to a balloon in a freezer?" (Shrinks — volume decreases) "What is 0°C in Kelvin?" (273 K)

Teacher: "Good. Now let's learn the Ideal Gas Law — the big equation that combines all three gas laws into one."

SEGMENT 2: The Ideal Gas Law — PV = nRT (12 minutes)

Teacher says: "The Ideal Gas Law combines Boyle's, Charles's, and Gay-Lussac's laws into one equation."

PV = nRT

Where:

P = pressure (usually in atmospheres, atm)

V = volume (in liters, L)

n = number of moles (mol)

R = ideal gas constant = 0.0821 L·atm/(mol·K)

T = temperature (in Kelvin, K)

Why Kelvin? *"Celsius can be negative. Negative temperatures would give negative volumes or pressures — which don't make sense. Kelvin starts at absolute zero (-273°C), where particles stop moving."*

The Gas Constant R:

R = 0.0821 L·atm/(mol·K)

Teacher: "This value of R works when pressure is in atmospheres (atm) and volume is in liters (L) ."

Pressure Unit Conversions (Reference)

Unit	Abbreviation	Conversion to atm
Atmosphere	atm	1 atm

Unit	Abbreviation	Conversion to atm
Pounds per square inch	psi	14.7 psi = 1 atm
Millimeters of mercury	mmHg	760 mmHg = 1 atm
Kilopascal	kPa	101.3 kPa = 1 atm

Worked Example 1: Finding Moles (n)

Problem: How many moles of gas are in a 2.00 L container at 25.0°C and 1.00 atm pressure?

Step 1: Convert temperature to Kelvin

Step	Calculation
T(K) = 25.0 + 273.15	= 298.15 K

Step 2: Rearrange PV = nRT to solve for n

Step	Calculation
n = PV ÷ (RT)	

Step 3: Plug in numbers

Step	Calculation
n = (1.00 atm × 2.00 L) ÷ (0.0821 L·atm/(mol·K) × 298.15 K)	
n = 2.00 ÷ (0.0821 × 298.15)	
n = 2.00 ÷ 24.47	= 0.0817 mol

Answer: 0.0817 moles of gas.

Worked Example 2: Finding Pressure (P)

Problem: A 5.00 L container holds 2.00 moles of gas at 300 K. What is the pressure in atm?

Step 1: Rearrange PV = nRT to solve for P

Step	Calculation
P = nRT ÷ V	

Step 2: Plug in numbers

Step	Calculation
P = (2.00 mol × 0.0821 L·atm/(mol·K) × 300 K) ÷ 5.00 L	
P = (2.00 × 0.0821 × 300) ÷ 5.00	

Step	Calculation
P = (49.26) ÷ 5.00	= 9.85 atm

Answer: 9.85 atm (about 9.85 × 14.7 = 145 psi — very high pressure).

Worked Example 3: Finding Volume (V)

Problem: What volume does 0.500 mol of gas occupy at 27.0°C and 2.00 atm?

Step 1: Convert temperature to Kelvin

Step	Calculation
T(K) = 27.0 + 273.15	= 300.15 K

Step 2: Rearrange PV = nRT to solve for V

Step	Calculation
V = nRT ÷ P	

Step 3: Plug in numbers

Step	Calculation
V = (0.500 mol × 0.0821 L·atm/(mol·K) × 300.15 K) ÷ 2.00 atm	
V = (0.500 × 0.0821 × 300.15) ÷ 2.00	
V = (12.32) ÷ 2.00	= 6.16 L

Answer: 6.16 liters.

Worked Example 4: Finding Temperature (T)

Problem: A 10.0 L container holds 2.00 moles of gas at 5.00 atm. What is the temperature in °C?

Step 1: Rearrange PV = nRT to solve for T

Step	Calculation
T = PV ÷ (nR)	

Step 2: Plug in numbers

Step	Calculation
T = (5.00 atm × 10.0 L) ÷ (2.00 mol × 0.0821 L·atm/(mol·K))	
T = 50.0 ÷ (0.1642)	= 304.5 K

Step 3: Convert to Celsius

Step	Calculation
°C = K - 273.15	= 304.5 - 273.15 = 31.4°C

Answer: 31.4°C.

Partner talk (1 minute): "Tell your partner: If you double the number of moles (n) in a container at constant T and V, what happens to pressure? (P doubles — because PV = nRT, so P ∝ n)."

SEGMENT 3: The Combined Gas Law — $P_1V_1/T_1 = P_2V_2/T_2$ (12 minutes)

Teacher says: "The Combined Gas Law is useful when gas conditions change — but the number of moles (n) stays constant."

Combined Gas Law: $P_1V_1/T_1 = P_2V_2/T_2$

Use when n is constant.

What each law looks like from the Combined Gas Law:

Law	What is constant	Formula
Boyle's Law	T constant	$P_1V_1 = P_2V_2$
Charles's Law	P constant	$V_1/T_1 = V_2/T_2$
Gay-Lussac's Law	V constant	$P_1/T_1 = P_2/T_2$

Worked Example 1: Tire Pressure Change (Gay-Lussac's Law)

Problem: Your car tire pressure is 32.0 psi at 20.0°C. After driving, the tire temperature rises to 50.0°C. What is the new pressure? (Assume volume constant.)

Step 1: Identify knowns

Variable	Initial	Final
P	32.0 psi	?
T	20.0°C = 293 K	50.0°C = 323 K
V	constant	constant

Step 2: Use $P_1/T_1 = P_2/T_2$ (since V constant)

Step	Calculation
$32.0 / 293 = P_2 / 323$	
$P_2 = 32.0 \times (323 / 293)$	

Step	Calculation
$P_2 = 32.0 \times 1.102$	= 35.3 psi

Answer: 35.3 psi — about a 10% increase. (That's why you check tire pressure when tires are COLD.)

Worked Example 2: Balloon in Freezer (Charles's Law)

Problem: A balloon has a volume of 2.50 L at 25.0°C. You put it in a freezer at -5.0°C. What is the new volume? (Assume pressure constant.)

Step 1: Identify knowns

Variable	Initial	Final
V	2.50 L	?
T	25.0°C = 298 K	-5.0°C = 268 K
P	constant	constant

Step 2: Use $V_1/T_1 = V_2/T_2$ (since P constant)

Step	Calculation
$2.50 / 298 = V_2 / 268$	
$V_2 = 2.50 \times (268 / 298)$	
$V_2 = 2.50 \times 0.899$	= 2.25 L

Answer: 2.25 L (balloon shrinks).

Worked Example 3: Syringe (Boyle's Law)

Problem: A syringe contains 10.0 mL of air at 1.00 atm. You push the plunger to compress the air to 5.00 mL. What is the new pressure? (Assume temperature constant.)

Step 1: Identify knowns

Variable	Initial	Final
P	1.00 atm	?
V	10.0 mL	5.00 mL
T	constant	constant

Step 2: Use $P_1V_1 = P_2V_2$ (since T constant)

Step	Calculation
$1.00 \times 10.0 = P_2 \times 5.00$	
$P_2 = (1.00 \times 10.0) \div 5.00$	= 2.00 atm

Answer: 2.00 atm (pressure doubles when volume halves).

Worked Example 4: Full Combined Gas Law (Everything Changes)

Problem: A sample of gas has V_1 = 2.00 L, P_1 = 1.00 atm, T_1 = 300 K. It is changed to P_2 = 2.00 atm, T_2 = 450 K. What is V_2?

Step 1: Use $P_1V_1/T_1 = P_2V_2/T_2$

Step	Calculation
$(1.00 \times 2.00) / 300 = (2.00 \times V_2) / 450$	
$2.00 / 300 = (2.00 \times V_2) / 450$	
$0.006667 = (2.00 \times V_2) / 450$	
$0.006667 \times 450 = 2.00 \times V_2$	
$3.00 = 2.00 \times V_2$	
$V_2 = 3.00 \div 2.00$	= 1.50 L

Answer: 1.50 L.

Partner talk (1 minute): *"Tell your partner: A gas at 1.00 atm and 300 K is heated to 600 K at constant volume. What happens to pressure? (P doubles to 2.00 atm — $P_1/T_1 = P_2/T_2$)."*

SEGMENT 4: Dalton's Law of Partial Pressures — Gas Mixtures (10 minutes)

Teacher says: "Air is a mixture of gases: about 78% N_2, 21% O_2, 1% Ar, and trace CO_2. Dalton's Law says the total pressure is the sum of the partial pressures of each gas."

Dalton's Law: $P_{total} = P_1 + P_2 + P_3 + ...$

Partial pressure: The pressure a gas would exert if it alone occupied the container.

Mole Fraction (χ)

Mole fraction (χ) = moles of gas A ÷ total moles of all gases

$P_A = \chi_A \times P_{total}$

Worked Example 1: Partial Pressure of Oxygen in Air

Problem: Air is 21% O_2 by volume (which is the same as 21% by moles). At sea level, P_{total} = 1.00 atm. What is the partial pressure of oxygen (P_{O_2})?

Step	Calculation
$\chi\_O_2 = 0.21$	
$P\_O_2 = 0.21 \times 1.00$ atm	= 0.21 atm

Answer: 0.21 atm. That's why at high altitudes (lower P_total), there's less oxygen available.

Worked Example 2: Collecting Gas Over Water (Common Lab Setup)

Problem: A student collects oxygen gas over water. The total pressure in the collection bottle is 1.00 atm (760 mmHg) at 25°C. The vapor pressure of water at 25°C is 24 mmHg. What is the partial pressure of oxygen?

Step	Calculation
$P\_total = P\_O_2 + P\_water$	
760 mmHg = P_O_2 + 24 mmHg	
P_O_2 = 760 - 24	= 736 mmHg (0.968 atm)

Answer: P_O_2 = 736 mmHg.

Worked Example 3: Scuba Diving (Partial Pressures at Depth)

Problem: At 30 meters underwater, the pressure is about 4.0 atm. Air is still 21% O_2. What is the partial pressure of oxygen at this depth? Why is this dangerous?

Step	Calculation
$P\_O_2 = 0.21 \times 4.0$ atm	= 0.84 atm

Teacher: "0.84 atm O_2 is safe. But below 50 meters, P_O_2 can exceed 1.4 atm, causing oxygen toxicity (seizures). That's why deep divers use special gas mixtures (like Trimix) with lower oxygen percentages."

Partner talk (1 minute): *"Tell your partner: At a depth where total pressure is 3.0 atm, what is P_O_2 in normal air? ($0.21 \times 3.0 = 0.63$ atm — safe)."*

SEGMENT 5: Gas Density and Molar Mass from Ideal Gas Law (10 minutes)

Teacher says: "You can rearrange the Ideal Gas Law to find density (ρ = m/V) or molar mass (M)."

Density (ρ) = (P × M) / (R × T)

Where M = molar mass in g/mol

Molar mass (M) = (m × R × T) / (P × V)

Where m = mass in grams

Worked Example 1: Density of Air

Problem: Calculate the density of air at 1.00 atm and 25°C. Assume average molar mass of air is 29.0 g/mol.

Step 1: Use ρ = (P × M) / (R × T)

Step	Calculation
T = 25°C + 273 = 298 K	
ρ = (1.00 atm × 29.0 g/mol) ÷ (0.0821 L·atm/(mol·K) × 298 K)	
ρ = 29.0 ÷ (0.0821 × 298)	
ρ = 29.0 ÷ 24.47	= 1.19 g/L

Answer: Air density = 1.19 g/L (water is 1000 g/L — air is much less dense, which is why we don't float away).

Worked Example 2: Density of Helium (Why Balloons Float)

Problem: Helium has M = 4.00 g/mol. Calculate its density at 1.00 atm and 25°C.

Step	Calculation
ρ = (1.00 × 4.00) ÷ (0.0821 × 298)	
ρ = 4.00 ÷ 24.47	= 0.163 g/L

Comparison:

Gas	Density (g/L)	Floats in Air?
Helium	0.163	Yes (less dense than air)
Air	1.19	No (reference)
Propane	~1.80	No (sinks — dangerous leak collects at floor)
CO_2	~1.84	No (sinks — dry ice fog stays low)

Safety note: Propane and CO_2 are heavier than air. Leaks collect at floor level — that's why propane detectors are placed low.

Worked Example 3: Finding Molar Mass of an Unknown Gas

Problem: A 0.500 g sample of an unknown gas occupies 0.250 L at 1.00 atm and 300 K. What is its molar mass?

Step 1: Use M = (m × R × T) / (P × V)

Step	Calculation
M = (0.500 g × 0.0821 L·atm/(mol·K) × 300 K) ÷ (1.00 atm × 0.250 L)	
M = (0.500 × 0.0821 × 300) ÷ 0.250	
M = (12.315) ÷ 0.250	= 49.3 g/mol

Answer: Molar mass ≈ 49.3 g/mol. Could be carbon dioxide (44) or a refrigerant gas.

Partner talk (1 minute): *"Tell your partner: Which is more dense at the same T and P — H_2 (M=2) or O_2 (M=32)? (O_2 — higher molar mass = higher density)."*

SEGMENT 6: Putting It All Together — Complete Gas Law Problem (8 minutes)

Teacher says: "Let's do one complete problem that uses everything from today's lecture."

Problem: A 5.00 L balloon contains helium at 25.0°C and 1.00 atm. The balloon rises to an altitude where the pressure is 0.500 atm and the temperature is -20.0°C.

How many moles of helium are in the balloon (using PV = nRT)?

What is the new volume of the balloon at that altitude (using $P_1V_1/T_1 = P_2V_2/T_2$)?

What is the density of helium at that altitude (using ρ = PM/RT)?

Part 1: Moles of helium at initial conditions

Step	Calculation
T_1 = 25.0 + 273 = 298 K	
n = PV / RT = (1.00 atm × 5.00 L) ÷ (0.0821 × 298)	
n = 5.00 ÷ 24.47	= 0.204 mol

Answer Part 1: 0.204 mol He.

Part 2: New volume at altitude

Step	Calculation
T_2 = -20.0 + 273 = 253 K	
$P_1V_1/T_1 = P_2V_2/T_2$	

Step	Calculation
$(1.00 \times 5.00) / 298 = (0.500 \times V_2) / 253$	
$5.00 / 298 = (0.500 \times V_2) / 253$	
$0.01678 = (0.500 \times V_2) / 253$	
$0.01678 \times 253 = 0.500 \times V_2$	
$4.245 = 0.500 \times V_2$	
$V_2 = 4.245 \div 0.500$	= 8.49 L

Answer Part 2: 8.49 L (balloon expands).

Part 3: Density of helium at altitude

Step	Calculation
M_He = 4.00 g/mol	
$\rho = (P \times M) \div (R \times T) = (0.500 \times 4.00) \div (0.0821 \times 253)$	
$\rho = 2.00 \div (20.77)$	= 0.0963 g/L

Answer Part 3: 0.0963 g/L (much lower than at sea level — 0.163 g/L).

CLOSING — The 60-Second Challenge (5 minutes)

Teacher says: *"Pair up. Person A: 60 seconds — write the Ideal Gas Law and explain what each variable means. Person B: 60 seconds — a gas at 1.00 atm and 300 K is compressed to 0.500 L at constant T. If the initial volume was 1.00 L, what is the new pressure? (Use Boyle's Law: $P_1V_1 = P_2V_2 \rightarrow 1.00 \times 1.00 = P_2 \times 0.500 \rightarrow P_2 = 2.00$ atm)."*

Final takeaway table (show on screen / read aloud):

You learned...	Household Example
Ideal Gas Law: $PV = nRT$	Find moles in a balloon, pressure in a tire
Combined Gas Law: $P_1V_1/T_1 = P_2V_2/T_2$	Tire pressure change with temperature
Boyle's Law (T constant): $P_1V_1 = P_2V_2$	Syringe, scuba diving (the bends)
Charles's Law (P constant): $V_1/T_1 = V_2/T_2$	Balloon in freezer
Gay-Lussac's Law (V constant): $P_1/T_1 = P_2/T_2$	Aerosol can in hot car
Dalton's Law: P_total = sum of partial pressures	Air (21% O_2 → P_{O_2} = 0.21 atm at sea level)
Gas density: $\rho = PM/RT$	He (0.163 g/L) floats in air (1.19 g/L)
Propane is denser than air	Leaks collect at floor — dangerous

Final line (preview of L2-11): *"Next session: Nuclear Chemistry (Advanced) — half-life calculations, binding energy, and why smoke detectors contain americium-241. See you then."*

SUPPLEMENTARY MATERIALS FOR L2-10 (No Grade)

Resource	Household Connection	Description	How to Find It
PhET "Gas Properties" simulation	Interactive gas laws	Pump gas, change T, see P and V change	Search "PhET gas properties"
Ideal Gas Law calculator	Practice problems	Online calculator with different units	Search "ideal gas law calculator"
Dalton's Law interactive	Partial pressures	Mix gases, see total P	Search "Dalton's law simulation"

"This week, check the tire pressure on a car (yours or a family member's). Do it in the morning when tires are cold. Then check again after driving for 30 minutes. Is the pressure higher or lower? (Higher — driving heats the tires, increasing pressure). Next time, tell us your results."