Essential Question

Why don’t real gases always follow *PV* = *nRT*?

The Ideal Gas Model

- **Particles have no volume.** They're treated like tiny points.
- **There are no attractions between particles.** They never pull on each other.

Real particles are not points, and they do attract each other. But at everyday conditions these assumptions are close enough, so $PV = nRT$ works well for most gases.

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Real Gas Particles

- Real particles take up space.
- Real particles attract each other through intermolecular forces (IMFs).

Particle volume and attractions are the only two reasons real gases deviate from ideal behavior. Every explanation you write for this topic will use one or both of them.

Measuring Deviation

- **= 1:** ideal behavior
- **< 1:** measured pressure is lower than predicted. Attractions win.
- **> 1:** measured pressure is higher than predicted. Particle volume wins.

$$\frac{PV}{nRT}$$

Attractions Lower Pressure

- Neighbors pull a particle back as it nears the wall.
- Wall hits are softer.
- Measured P < ideal P, so PV/nRT < 1.

Pressure comes from collisions with the walls. When attractions pull a particle back, it hits the wall with less force, and softer collisions mean lower pressure.

When Do Attractions Matter?

- **Low T:** slow particles can’t pull away from each other quickly, so attractions have more time to act.
- **High P:** particles are squeezed close together, and attractions are much stronger at short distances.
- **Worst near condensation:** low T and high P are the conditions that turn a gas into a liquid.

A gas deviates the most when it’s close to condensing. That makes sense, because condensing happens when attractions take over.

Particle Volume Matters Too

- At extremely high *P*, particles fill much of the container.
- Less free space → more wall hits.
- Measured *P* > ideal *P*, so *PV*/*nRT* > 1.

At low pressure a gas is mostly empty space, so particle volume can be ignored.

PV/nRT vs. Pressure

- Moderate *P*: the curve dips below 1 (attractions).
- Extremely high *P*: the curve rises above 1 (particle volume).

A real gas at constant temperature starts near 1 at very low pressure.

Most Ideal vs. Least Ideal

- **Most ideal: high T + low P.** Particles are fast and far apart, so attractions barely matter and particle volume is tiny compared to the container.
- **Least ideal: low T + high P.** Particles are slow and close together, so attractions are strong and particle volume is a bigger part of the container.

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Comparing Different Gases

- Stronger IMFs → more deviation, because the particles attract each other more.
- Larger particles → more deviation, because they take up more space.

Exam tip: when you're asked which gas is least ideal, start by comparing the IMFs, and name the specific IMF in your answer.

Example: Ideal vs. Measured Pressure

Try it first!

A 1.00 mol sample of CO_2 is in a 0.500 L container at 300. K. Its measured pressure is 39.5 atm. Calculate the ideal gas pressure and explain the difference.

Example: Choosing Conditions

Try it first!

A sample of N_2 gas can be studied under any of these conditions. Under which conditions will N_2 behave most like an ideal gas? Justify your answer.

- **A:** 100 K, 50 atm
- **B:** 100 K, 1 atm
- **C:** 500 K, 50 atm
- **D:** 500 K, 1 atm

Example: Which Gas Is Least Ideal?

Try it first!

Under the same conditions, which gas deviates most from ideal behavior: Ne (20.18 g/mol), CH_4 (16.04 g/mol), or NH_3 (17.03 g/mol)? Justify your answer.

Example: Explaining a Graph

Try it first!

The graph shows *PV*/*nRT* vs. pressure for a real gas at constant temperature. Explain why *PV*/*nRT* is less than 1 at point A but greater than 1 at point B.

Diagram: Graph of PV/nRT versus pressure for a real gas. Point A is at the bottom of the dip below the dashed line at 1. Point B is at high pressure, above the line.

Big Ideas

- Real particles have volume and attract each other.
- Attractions → *P* lower than ideal.
- Particle volume → *P* higher than ideal.
- Gases are most ideal at high T and low P.

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