ANSWER KEY. Gauge readings vary a little from run to run; Gas law values are exact.

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Gas Laws and KMT Virtual Lab

Chemistry and AP Chemistry | ANSWER KEY

Open **activities.zornscience.org/gas-lab/** and follow each step. To type a value, click the number box, type, and press Enter. Show your work with units.

Objective

Use a particle model of a gas to find how **pressure**, **volume** and **temperature** are related, and how kelvin temperature is connected to particle **kinetic energy** and **speed**.

NGSS alignment: HS-PS3-2. Develop and use models to illustrate that energy at the macroscopic scale can be accounted for as a combination of energy associated with the motions of particles (objects) and energy associated with the relative positions of particles (objects).
Practices and concepts: Developing and using models; Analyzing and interpreting data; Using mathematics; Scale, proportion, and quantity; Cause and effect.
AP alignment: 3.4 Ideal Gas Law; 3.5 Kinetic Molecular Theory; 3.6 Deviation from Ideal Gas Law. Skills 5.C, 4.A and 6.E.

Meet the simulation

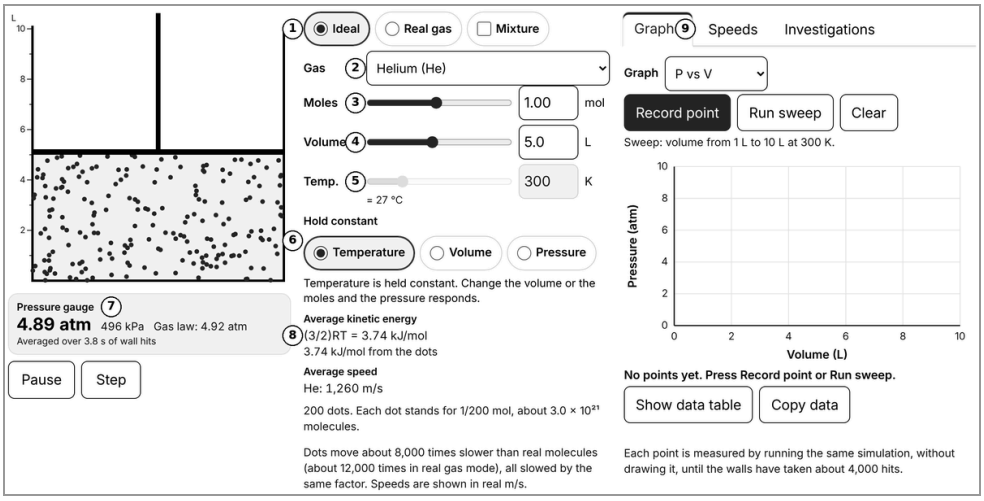


Figure 1. The Gas Laws and KMT Lab.

1. Model: Ideal, Real gas or Mixture.
2. Gas menu.
3. Moles (each dot is 1/200 mol).
4. Volume (moves the piston).
5. Temperature in K.
6. Hold constant. The held control turns gray.
7. Pressure gauge. "Gas law" is the PV = nRT value.
8. Average kinetic energy and speed.
9. Tabs, including Speeds.

Pre-lab

1. According to kinetic molecular theory, what causes the pressure of a gas? What happens to the pressure if the particles hit the walls more often?

Pressure is caused by gas particles colliding with the walls of the container (force per area). More hits each second means more force on the walls, so the pressure increases.

2. Use PV = nRT (R = 0.08206 L·atm/(mol·K)) to calculate the pressure of 2.00 mol of gas at 400 K in 8.00 L.

P = nRT / V = (2.00 mol)(0.08206 L·atm/(mol·K))(400 K) / (8.00 L) = <u>8.21 atm</u>

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Part A. Pressure and volume

1. Select **Ideal**, gas **Helium (He)**, **1.00** mol, **300** K. Hold constant: **Temperature**.
2. Type each volume below. Wait about 5 s for the gauge to settle, then record the gauge and the **Gas law** value.

Volume (L)	Gauge pressure (atm)	Gas law pressure (atm)	P × V (atm·L)
10.0	**about 2.5**	**2.46**	**24.6**
5.0	**about 4.9**	**4.92**	**24.6**
2.5	**about 9.8**	**9.85**	**24.6**

3. What happens to the pressure when the volume is cut in half? What stays the same in the last column? Name this gas law and write its equation.

The pressure doubles (4.92 → 9.85 atm). P × V stays the same (24.6 atm·L). This is Boyle's law: $P_1V_1 = P_2V_2$ (pressure and volume are inversely proportional).

4. Watch the dots as you change from 10.0 L to 2.5 L. The temperature (and particle speed) did not change. Explain with particles why the pressure went up.

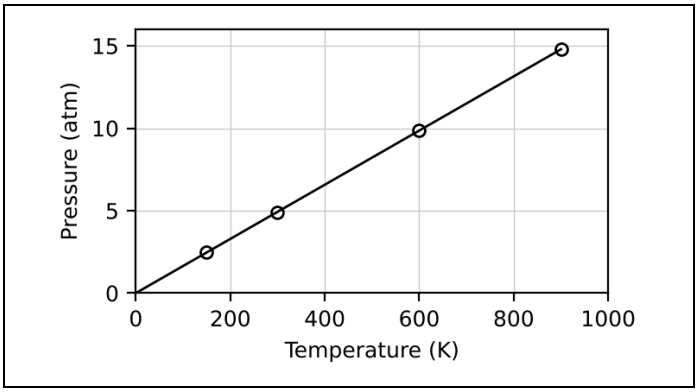
The same number of particles is packed into a smaller space at the same speed, so they hit the walls more often. More collisions with the walls each second means a greater pressure.

Part B. Temperature, energy and speed

1. Keep He, **1.00** mol. Type a volume of **5.0** L, then select Hold constant: **Volume**.
2. Type each temperature below and record the Gas law pressure, (3/2)RT and the average speed.

Temperature (K)	Temperature (°C)	Gas law pressure (atm)	Average kinetic energy (3/2)RT (kJ/mol)	Average speed of He (m/s)
150	**−123**	**2.46**	**1.87**	**891**
300	**27**	**4.92**	**3.74**	**1,260**
600	**327**	**9.85**	**7.48**	**1,781**
900	**627**	**14.8**	**11.2**	**2,182**

5. Graph pressure (y-axis) against kelvin temperature (x-axis). Label both axes with units and draw a line of best fit. Where does the line meet the x-axis?



Straight line through the origin (0 K, absolute zero). Slope = 14.8 atm / 900 K = 0.0164 atm/K. Gay-Lussac's law: $P_1/T_1 = P_2/T_2$.

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6. When the kelvin temperature doubled from 300 K to 600 K, by what factor did the average kinetic energy change? The average speed? Show the speed ratio with numbers.

Kinetic energy doubled (3.74 → 7.48 kJ/mol), because kelvin temperature is proportional to average kinetic energy. Speed did not double: 1,781 ÷ 1,260 = 1.41 (= √2), because KE = ½mv^2, so v changes with the square root of T.

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7. A student says, "Heating from 27 °C to 327 °C makes the particles 12 times more energetic." Use your table to correct the student.

The student used °C. In kelvins, 27 °C = 300 K and 327 °C = 600 K, so the average kinetic energy only doubles (3.74 → 7.48 kJ/mol). Kinetic energy is proportional to kelvin temperature, not °C.

Part C. Different gases at the same temperature

1. Keep **1.00** mol, **300** K, **5.0** L. Choose each gas below from the Gas menu and record the readouts.
2. Open the **Speeds** tab and select **He vs Ar** (Figure 2).

Gas	Molar mass (g/mol)	(3/2)RT (kJ/mol)	Gas law P (atm)	Average speed (m/s)
He	4.00	**3.74**	**4.92**	**1,260**
N_2	28.01	**3.74**	**4.92**	**476**
Ar	39.95	**3.74**	**4.92**	**399**
Xe	131.29	**3.74**	**4.92**	**220**

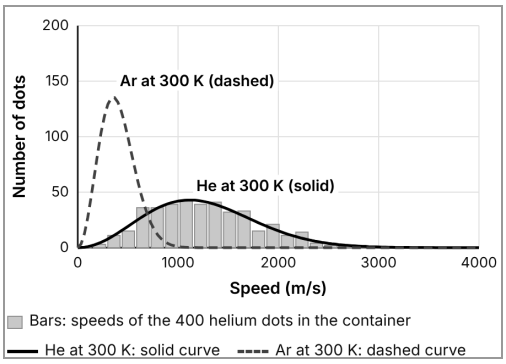


Figure 2. Speed curves for He and Ar at 300 K (Speeds tab).

8. Which readouts were the same for every gas? Which changed? Explain why heavier particles move more slowly at the same temperature.

The average kinetic energy (3.74 kJ/mol) and the pressure (4.92 atm) were the same. The average speed changed. All gases at the same temperature have the same average KE = ½mv^2, so a particle with a larger mass must have a smaller speed.

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9. In Figure 2, describe two ways the argon curve differs from the helium curve. Use particle mass to explain the difference.

The argon peak is farther left (most probable speed 353 m/s vs 1,116 m/s for He), and the curve is taller and narrower (speeds bunched closer together). Argon atoms are about 10 times heavier, so at the same average KE they move more slowly.

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10. (AP extension) Show that the ratio of average speeds of He and Xe equals $\sqrt{(M_{Xe}/M_{He})}$.

Speed ratio = 1,260 m/s ÷ 220 m/s = 5.73. √(131.29 g/mol ÷ 4.00 g/mol) = √32.8 = 5.73. They match, because $v_{avg} = \sqrt{(8RT/\pi M)}$.

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Part D. (AP extension) Real gases

1. Check that the temperature is **300** K. Select **Real gas**, choose **Carbon dioxide (CO_2)**, type **1.00** mol, and select Hold constant: **Temperature**.
2. Type each volume and record the van der Waals pressure, the ideal gas law pressure and PV/nRT. Then repeat with **Helium (He)**.

Gas	Volume (L)	Real pressure (atm)	Ideal gas law (atm)	PV/nRT
CO_2	2.50	**9.44**	**9.85**	**0.959**
CO_2	1.00	**22.1**	**24.6**	**0.899**
CO_2	0.50	**39.5**	**49.2**	**0.801**
He	2.50	**9.94**	**9.85**	**1.009**
He	0.50	**51.6**	**49.2**	**1.047**

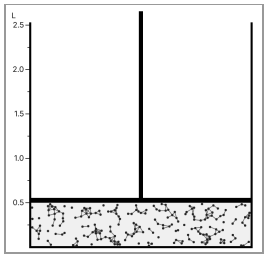


Figure 3. CO_2 at 0.50 L. Lines join dots close enough to attract.

11. (AP extension) For CO_2, how does PV/nRT change as the gas is squeezed? Is the real pressure higher or lower than ideal? Explain using attractions between particles.

PV/nRT drops further below 1 as the volume shrinks (0.959 → 0.801), so the real pressure is lower than ideal (39.5 vs 49.2 atm at 0.50 L). Close together, CO_2 molecules attract each other, so they hit the walls less often and less hard.

12. (AP extension) Helium's PV/nRT is above 1 at 0.50 L. What causes this, and why does helium behave so differently from CO_2?

The particles' own volume: at high pressure the molecules take up a real share of the container, so there is less free space and the pressure is higher than ideal. Helium is small with few electrons, so its dispersion attractions are very weak and the volume effect wins. CO_2 has stronger attractions, so the attraction effect wins.

Analysis

13. In Part A the gauge readings were close to, but not exactly equal to, the Gas law values. Explain why. Then give one way a real gas law lab would differ from the ideal simulation.

The gauge adds up the actual, random hits on the walls over a few seconds, so it varies a little around the average that PV = nRT predicts. Real lab (any one): real gas particles have volume and attract each other, so real gases are not perfectly ideal; squeezing a real gas warms it; a syringe can leak; pressure sensors and volume marks have measurement error.

Conclusion

14. Write a claim that answers the objective. Support it with evidence from Parts A, B and C (specific numbers), then explain your reasoning with kinetic molecular theory.

Claim: Pressure increases when volume decreases or temperature increases, and kelvin temperature is proportional to the particles' average kinetic energy.
Evidence: Halving V from 5.0 to 2.5 L doubled P from 4.92 to 9.85 atm. Doubling T from 300 to 600 K doubled P (4.92 to 9.85 atm) and doubled KE (3.74 to 7.48 kJ/mol). At 300 K every gas had KE 3.74 kJ/mol, but He moved at 1,260 m/s and Xe at 220 m/s.
Reasoning: Pressure comes from particles hitting the walls. A smaller volume means more frequent hits; faster particles at a higher temperature hit harder and more often. Same temperature means same average KE, so heavier particles move slower.

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