Zumdahl | Zumdahl | DeCoste

World of CHEMISTRY

Chapter 13

- **Gases**

Chapter 13 Overview

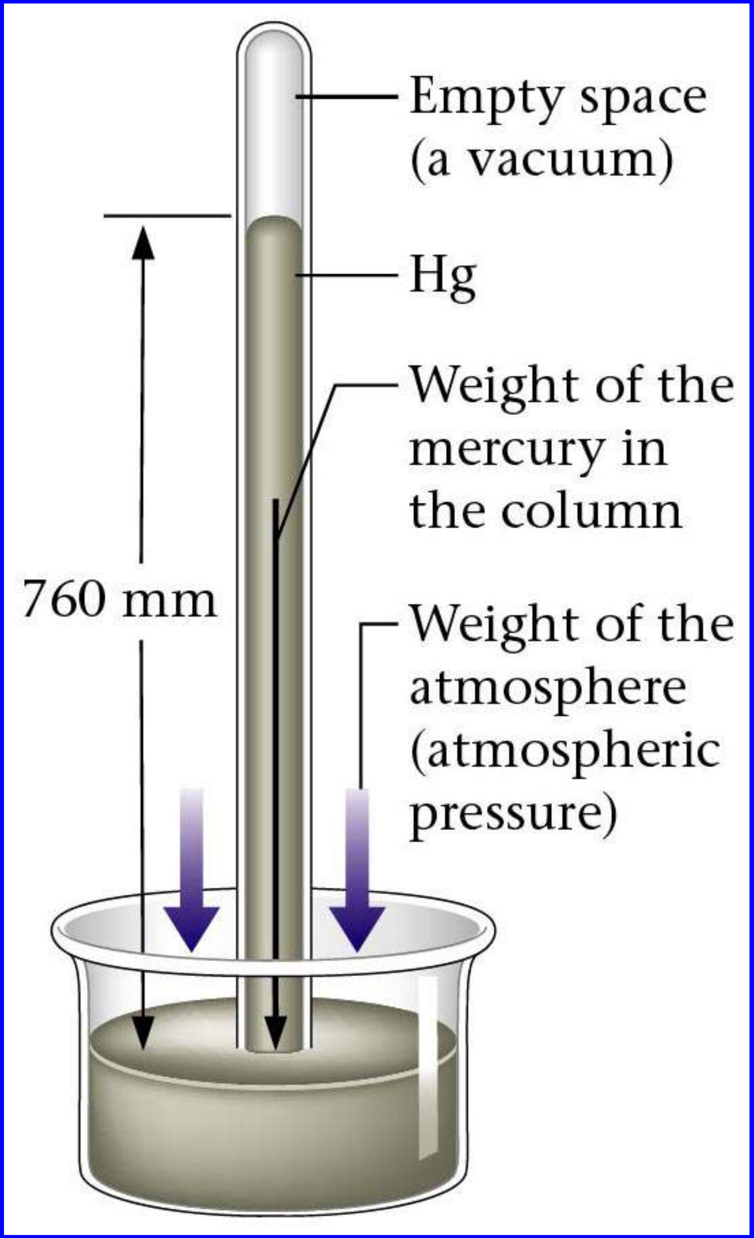
- Atmospheric Pressure/Barometers
- Units of pressure
- Boyle's Law
- Charles' Law/Absolute Zero
- Avogadro's Law
- Ideal Gas Law
- Partial/Total Pressure
- Relationship between laws and models
- Kinetic Molecular Theory
- Meaning of Temperature
- Properties of real gases
- Molar Volume of Ideal Gas/STP

Properties of Gases

- Uniformly fills any container
- Easily compressed
- Mixes completely with other gases
- Exerts pressure on surroundings
  – Blow up balloon, pushes on elastic sides keeping it firm
  – Atmosphere exerts pressure

Torricelli's Barometer

Figure 13.2: A glass tube is filled with mercury and inverted in a dish of mercury at sea level.



Barometer: device that measures atmospheric pressure

Atmospheric pressure results from the mass of the air being pulled toward the center of the earth by gravity (in other words, the weight of the air above us)

Units of Pressure

- torr = mm of Hg (named after Torricelli)
- 1 std atmosphere = 1 atm = 760 mm Hg
- SI unit for pressure = Pascal (Pa)
- 1 atm = 101,325 Pa = 101.325 kPa
- In engineering, use psi: 1 atm = 14.69 psi

Figure 13.3: Gas pressure = atmospheric pressure – *h*.

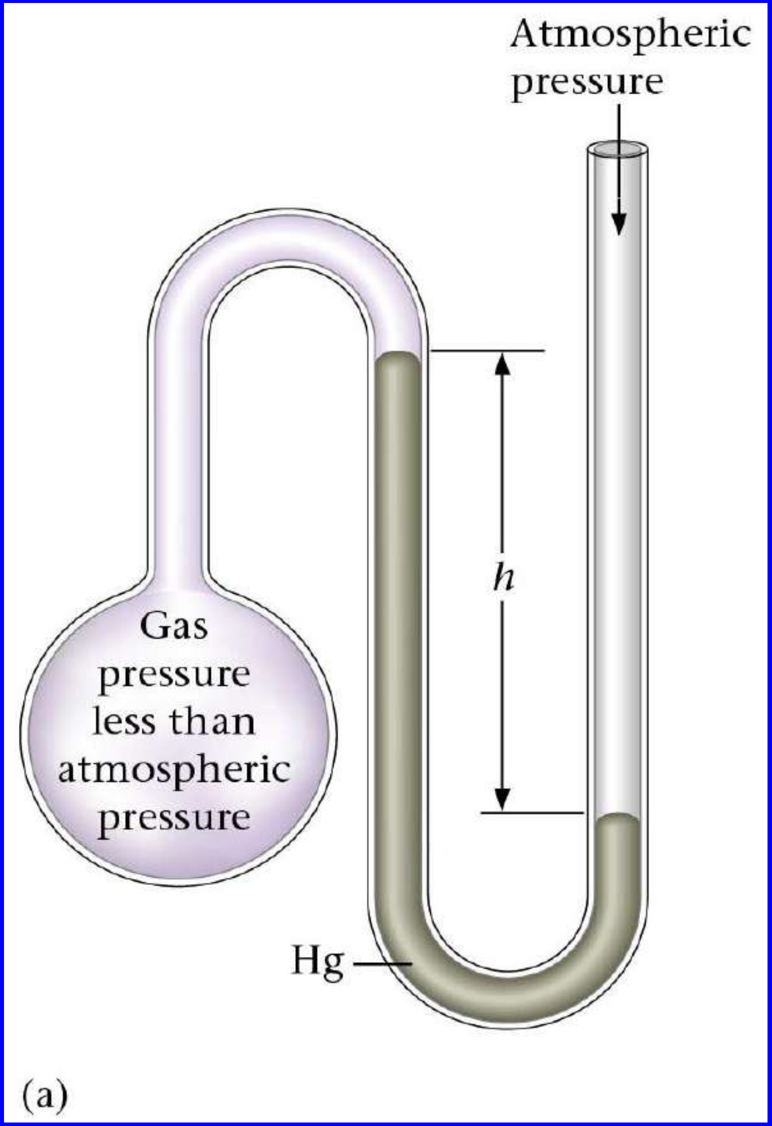
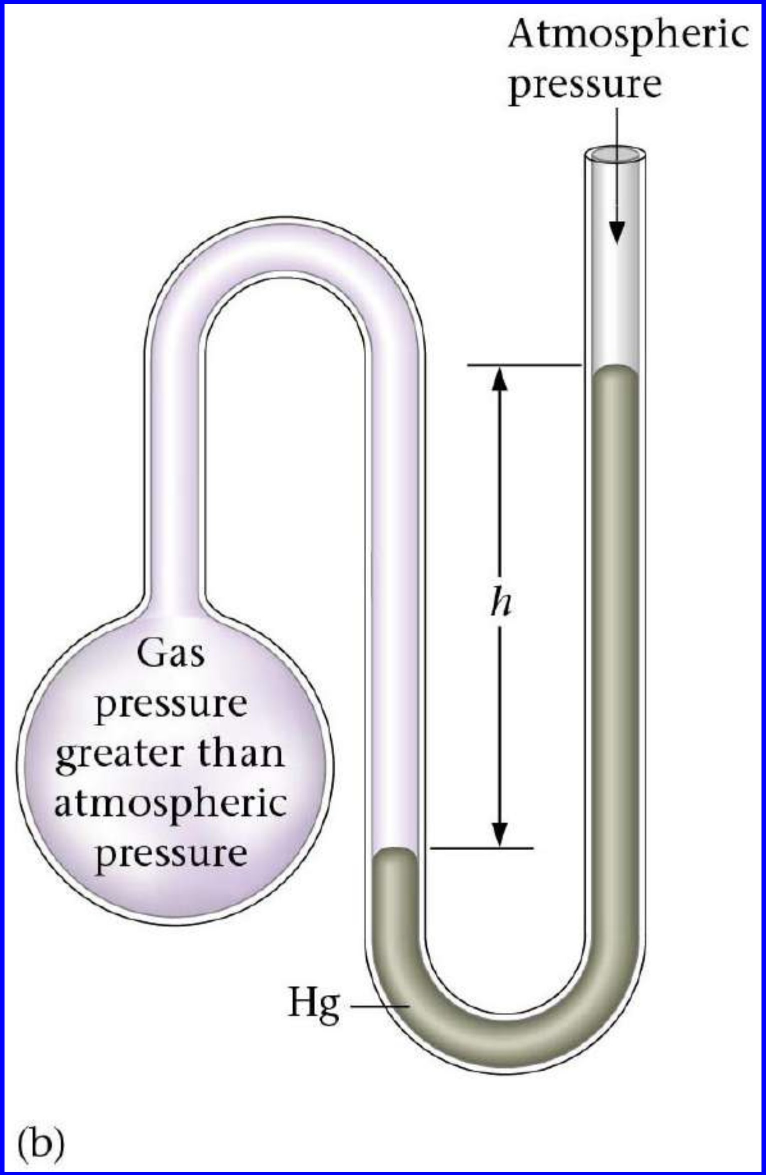


Figure 13.3: Gas pressure = atmospheric pressure + h.



Boyle's Law

- Boyle was first to perform careful experiments on gases
- Studied relationship between the pressure of a trapped gas and its volume
- As pressure increased, volume decreased
- Analyzing data on next slide, as pressure doubles, volume is cut in half

Table 13.1

TABLE 13.1

A Sample of Boyle's Observations (moles of gas and temperature both constant)

Experiment	Pressure (in. Hg)	Volume (in.^{3})	Pressure × Volume (in. Hg) × (in.^{3}) Actual	Rounded*
1	29.1	48.0	1396.8	1.40×10^3
2	35.3	40.0	1412.0	1.41×10^3
3	44.2	32.0	1414.4	1.41×10^3
4	58.2	24.0	1396.8	1.40×10^3
5	70.7	20.0	1414.0	1.41×10^3
6	87.2	16.0	1395.2	1.40×10^3
7	117.5	12.0	1410.0	1.41×10^3

*Three significant figures are allowed in the product because both of the numbers that are multiplied together have three significant figures.

Figure 13.4: A J-tube similar to the one used by Boyle.

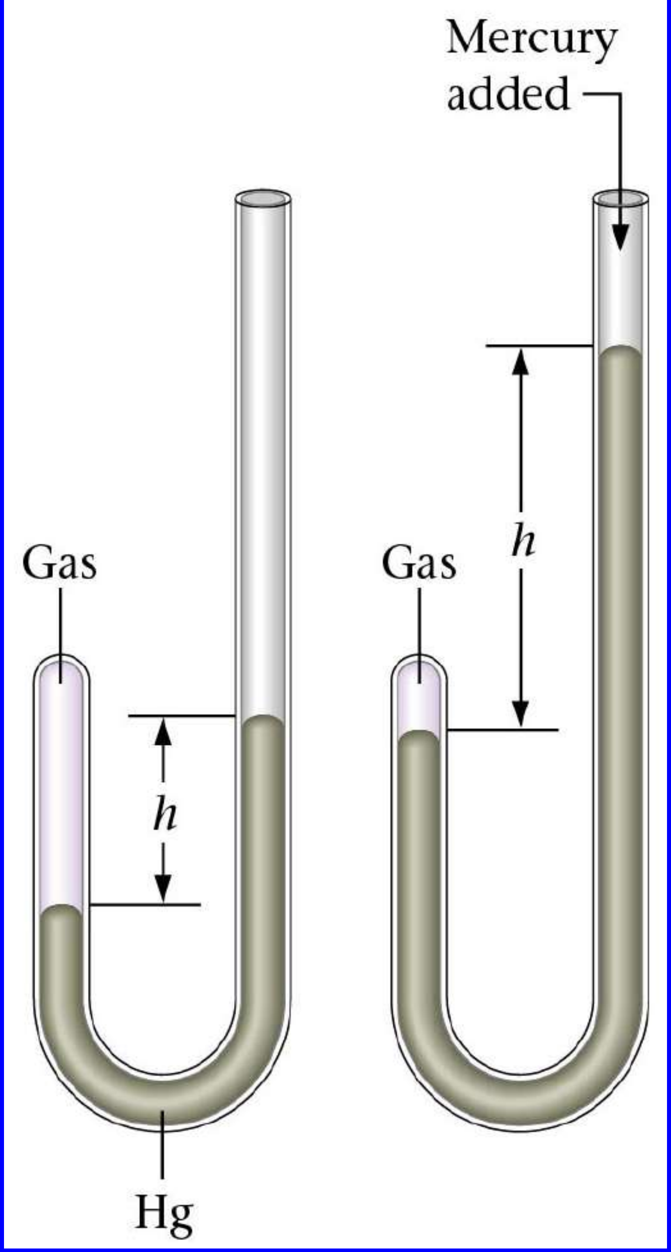


Figure 13.5: A plot of P versus V from Boyle's data.

PV = k

Pressure x volume = a constant

Pressure and volume are inversely proportional (when one increases, the other decreases)

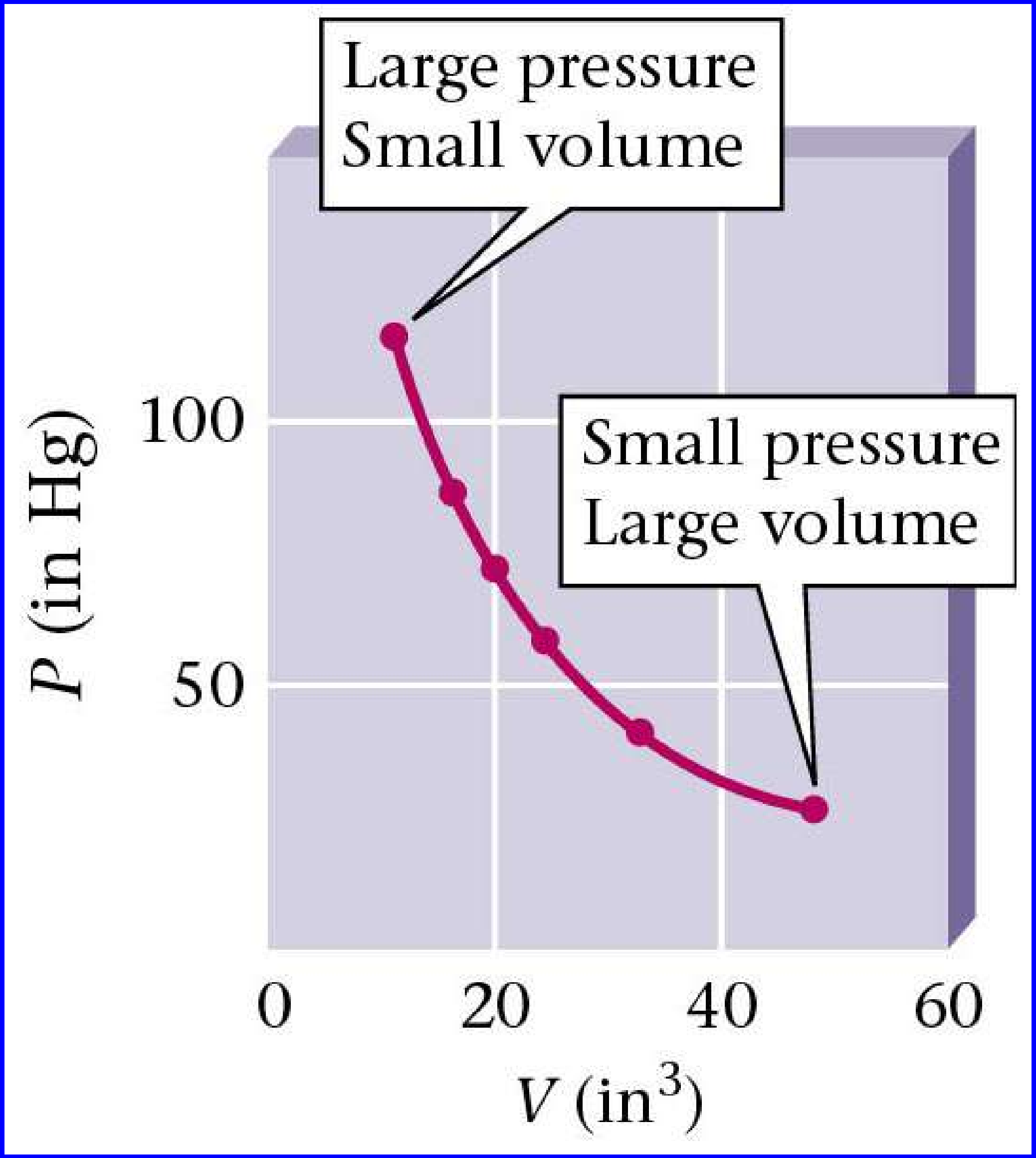
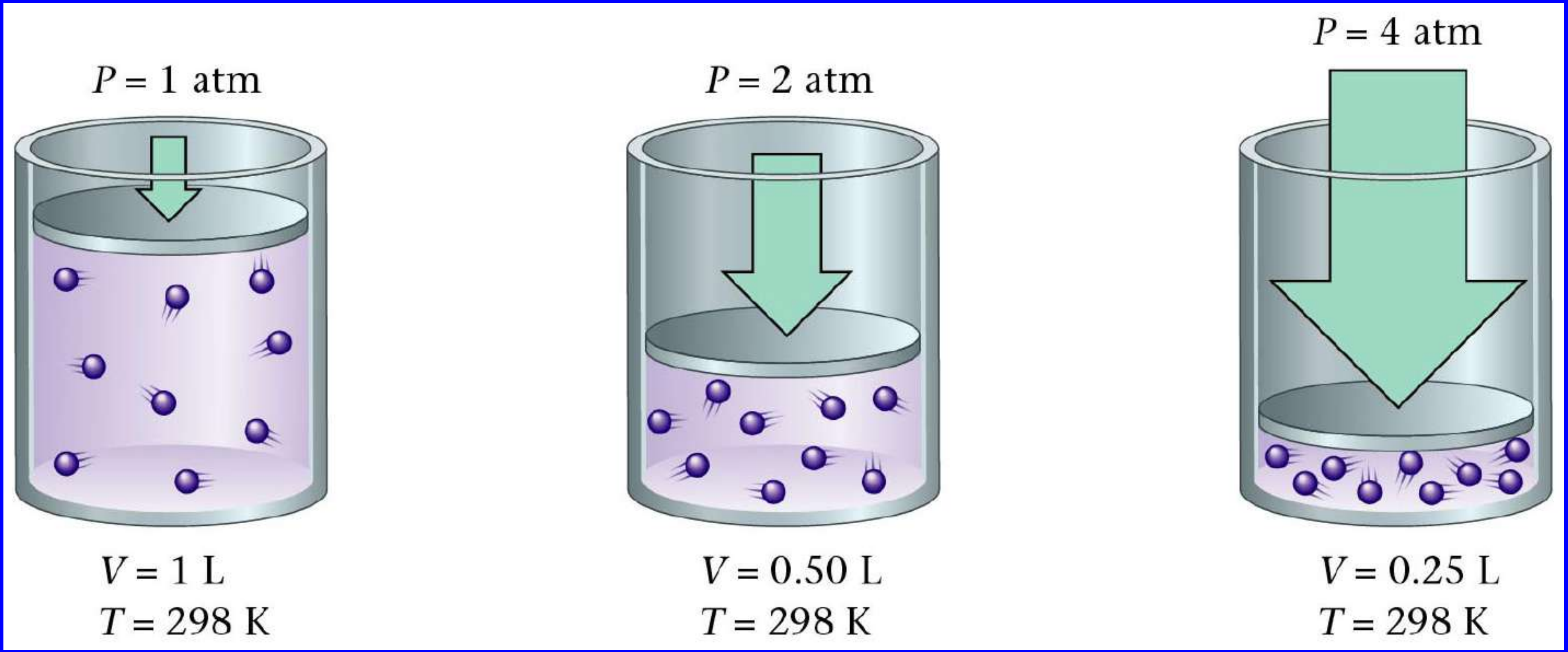


Figure 13.6: Illustration of Boyle's law.



Note: Temperature & amount of gas remain constant

Calculations using Boyle's Law

- If we have a gas at a certain volume and pressure (V_1 and P_1), we can find the new volume (V_2) if the pressure changes (P_2) or the new pressure if the volume changes

Since $V_1P_1 = k$ and $V_2P_2 = k$ then $V_1P_1 = V_2P_2$

See Examples 13.2 and 13.3

Charles' Law: Volume vs. Temperature

- Jacques Charles – first to fill balloon with hydrogen gas & made 1st solo balloon flight
- Showed volume of a given gas (at constant pressure) increases as the temperature increases
- Plot of volume of given gases versus temperature gives straight line = linear relationship (see next slide)

Figure 13.7: Plots of V (L) versus T (°C) for several gases.

All lines extrapolate back to -273°C = ***absolute zero***

Lower temperature would result in negative volume which is impossible

Cannot cool matter below this temperature

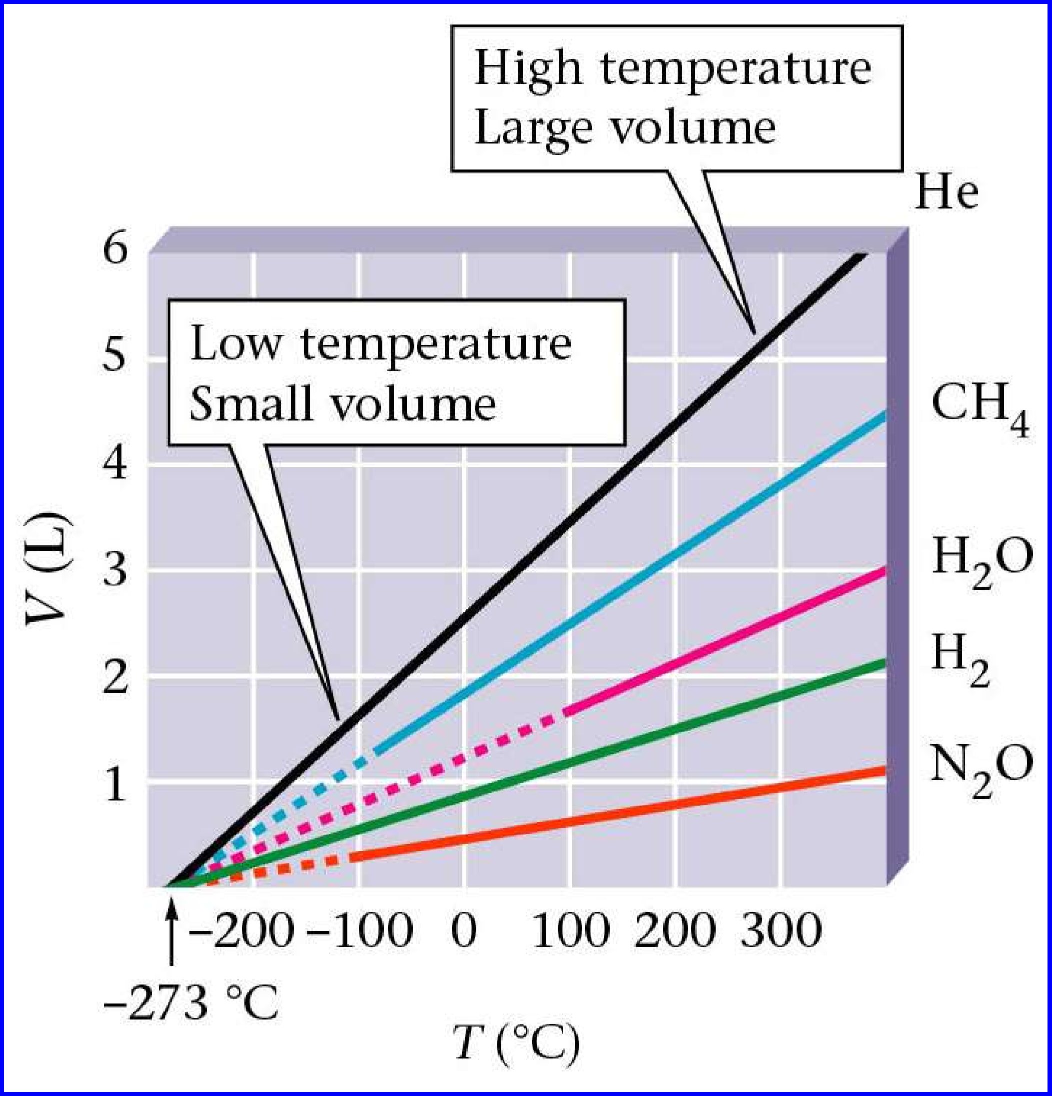


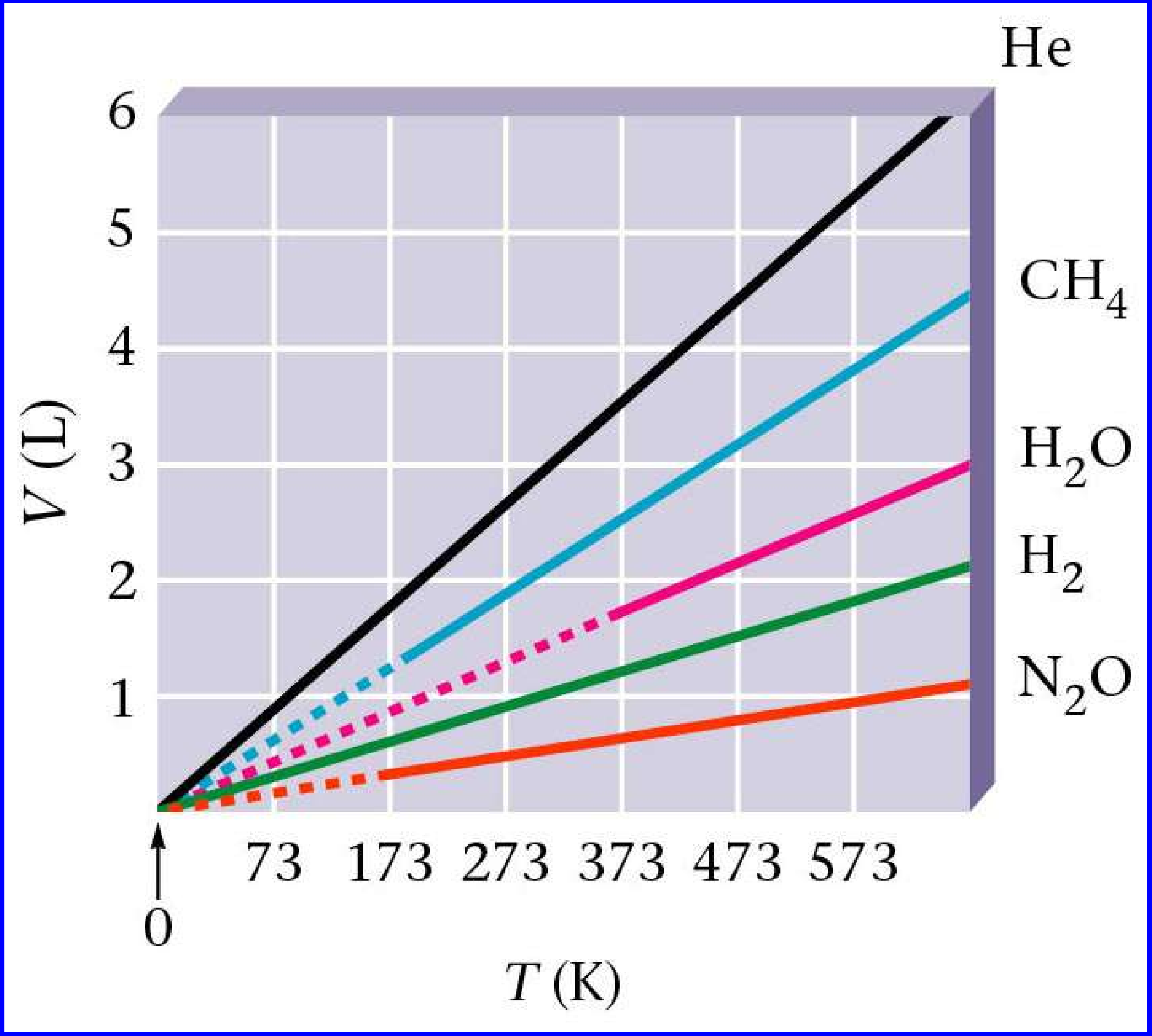
Figure 13.8: Plots of V versus T using the Kelvin scale for temperature.

Volume is directly proportional to temperature (if temp doubles, volume also doubles)

$V = bT$

b = constant

T = temp (K)



Calculations using Charles' Law

Since $V = bT$, $V/T = b$ (a constant)

So, if $V_1/T_1 = b$

and $V_2/T_2 = b$

then $V_1/T_1 = V_1/T_2$

See examples 13.4, 13.5, and 13.6

Avogadro's Law

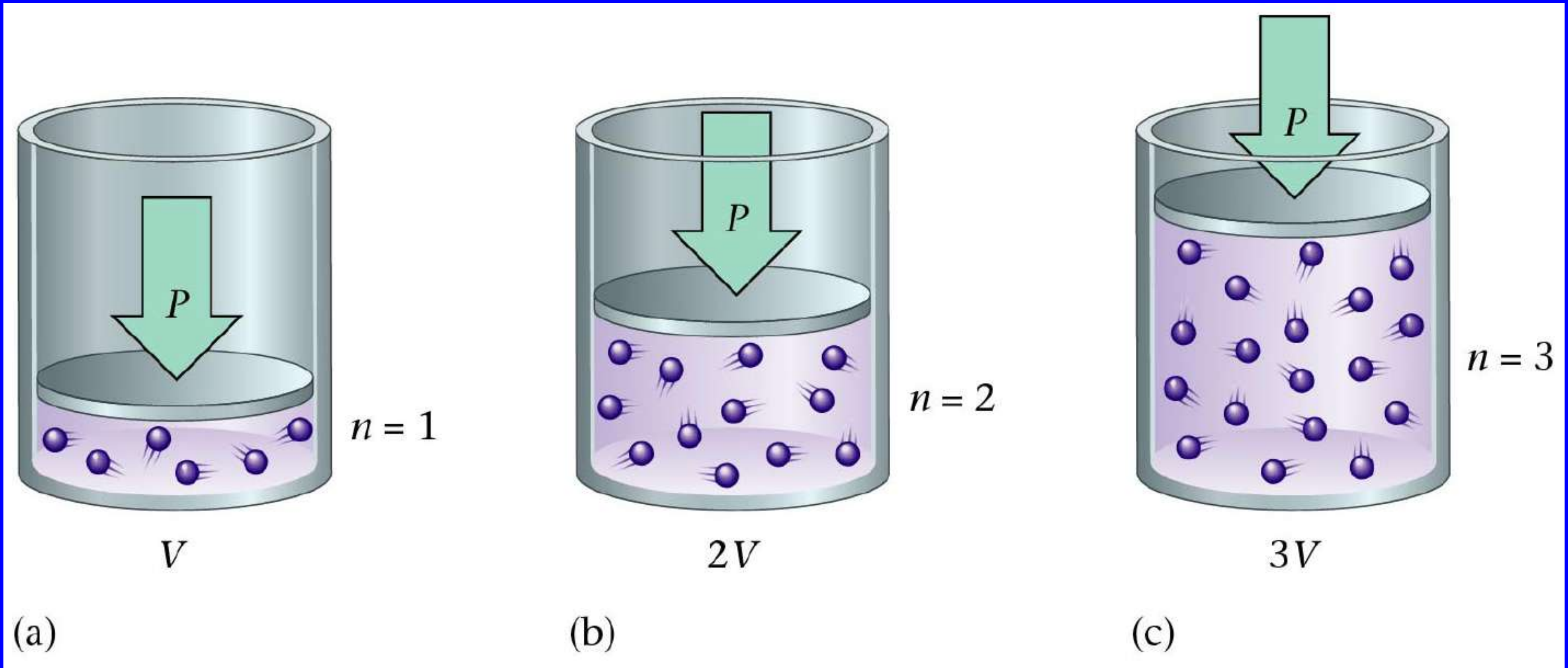
- The volume of gas is directly proportional to the number of moles if the temperature and pressure are held constant (if volume doubles, number of moles also doubles)

$V = an$ $\quad$ V = volume, a = constant, n = # moles

Rearranging, $V/n = a$

$V_1/n_1 = a$ & $V_2/n_2 = a$ so $V_1/n_1 = V_2/n_2$

Figure 13.9: The relationship between volume V and number of moles n.



The Ideal Gas Law

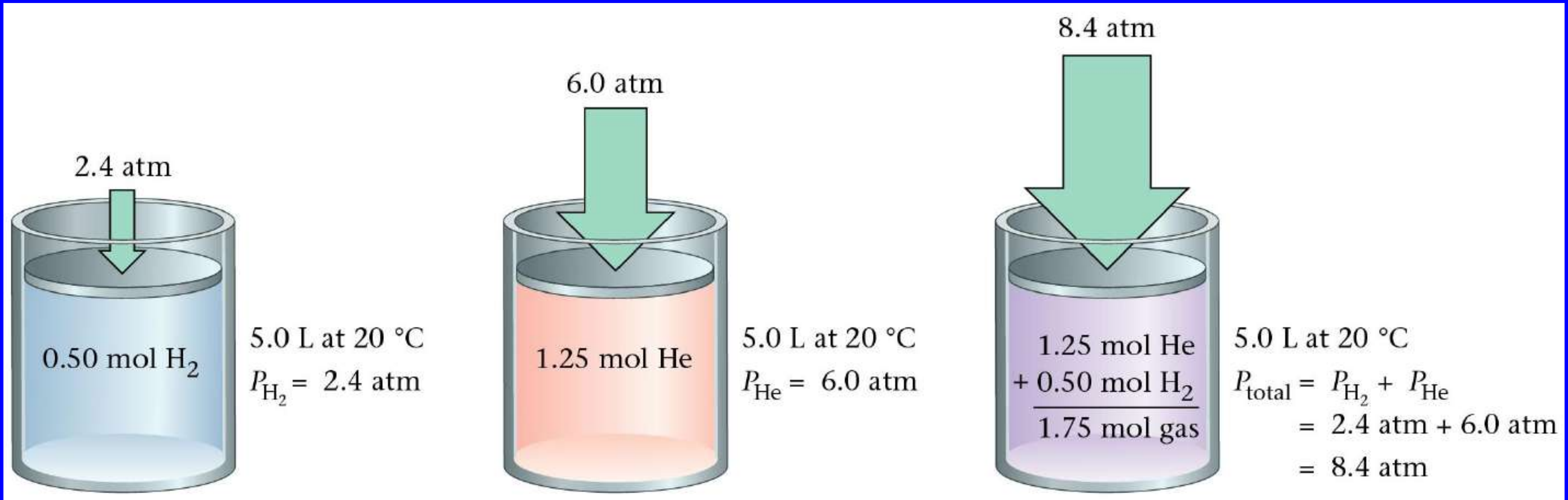
- By combining Boyle's, Charles', and Avogadro's Laws, we obtain the *Ideal Gas Law: $PV = nRT$*
- The constants in each of these laws are combined to form the universal gas constant, *R*
- $R = 0.08206$ L atm/K mol
- Only applies to ideal gases (P is approximately 1 atm or lower & temp is 0°C or higher)
- See examples 13.9, 13.10, and 13.11

Combined Gas Law

- By combining Boyle's and Charles' Laws, we get the combined gas law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Figure 13.10: When two gases are present, the total pressure is the sum of the partial pressures of the gases.



Dalton's Law of Partial Pressures

- Many important gases are mixtures of components (ex: air, scuba diver tanks)
- Summary of Dalton's observations: for a mixture of gases in a container, the total pressure is the sum of the partial pressures of the gases present
- Partial pressure: pressure that the gas would exert if it were alone in the container

$$P_{total} = P_1 + P_2 + P_3$$

$$P_{total} = n_{total}\ (RT/V)$$

n = sum of the number of moles of gases in mixture

Identity of gas particles is not important

Figure 13.11: The total pressure of a mixture of gases depends on the number of moles of gas particles present.

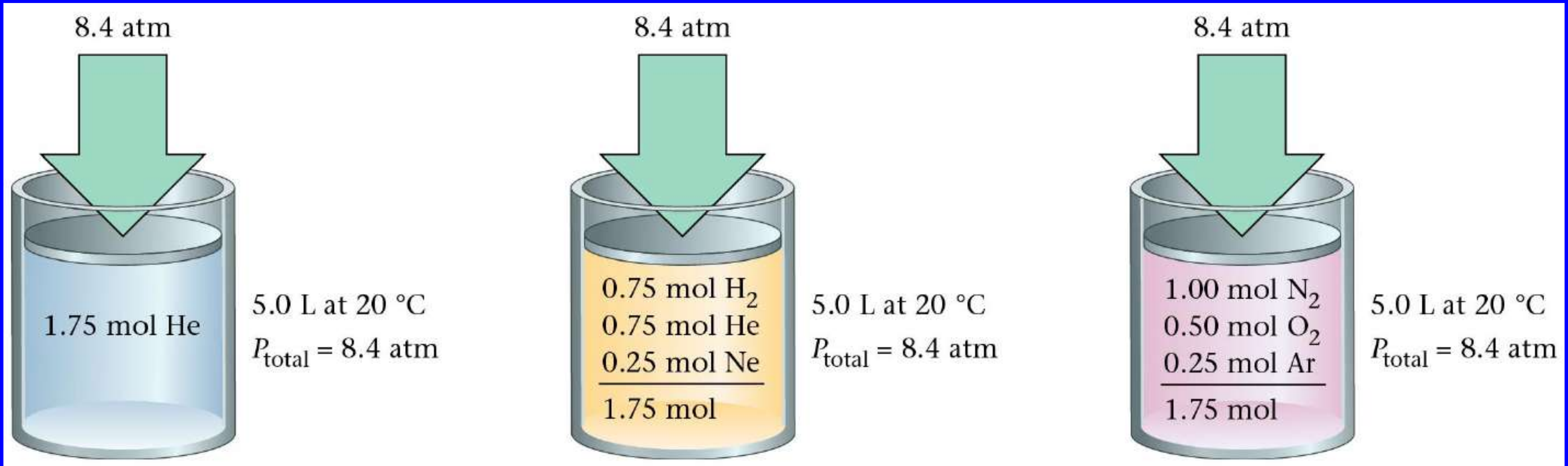
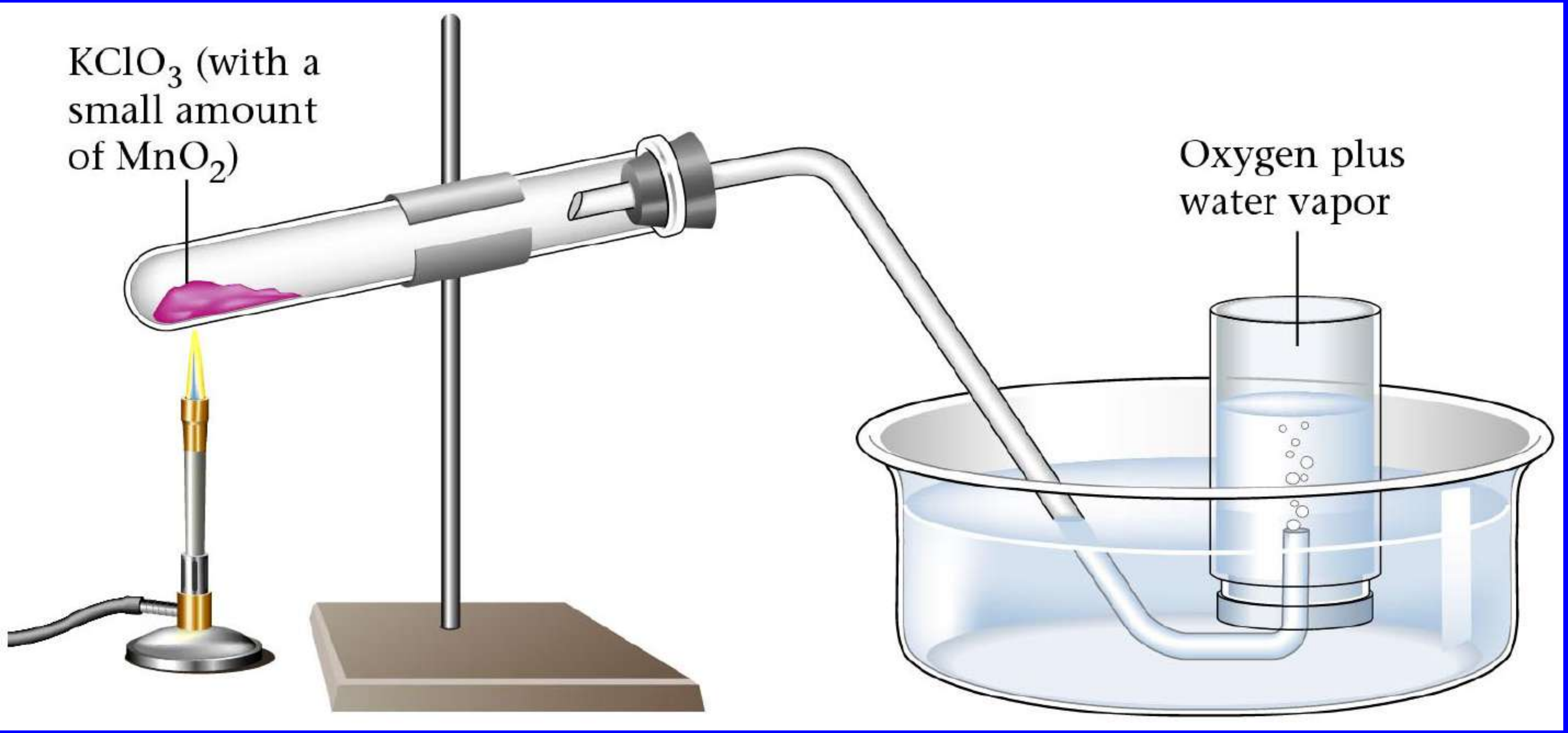


Figure 13.12: The production of oxygen by thermal decomposition.



Ideal vs. Real

- At high pressures and/or low temperatures, gases deviate significantly from ideal gas law
- At low pressures and/or high temperatures, real gases approach behavior of ideal gas
- Ideal gas is hypothetical substance

Postulates of KMT.

Postulates of the Kinetic Molecular Theory of Gases

1. Gases consist of tiny particles (atoms or molecules).
2. These particles are so small, compared with the distances between them, that the volume (size) of the individual particles can be assumed to be negligible (zero).
3. The particles are in constant random motion, colliding with the walls of the container. These collisions with the walls cause the pressure exerted by the gas.
4. The particles are assumed not to attract or to repel each other.
5. The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature of the gas.

Kinetic Molecular Theory

- Qualitative analysis (not mathematical)
- Temperature reflects how fast particles move
  - High temp: particles move fast, hit sides of container often
  - Low temp: particles move slower, hit sides less often
  - KM theory: temp is directly proportional to kinetic energy
  - Pressure due to collisions with wall, as temp increases, pressure increases

Figure 13.13: (a) A gas confined in a cylinder with a movable piston.
(b) The temperature of the gas is increased at constant pressure P_{ext}.

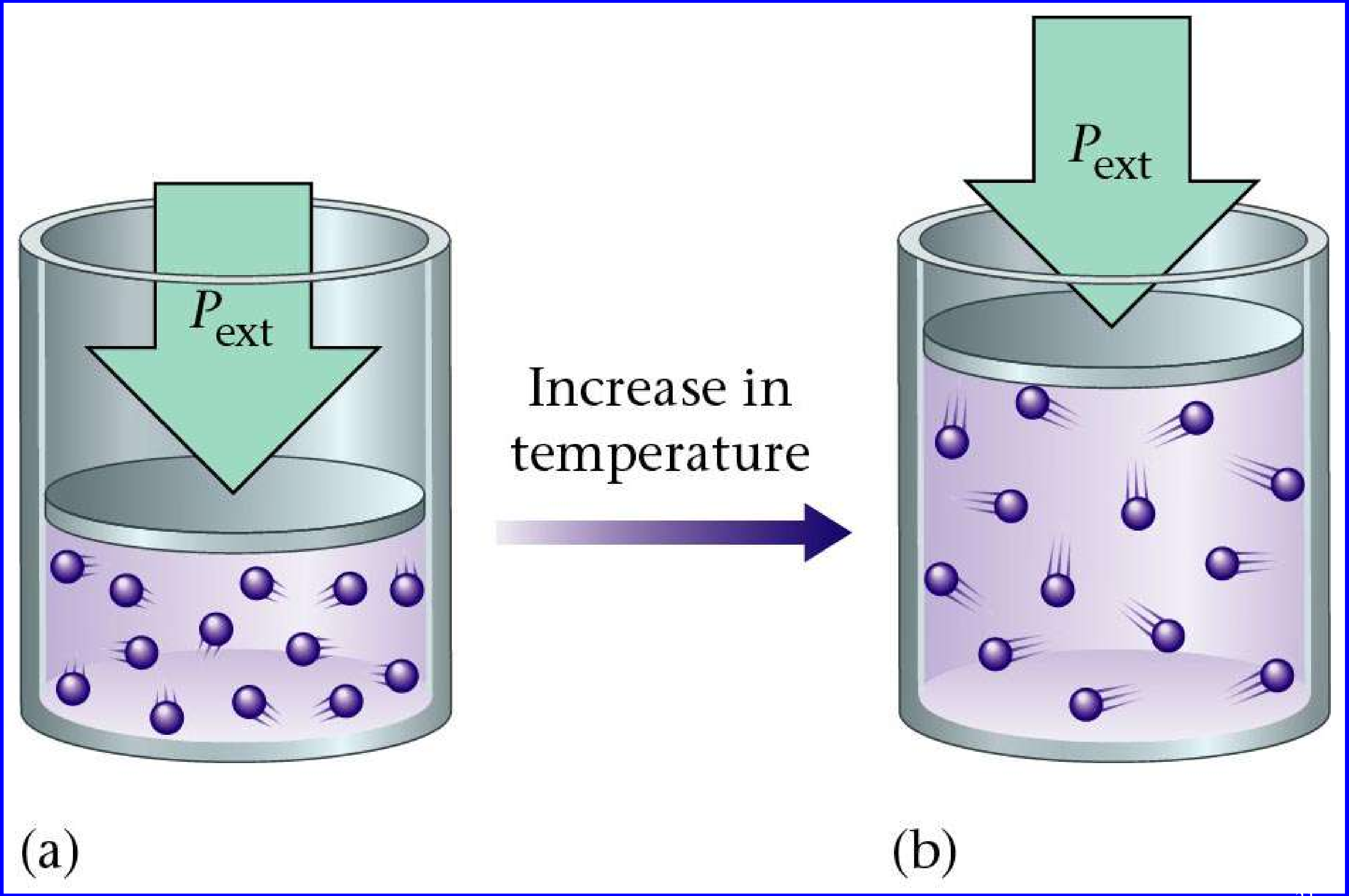
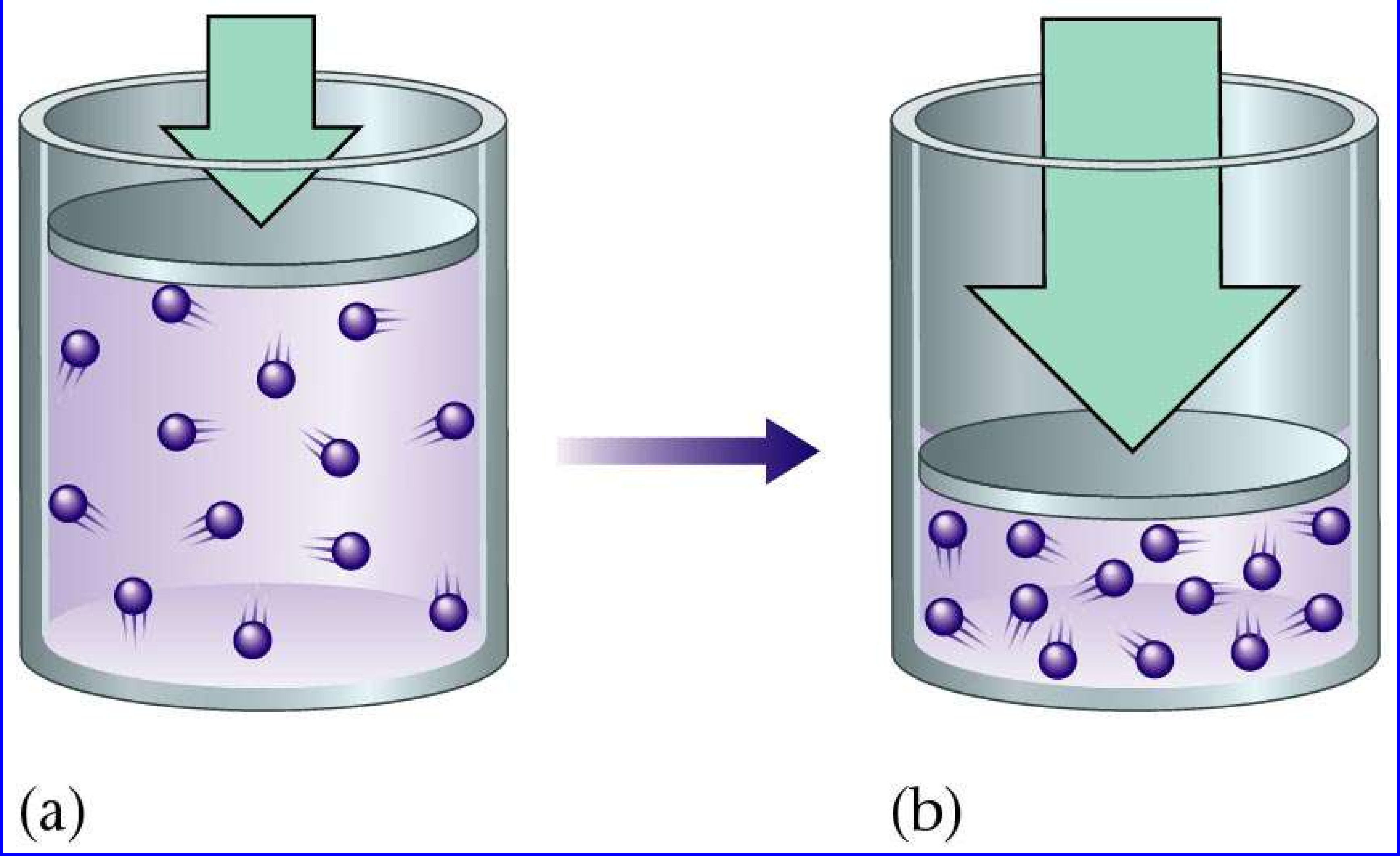


Figure 13.14: A gas sample is compressed.



Real Gases

- Ideal gas: hypothetical substance consisting of particles with zero volume and no attractions for one another
- Behave like ideal gases under many conditions
- High pressure: molecules closer and more likely to attract, does not obey ideal gas law

Gas Stoichiometry

- Can use ideal gas law to convert from moles to volumes
- One mole of any gas at STP (standard temperature and pressure) occupies 22.4 L = molar volume
  - Standard temperature = 0°C = 273 K
  - Standard pressure = 1 atm