

**AP Chemistry
Assignment Packet**

States of Matter & Gas Laws

Introduction: This unit will cover states of matter and gas laws. All of chapter 5 and portions of chapter 11 (pages 458-476) will be included in this unit. It is imperative that you do not misplace or lose this packet. Should you do so, please see my schoolnotes.com website.

AP Chemistry WS 5.1 Introduction to Phase Diagrams & States of Matter

1. What is meant by heat of fusion?
2. Which phase changes are endothermic?
3. Explain the following observations:
 - a. During the cold winter months, snow often gradually disappears without melting
 - b. The heat of fusion of any substance is generally lower than its heat of vaporization
4. On a phase diagram why does the line that separates the gas and liquid phases end rather than go to infinite pressure and temperature?
5. What is the significance of the triple point in a phase diagram?
6. Explain why energy is absorbed during phase change without the temperature changing.
7. What two factors affect states of matter?
8. What causes pressure?

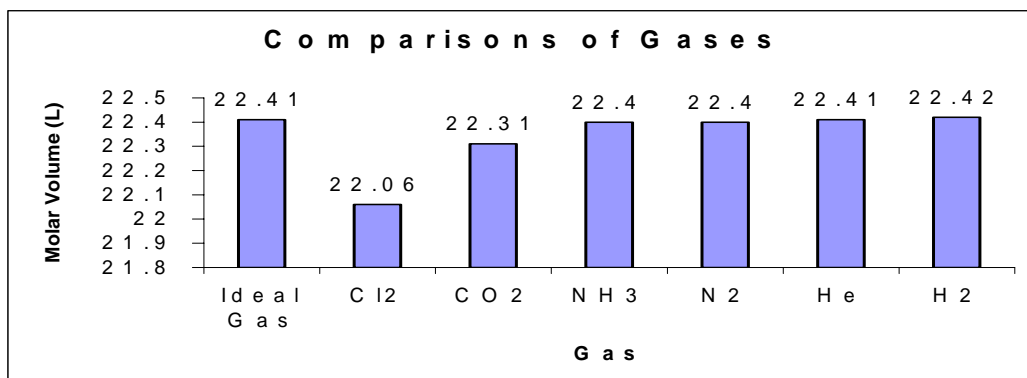
AP Chem WS 5.2 Gas Laws – Combined Gas Law & Ideal Gas Law

1. When 2.00 L of gas at 760 torr and 27°C is expanded to 3.00 L at 77°C, what will the final pressure be?
2. A 20.0 L sample carbon monoxide at 3.4 atm and 45°C is heated until the temperature is doubled and the pressure is cut by three-fourths. What volume will the gas occupy at these conditions?
3. A certain flexible weather balloon contains helium gas. Initially, the balloon is at sea level, where the temperature is 25°C and the barometric pressure is 720 torr and occupies a volume of 5.0 L. The balloon then rises to an altitude of 6000 feet where the pressure is 605 torr and the temperature is 15°C. What is the change in the volume of the balloon as it ascends from sea level to 6000 ft?
4. The steel reaction vessel of a bomb calorimeter, which has a volume of 75 mL, is charged with oxygen gas to a pressure of 145 atm at 22°C. Calculate the moles of oxygen gas in the reaction vessel.
5. A 3.5 L flask contains 0.90 g Cl₂ at a temperature of 34°C. What is the pressure (in mm Hg) inside the flask?
6. A steel cylinder contains 23.4 g H₂ gas at a pressure of 45 MPa and temperature of 34°C. Some of the hydrogen is removed to react with oxygen to produce water. The pressure of the gas after removal is found to be 34 MPa at a temperature of 14°C. What mass of gas remains in the tank?

Worksheet 5.3 Ideal Gas Law

1. Calculate the volume of 0.3000 moles of a gas at 60°C and 0.821 atm.
2. Calculate the volume of 8.40 grams N₂ at 100°C and 800 Torr.
3. What volume will 1.216 g of NO₂ occupy at 19°C and 654 torr?

- A 10.0 L cylinder of oxygen at 4.00 atm pressure and 17°C developed a leak. When the leak was repaired, 2.50 atm of oxygen remained in the cylinder, still at 17°C. How many moles of gas escaped?
- A gas cylinder contains 370 g of oxygen gas at 30.0 atm and 25°C. What mass of oxygen will escape if first the cylinder is heated to 75°C and the valve is held open until the gas pressure is 1.00 atm while the temperature being maintained at 75°C?
- Below is a chart comparing an Ideal Gas at STP to other gases. Explain why the other gases have volumes different from the Ideal Gas?



Ideal Gas Law Problems

(Guided & Independent Practice)

- What is the molecular mass of a gas at 27.0°C and a pressure of 506.5 kPa. The volume of the flask is 1270 mL. The mass of the flask after the gas was added was 16.55 g. The mass of the flask by itself is 8.30 g.
- The density of a gas was measured at 1.50 atm and 27.0°C and found to be 1.95 g/L. Calculate the molar mass of the gas.
- Calculate the density of SiCl₄ vapor at 85.0°C and 758 torr.
- Calculate the density of ammonia gas at 27.0°C and 635 torr.
- Find the molecular mass of a gas whose density at 40.0 °C and 758 torr is 1.386 kg/m³.
- At 5 atm pressure and 70.0 °C, how many grams of O₂ gas are present?

Worksheet 5.4 Gas Stoichiometry and Dalton's Law of Partial Pressures

- Quicklime (CaO) is produced by the decomposition of calcium carbonate (CaCO₃). Calculate the volume of CO₂ at STP produced from the decomposition of 152 g CaCO₃. Note: When carbonates decompose, an oxide and carbon dioxide are formed.
- The method used by Joseph Priestley to obtain oxygen made use of the thermal decomposition of mercury (II) oxide: 2 HgO (s) → 2 Hg (l) + O₂ (g). What volume of oxygen gas, measured at 30°C and 725 torr, can be produced from the complete decomposition of 4.1 g mercury (II) oxide?
- Explain Dalton's Law of Partial Pressures.
- What is meant by mole fraction?
- Calculate the mole fraction for each element:
 - 2.00 mole N₂, 1.50 mole O₂, 2.34 mole H₂O
 - 5.00 g H₂, 3.4 L N₂ (@STP), and 4.5 x 10²⁴ molecules O₂
- The mole fraction of nitrogen in the air is 0.7808. Calculate the partial pressure of N₂ in air when the atmospheric pressure is 760. Torr.

- A 250. mL of oxygen is collected over water at 25°C and 760.0 torr pressure. What is the pressure of the dry gas alone? (Vapor pressure of water at 25°C = 23.8 torr)
- A 32.0 mL sample of hydrogen gas is collected over water at 20.°C and 750.0 torr pressure. What is the volume of the dry gas at STP? (Vapor pressure of water at 20.°C = 17.5 torr).
- A 54.0 mL sample of oxygen gas is collected over water at 23°C and 770.0 torr pressure. What is the volume of the dry gas at STP? (Vapor pressure of water at 23°C = 21.1 torr)
- A mixture of 2.00 moles H₂, 3.00 moles NH₃, 4.00 moles of CO₂, and 5.00 moles of N₂ exert a total pressure of 800. Torr. What is the partial pressure of each gas?
- The partial pressure of F₂ in a mixtures of gases where the total pressure is 1.00 atm is 300. Torr. What is the mole fraction of F₂?

AP Chemistry Review Problems (Mini-Set) Gas Laws (Chapter 5)

- You vaporize a sample of a liquid using the Dumas Method. Calculate the molar mass of the unknown liquid if you determine the mass to be 1.012 g, volume of the flask to be 354 cm₃ at a pressure of 742 torr and temperature of 99.0°C.
- Hydrogen gas is produced when zinc reacts with sulfuric acid: Zn (s) + H₂SO₄ (aq) → ZnSO₄ (aq) + H₂ (g). If 159 mL of wet H₂ is collected over water at 24.0°C and a barometric pressure of 738 torr, how many grams of Zn have been consumed? (hint P_{H₂O} at 24.0°C = 22.38 torr)
- Suppose you have two 1-Liter flask, one container N₂ at STP, the other containing CH₄ at STP. How do these systems compare with respect to:
 - number of molecules
 - density
 - average kinetic energy of the molecules
 - rate of effusion through a pinhole leak
- Explain why He would travel faster than O₂ and CO.
- Explain the conditions at which gases deviate from ideality.
- Calculate the relative rate of effusion of SO₃ to SO₂

Study Guide-Gas Laws Intro

Useful Formulas:

Formulas	Meaning of Variables
$E = n * \Delta H_{\text{vaporization}}$; $E = n * \Delta H_{\text{fusion}}$	E = energy (kilojoules (kj)); $\Delta H_{\text{vaporization}}$ = heat of vaporization (kj/mol); ΔH_{fusion} = heat of fusion (kj/mol) {heat of fusion, heat of vaporization are values that will be provided to you if needed}
$P = F / A$	P = pressure (N * m ² or Pa), F = force (in Newtons), A = area (m ²)
$P = g * d * h$	P = Pressure, g = 9.8 m/s ² , d = density of liquid (g/mL or g/cm ³), h = height of liquid (m)
$P_1 * V_1 = P_2 * V_2$	P = pressure, V = volume (Units must be same on both sides) Boyle's Law
$V_1 / T_1 = V_2 / T_2$	T must be in Kelvin; Charles's Law
$P_1 / T_1 = P_2 / T_2$	T must be in Kelvin, Gay-Lussac's Law
$V_1 / n_1 = V_2 / n_2$	n = number of Moles, Avogadro's Law
$(P_1 * V_1) / (n_1 * T_1) = (P_2 * V_2) / (n_2 * T_2)$	Combined Gas Law (Realize that if something is constant, it cancels)

	out in the combined gas law)
$P * V = n * R * T$	Ideal Gas Law, R = Ideal Gas Law Constant, 0.821 (L * atm) / (mol * K), 8.314 J / (mol * K), 8.31 (L * kPa) / (mol * K)
$n = m / M$	n = number of moles, m = actual mass of the gas, M = molar mass (gfm)
$P * V = (m / M) * R * T$ OR $M = (m * R * T) / (P * V)$ OR $M = (R * T * d) / P$	d = density (g / L)
1 mol = 22.4 L (@STP)	STP = 1 atm and 0°C
$P_{total} = P_1 + P_2 + P_3 + \dots P_n$	Dalton's Law of Partial Pressures
$\chi = n_1 / n_{total}$	χ = mole fraction; no units
$\chi_n = P_n / (P_1 + P_2 + P_3 + \dots P_n)$	Dalton's Law using Pressures and Mole Fraction
$P_n = \chi_n * P_{total}$	Different Variation of Dalton's Law
$P_{total} = P_{gas} + P_{water\ vapor}$	Collection of a gas over water (Partial Pressure of water provided to you)
$u_{rms} = \sqrt{3 * R * T / M}$	u_{rms} = root mean square speed of the molecules; M = molar mass (gfm) in kg / mol (so units cancel out); Use R = 8.314 J / (mol * K) so units cancel out
$\frac{\text{Rate Gas A}}{\text{Rate Gas B}} = \sqrt{\frac{\text{molar mass gas B}}{\text{molar mass gas A}}}$	Graham's Law of Effusion; Interpretation is based on Gas A to Gas B (be careful)

Conversions Factors
1 mL = 1 cm ³
1 L = 1 dm ³
1 atm = 101.3 kPa
1 atm = 760 torr
1 torr = 1 mm Hg

This test will cover portions of chapter 11 (pages 458-476) and chapter 5 (pg 177-200)

- What two factors determine the state of matter for a substance?
- Define the following phase changes:
 - freezing
 - melting
 - evaporation
 - condensation
 - sublimation
 - deposition
 - vaporization
- Describe the arrangement of particles in solids, liquids, and gases.
- Describe the following phases changes as exothermic or endothermic:
 - melting
 - evaporation
 - freezing
 - condensation
- Draw a heating curve. Label the x and y axis. On the curve, label the following: solid, liquid, gas, heat of vaporization, heat of fusion.
- The heat of vaporization of methanol, CH₃OH, is 35.3 kJ/mol at 64.6°C. How much energy is required to evaporate 1.00 kg of this alcohol?
- How much energy is needed to melt 1 pound of ice?
- Define the following terms:
 - triple point
 - supercritical fluid
 - critical point
 - critical pressure
- Draw a phase diagram and label the following the points: solid, liquid, gas, triple point, and critical point. (Remember to label the units)
- What are the characteristics of an ideal gas?