**Ch 6 Gases Name:**

*6.1 Supersonic Skydiving*

How do gases create pressure?

How would uncontrolled decompression affect Eustace?

*6.2 Pressure: The Result of Molecular Collisions P = \_\_\_ / \_\_\_*

What do notice about the isobars *(lines of \_\_\_\_\_\_\_\_ pressure)* in the NE US? What do you think this tells you about the winds in that area?

Why do your ears hurt when you go up a mountain?

Briefly explain how the mercury barometer in *Figure 6.4* was made. How does it measure air pressure?

*Note: The pressure of the atmosphere is supporting the weight of the Hg in the column. Normal sea level pressure*

*supports a column \_\_\_\_\_\_\_mm (\_\_\_\_\_\_\_\_in) high. This is equal to \_\_\_atm and also equal to \_\_\_\_\_\_psi.*

*The unit torr is the same as mmHg.*

The SI unit of pressure is the \_\_\_\_\_\_\_\_\_\_\_ 1 Pa = \_\_\_\_\_\_\_\_\_\_ 1atm = \_\_\_\_\_\_\_\_\_\_Pa (or 101.325 kPa)

*In*[*chemistry*](https://en.wikipedia.org/wiki/Chemistry)*and in various industries, the reference pressure referred to in "*[*standard temperature and pressure*](https://en.wikipedia.org/wiki/Standard_temperature_and_pressure)*" (STP) was commonly 1 atm (101.325 kPa) but standards have since diverged; in 1982, the*[*International Union of Pure and Applied Chemistry*](https://en.wikipedia.org/wiki/International_Union_of_Pure_and_Applied_Chemistry)*(IUPAC) recommended that for the purposes of specifying the physical properties of substances, "standard pressure" should be precisely 100*[*kPa*](https://en.wikipedia.org/wiki/Pascal_(unit))*(1*[*bar*](https://en.wikipedia.org/wiki/Bar_(unit))*). Millibars are used in meteorology. kPa units are on tires.*

*Practice 6.1 More Practice 6.1*

How do you know the gas in the manometer is at higher pressure than atmospheric pressure by the position of the Hg?

Read *Chemistry and Medicine: Blood Pressure*

*6.3 The Simple Gas Laws: Boyle’s Law, Charles’s Law, and Avogadro’s Law*

Why do scuba divers need Boyle’s Law?

*Figure 6.6 & 6.7* show that \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_ are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ related. (T,n is constant)

*Note how the equation is derived on pg 217-18*

*Practice 6.2*

*Figure 6.10* shows that \_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ are \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ related. (P, n is constant)

When all of the lines in the graph are extrapolated towards lower T it is shown that a gas would have \_\_\_\_\_ volume

at\_\_\_\_\_\_\_\_oC (\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ zero). Why would a gas have “no volume” at this temperature?

Read *Chemistry in Your Day: Extra-Long Snorkels* ([link](https://www.scubadoctor.com.au/scorkl-dangers.htm%20%20https:/aquazealots.com/maximum-snorkel-length-and-why-they-are-not-longer/))

*Note how the equation is derived on pg 220. Practice 6.3 CC6.1 \_\_\_\_*

*The textbook does not introduce Gay-Lussac’s Law which shows that \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ are directly related (V, n constant). P1/T1 = P2/T2*

*Figure 6.12* shows that \_\_\_\_\_\_\_\_\_\_\_\_\_\_ is directly related to the number of \_\_\_\_\_\_\_\_\_. (\_\_\_ and \_\_\_ constant)

*Note how the equation is derived on pgs 221-2. Practice 6.4*

*6.4 The Ideal Gas Law* If Boyle’s, Charles’s, and Avogadro’s Laws are combined you obtain:

*P1V1 = P2V2 = CONSTANT* This constant is called the \_\_\_\_\_\_\_ gas constant and is the letter \_\_\_\_.

*n1T1 n2T2*

So this gives *PV* *= R or PV = nRT* the \_\_\_\_\_\_\_\_\_\_ gas law. *Solve for the value of R by using P=1atm,*

*nT V=22.4L, n= 1mol, and T = 273.*

R =

*Practice 6.5(use R = 0.08206atm•L/mol•K) Practice 6.6 (use R = 62.4mmHg•L/mol•K)*

*6.5 Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas*

The calculation at the bottom of pg 225 shows how molar volume at STP (\_\_\_oC and \_\_\_\_\_atm)is determined to be \_\_\_\_\_\_L. The image shows that \_\_ mole of any \_\_\_\_\_ will have this volume.

*Note: The section on Density of a Gas shows how you can use molar volume to determine density of a gas. I feel that it it is easier to use the ideal gas law to solve for volume and then divide by the molar mass of the gas D=m/v.*

*Practice 6.8*

*6.6 Mixtures of Gases and Partial Pressures*

Because molecules of an \_\_\_\_\_\_\_\_\_\_\_\_\_ gas do NOT \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_, each of the components in a mixture acts

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of the others. What % of the total atmospheric pressure is due to CO2? \_\_\_\_\_\_

What does Dalton’s law of partial pressures say about the gas pressures in a mixture?

*Practice 6.9 Convert your answer to grams of H2 gas. CC6.5 \_\_\_*

When mtn climbing the partial pressure of O2 becomes too \_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_ can occur. When diving the

partial pressure of O2 becomes too \_\_\_\_\_\_ and can lead to \_\_\_\_\_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_\_\_. If the partial pressure

of N2increases above \_\_\_\_atm, nitrogen \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ can occur. (also the “bends”)

*Practice 6.10*

When you collect gas “over water” *(Figure 6.14)* how do you find the partial pressure of the collected gas?

*Practice 6.11*

*6.7 Gases in Chemical Reactions: Stoichiometry Revisited.*

*Practice 6.12 (solve for mole O2 first using PV=nRT with R=62.4mmHg•L/mol•K then use stoich to find g Ag2O)*

*Practice 6.13 CC6.6 \_\_\_\_*

*6.8 Kinetic Molecular Theory: A Model for Gases*

How does Charles’s Law differ from KMT?

Briefly *summarize* the 3 basic postulates of KMT.

Atoms of a sample of He and a sample of Ar at the same \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ have the same \_\_\_\_\_\_\_\_\_\_\_\_\_\_ KE

NOT the same average \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. (Which is faster?)

*Carefully read over how pressure and the gas laws can be explained by KMT.*

*Note: We will discuss rms (root mean square velocity) using a demonstration.*

What is shown by *Figure 6.18 and by Figure 6.19?*

*Practice 6.14 CC6.8 ­­­\_\_\_\_*

6*.9 Mean Free Path, Diffusion, and Effusion of Gases CC6.9 \_\_\_\_*

Why do gas particles have longer mean free paths when they are at lower pressures?

What is the difference between diffusion and effusion?

*6.10 Real Gases: The Effects of Size and Intermolecular Forces*

Under what conditions do gases not behave ideally?

Johannes van der Waals showed that as pressure on a gas \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ the particles start to occupy a

significant portion of the total volume. Also, as the temperature of a gas \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ the

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ forces (IMFs) between particles become more important. *Notice in the van der waals equation*

*the correction factor for IMFs is added to the pressure while the correction factor for volume is subtracted.*

What does *Figure 6.26* show about water vs helium as “real” gases?

*Exercises*(pgs 254-258) #27(a,b), 29, 31, 33, 35, 37, 41, 43, 51, 56, 63, 67, 70, 81, 89, 97, 107, 111