Chapter 12 Section 1

Pressure

A gas exerts pressure on its surroundings.
    Blow up a balloon.
The gas we are most familiar with is the atmosphere, a mixture of mostly elemental nitrogen and oxygen.

Pressure of the atmosphere varies with elevation and weather conditions.

Barometer- device used to measure atmospheric pressure.

At sea level the average barometric pressure is 760 mm of mercury.

The original barometer was a column of mercury in a dish of mercury (see page 360 in textbook for a diagram). Be able to explain how a barometer works.

At sea level the pressure of the atmosphere holds the column of mercury at approximately 760 mm.

The pressure of the atmosphere is caused by the pull of gravity on the air, i.e. the weight of the air.

Units of Pressure

mm of mercury and torr are two different units of pressure that can be used interchangeably. 1 mm of mercury = 1 torr

Standard atmosphere (atm) is another common unit. 1 atm = 760 torr or 760 mm of mercury

The SI unit for pressure is the pascal (abbreviated Pa). 1 atm = 101325 Pa

Another unit of pressure is pounds per square inch or psi. 1 atm = 14.69 psi This unit of pressure is commonly used in the engineering field.

You need to be able to convert from one unit of pressure to another when given the above equivalency statements!

Manometer- a device used to measure the pressure of a gas in a container. The pressure of the contained gas is equal to the difference in mercury levels. See diagram on page 361 and be able to explain how a manometer works.

Chapter 12 Section 2

Pressure and Volume Relationship in Gases

Robert Boyle performed careful studies of gases in the 1600's

Determined a relationship between pressure on a gas and its volume.

Pressure x Volume = k (constant)

This means that Pressure and Volume are inversely related

As pressure increases, volume decreases; As pressure decreases volume increases.

Graphed data creates a hyperbola (half of one) indicating an inverse relationship

Boyle's experiments assume that temperature and the amount of gas remain unchanged!

We can use Boyle's law to determine the new pressure or volume of a gas when we are told a change to initial pressure or volume.

$P_1 \times V_1 = P_2 \times V_2$ can be rearranged to solve for any part when the other three parts are known.

Be able to perform calculations using Boyle's Law!

Chapter 12 Section 3

Temperature and Volume Relationship in Gases

Jacques Charles studied gases about a century after Boyle (late 1700's to early 1800's)

Volume of a gas increases as Temperature increases if pressure and amount of gas remains constant. A plot of this data gives a straight line indicating a linear relationship. Relationship between temperature and volume of a gas is directly proportional.

As gases are cooled they eventually turn to liquid. If experimental data is extrapolated (extension of a straight line) all of the lines meet at the same "zero volume point" at -273 degrees C. Since negative volume is impossible, this temperature represents the lowest possible temperature. This temperature is defined as **absolute zero**.

Charle's law : Volume= b(constant) x Temperature which can be rearranged to be Volume / Temperature = b (constant) NOTE: Temperature must be in Kelvins!!!
Same initial to final condition rules apply.

Be able to perform calculations with Charle's law!

Chapter 12 Section 4

Number of Moles (Amount of Gas) and Volume Relationship in Gases

Amadeo Avogadro, 1811, another gas relationship.

Volume of a gas is directly proportional to the number of moles of gas (amount) present at constant temperature and pressure.

Avogadro's law: Volume = number of moles (n) x a (constant) or Volume/ number of moles = constant.

Same initial to final condition rules apply.

Be able to perform calculations with Avogadro's law!

Chapter 12 Section 5

So far we have discussed and performed calculations with three different laws pertaining to the behavior of gases.

These three laws can be combined into a single equation the **IDEAL GAS LAW**.

Volume = R(constant) x (Temperature x Number of moles / Pressure)

R is the combined constant known as the universal gas constant and has a value of 0.0821 Liter x atm/K x moles

PV= nRT is the usual form for the Ideal Gas law and simply represents the same formula as above, just rearranged algebraically.

This equation predicts the behavior of a gas that is behaving ideally. Most gases obey these predictions near 1 atm and zero degrees C or higher. Most texts assume ideal behavior and use this equation liberally.

The Ideal Gas Law also be used to predict initial and final conditions when changes are given.

When solving this type of problem, you must arrange the equation so all changing components are on one side of the equals sign and the unchanged components are on the other.

Be able to perform calculations with The Ideal Gas Law!

Chapter 12 Section 6

Dalton's Law of Partial Pressures

Mixtures of gases behave independently of each other.

For a mixture of gases in a container, the total pressure exerted is the sum of the partial pressures of the gases present. – DALTON'S LAW OF PARTIAL PRESSURES

Partial Pressure- the pressure that a gas would exert if it were alone in the container

Two important points:

1. The volume of an individual gas particle (atom or molecule) is not very important
2. The forces between particles is not very important

Pressure is dependent simply on the number of particles, it does not matter what kind.

A mixture of gases is made when you collect a gas over water in the lab.

Vapor pressure of water- table given to you which is temperature dependent, gives the partial pressure of the water in your gas sample.

Finding the pressure of a mixture of gases in all about finding the individual pressure of each.

Examples: See text Page 379 – 382 - these are difficult to show in Word. I will also be doing examples in class so take notes then also.

Chapter 12 Section 7

Laws and Models Review

Laws- allow prediction of behavior
Models- theories that attempt to explain the behavior, approximation, may need modifications and may not hold perfectly true.

Chapter 12 Section 8

Kinetic Molecular Theory of Gases

A Model that attempts to explain the behavior of ideal gases.

Kinetic molecular theory-

1. Gases consist of tiny particles. These particles can be atoms or molecules.
2. The particles in a gas are so tiny compared to the space in between the particles that the size of the particle (volume of an individual particle) can be assumed to be zero.
3. The particles of a gas are in constant random motion and the collide with the sides of the container causing the pressure in the container.
4. The particles are assumed to not attract or repel each other.
5. The average kinetic energy of the gas particles is directly proportional to the temperature in Kelvin. This statement means that the more you heat a gas the faster the particles are moving and therefore the more often they collide with the container.

Chapter 12 Section 9

Implications of the Kinetic Molecular Theory of Gases-Qualitative relationships- no math

Meaning of Temperature- temperature is an indication of molecular motion
- in a gas particle motion and Kelvin temperature are directly related

*Relationship between Temperature and Pressure- Pressure of a gas increases as temperature is increased because an increase in temperature increases the number and magnitude of the impacts thus increasing the pressure.

*Relationship between Volume and Temperature- As temperature is increased the volume will increase because in order to keep the pressure the same more space is needed for the molecules to move around in.

*Relationship between Pressure and Volume- If we make the volume less (think smaller container) the particles have less space to travel before hitting the walls of the container. This will increase the number of impacts and therefore increase the pressure.
Chapter 12 Section 10

Gas Stoichiometry

Very similar to other Stoichiometry problems.

Step 1: Convert Grams to Moles
Step 2: Convert Moles of one substance to Moles of another substance
Step 3: Convert Moles to Grams

For gases most often a volume measure, not grams is used in the lab. The ideal gas law allows you to calculate moles from volume when temperature and pressure are known. We will be simply applying this to Stoichiometry problems.

A Few Terms

*Standard Temperature and Pressure (STP) = 1 atm of Pressure and Zero Degrees Celsius

* Molar Volume of an Ideal Gas = 22.4 L at Standard Temperature and Pressure (STP)

Example:

What volume of oxygen is produced at 1 atm of pressure and 25°C by the complete reaction of 10.5 grams of potassium chlorate ($KClO_3$) according to the following balanced equation: 2 $KClO_3$ (s) → 2KCl (s) + 3O_2 (g)?

Step one: Grams to moles- 10.5 grams $KClO_3$ to moles $KClO_3$ by dividing by molar mass of $KClO_3$. 10.5/122.6 = .0856 mol $KClO_3$

Step two: Moles to Moles- Convert moles of $KClO_3$ to moles of O_2 (g) using the mole ratio from the balanced equation. Ratio of $KClO_3$ to O_2 is 2 to 3.

.0856 mol $KClO_3$	3 mol O_2	= .128 mol O_2
	2 mol $KClO_3$	

Step three: Moles of O_2 to volume of O_2 using the ideal gas law. PV=nRT

1atm(v)= (.128 mol)(0.0821)(298)

V= 3.13 L of O_2 will be produced.