Name:

Stoichiometry, Formulas and Equations	**Objectives**
10. Properties of Gases	-State and explain Boyle's law -describe the significance of Boyle' air pump -state and explain Charles's law **-state and explain Gay-Lussac's law of combining volumes** **-state and explain Avogadro's law** -carry out simple calculations using the combined gas law $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$ = constant **-define ideal gases** **-list the assumptions of the kinetic theory of gases** **-explain why gases deviate from ideal gas behavior** **-carry out simple calculations involving PV = nRT (units: Pa m^3,K)**

Def^n: A **gas** is a substance that has no well-defined boundaries and diffuses rapidly to fill any container in which it is placed.

Temperature:

We use the Kelvin scale of temperature often for calculations.

- To convert ^{o}C to K, add 273.
- 0 ^{o}C = 273 K = **standard temperature**

Pressure:

Units used are Pascals (Pa).

- 1 kPa = 1,000 Pa
- 1 hPa = 100 Pa
- 1 x 10^5 Pa = 100 kPa = normal atmospheric pressure = **standard pressure**

Volume:

Units used are m^3

- 1 L = 1 dm^3 = 1,000 cm^3
- 1 m^3 = 1,000 L = 1 x 10^6 cm^3 = 1,000,000 cm^3

Standard Temperature and Pressure (s.t.p.)

The conditions of s.t.p. are:

1. 100,000 Pa of pressure
2. 273 K in temperature

<u>Gas Laws:</u>

1. **Boyle's Law:**
   At constant temperature, the volume of a fixed mass of a gas is inversely proportional to its pressure.
2. **Charles' Law:**
   At constant pressure, the volume of a fixed mass of a gas is directly proportional to its Kelvin temperature.

3. **The Combined Gas Law:**

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

P_1 = Initial Pressure	P_2 = Final Pressure
V_1 = Initial Volume	V_2 = Final Volume
T_1 = Initial Temperature (K)	T_2 = Final Temperature (K)

This equation is used when the initial conditions of Pressure, Volume and Temperature of a gas are given. Also given are two final conditions. You are asked to find the missing 3rd final condition (either a P, V or T)

4. **Gay-Lussac's Law of Combining Volumes:**
   In a reaction between gases at the same temperature and pressure, the volumes of the reacting gases, and gaseous products, are in simple whole number ratios.

5. **Avogadro's Law:**
   Equal volumes of gases, at the same temperature and pressure, contain equal numbers of molecules.

   Notes: 1 mole of any gas at s.t.p. occupies a volume of 22.4 L.
   1 mole of any gas at room temperature and pressure occupies a volume of 24 L.

The Kinetic Theory of Gases

Def^n: An **ideal gas** is one that obeys all of the assumptions of the kinetic theory of gases under all conditions of temperature and pressure.

Assumptions of the Kinetic Theory of Gases (Properties of an Ideal Gas):

1. Gases are made up of particles which are always randomly moving, colliding with other particles and with the walls of the container.
2. There are no attractive or repulsive forces between the gas molecules.
3. The volumes of the gas molecules are negligible compared to the distance between the molecules.
4. Collisions between the molecules are perfectly elastic.
5. The average kinetic energy of the molecules is proportional to the Kelvin temperature of the gas.

Limitations to the Kinetic Theory of Gases (Properties of a Real Gas):

1. Attractive forces **do** exist between gas molecules due to Van-der-Waals forces, dipole-dipole forces and hydrogen bonding.
2. The volume of the gas molecules is **not** negligible compared to the distances between them.

Ideal Gas Equation:

$$PV = nRT$$

P = Pressure (**Pa**) →**Pa must be used**

V = Volume (**m^3** = cm^3 ÷ 1,000,000 = L ÷ 1,000) →**m^3 must be used**

n = number of Moles of gas (**mol**)

R = 8.3 (**J K^{-1} mol^{-1}**) Universal Gas Constant

T = Temperature (**K**) →**K must be used**

We use this equation when we are asked to find the volume, Pressure, Temperature, number of moles, or relative molecular mass of a gas (or volatile liquid) and we are given only **one** set of conditions.

Experiment: *To measure the relative molecular mass of a volatile liquid.*

1. Find the mass of a clean, dry conical flask, some aluminium foil and a rubber band.
2. Place some of the volatile liquid into the conical flask.
3. Seal liquid in flask using the foil as a cap, securing with the aluminium foil.
4. Make a small pinhole in the top of the foil. This allows the excess vapour to exit the flask, and keeps the contents at the same pressure as the room.
5. Submerge the flask to the neck in boiling water until the liquid is fully vapourised.
6. Remove flask from boiling water and allow to cool. Vapour condenses back into a liquid. Dry the outside of the flask.
7. Find the mass of the conical flask, foil cap, rubber band and condensed vapour. Subtract the mass found in step 1 from this mass to get the mass of the condensed vapour.

Calculations:

1. Use PV=nRT to find the number of moles of vapour which was present in the flask.
   - P = Pressure in the room found using a barometer.
   - V = Volume (in m^3) of the conical flask. Find this by filling conical flask to the brim with water and emptying into a graduated cylinder.
   - n = ?
   - R = 8.3
   - T = Temperature of the boiling water (use a thermometer and convert to K)
2. Find the Relative Molecular Mass of the volatile liquid using the formula:

$$M_r = \frac{\text{Mass of Condensed Liquid}}{n}$$

Mass of the condensed liquid comes from step 7 of the procedure

n is the number of moles of vapour in the flask. This number is your answer to the equation above in step 1 of the calculations.

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