Chemistry

The Behavior of Gases
Lesson 11
Lesson Plan
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The Properties of Gases

Objectives: Describe the properties of gas particles; Explain how the kinetic theory of gas particles relates to Kelvin temperature.

- Kinetic Theory Revisited
  - While we have talked about how heat affects the movement of gas particles, you have probably seen it with a sealed bag that got placed in the sun. The bag will expand and become tight as the temperature and thus the pressure of the gas inside the bag goes up.
  - When we used the kinetic theory to explain the movement of gas particles, we made several assumptions that we will soon find are not totally correct.
    - We stated that gas particles are very small hard spherical objects (usually atoms or molecules), with large distances between them – this means that gases are compressible since there were large distances between the particles and the particles are small compared to the space between them.
      - The compression of gases is used in airbags to absorb the impact of a person during a crash – thus slowing down how fast a person stops.
    - The second statement regarding gases is they have no attraction or repulsion forces between the particles – as a result the particles are free to take up the entire space they are confined in.
    - The third property of a gas is that the gas particles move at high speeds and in straight lines until they collide with another particle or with the sides of the container. The particles rebound with perfect elasticity – that is momentum is conserved. Also, the speed of the particles is directly proportional to their temperature – often expressed in Kelvin.
  - We use four variables to describe a gas
    - Pressure abbreviated as P with units of kPa
    - Temperature abbreviated as T with units of Kelvin
    - Volume abbreviated as V in units of liters
    - Moles abbreviated as n in units of moles

Factors Affecting Gas Pressure

Objectives: Explain how the amount of gas and the volume of the container affect gas pressure; Infer the effect of temperature changes on the pressure exerted by a contained gas.

- Amount of Gas
  - With kinetic theory we can predict that in a container of a given volume, if we add more gas particles that the pressure would increase due to their being more gas particles striking the inside of the container.

- As long as the temperature remains the same, when you double the number of gas particles in a given volume, you double the pressure. If you tripled the number of particles you would triple the pressure, etc – at least until the container burst.
- As long as the temperature remains the same, if you half the number of gas particles you half the pressure.
- When you use an aerosol can, the can contains a liquid to be sprayed along with a volume of gas under pressure. When you depress the push button, the gas pressure pushes the liquid out – until you run out of gas. If you turn a aerosol can upside down and press the button you will release the gas and no liquid

- If you could reduce the volume of a container of gas by ½, then the pressure would double since you have the same number of gas particles in ½ of the space.
  - Likewise, doubling the volume would reduce the pressure by ½.
- Raising the temperature of a constant volume container will raise the temperature. Specifically, if you double the temperature measured in Kelvin, you will double the pressure as you increase the kinetic energy of the particles.
  - Likewise, reducing the temperature by ½ will reduce the pressure by ½.

The Gas Laws

Objectives: State Boyle's law, Charle's La, Gay-Lussac's law; Apply the gas laws to problems involving the temperature, volume and pressure of a contained gas.

- Boyle's Law – Pressure-Volume Relationship
  - In the mid 1600's Robert Boyle studied the relationship of gases as they related to pressure and volume. He noted that if you start off with a certain volume and pressure of gas, and halved the volume, while maintaining constant temperature, then the pressure doubled. He expressed this in a simple relationship where $P_1 \; x \; V_1 = P_2 \; x \; V_2$, where the pressure times the volume before a change would equal the pressure times the volume after a change of either P or V

Sample Problem

Suppose a high altitude balloon contains 30.0L of helium at 103 kPa. When the balloon ascends to an altitude where the pressure is only 25.0 kPa what is the volume?

$$P_1 \; x \; V_1 = P_2 \; x \; V_2 \qquad V_2 = \frac{P_1 \; x \; V_1}{P_2} = \frac{103 \; kPa \bullet 30.0L}{25.0 \; kPa} = 1.24 \; x \; 10^2 \; L$$

- Charles's Law – Temperature and Pressure
  - In 1787 French physicist Jacques Charles investigated the actions of gases and temperatures.
  - Charles was limited with the range of temperatures he could work with since when you lowered the temperature of a gas too much it would become a liquid

- But from what Charles observed, that a graph of a gas at constant pressure, over a range of temperatures, a straight line was produced. If was Charles who noticed that if you extrapolated these graphs toward lower temperatures that they all met at one temperature –273.15ºC. It was not until William Thompson (Lord Kelvin) came along and identified the importance of this temperature as being absolute zero – this was the basis for the temperature scale established by Kelvin with 0ºK being the temperature at which all motion would cease. Typically it is rounded off to be -273ºC.
- Charles stated his findings in the equation $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Sample Problem

A balloon inflated in a room at 24ºC has a volume of 4.00L. If the balloon is then heated to a temperature of 58ºC, what is the new volume if the pressure remains constant?

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad V_2 = \frac{V_1 \bullet T_2}{T_1} = \frac{4.00L \bullet 58^{\circ}C}{24^{\circ}C} = \frac{4.00L \bullet 331^{\circ}K}{297^{\circ}K} = 4.46L$$

- Gay-Lussac's Law – Temperature-Pressure Relationship
  - In 1082 French chemist Joseph Gay-Lussac discovered the relationship between temperature and pressure with constant volume
  - He stated that $\frac{P_1}{T_1} = \frac{P_2}{T_2}$

Sample Problem

If an aerosol can is used up so that the pressure inside the can is 103kPa at 25ºC. If the can was thrown into a fire and heated to 928ºC, what would be the pressure inside the can?

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \quad P_2 = \frac{P_1 \bullet T_2}{T_1} = \frac{103kPa \bullet 25^{\circ}C}{928^{\circ}C} = \frac{103kPa \bullet 298^{\circ}K}{1201^{\circ}K} = 4.15\, x\, 10^2\, kPa$$

- The Combined Gas Law
  - All three gas laws can be put together into one statement
    - $\frac{P_1 \bullet V_1}{T_1} = \frac{P_2 \bullet V_2}{T_2}$
- When any one variable remains constant, then it cancels out

Sample Problem

The volume of a gas-filled balloon is 30.0L at 40ºCand 153kPa. What volume would the balloon have at STP (0ºC, 103kPa)

$$\frac{P_1 \bullet V_1}{T_1} = \frac{P_2 \bullet V_2}{T_2} \quad V_2 = \frac{P_1 \bullet V_1 \bullet T_2}{T_1 \bullet P_2} = \frac{153kPa \bullet 30.0L \bullet 273^{\circ}K}{313^{\circ}K \bullet 103kPa} = 39.5L$$

Ideal Gases

Objectives: Calculate the amount of gas at any specified conditions of pressure, volume, and temperature; Distinguish between ideal and real gases.

Ideal Gases

- Ideal Gas Law
  - So far we have described the behavior of gases using the kinetic theory and three variables – but in real life there is a problem with using this – it does not work just right! Not to fear, this can be fixed with what scientist call a "fudge factor".
  - We first need to add another variable to the mix – quantity measured in moles.
  - Our combined gas law becomes $\frac{P_1 \bullet V_1}{T_1 \bullet n_1} = \frac{P_2 \bullet V_2}{T_2 \bullet n_2}$
  - If we multiply this out we get $\frac{P \bullet V}{T \bullet n}$ which would work in an ideal world, and does under certain conditions.
  - Solving the above equation using that 1mol of gas occupies 22.4L at STP, we would get 8.31 (LxkPa)/(Kxmol), which is given the symbol R.
  - As it turns out, we can use R as the **ideal gas constant** which gives us the **ideal gas law** $R = \frac{P \bullet V}{T \bullet n}$. With this we can solve for any one variable if we know the other three.

Sample Problem

You fill a rigid steel cylinder that has a volume of 20.0L with nitrogen gas (N_2) to a final pressure of 2.00 x 10^4kPa at 28ºC. How many moles of $N_2(g)$ does the cylinder contain?

$$R = \frac{P \bullet V}{T \bullet n} \quad n = \frac{P \bullet V}{T \bullet R} = \frac{2.00x10^4 kPa \bullet 20.0L}{301^\circ K \bullet 8.31 \frac{L \bullet kPa}{K \bullet mol}} = 160 mol\ N_2(g)$$

- The Ideal Gas Law and Kinetic Theory
  - In the previous discussions of the kinetic theory we were dealing with an ideal gas, there are several points that a real gas deviates from. But while real gases deviates from the ideal gas theory, under certain conditions real gases behave very much like an ideal gas.
    - The particles of a real gas have volume, and ideal gas does not.
    - Real gases turn into a liquid – ideal gases do not
    - There is some intermolecular forces at work on real gas particles that tend to make them attract each other.

Gas Molecules: Mixtures and Movements

Objectives: State Avogadro's hypothesis, Dalton's Law, and Graham's laws; Calculate moles ,masses and volumes of gases at STP; Calculate partial pressures and rates of effusion.

- Avogadro's Hypothesis

- We know that all gas particles are not the same size – a hydrogen molecule only has two protons and two electrons – where as a molecule of chlorine gas has 34 protons, 36 neutrons plus 34 electrons. Thus when Avogadro purposed that equal volumes of gases at the same temperature and pressure contain the same number of particles, it was met with a lot of disbelief
    - It was sort of like saying that you could fill two identical rooms with the same number of objects regardless of their size – basketballs or marbles.
    - But Avogadro believed that the gas particles were very small and far apart from each other – thus the difference of a large gas particle and a small gas particle we not much compared to the distance between them.
    - Thus today we recognize that 1 mole of gas contains 6.02×10^{23} particles and occupies 22.4L

Sample Problem

Determine the volume (L) occupied by .202 mol of gas at STP

1mol = 22.4L

$$.202mol \bullet \frac{22.4L}{1mol} = 4.52L$$

- Dalton's Law
    - The air we breathe is not pure oxygen, but rather is made up of a mixture of gases. Each gas contributes to the total pressure that makes up atmospheric pressure.

Nitrogen	78.08%	79.11 kPa
Oxygen	20.95%	21.22 kPa
Carbon dioxide	0.04%	0.04 kPa
Argon and Others	0.93%	0.95 kPa
	100.00%	101.32 kPa

    - The contribution each gas males to the total pressure is called the **partial pressure**.
        - $P_{total} = P_1 + P_2 + P_3 + \ldots\ldots$
        - This is a mathematical form of **Dalton's law of partial pressures** – the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases
        - The fractional component of each gas does not change as the temperature, pressure or volume changes.
            - This is why above 10,000 feet why people must have supplemental oxygen; the oxygen is not sufficient to maintain the human respiratory system.
            - As seen from the table, at sea level the partial pressure for oxygen is 21.22 kPa. On top of Mount Everest the partial pressure of oxygen is 7.06 kPa – the human body needs at least 10.67 kPa.

Sample Problem

Air contains oxygen, nitrogen, carbon dioxide and trace amounts of other gases. What is the partial pressure of oxygen (P_{O2}) at 101.30 kPa of total pressure if the partial pressures of nitrogen, carbon dioxide and other gases are 79.10 kPa, 0.040 kPa, and 0.94 kPa respectively?

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

$101.30\text{kPa} = P_{O2} + 79.10 \text{ kPa} + 0.040 \text{ kPa} + 0.94 \text{ kPa}$

$P_{O2} = 101.30\text{kPa} - (79.10 \text{ kPa} + 0.040 \text{ kPa} + 0.94 \text{ kPa})$

$P_{O2} = 21.22 \text{ kPa}$

- Graham's Law
    - When you open a bottle of perfume, it soon fills the entire room with its scent. Yet if you open a can of soda, you do not smell it across the room – why.
    - Particles of gas move at high speed, until they collide with another object – maintaining momentum. In physics we learn that a component of momentum is kinetic energy; $ke = 1/2mv^2$. So this means that for gas particles of different sizes (mass) to have the same kinetic energy the smaller particles must move at a high velocity. It is this velocity that allows a particular gas particle to quickly travel from one side of a room to the other side.
    - When a gas fills a room, it is called **diffusion**.
    - In the 1840's Thomas Graham measured the rate of **effusion,** how fast gas escaped through a tiny hole - he noticed that gases of lower molar mass effused faster than gases with higher molar mass.
    - Graham purposed his law of effusion – the rate of effusion of a gas is inversely proportional to the square root of the gas's molar mass.
    - This can be seen with a balloon filled with helium and one with nitrogen
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        - $\frac{Rate_A}{Rate_B} = \frac{\sqrt{molar\ mass_B}}{\sqrt{molar\ mass_A}}$