

Phy 212: General Physics II

Chapter 19: Ideal Gases & Thermodynamics Lecture Notes

Counting Atoms

1. 1 mole = 6.022×10^{23} units
2. The number 6.022×10^{23} is called Avogadro's number (N_A)
3. Why the mole? *Because, a one mole quantity of any element has a mass (in grams) equal to its atomic mass.*

- a. It's the relationship between mass and numerical quantity of any element or compound

e.g. Molar mass of ____.

H = 1.008 g/mol {1 mole of H atoms has a mass of 1.008 grams}

O = 16.00 g/mol {1 mole of O atoms has a mass of 16.00 grams}

e.g. Molar mass of ____.

H₂ = $2 \times 1.008 \text{ g/mol} = 2.016 \text{ g/mol}$

{1 mole of H₂ has a mass of 2.016 grams}

H₂O = $2.016 + 16.00 \text{ g/mol} = 18.016 \text{ g/mol}$

{1 mole of H₂O has a mass of 18.016 grams}

The Ideal Gas Law

- The measurable physical parameters that describe the state of a simple gas are:
 - Pressure (p)
 - Volume (V)
 - Number of gas particles or molecules (n or N)
 - Temperature (T) in K
- A relation that describes how these parameters are related is called the **Ideal Gas Law**, which takes 2 forms:

$$(i) \frac{pV}{nT} = R = 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}$$

where n is # of moles & R is Universal Gas Constant

Alternatively:

$$(ii) \frac{pV}{NT} = k = 1.38 \times 10^{-23} \frac{\text{J}}{\text{K}}$$

where N is # of gas particles & k is Boltzmann's Constant

Note: $k = R/N_A$

Kinetic Theory of Gases

- Gas pressure is due to molecular collisions between gas particles and the walls of the container
- The average kinetic energy (K_{avg}) of a single gas particle and is proportional to the bulk gas temperature in K

where f is the # of degrees of freedom

$$a. v_{\text{rms}} = \sqrt{\frac{2K_{\text{avg}}}{m_{\text{particle}}}} = \sqrt{\frac{f(kT)}{m_{\text{particle}}}}$$

$$b. f = 5 \text{ for diatomic gas}$$

$$c. f = 6 \text{ for polyatomic gas}$$

$$E_{\text{int}} = N \cdot K_{\text{avg}} = \frac{f}{2} NkT = \frac{f}{2} nRT$$

- The root-mean-squared speed of the gas particle

Laws of Thermodynamics

0th Law: when 2 objects are in thermodynamic equilibrium independently with a 3rd object, they are in fact in thermodynamic equilibrium with each other or in other words,

You can use a thermometer to measure the temperature of something

1st Law: *conservation of energy*

$$\Delta E_{\text{int}} = Q - W \quad \{\text{where } dE_{\text{int}} = dQ - dW\}$$

2nd Law: *thermodynamic limit of heat engine efficiency*

1. In nature, heat only flows spontaneously from high T to cold T
2. A heat engine can never be more efficient than a "Carnot" engine operating between the same hot & cold temperature range
3. The total entropy of the universe never decreases

3rd Law: *it is not possible to lower the temperature of a system to absolute in a finite number of steps*

It is impossible to reach the temperature of absolute zero

Laws of Thermodynamics (alternative interpretation)

{Also known as Ginsberg's Theorem}

0th: There is a game and you can play it

1st: You can't win the game.

2nd: You can't break even.

3rd: You can't get out of the game.

4th: *The perversity of the universe tends towards a maximum. {The Law of Entropy}*

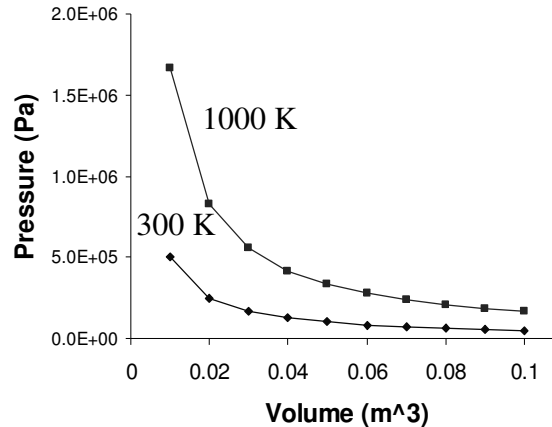
An interesting comment on this interpretation (source: some guy named Freeman):

*"Every major philosophy that attempts to make life seem meaningful is based on the negation of one part of Ginsberg's Theorem. **To wit:***

1. *Capitalism is based on the assumption that you can win.*
2. *Socialism is based on the assumption that you can break even.*
3. *Mysticism is based on the assumption that you can quit the game."*

Pressure-Volume Graphs

1. Useful visual display of a gas system
2. Demonstrate the work performed during a thermodynamic cycle
 - a. Area under the PV curve is the work performed
3. The graphs on the right are PV curves performed at constant temperature (T)



Important Thermodynamic Processes (for gases)

1. Isothermic: constant temperature
 $\Delta T = 0$ means $\Delta E_{\text{int}} = 0$ (E_{int} is constant!)
2. Isochoric (or isovolumetric): constant volume
 $\Delta V = 0$ means $W = 0$ (no work performed on/by system)
3. Isobaric: constant pressure
 $\Delta P = 0$ means $W = P\Delta V$
4. Adiabatic: no gain/loss of heat energy (Q)
 $Q = 0$ means $\Delta E_{\text{int}} = W$

Isothermal Processes

1. Process where the temperature (T) of the working substance remains constant
2. There are 2 possible outcomes
 - a. As work is performed on the working substance, it releases the energy as heat (Q) {Q is released}
 - b. All heat energy (Q) absorbed is converted to work by the working substance {Q is absorbed}
3. Since $\Delta T = 0$, then $\Delta E_{\text{int}} = 0$ so

$$\Delta E_{\text{int}} = Q - W = 0$$

becomes

$$W = Q = (nRT)\ln(V_f/V_i)$$

{work performed by working substance}

Isochoric Processes

1. Process where the volume (V) of the working substance remains constant
2. No work is performed during this process since:

$$\Delta V = 0 \text{ \{area under PV curve is zero\}}$$
3. Two outcomes:
 - a. Internal energy (E_{int}) increases as heat energy (Q) is absorbed by the working substance
 - b. Internal energy (E_{int}) decreases as heat (Q) is released by the working substance
4. Since $\Delta V = 0$ & $W = 0$, then all internal energy changes are due to thermal transfer:

$$dE_{\text{int}} = dQ \text{ \underline{and} } \Delta E_{\text{int}} = Q = nc_v\Delta T$$

Isobaric Processes

1. Process where the pressure (P) of the working substance remains constant
2. Since $\Delta P = 0$, then the area under the PV curve (the work performed) is equal to $P \cdot \Delta V$ or

$$W = P \cdot \Delta V$$

{work performed by working substance}

3. The internal energy change of the system is equal to the difference between heat energy absorbed & work performed, or

$$\Delta E_{\text{int}} = Q - W = Q - P \cdot \Delta V \quad \{\text{where } Q = n c_p \Delta T\}$$

4. The internal energy change (ΔE_{int}) is also related to the change in temperature (ΔT):

$$\Delta E_{\text{int}} = (3/2)nR \cdot \Delta T \quad \{\text{monatomic gas}\}$$

Or

$$\Delta E_{\text{int}} = (5/2)nR \cdot \Delta T \quad \{\text{diatomic gas}\}$$

Molar Heat Capacity & 1st Law of Thermodynamics (for monatomic gases)

1. The heat (Q) absorbed by a gas can be expressed as:

$$Q = c n \Delta T$$

where c is the molar heat capacity (J/mol·K)

2. Since gases it is necessary to distinguish between molar heat capacity at constant pressure (c_p) and constant volume (c_v)

3. Let's begin with the 1st Law of Thermodynamics:

$$Q = \Delta E_{\text{int}} + W$$

4. At constant volume: $\{\Delta E_{\text{int}} = (3/2)nRT \text{ \& } W = 0\}$

$$Q = (3/2)nR\Delta T + 0 = (3/2)nR\Delta T \rightarrow c_v = (3/2)R$$

5. At constant pressure: $\{\Delta E_{\text{int}} = (3/2)nRT \text{ \& } W = P\Delta V = nR\Delta T\}$

$$Q = (3/2)nR\Delta T + nR\Delta T = (5/2)nR\Delta T \rightarrow c_p = (5/2)R$$

Adiabatic Processes

1. Process where no heat energy (Q) is allowed to enter or leave the system, $Q = 0$ ($\Delta E_{\text{int}} = Q - W = -W$)
2. Two outcomes:
 - a. Work performed on the working substance increases its internal energy by exactly the same amount
 - b. Work performed by the working substance decreases its internal energy by exactly the same amount
3. Since $\Delta E_{\text{int}} = -W$ & ΔE_{int} depends on ΔT ,
 - a. For a monatomic gas:

$$dW = -(3/2)nRdT = pdV \quad \text{and} \quad W = - (3/2)nR\Delta T$$
 - b. For a diatomic gas:

$$dW = -(5/2)nRdT = pdV \quad \text{and} \quad W = - (5/2)nR\Delta T$$

P-V & T-V relations for Adiabatic Processes

1. When a gas undergoes an adiabatic expansion/contraction the relationship between pressure and volume is

$$P_i V_i^\gamma = P_f V_f^\gamma$$

or

$$T_i V_i^{\gamma-1} = T_f V_f^{\gamma-1}$$

where γ is c_P/c_V (the ratio the molar heat capacities at constant P and V, respectively)

2. For monatomic gases, γ is 5/3
3. For diatomic gases, γ is 7/5