Physics 7B Midterm 1

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17 Temperature, Thermal Expansion, and Ideal Gas Law

(17.1)
atomic mass (molecular mass) [u]: relative masses of atoms/molecules, numerically same as molar mass

Conversions
$1u = 1.6605 \times 10^{-27}$ kg
$C = \frac{5}{9}(F-32)$
$F = \frac{9}{5}(C)+32$
$K = C + 273.15$

(17.2)
Temperature [C°/F°/K]: how hot or cold something is

	Freezing	Boiling
	0° C	100° C
	32° F	212° F

(17.3)
Zeroth Law of Thermodynamics: If two systems are in thermal equilibrium w/ a third system, they are in thermal equilibrium w/ each oth

(17.4)
Linear Expansion: $\boxed{\Delta l = \alpha l_0 \Delta T}$ $\boxed{l = l_0(1+\alpha \Delta T)}$

$\Delta T = T - T_0$

coefficient of linear expansion $\left[\frac{1}{C°}\right]$

coefficient of volume expansion $\left[\frac{1}{C°}\right]$

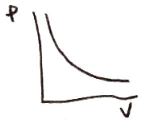
NOTE: ring expands linearly by diameter

Volume Expansion: $\boxed{\Delta V = \beta V_0 \Delta T}$ $\boxed{V = l_0(1+\alpha\Delta T)W_0(1+\alpha\Delta T)H_0(1+\alpha\Delta T)}$

for isotropic: $\Delta V \approx (3\alpha)V_0\Delta T$

(17.6)

Boyle's Law
$V \propto \frac{1}{P}$



Charles's Law
$V \propto T$

Gay-Lussac's Law
$P \propto T$

(17.7)
mole [n]: amount of substance containg same atoms in 12g of C

$$n \text{ (moles)} = \frac{\text{mass (g)}}{\text{molecular mass } \left(\frac{g}{mol}\right)}$$

Ideal Gas Law: $\boxed{PV = nRT}$ ← temperature [K]

moles

universal gas constant: $8.314 \frac{J}{mol \cdot K} = 1.99 \frac{cal}{mol \cdot K}$

Standard Temperature Pressure (STP): $T = 273$ K $P = 1$ atm $= 101.3$ kPa

(17.9)
Avagadro's number: number of molecules in one mole

$N_A = 6.022 \times 10^{23}$

$\boxed{PV = nN_A kT = NkT = \frac{N}{N_A}RT}$

$N = nN_A$

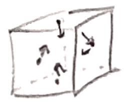
Boltzmann Constant $\left(\frac{R}{N_A}\right) = 1.38 \times 10^{-23} \frac{J}{K}$

18 Kinetic Theory of Gases

(18.1)

Ideal Gas Law assumptions:

1. Large number of molecules N, mass m, moving in random directions different speeds
2. Molecules on average far apart from each other
3. Obey classical mechanics, interact when they collide
4. Collisions w/ wall or molecules elastic.



Average translational kinetic energy is proportional to temperature

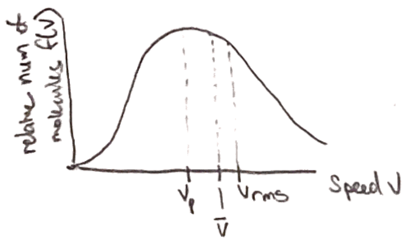
$$\boxed{\bar{K} = \frac{1}{2} m \bar{v^2} = \frac{3}{2} kT}$$

root-mean-square speed: how fast molecules move on avg

$$V_{rms} = \sqrt{\bar{v^2}} = \sqrt{\frac{3kT}{m}}$$

(18.2)

Maxwell Distribution of speeds: probable distribution of speeds in gas w/ N molecules



$$\boxed{f(v) = 4\pi N \left(\frac{m}{2\pi kT}\right)^{\frac{3}{2}} v^2 e^{-\frac{1}{2}\frac{mv^2}{kT}}}$$

$$\boxed{\int_0^{\infty} f(v)\, dv = N}$$

most probable $V_p = \sqrt{2\frac{kT}{m}} \approx 1.41\sqrt{\frac{kT}{m}}$

avg speed $\bar{V} = \sqrt{\frac{8}{\pi}\frac{kT}{m}} \approx 1.60\sqrt{\frac{kT}{m}}$

rms speed $V_{rms} = \sqrt{3\frac{kT}{m}} \approx 1.73\sqrt{\frac{kT}{m}}$

Expected Value of F

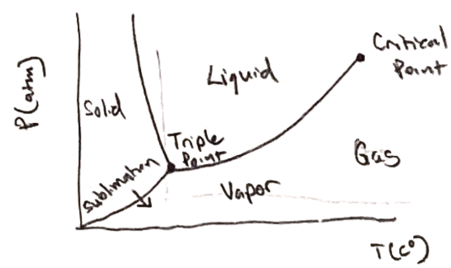
$$\langle F(x) \rangle = \int F(x) p(x)\, dx$$

(18.3)

Liquid crystals between liquid and solid

Sublimation: solid → vapor

Phase diagram:



(18.4)

Evaporation: liquid → gas

Condensation: gas → liquid

Saturated vapor pressure: equilibrium between liquid, vapor

Boiling when saturated vapor pressure = external pressure

relative humidity: $\frac{\text{partial pressure } H_2O}{\text{saturated vapor pressure } H_2O} \times 100$

Van der Waals

(18.5) Van der Waals takes into account 1) finite size of molecules, 2) range of forces between molecules > size molecules

$$\boxed{P = \frac{RT}{\left(\frac{V}{n}\right) - b} - \frac{a}{\left(\frac{V}{n}\right)^2}}$$

$$\boxed{\left(P + \frac{a}{\left(\frac{V}{n}\right)^2}\right)\left(\frac{V}{n} - b\right) = RT}$$

a, b different for different gases

19 Heat and First Law of Thermodynamics

Conversions

1000 cal = 1 kcal = 4184 J

1 L·atm = 101.33 J

$F = PA$

19.1

Unit of heat: calorie (cal), kilocalorie (kcal): amount of heat necessary to raise 1 g of water by 1°C

heat: energy transferred from one object to another bc difference in temperature

19.2

Internal Energy: sum total of all energy of all molecules in an object

Temperature	vs. Internal Energy	vs. Heat
avg kinetic energy of individual molecules	total energy of all the molecules	transfer of energy from one object to another

Internal Energy (E_{int}): $\boxed{E_{int} = \frac{3}{2} nRT}$ for monotomic gas

19.3

Heat Q to change temp of material: $\boxed{Q = mc\,\Delta T}$

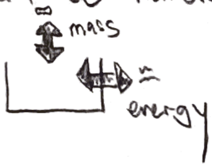
specific heat $\left[\frac{J}{kg \cdot C}\right]$: specific to material

19.4

Systems:

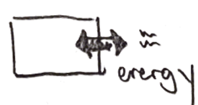
Open System

mass can be transferred
energy can be transferred



Closed System

mass cannot be transfered
energy can be trasfered



Closed (Isolated)

mass cannot be transfered
energy canot be transfered

In isolated environments: heat lost = heat gained

$$Q_1 = Q_2$$

19.5

Energy is involved in change of phase

Heat of fusion (L_F): heat to change 1 kg substance solid → liquid
Heat of vaporization (L_V): heat to change 1 kg substance liquid → vapor $\left[\frac{kJ}{kg}\right]$

$\boxed{Q = mL}$ ← Latent heat

energy is needed to break attractive forces

(19.6)

First Law of Thermodynamics: $\boxed{\Delta E_{int} = Q - W}$

net heat added to system; net work done by system

Heat added : +
Heat lost : −
Work on System : −
Work by System : +

Thermodynamics: $\boxed{\Delta E_{int} = \frac{d}{2} nR \Delta T}$ — degrees of freedom

$$\boxed{W = \int_{V_1}^{V_2} P\,dV}$$

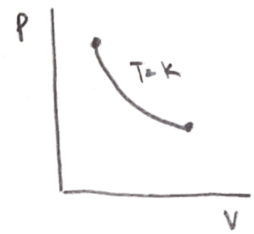
$$\boxed{\Delta E_{int} = Q - W}$$

(19.7)

1) Isothermic ($\Delta T = 0$)

$\Delta E_{int} = 0$, so $Q = W$

$$W = nRT \ln\left(\frac{V_2}{V_1}\right)$$

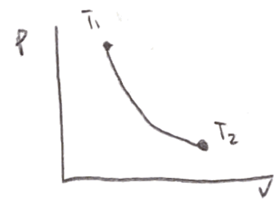


2) Adiabatic ($Q = 0$)

$$\Delta E_{int} = -W = \frac{d}{2} nR \Delta T$$

$$PV^{\gamma} \Rightarrow \frac{P_1}{P_2} = \left[\frac{V_2}{V_1}\right]^{\gamma} = \left[\frac{T_1}{T_2}\right]^{\frac{\gamma}{\gamma-1}}$$

Monotomic : $\gamma = 5/3$ $d=3$
Diatomic : $\gamma = 1.4$ $d=5$
Triatomic : $\gamma \approx 1.31$ $d=7$

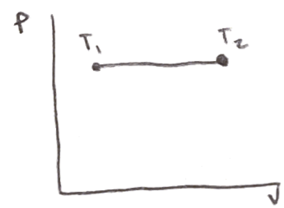


3) Isobaric ($\Delta P = 0$)

if ideal: $W = nRT_2\left(1 - \frac{V_1}{V_2}\right)$

else: $W = P\Delta V$

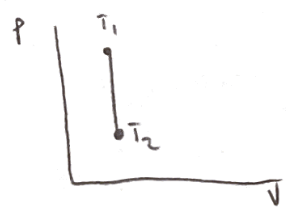
$$Q = \Delta E_{int} + P\Delta V$$



4) Isovolumetric ($\Delta V = 0$)

$W = 0$

$$\Delta E_{int} = Q = \frac{d}{2} nR \Delta T$$



x) For problems, set up table

	a	b	c
P			
V			
T			

	ab	bc	ca
ΔE_{int}			
Q			
W			

(19.8)

Molar Specific Heats (C_v, C_p): heat required to raise 1 mol of gas by 1 C° const volume, temp

Constant Volume

$$\boxed{Q = n\,C_v\,\Delta T}$$

$C_v = M\,c_v$ ← specific heat @ const volume; molecular mass $\left[\frac{g}{mol}\right]$

$\Delta E_{int} = Q$

$$C_v = \frac{3}{2} R$$

Constant Pressure

$$\boxed{Q = n\,C_p\,\Delta T}$$

$C_p = M\,c_p$

- More heat is required, need work

$Q_p - Q_v = P\Delta V$

$$C_p - C_v = R$$

19.13

• Heat transfer via:

1) Conduction: hot to cold via molecular collisions

$$\boxed{\frac{dQ}{dt} = -kA\frac{dT}{dx}}$$

thermal conductivity constant (specific to metal)

2) Convection: heat flows by mass movement of molecules

3) Radiation: heat by electromagnetic waves

Stephan-Boltzmann eq: $\boxed{\frac{\Delta Q}{\Delta t} = \epsilon \sigma A T^4}$

area of emitting object

emissivity: [0,1]
characteristic of surface

Stephan-Boltzmann constant:
$\sigma = 5.67 \times 10^{-8} \frac{W}{m^2 \cdot K^4}$

Sun radiation heat: $\frac{\Delta Q}{\Delta t} = \left(1000 \frac{W}{m^2}\right) \epsilon A \cos\theta$

20 Second Law of Thermodynamics

20.2

Heat Engines produce work from thermal energy

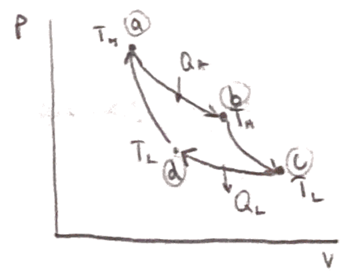
$\boxed{Q_H = W + Q_L}$ $\quad \Delta E_{int} = 0$

Efficiency (e): ratio of work done to heat input

$$\boxed{e = \frac{W}{Q_H} = 1 - \frac{Q_L}{Q_H}}$$

20.3

Carnot's Engine is an idealized reversible cycle



ab 1) expanded isothermally, Q_H added

bc 2) expanded adiabatically, temperature reduced to T_L

cd 3) compressed isothermally, Q_L removed

da 4) compressed adiabatically, temperature raised to T_H

$\boxed{\frac{Q_L}{Q_H} = \frac{T_L}{T_H}}$ $\quad \boxed{e_{ideal} = 1 - \frac{T_L}{T_H}}$ ← [K]

20.4

Coefficient of Performance (COP): heat removed for work done refridgerator

$$COP = \frac{Q_L}{W} = \frac{Q_L}{Q_H - Q_L}$$

$$COP_{ideal} = \frac{T_L}{T_H - T_L}$$

20.5

Entropy [S] is a state variable, measure of order or disorder

20.6

$$\Delta S = \Delta S_H + \Delta S_L = -\frac{Q}{T_{HM}} + \frac{Q}{T_{LM}}$$

For hot → cold, $\Delta S > 0$

$$\boxed{\Delta S = \frac{Q}{T} = \int_{T_1}^{T_2} \frac{mc\,dT}{T} = mc \ln\left(\frac{T_2}{T_1}\right)}$$

Entropy of isolated system never decreases.

$$\Delta S = \Delta S_{syst} + \Delta S_{env} > 0$$

20.7

Second Law of Thermodynamics: Natural Processes tend to move toward a state of greater disorder

20.8

Energy eventually becomes degraded and unavailable to do useful work

Tips

* Remember to change T to Kelvins

Units / Conversions

Force	Newton [N]	$\frac{m \cdot kg}{s^2}$	$F = PA$ (Pressure) $\quad F = \frac{p}{t}$ (momentum, $p = mv$)
Pressure	Pascal [Pa]	$\frac{kg}{m \cdot s^2}$	101,325 Pa = 1 atm = 760 mmHg = 14.7 psi
Energy	Joule [J]	$\frac{m^2 \cdot kg}{s^2}$	1000 J = 1 kJ , 101.33 J = 1 L·atm
Energy	Calorie [cal]	$\frac{m^2 \cdot kg}{s^2}$	1000 cal = 1 kcal , 1 cal = 4.184 J
Power	Watt [W]	$\frac{J}{s} = \frac{kg \cdot m^2}{s^3}$	745.7 W = 1 hp , power → $P = \frac{W}{\Delta t}$
Gas constant	[R]		$8.314 \frac{J}{mol \cdot K} = 1.99 \frac{cal}{mol \cdot K}$

$KE = \frac{1}{2}mv^2$ $\qquad \Delta K + \Delta U + \Delta E_{int} = Q - W$