

Review for Exam 4: Chapters 17 – 20

Physics 1425, Section 1 (Dr. Yost)

Exam 4 will cover chapters 17 through 20 except chapter 18. You may use any calculator for this exam. You will be given any constants you need, and you may bring this set of notes. You should remember any other relations you may need, including geometric relationships and trigonometric identities. Physics is cumulative, so you may need concepts from earlier chapters as well.

Chapter 17: Temperature, Thermal Expansion, and the Ideal Gas Law

Concepts: microscopic states, macroscopic states, state variables, temperature, thermometers, Celcius, Fahrenheit, thermal equilibrium, zeroth law of thermodynamics, thermal expansion, equation of state, absolute zero, Kelvin scale, Ideal Gas Law, atoms, molecules, atomic or molecular mass, Avogadro's number, moles

Equations:

Zeroth law of thermodynamics: two systems in thermal equilibrium with a third system are in thermal equilibrium with each other.

Linear expansion: $\Delta L = \alpha L_0 \Delta T$.

Volume expansion: $\Delta V = \beta V_0 \Delta T$.

The factors α and β are the coefficients of linear and bulk expansion, respectively. Normally, $\beta \approx 3\alpha$. These coefficients have dimensions of $(C^\circ)^{-1}$.

The ideal gas law is $PV = nRT = NkT$, where P and V are the pressure and volume of the gas, T is the absolute temperature, n is the number of moles, N is the number of atoms or molecules, and the constants R and k are universal:

Ideal gas constant: $R = 8.315 \text{ J}/(\text{mol} \cdot \text{K})$,

Boltzmann constant: $k = R/N_A = 1.38 \times 10^{-23} \text{ J/K}$.

The number of atoms or molecules in a mole is Avogadro's number, $N_A = 6.02 \times 10^{23}$.

The ideal gas law applies to any gas which is sufficiently dilute that the atoms or molecules can be considered as non-interacting.

Units:

Temperature is usually measured in Celcius and Kelvin scales, with T (Kelvin) =

T (Celcius) + 273.15K.

Atomic mass units (u) are the number of grams in a mole of a substance. This is defined so that ^{12}C has an atomic mass of exactly 12 u. This determines the number of atoms in a mole to be $N_A = 6.02 \times 10^{23}$. In terms of kilograms, $1 \text{ u} = 1.6605 \times 10^{-27} \text{ kg}$.

Chapter 18: Kinetic Theory of Gasses

We skipped this chapter.

Chapter 19: Heat and the First Law of Thermodynamics

Concepts: Heat, energy transfer, calories, thermal energy, internal energy, specific heat, calorimetry, latent heat, first law of thermodynamics, closed system, open system, isolated system, isothermal, quasistatic, isochoric, isobaric, adiabatic, molar specific heats of gases, equipartition of energy, heat transfer, conduction, convection, radiation, solar constant

Equations:

The internal energy of an ideal gas is $U = \frac{d}{2}nRT$, where d is the number of degrees of freedom of the gas. For a monatomic gas, $d = 3$ always. For a diatomic gas, rotations are normally possible at room temperature, raising d to 5, but at very low temperatures, the rotational degrees of freedom can freeze out, leaving $d = 3$ again. At very high temperatures, vibrational modes may raise d to 7.

The amount of heat Q needed to change the temperature of a mass m of a material by ΔT is given by $Q = mc\Delta T$, where c is the specific heat.

The specific heat of water is $1 \text{ cal/g}^\circ\text{C}$.

The molar specific heats are defined by $C = Mc$, where M is the molecular mass (grams per mole). The molar specific heat of an ideal gas at constant volume is $C_v = \frac{d}{2}R$, while at constant pressure it is $C_p = (\frac{d}{2} + 1)R$, where d is the number of degrees of freedom ($d = 3$ for monatomic gases). The internal energy of a gas is $U = nC_vT$. The internal energy always depends only on the temperature. Note that it is always proportional to C_v , regardless of the process. The first law of thermodynamics states that for a closed system (no material enters or leaves, but energy may), the change in internal energy is

$$\Delta U = Q - W. \tag{1}$$

where Q is the heat flow into the system, and W is the work done *by* the system.

The work done by a gas in a closed system is $W = \int PdV$. This is the area under the curve on a PV diagram. The net work done in a closed cycle on a PV diagram is the area inside the curve traversed by the system through the cycle.

For ideal gasses:

Process	Definition	ΔU	W	Q
Isochoric	$\Delta V = 0$	$nC_v\Delta T$	0	ΔU
Isobaric	$\Delta P = 0$	$nC_v\Delta T$	$P\Delta V = nR\Delta T$	$nC_p\Delta T$
Isothermal	$\Delta T = 0$	0	$nRT \ln(V_2/V_1)$	W
Adiabatic	$Q = 0$	$-W$	$\frac{P_1V_1}{\gamma-1} [1 - (V_1/V_2)^{\gamma-1}]$	0

- Isothermal process: $PV = \text{constant}$.
- Adiabatic process: $PV^\gamma = \text{constant}$ where $\gamma = C_P/C_V$.

If a system is taken through a series of PV transformations that make a closed cycle, ending at the initial conditions, then $\Delta U = 0$ for the complete cycle, and the net heat input equals the net work done by the system. $Q = W$ is the area enclosed by the curve on a PV diagram.

The heat conducted per unit time across a distance l separating two temperatures T_1 and T_2 , through a cross-sectional area A is

$$\frac{dQ}{dt} = \frac{kA}{l}(T_1 - T_2) \quad (2)$$

where k is the thermal conductivity.

Heat can be transferred by electromagnetic radiation. An object with surface area A at absolute temperature T radiates energy

$$\frac{dQ}{dt} = e\sigma AT^4 \quad (3)$$

where e is a number between 0 and 1 called the “emissivity”, and $\sigma = 5.67 \times 10^{-8} \text{ W/m}^2 \cdot \text{K}^4$ is the Stefan-Boltzmann constant. Objects with high emissivity are the best absorbers and also the best emitters. Such objects are black, and perfect absorbers ($e = 1$) are called blackbodies. The radiation at room temperature is mostly in the infrared wavelengths. Shiny objects are the worst absorbers since they reflect most of the radiation striking them. Therefore, they are also the worst emitters. A perfect reflector would have emissivity 0.

If the body at temperature T_1 is surrounded by an environment at temperature T_2 , then the energy radiated into the environment is

$$\frac{dQ}{dt} = e\sigma A(T_1^4 - T_2^4). \quad (4)$$

The Earth receives a power per unit area of 1350 W/m^2 from the sun. This is called the solar constant. When the sun is at an angle θ in the sky with respect to

the zenith (overhead), the power received is $1350 \text{ W/m}^2 \cos \theta$. About 74% of this power actually reaches the surface of the earth (1000 W/m^2), and a fraction e will be absorbed by an object with emissivity e .

Units:

1 calorie (cal) = 4.186 J is the amount of heat needed to raise the temperature of 1 g of water by 1°C . 1 kcal is also called a “large calory” (Cal). Food calories are kilocalories (the amount of energy which may be obtained by breaking the chemical bonds in the food).

Chapter 20: Second Law of Thermodynamics

Concepts: heat engines, direction of heat flow, efficiency, second law of thermodynamics, reversible and irreversible processes, Carnot cycle, refrigeration, entropy, order and disorder, availability of energy, statistical nature of entropy, third law of thermodynamics, absolute zero

Equations:

A heat engine takes in heat Q_H at a high temperature, performs some work W , and exhausts heat Q_L at a lower temperature. Conservation of energy (the first law of thermodynamics) shows that $W = Q_H - Q_L$. The efficiency of the engine is defined to be $e = W/Q_H$.

The Carnot cycle has the maximum efficiency possible. It consists of only reversible processes. If it operates between absolute temperatures T_H and T_L , the efficiency is $e_{\text{Carnot}} = 1 - T_H/T_L$.

A refrigerator or heat pump is a heat engine operated in reverse. Work is put in instead of taken out. The coefficient of performance $\text{CP} = Q_L/W$, is the measure of efficiency. The best possible coefficient of performance is obtained with a Carnot cycle: $\text{CP}_{\text{Carnot}} = T_L/(T_H - T_L)$.

If an infinitesimal quantity of heat dQ flows into a system at temperature T , the entropy S of the system changes by dQ/T . For any *reversible* process between states a and b , the change in entropy is

$$\Delta S_{ab} = \int_a^b \frac{dQ}{T}. \quad (5)$$

This does not depend on the process, so the entropy is purely a property of the states, *i.e.* a state variable. To calculate the change of entropy during an irreversible process, you may apply the previous equation to any reversible process which connects the same initial and final state. In any natural process, $\Delta S > 0$ when the entire environment is included. Adiabatic processes are also isentropic, $\Delta S = 0$.

For ideal gasses:

Process	ΔS
Isochoric	$nC_v \ln \left(\frac{T_2}{T_1} \right)$
Isobaric	$nC_p \ln \left(\frac{T_2}{T_1} \right)$
Isothermal	$nR \ln \left(\frac{V_2}{V_1} \right)$
Adiabatic	0

The second law of thermodynamics can be stated in a variety of equivalent ways:

1. Heat flows spontaneously from a hot object to a cold one, but not in reverse.
2. No device can transform a given amount of heat completely into work.
3. No device is possible whose sole effect is to transfer heat from a system at a lower temperature to a system at a higher temperature.
4. The entropy of an isolated system never decreases.
5. Natural processes move toward a state of greater disorder, when the entire environment is taken into account.

The third law of thermodynamics states that absolute zero cannot be reached by any finite number of thermodynamic processes. A consequence of the second and third law together is that there can be no 100% efficient heat engine or refrigerator. An alternative statement of the third law is that $S = 0$ at $T = 0$.