



# General Physics 1 12 Q2 Mod7 Ideal Gases and the Laws of Thermodynamics Version 2

Science Technology Engineering Mathematics (Titay National High School)



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# GENERAL PHYSICS 1

## Quarter 2 - Module 7

### Ideal Gases and The Laws of Thermodynamics



SENIOR HIGH SCHOOL



Department of Education • Republic of the Philippines



## General Physics 1 – Grade 12

### Alternative Delivery Mode

#### Quarter 2 – Module 7: Ideal Gases and The Laws of Thermodynamics First Edition, 2020

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Published by the Department of Education – Division of Cagayan de Oro

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Printed in the Philippines by  
Department of Education – Bureau of Learning Resources (DepEd-BLR)  
Office Address : Fr. William F. Masterson Ave Upper Balulang, Cagayan de Oro  
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**Senior High School**

# **General Physics 1**

**Quarter 2 - Module 7:**

## **Ideal Gases and The Laws of Thermodynamics**

This instructional material was collaboratively developed and reviewed by educators from public and private schools, colleges, and or/universities. We encourage teachers and other education stakeholders to email their feedback, comments, and recommendations to the Department of Education at [action@deped.gov.ph](mailto:action@deped.gov.ph).

We value your feedback and recommendations.

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## Module 7

# Ideal Gases and The Laws of Thermodynamics

### What This Module is About

This module provides you with scientific knowledge and skills about Ideal Gases and The Laws of Thermodynamics. An Ideal gas is one in which the molecules of the gas are treated as point particles that do not interact with one another. The Ideal Gas Law gives the relationship among pressure, volume, temperature, and number of molecules in an ideal gas. The key properties of gases are the thermodynamic quantities of temperature, pressure, and volume, which is why this chapter is with the others on Thermodynamics. Thermodynamics is a branch of physics that deals with heat, work, and temperature, and their relation to energy and other physical properties of matter. The laws of thermodynamics describe how the energy in a system changes and whether the system can perform useful work on its surroundings. The lessons in this module are necessary in studying other concepts of thermal physics in the next modules.

The following are the lessons contained in this module:

-  **Lesson 1- Ideal Gas Law**
-  **Lesson 2- First Law of Thermodynamics**
-  **Lesson 3- Second Law of Thermodynamics**



### What I Need to Know

After going through this module, you are expected to:

1. Enumerate the properties of an ideal gas. **(STEM\_GP12GLTIh-57)**
2. Solve problems involving ideal gas equations in contexts such as, but not limited to, the design of metal containers for compressed gases. **(STEM\_GP12GLTIh-58)**
3. Interpret PV diagrams of a thermodynamic process. **(STEM\_GP12GLTIh-60)**
4. Compute the work done by a gas using  $dW = PdV$ . **(STEM\_GP12GLTIh-61)**
5. State the relationship between changes internal energy, work done, and thermal energy supplied through the First Law of Thermodynamics. **(STEM\_GP12GLTIh-62)**
6. Differentiate the following thermodynamic processes and show them on a PV diagram: isochoric, isobaric, isothermal, adiabatic, and cyclic. **(STEM\_GP12GLTIh63)**
7. Calculate the efficiency of a heat engine. **(STEM\_GP12GLTIh-67)**
8. Describe reversible and irreversible processes. **(STEM\_GP12GLTIh-68)**
9. Explain how entropy is a measure of disorder. **(STEM\_GP12GLTIh-69)**







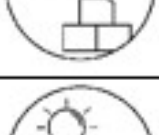

10. State the 2nd Law of Thermodynamics. **(STEM\_GP12GLTIi-70)**
11. Calculate entropy changes for various processes e.g., isothermal process, free expansion, constant pressure process, etc. **(STEM\_GP12GLTIi-71)**

## How to Learn from this Module

To achieve the learning competencies cited above, you are to do the following:

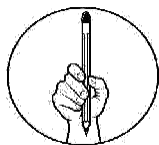
- Take your time reading the lessons carefully.
- Follow the directions and/or instructions in the activities and exercises diligently.
- Answer all the given tests and exercises.

## Icons of this Module

	What I Need to Know	This part contains learning objectives that are set for you to learn as you go along the module.
	What I know	This is an assessment as to your level of knowledge to the subject matter at hand, meant specifically to gauge prior related knowledge
	What's In	This part connects previous lesson with that of the current one.
	What's New	An introduction of the new lesson through various activities, before it will be presented to you
	What is It	These are discussions of the activities as a way to deepen your discovery and understanding of the concept.
	What's More	These are follow-up activities that are intended for you to practice further in order to master the competencies.
	What I Have Learned	Activities designed to process what you have learned from the lesson
	What I can do	These are tasks that are designed to showcase your skills and knowledge gained, and applied into real-life concerns and situations.







## What I Know

**MULTIPLE CHOICE: Directions:** Read and understand each item and choose the letter of the correct answer. Write your answers on a separate sheet of paper.

1. Which one of the following statements is not consistent with the kinetic-molecular theory of gases?
  - a. Individual gas molecules are relatively far apart.
  - b. The actual volume of the gas molecules themselves is very small compared to the volume occupied by the gas at ordinary temperatures and pressures.
  - c. The average kinetic energies of different gases are different at the same temperature.
  - d. There is no net gain or loss of the total kinetic (translational) energy in collisions between gas molecules.
  - e. The theory explains most of the observed behavior of gases at ordinary temperatures and pressures.
  
2. Which statement best describe gases?
  - a. Gases consist of molecules that have enough spatial separation that the intermolecular bonding characteristics of liquids and solid is absent.
  - b. Gases are almost incompressible. In liquids molecules are close to each other.
  - c. Gases are characterized by structural rigidity and resistance to a force applied to the surface.
  - d. Gases are in the state in which matter adapts to the shape of its container but varies only slightly in volume
  
3. Which of the following term does not involve in ideal gas law?
  - a. Pressure
  - b. Volume
  - c. Temperature
  - d. Time
  
4. Which mathematical formula summarizes the Ideal Gas Law?
  - a.  $PV = Dt/r$
  - b.  $PV = ma$
  - c.  $PV = nRT$
  - d.  $PV = mgh$
  
5. A real gas most closely approaches the behavior of an ideal gas under conditions of:
  - a. high P and low T
  - b. low P and high T
  - c. low P and T
  - d. STP
  
6. Which of the following statements best describe the first law of thermodynamics?
  - a. Energy can be changed from one form to another, but it cannot be created or destroyed.
  - b. Processes that involve the transfer or conversion of heat energy are irreversible.
  - c. It is concerned with the limiting behavior of systems as the temperature approaches absolute zero.
  - d. The total entropy of an isolated system can never decrease over time.

7. All the following statements define the second law of thermodynamics except,
- Processes that involve the transfer or conversion of heat energy are irreversible.
  - The total entropy of an isolated system can never decrease over time.
  - Energy can be changed from one form to another, but it cannot be created or destroyed.
  - Heat is a form of energy, and thermodynamic processes are therefore subject to the principle of conservation of energy.

8. Which one of the following definitions best describes the concept of work?
- the flow of energy from one object or substance to another due to a difference in temperature
  - the flow of energy from one body to another through uniform molecular motion
  - the force associated with molecular motion
  - the random motion of molecules in a gas at low pressure

9. Consider the following thermodynamic properties.
- work done on a system
  - heat absorbed
  - entropy
  - enthalpy

Which of these properties are state functions?

- a. I and II only                      b. I and III only                      c. I only                      d. III and IV only

10. Which of the following statements will always apply when a reversible chemical reaction has attained equilibrium?
- All reactants will convert to products
  - The reaction proceeds alternately in the forward and reverse directions
  - The Gibbs free energy of the system reaches a minimum
  - The forward reaction will dominate over the reverse reaction.

## Lesson

## 1

## Ideal Gas Law

**What's In**

We are surrounded by an ocean of gas—the atmosphere—and many of the properties of gases are familiar to us from our daily activities. Heated gases expand, which can make a hot air balloon rise or cause a blowout in a bicycle tire left in the sun on a hot day. Gases have played an important part in the development of chemistry. In the seventeenth and eighteenth centuries, many scientists investigated gas behavior, providing the first mathematical descriptions of the behavior of matter. In this chapter, we will examine the relationships between gas temperature, pressure, amount, and volume. We will study a simple theoretical model and use it to analyze the experimental behavior of gases.

**What I Need to Know**

After this lesson, you should be able to:

1. Enumerate the properties of an ideal gas.
2. Solve problems involving ideal gas equations in contexts such as, but not limited to, the design of metal containers for compressed gases.
3. Interpret PV diagrams of a thermodynamic process.
4. Compute the work done by a gas using  $dW = PdV$ .

**What's New****Activity 7.1 Learning about the Ideal Gas Law****I. Boyle's Law: Balloon and Syringe**

**Objectives:** To be able to understand Ideal Gas Law concept by demonstrating Boyle's Law.

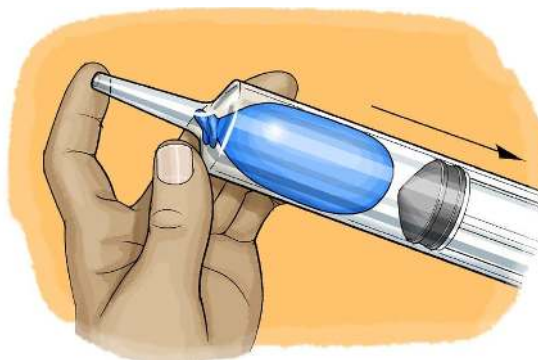
**Material:** Balloon, 50-milliliter (mL) syringe or medicine dropper

Situation:

As the ideal gas law is based on other laws, these activities will look at Avogadro's, Boyle's, and Charles laws. **Boyle's law** states that pressure and volume are inversely proportional (assuming the number of molecules and temperature are constant). So, if you increase the pressure, the volume is reduced. This first activity will help demonstrate Boyle's law.

Directions:

- Add a tiny bit of air to the balloon (not too much, it will need to fit in the syringe) and tie the balloon.
- Place the balloon into the syringe.
- Add the plunger to the syringe and then push the piston towards the balloon.
- Students should write down their observations. (The balloon does not change size because the atmospheric pressure remains the same.)
- Now, place a finger over the opening of the syringe and push the piston towards the balloon.
- Record your observations. (When the air is prevented from escaping, the pressure increases as the plunger is pushed down. This results in a reduction in the balloon's volume.)



## II. **Charles' Law: The Can Crasher**

Objective: To be able to understand Ideal Gas Law concept by demonstrating Boyle's Law.

Materials: Small Soda Can, Heating Pan, Tongs/Pot holder

Situation: **Charles' law** says that as the temperature increases, the volume will increase (at a constant pressure). The can crusher experiment demonstrates Charles's Law, the basic principle that gases expand when heated and contract when cooled. This activity helps students visualize this law.

Direction:

1. You will need a small soda can; fill it with about half an ounce of water.
2. Boil the can in a pan of water for about a minute, and you will notice vapor steaming from the opening of the soda can.
3. Using tongs, grab the can and place it upside down in a bowl of cold water. The can will crush immediately.
4. The water vapor exits the can immediately, and cold water condenses the vapor, leaving the can at very low pressure inside. It happens so quickly that the normal air pressure outside the can crushes the exterior of the can.

Hint: As the water heats, it will cause the balloon to expand. (An increase in temperature results in an increase in volume per the law.)

### **III. Avogadro's Law: Water Bottle Balloon**

Objective: To be able to understand Ideal Gas Law concept by demonstrating Avogadro's Law.

Materials: Empty Soda Bottle, Balloon

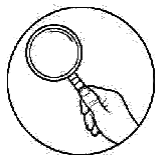
Situation:

Avogadro's Law: The volume of a gas is directly proportional to the amount of gas. This experiment will show you Avogadro's Law.

Direction:

1. Find an empty glass bottle, such as a soda bottle, and fill it with about an ounce of water.
2. In a pan of water, heat the bottle until the water inside reaches a boil. Stretch balloon over the mouth of the bottle.
3. Record your observation.





## What Is It

Activity 7.1 walks us through the basic concepts of the different properties and characteristics of gases that will help us understand its behavior following then our discussion on the unified idea of the Ideal Gas Law.

### Properties of Gases

1. Gases have no definite shape and volume.
2. Gases diffuse rapidly.
3. Gases can be readily compressed.
4. Gases have densities much lower than solids and liquids.

### The Kinetic Molecular Theory

The word “kinetic” means motion. Objects in motion have energy called kinetic energy. The Kinetic Molecular Theory describes behavior of matter in terms of particles in motion. When applied to gases, the basic postulates are the following:

1. Gases are composed of atoms or molecules which are widely separated from one another.
2. Gas molecules move at a very high speeds, traveling in straight paths but in random directions.
3. Gas molecules collide with one another and with the walls of its container.
4. The average kinetic energy of the molecules is determined by the gas temperature.

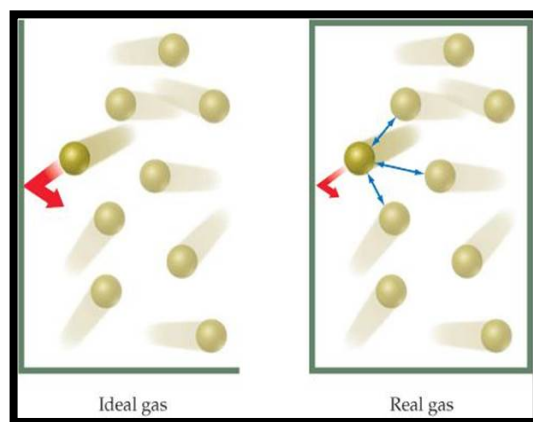
### Ideal Gases versus Real Gases

The postulates of the kinetic molecular theory apply to all gases. However, they are true only for an ideal gas. Hence, an **ideal gas** is one in which intermolecular forces are negligible and gas molecules occupy a negligible volume.

Real gases behave almost as ideal gas at ordinary temperature and pressures.

### Ideal Gases vs. Real Gases

Ideal gases	Real Gases
Ideal gases obey all gas laws under all conditions of temperature and pressure.	Real gases obey gas laws only at low pressures and high temperature.
The volume occupied by the molecules is negligible as compared to the total volume occupied by the gas.	The volume occupied by the molecules is not negligible as compared to the total volume of the gas.
The force of attraction among the molecules are negligible.	The force of attraction are not negligible at all temperatures and pressures.



- The relationship between Pressure, Temperature, Volume, and Amount of a gas are described and how these relationships can be combined to give a general expression that describes the behavior of a gas – the ideal gas law.
- The Ideal Gas Law is very simply expressed:  **$PV=nRT$** .
- From which simpler gas laws such as Boyle's, Charles's, Avogadro's and Amonton's law be derived.
- The four gas variables are: *pressure (P)*, *volume (V)*, *number of moles of gas (n)*, and *temperature (T)*. Lastly, the constant in the equation shown below is R, known as the *gas constant*.
- The universal value of STP is 1 atm (pressure) and 0° C. Note that this form specifically stated 0° C degree, not 273 Kelvin, even though you will have to convert into Kelvin when plugging this value into the Ideal Gas equation or any of the simple gas equations.
- In STP, 1 mole of gas will take up 22.4 L of the volume of the container.

### Units of P, V and T

The table below lists the different units for each property.

Factor	Variable	Units
Pressure	P	atm Torr Pa mmHg
Volume	V	L m <sup>3</sup>
Moles	n	mol
Temperature	T	K
Gas Constant	R*	see Values of R table below

### Pressure Units

Use the following table as a reference for pressure.

Common Units of Pressure		
Unit		Equivalent to 1 atm
Atmosphere	This document is available on	1 atm
Millimeter of Mercury	Downloaded by aldrin alciso (aldrinalciso2@gmail.com)	760 mmHg
Torr		760 Torr



Kilopascal	kPa*	101.326 kPa
Bar	bar	1.01325 bar
Millibar	mb	1013.25 mb

### The Gas Constant (R)

Value of R *WILL* change when dealing with different unit of pressure and volume (Temperature factor is overlooked because temperature will always be in Kelvin instead of Celsius when using the Ideal Gas equation). Only through appropriate value of R will you get the correct answer of the problem.

It is simply a constant, and the different values of R correlates accordingly with the units given. When choosing a value of R, choose the one with the appropriate units of the given information (sometimes given units must be converted accordingly).

Here are some commonly used values of R:

Values of R
0.082057 L atm mol <sup>-1</sup> K <sup>-1</sup>
62.364 L Torr mol <sup>-1</sup> K <sup>-1</sup>
8.3145 m <sup>3</sup> Pa mol <sup>-1</sup> K <sup>-1</sup>
8.3145 J mol <sup>-1</sup> K <sup>-1</sup> *
*note: This is the SI unit for the gas constant

### Example Problems:

1. A 5.0 g of neon is at 256 mm Hg and at a temperature of 35° C. What is the volume?

A. Write down your given information:

$$P = 256 \text{ mmHg}, m = 5.0 \text{ g}, R = 0.0820574 \text{ L} \cdot \text{atm} \cdot \text{mol}^{-1} \text{K}^{-1}, T = 35^\circ \text{ C}$$

B. Unknown:  $V = ?$

C. Convert units as necessary:

$$P: 256 \text{ mmHg} \times (1 \text{ atm} / 760 \text{ mmHg}) = 0.3368 \text{ atm} \quad 256 \text{ mmHg} \times (1 \text{ atm} / 760 \text{ mmHg}) = \underline{0.3368 \text{ atm}}$$

Moles:  $5.0\text{g Ne} \times (1\text{mol}/20.1797\text{g}) = 0.25\text{mol Ne}$   $5.0\text{g Ne} \times (1\text{mol}/20.1797\text{g}) = \mathbf{0.25\text{mol Ne}}$

Temperature:  $35^\circ\text{C} + 273 = 308\text{K}$   $35^\circ\text{C} + 273 = \mathbf{308\text{K}}$

D. Plug in the variables into the appropriate equation.

$$V = (nRT/P)$$
$$V = \frac{(0.25\text{ mol})(0.08206\text{ Latm/Kmol})(308\text{ K})}{(0.3368\text{ atm})}$$
$$V = 19\text{ L}$$

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2. What is a gas's temperature in Celsius when it has a volume of 25 L, 203 mol, 143.5 atm?

A. Given  $P = 143.5\text{ atm}$ ,  $V = 25\text{ L}$ ,  $n = 203\text{ mol}$ ,  $R = 0.0820574\text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\text{ K}^{-1}$

B. Unknown:  $T = ?$

C. Solution:

$$T = \frac{PV}{nR}$$
$$T = \frac{(143.5\text{ atm})(25\text{ L})}{(203\text{ mol})(0.08206\text{ L}\cdot\text{atm/Kmol})}$$
$$T = 215.4\text{ K}$$

$$215.4\text{ K} - 273 = -57.4^\circ\text{C}$$

3. A cylinder of argon gas contains 50.0 L of Ar at 18.4 atm and  $127^\circ\text{C}$ . How many moles of argon is in the cylinder?

A. Given:  $V = 50\text{L}$ ,  $P = 18.4\text{atm}$ ,  $T = 127\text{ degrees Celsius}$

B. Unknown:  $n = ?$

C. Convert units necessary:

$$T = ^\circ\text{C} + 273$$

$$T = 127^\circ\text{C} + 273$$

$$T = 400\text{ K}$$

$$R = 0.0821\text{ liter}\cdot\text{atm/mol}\cdot\text{K}$$

D. Solution:

$$n = \frac{PV}{RT}$$

plug in our values

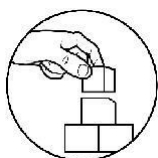
$$n = \frac{(18.4 \text{ atm})(50.0 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(400 \text{ K})}$$

$$n = 28.0 \text{ mol}$$

**Answer**

There are 28.0 moles of argon in the cylinder.

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## What's More

### Activity 7.2: More on Solving Ideal Gas Law Problems

**Directions:** Use the ideal gas law, “PV=nRT”, and the universal gas constant  $R = 0.0821 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol}$  to solve the following problems: K\*mol If pressure is needed in kPa then convert by multiplying by  $101.3 \text{ kPa} / 1 \text{ atm}$  to get  $R = 8.31 \text{ kPa}\cdot\text{L} / (\text{K}\cdot\text{mole})$ . Show your answers legibly, concisely, and completely. Use separate sheet of paper.

1. If I have 4 moles of a gas at a pressure of 5.6 atm and a volume of 12 liters, what is the temperature?
2. If I have an unknown quantity of gas at a pressure of 1.2 atm, a volume of 31 liters, and a temperature of  $87^\circ \text{C}$ , how many moles of gas do I have?
3. If I have 7.7 moles of gas at a pressure of 0.09 atm and at a temperature of  $56^\circ \text{C}$ , what is the volume of the container that the gas is in?
4. If I contain 3 moles of gas in a container with a volume of 60 liters and at a temperature of 400 K, what is the pressure inside the container?



## What I Have Learned

### Activity 7.3: Sum it Up!

Direction: Answer the following questions. Choose the best letter of your answer.

1. Which of the following is not a property of gases?
  - a. Gases have no definite shape and volume.
  - b. Gases diffuse rapidly.
  - c. Gases can be readily compressed.
  - d. Gases have densities much higher than solids and liquids.

2. The word "kinetic" means?
  - a. Stable
  - b. Electric
  - c. Motion
  - d. Compress
3. Which statement does not describe Kinetic Molecular Theory?
  - a. Gases are composed of atoms or molecules which are widely separated from one another.
  - b. Gas molecules move at a very high speeds, traveling in straight paths but in random directions.
  - c. Gas molecules collide with one another and with the walls of its container.
  - d. The average kinetic energy of the molecules is determined by the gas temperature.

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4. Which mathematical equation summarizes the Ideal Gas Law?
  - a.  $PV = mgh$
  - b.  $PV = nRT$
  - c.  $P = \frac{1}{2}mv$
  - d.  $P = ma$
5. If I have 72 liters of gas held at a pressure of 3.4 atm and a temperature of 225 K, how many moles of gas do I have?
  - a. 13.0mol
  - b. 0.4587mol
  - c. 0.0982mol
  - d. 12.6mol
6. If I have 0.275 moles of gas at a temperature of 75 K and a pressure of 1.75 atmospheres, what is the volume of the gas?
  - a. 0.4587L
  - b. 0.0982L
  - c. 0.0567L
  - d. 0.9676L
7. If I have an unknown quantity of gas held at a temperature of 1195 K in a container with a volume of 25 liters and a pressure of 560 atm, how many moles of gas do I have?
  - a. 143.0mol
  - b. 0.01004mol
  - c. 146.0mol
  - d. 0.3457mol
8. If I have 21 moles of gas held at a pressure of 78 atm and a temperature of 900 K, what is the volume of the gas?
  - a. 18.89L
  - b. 20.0L
  - c. 17.89L
  - d. 16.89
9. Which of the following variable is not part of the ideal gas equation?
  - a. Pressure
  - b. volume
  - c. Acceleration
  - d. Temperature
10. All the following describes the properties of an ideal gas except,
  - a. Ideal gas law obeys all gas laws under all conditions of temperature and pressure.

total volume

3.



# First Law of Thermodynamics



## What's In

Earth's weather is driven by thermal energy in the atmosphere. The equatorial regions receive more of the sun's radiation than the polar regions do; so, air moves north to south from the equator toward the poles to distribute the thermal energy more evenly. This transfer of energy, called **convection**, sets up wind currents around the world, carrying clouds and rain as well as air. In the previous lesson, we learned about the different gas laws that makes up the ideal gas law and its characteristics. In this lesson we will examine the nature of heat and the mechanisms of thermal energy transfer. Heat is another form of energy that is transferred into or out of a system. Thus, heat is governed by a more general form of the Law of conservation of energy, known as the **First Law of Thermodynamics**.



## What I Need to Know

After this lesson, you should be able to:

- State the relationship between changes internal energy, work done, and thermal energy supplied through the First Law of Thermodynamics. (STEM\_GP12GLTIh-62)

- 

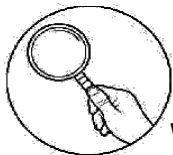
**10**

1. Study Concepts in a Box for three minutes. There is no right or wrong answers for now. Take note of your answers and validate basic concepts you mentioned.

- First Law of Thermodynamics**

  - tells us
    - ENERGY**
      - CANNOT** be Created or Destroyed
        - therefore **TOTAL ENERGY OF AN ISOLATED SYSTEM** is **CONSTANT**
      - can only be converted from **CONSERVED**
  - represents **"Good News"**
    - Deepest & most widely applicable Principles in Science
    - Great Integrating Principles of Science
  - is among
    - One Type to Another
      - e.g. Riding a Bicycle
        - you use the **Chemical Energy in Your Cells (P.E.)**
          - to create
            - Heat in Your Muscles**
            - Motion of Your Legs (K.E.)**
              - leads to **Motion of the Bicycle and Rider (K.E.)**
                - allows you to move **Bicycle & Rider to Top of Hill (P.E.)**
                  - creates **Friction**
                    - Tires** and **Gears** in **Noise**
                    - represent **Loss of Useful Energy**
                  - and then coast down **Fastest (Max K.E.)**
                    - is where you travel **Bottom of Hill**
                      - and then coast up to **Top of a Second, Hill**
                        - is **Lower** than **First Hill**
                          - because
      - applies e.g. to
        - Largest Supercluster of Galaxies**
        - Cells in your Body**
        - Fundamental Particles like Quarks**
      - IN TOTAL remain **Constant**
        - Cellular Energy**
        - Kinetic Energy KE**
        - Potential Energy PE**
        - Heat (Rider, Tires, Brakes, etc.)**





## What Is It

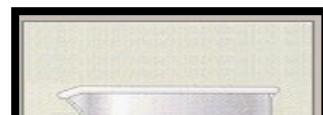
**Thermodynamics** is the branch of physics that deals with the relationships between heat and other forms of energy. In particular, it describes how thermal energy is converted to and from other forms of energy and how it affects matter.

**Thermal energy** is the energy a substance or system has due to its temperature, i.e., the energy of moving or vibrating molecules.

**Heat** is energy transferred between substances or systems due to a temperature difference between them.

As a form of energy, **heat is conserved, it cannot be created or destroyed**. It can, however, be transferred from one place to another. Heat can also be converted to and from other forms of energy. For example, a steam turbine can convert heat to kinetic energy to run a generator that converts kinetic energy to electrical energy. A light bulb can convert this electrical energy to electromagnetic radiation (light), which, when absorbed by a surface, is converted back into heat.

## The System and Surroundings

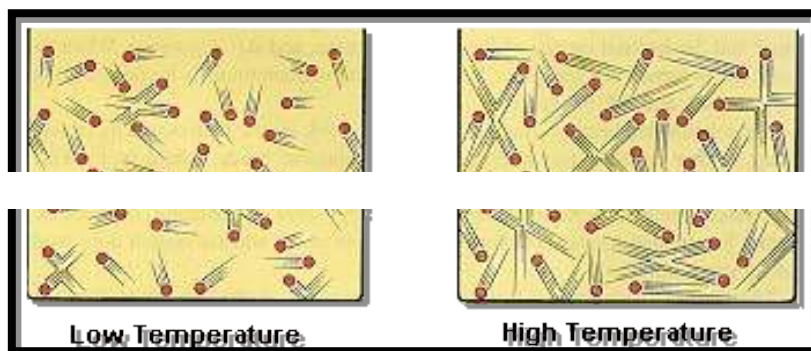


One of the basic assumptions of thermodynamics is the idea that we can arbitrarily divide the universe into a **system** and its **surroundings**. The boundary between the system and its surroundings can be as real as the walls of a beaker that separates a solution from the rest of the universe (as in the figure below).

## Internal Energy

One of the thermodynamic properties of a system is its **internal energy**,  $E$ , which is the sum of the kinetic and potential energies of the particles that form the system. The internal energy of a system can be understood by examining the simplest possible system: an ideal gas. Because the particles in an ideal gas do not interact, this system has no potential energy. The internal energy of an ideal gas is therefore the sum of the kinetic energies of the particles in the gas.

The kinetic molecular theory assumes that the temperature of a gas is directly proportional to the average kinetic energy of its particles, as shown in the figure below.





The internal energy of an ideal gas is therefore directly proportional to the temperature of the gas.

$$E_{\text{sys}} = \frac{3}{2} RT$$

In this equation,  $R$  is the ideal gas constant in joules per mole kelvin (J/mol-K) and  $T$  is the temperature in kelvin.

The internal energy of systems that are more complex than an ideal gas can't be measured directly. But the internal energy of the system is still proportional to its temperature. We can therefore monitor changes in the internal energy of a system by watching what happens to the temperature of the system. Whenever the temperature of the system increases, we can conclude that the internal energy of the system has also increased.

Assume, for the moment, that a thermometer immersed in a beaker of water on a hot plate reads 73.5°C, as shown in the figure on the right. This measurement can only describe the state of the system at that moment in time. It can't tell us whether the water was heated directly from room temperature to 73.5°C or heated from room temperature to 100°C and then allowed to cool.

Temperature is therefore a **state function**. It depends only on the state of the system at any moment in time, not the path used to get the system to that state. Because the internal energy of the system is proportional to its temperature, internal energy is also a state function. Any change in the internal energy of the system is equal to the difference between its initial and final values.

$$\Delta E_{\text{sys}} = E_f - E_i$$

### ***The First Law Of Thermodynamics***

The first law of thermodynamics can be captured in the following equation, which states that the energy of the universe is constant. Energy can be transferred from the system to its surroundings, or vice versa, but it can't be created or destroyed.

$$\text{First Law of Thermodynamics: } \Delta E_{\text{univ}} = \Delta E_{\text{sys}} + \Delta E_{\text{surr}} = 0$$

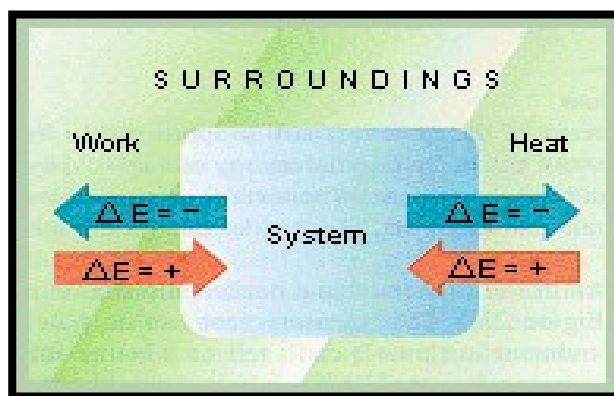
A more useful form of the first law describes how energy is conserved. It says that the change in the internal energy of a system is equal to the sum of the heat gained or lost by the system and the work done by or on the system.

$$\text{First Law of Thermodynamics: } \Delta E_{\text{sys}} = q + w$$

The sign convention for the relationship between the internal energy of a system and the heat gained or lost by the system can be understood by thinking about a concrete example, such as a beaker of water on a hot plate. When the hot plate is turned on, the system gains heat from its surroundings. As a result, both the temperature and the internal energy of the system increase, and  $\Delta E$  is *positive*. When the hot plate is turned off, the water loses heat to its surroundings as it cools to room temperature, and  $\Delta E$  is *negative*.

The relationship between internal energy and work can be understood by considering another concrete example: the tungsten filament inside a light bulb. When work is done on this system by driving an electric current through the tungsten wire, the system becomes hotter and  $\Delta E$  is therefore *positive*. (Eventually, the wire becomes hot enough to glow.) Conversely,  $\Delta E$  is *negative* when the system does work on its surroundings.

The sign conventions for heat, work, and internal energy are summarized in the figure below. The internal energy and temperature of a system decrease ( $\Delta E < 0$ ) when the system either loses heat or does work on its surroundings. Conversely, the internal energy and temperature increase ( $\Delta E > 0$ ) when the system gains heat from its surroundings or when the surroundings do work on the system.



### **The System and Work**

The system is usually defined as the chemical reaction and the boundary is the container in which the reaction is run. In the course of the reaction, heat is either given off or absorbed by the system. Furthermore, the system either does work on its surroundings or has work done on it by its surroundings. Either of these interactions can affect the internal energy of the system.

$$\Delta E_{\text{sys}} = q + w$$

Two kinds of work are normally associated with a chemical reaction: *electrical work* and *work of expansion*. Chemical reactions can do work on their surroundings by driving an electric current through an external wire. Reactions also do work on their surroundings when the volume of the system expands during the course of the reaction. The amount of work of expansion done by the reaction is equal to the product of the pressure against which the system expands times the change in the volume of the system.

$$w = -P\Delta V$$

The sign convention for this equation reflects the fact that the internal energy of the system decreases when the system does work on its surroundings.

### **Enthalpy Versus Internal Energy**

What would happen if we created a set of conditions under which no work is done by the system on its surroundings, or vice versa, during a chemical reaction? Under these conditions, the heat given off or absorbed by the reaction would be equal to the change in the internal energy of the system.

$$\Delta E_{\text{sys}} = q \text{ (if and only if } w = 0\text{)}$$

The easiest way to achieve these conditions is to run the reaction at constant volume, where no work of expansion is possible. At constant volume, the heat given off or absorbed by the reaction is equal to the change in the internal energy that occurs during the reaction.

$$\Delta E_{\text{sys}} = q_v \text{ (at constant volume)}$$

The figure below shows a calorimeter in which reactions can be run at constant volume. Most reactions, however, are run in open flasks and beakers. When this is done, the volume of the system is not constant because gas can either enter or leave the container during the reaction. The system is at constant pressure, however, because the total pressure inside the container is always equal to atmospheric pressure.

If a gas is driven out of the flask during the reaction, the system does work on its surroundings. If the reaction pulls a gas into the flask, the surroundings do work on the system. We can still measure the amount of heat given off or absorbed during the reaction, but it is no longer equal to the change in the internal energy of the system, because some of the heat has been converted into work.

$$\Delta E_{\text{sys}} = q + w$$

We can get around this problem by introducing the concept of **enthalpy** ( $H$ ), which is the sum of the internal energy of the system plus the product of the pressure of the gas in the system times the volume of the system.

$$H_{\text{sys}} = E_{\text{sys}} + PV$$

For the sake of simplicity, the subscript "sys" will be left off the symbol for both the internal energy of the system and the enthalpy of the system from now on. We will therefore abbreviate the relationship between the enthalpy of the system and the internal energy of the system as follows.

$$H = E + PV$$

The change in the enthalpy of the system during a chemical reaction is equal to the change in its internal energy plus the change in the product of the pressure times the volume of the system.

$$\Delta H = \Delta E + \Delta(PV)$$

Let's assume that the reaction is run in a Styrofoam cup, as shown in the figure right.

Because the reaction is run at constant pressure, the change in the enthalpy that occurs during the reaction is equal to the change in the internal energy of the system plus the product of the constant pressure times the change in the volume of the system.

$$\Delta H = \Delta E + P\Delta V \text{ (at constant pressure)}$$

Substituting the first law of thermodynamics into this equation gives the following result.

$$\Delta H = (q_p + w) + P\Delta V$$

Assuming that the only work done by the reaction is work of expansion gives an equation in which the  $P\Delta V$  terms cancel.

$$\Delta H = (q_p - P\Delta V) + P\Delta V$$

Thus, the heat given off or absorbed during a chemical reaction at constant pressure is equal to the change in the enthalpy of the system.

$$\Delta H = q_p \text{ (at constant pressure)}$$

The relationship between the change in the internal energy of the system during a chemical reaction and the enthalpy of reaction can be summarized as follows.

1. The heat given off or absorbed when a reaction is run at *constant volume* is equal to the change in the internal energy of the system.

$$\Delta E_{\text{sys}} = q_v$$

2. The heat given off or absorbed when a reaction is run at *constant pressure* is equal to the change in the enthalpy of the system.

$$\Delta H_{\text{sys}} = q_p$$

3. The change in the enthalpy of the system during a chemical reaction is equal to the change in the internal energy plus the change in the product of the pressure of the gas in the system and its volume.

$$\Delta H_{\text{sys}} = \Delta E_{\text{sys}} + \Delta(PV)$$

4. The difference between  $\Delta E$  and  $\Delta H$  for the system is small for reactions that involve only liquids and solids because there is little if any change in the volume of the system during the reaction. The difference can be relatively large, however, for reactions that involve gases, if there is a change in the number of moles of gas in the course of the reaction.

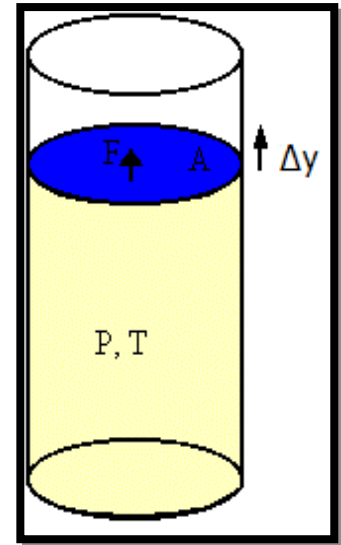
## Thermodynamic processes

Often the system from which we want to extract heat to do work is a gas. When a gas expands, how much work does it do on its surroundings?

Consider an ideal gas in a cylinder with a movable piston. The gas occupies a volume  $V$ , the temperature is  $T$ , and the pressure is  $P$ . Assume that the piston has a cross-sectional area  $A$ . Assume the gas expands by a small amount and the piston moves out by a small amount  $\Delta y$ .

The gas exerts a force  $F = PA$  on the piston and therefore does work  $\Delta W = F\Delta y$  on the piston.

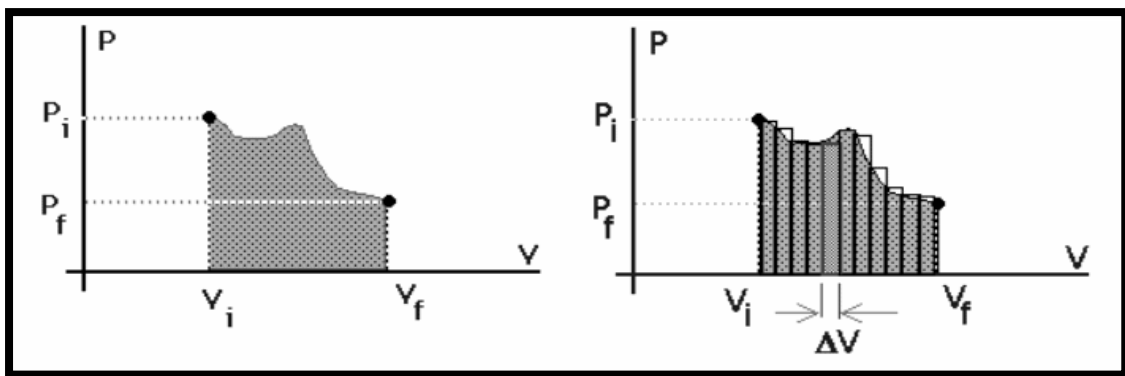
$$\Delta W = F\Delta y = PA\Delta y = P\Delta V$$



$\Delta V$  is the change in volume of the gas.  $\Delta V$  is positive because the gas expands. (If the gas is being compressed, then  $\Delta V$  is negative and work is done on the gas.) If the volume of the gas changes from  $V_1$  to  $V_2$ , then the total work done by the gas depends on how exactly the pressure varies during the expansion process.

$$W = \sum_j P_j \Delta V_j.$$

We sum the work done during a large number of small changes of the volume.



**A. Adiabatic process** is a process during which no heat enters or leaves the system. We then have  $\Delta U = -\Delta W$ , i.e.  $\Delta W$  equals the change in a physical property of the system.

In practice there are two different ways to prevent the transfer of heat.

(a) Provide very good thermal insulation of the system.

(b) Complete the process in a very short time interval, so that there is no time for appreciable heat transfer. The combustion process inside a car engine is essentially adiabatic for this reason.

**B. Isobaric process** is a process that occurs at constant pressure.

We then have  $W = P(V_2 - V_1)$ . If the pressure of an ideal gas is kept constant, then the temperature must increase as the gas expands. ( $PV/T = \text{constant}$ .) Heat must be added during the expansion process.

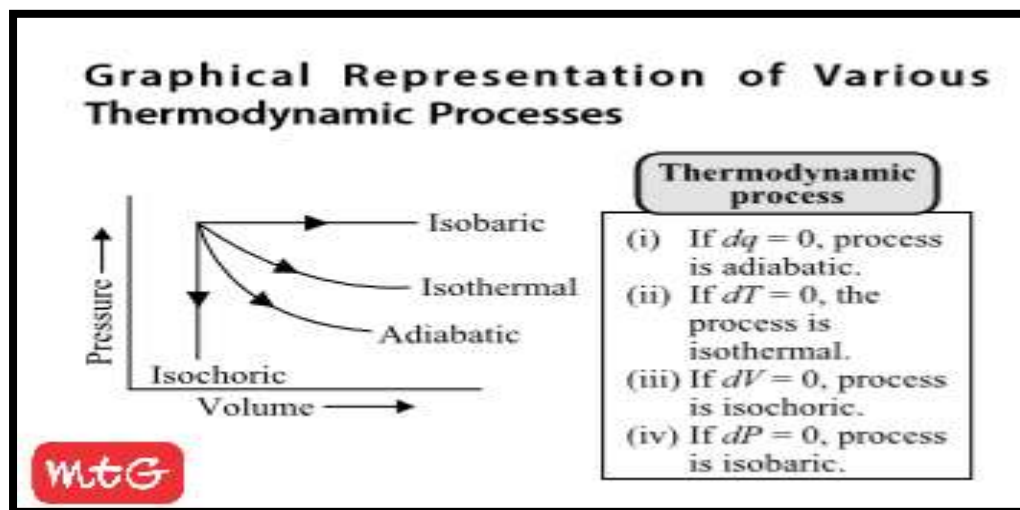
For an **isobaric process** we write  $\Delta U = \Delta Q - \Delta W = \Delta Q - P(V_2 - V_1)$ , or, rearranging the terms,  $\Delta H = \Delta Q$ . Chemical reactions (including biological ones) often take place at constant pressure, and then  $\Delta Q$  is equal to the change in a physical property of the system.

**C. Isovolumetric or isometric process** takes place at constant volume.

Then  $W = 0$  and  $\Delta U = \Delta Q$ . All the heat added to the system goes into increasing its internal energy. An isothermal process occurs at constant temperature. Since the internal energy of a gas is only a function of its temperature,  $\Delta U = 0$  for an isothermal process. For the isothermal expansion of an ideal gas we have  $W = nRT \ln(V_2/V_1)$ .  $W$  is positive if  $V_2 > V_1$ . Since  $\Delta U = 0$ , the heat transferred to the gas is  $\Delta Q = W$ .

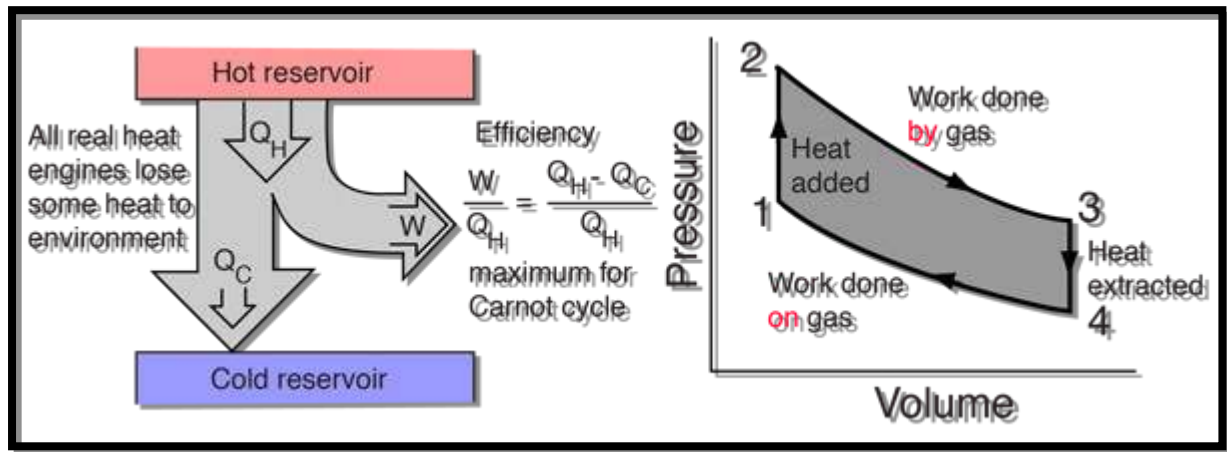
**D. Isothermal process** occurs at constant temperature. Since the internal energy of a gas is only a function of its temperature,  $\Delta U = 0$  for an isothermal process. For the isothermal expansion of an ideal gas we have  $W = nRT \ln(V_2/V_1)$ .  $W$  is positive if  $V_2 > V_1$ . Since  $\Delta U = 0$ , the heat transferred to the gas is  $\Delta Q = W$ .

The illustration below shows the graphical representation of Various Thermodynamic Processes.



## Heat Engines

A heat engine typically uses energy provided in the form of heat to do work and then exhausts the heat which cannot be used to do work. Thermodynamics is the study of the relationships between heat and work. The first law and second law of thermodynamics constrain the operation of a heat engine. The first law is the application of conservation of energy to the system, and the second sets limits on the possible efficiency of the machine and determines the direction of energy flow.



General heat engines can be described by the reservoir model (left) or by a PV diagram (right)

### Sample Problems:

1. During a cyclic process, a heat engine absorbs 500 J of heat from a hot reservoir, does work and ejects an amount of heat 300 J into the surroundings (cold reservoir). Calculate the efficiency of the heat engine?

Solution:

The efficiency of heat engine is given by

$$\eta = 1 - 0.6 = 0.4$$

The heat engine has 40% efficiency, implying that this heat engine converts only 40% of the input heat into work.

$$\eta = 1 - \frac{Q_L}{Q_H}$$

$$\eta = 1 - \frac{300}{500} = 1 - \frac{3}{5}$$

$$\eta = 1 - 0.6 = 0.4$$

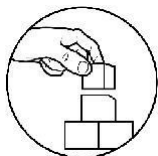
2. A steam engine boiler is maintained at 250°C and water is converted into steam. This steam is used to do work and heat is ejected to the surrounding air at temperature 300K. Calculate the maximum efficiency it can have?

Solution:

The steam engine is not a Carnot engine, because all the process involved in the steam engine are not perfectly reversible. But we can calculate the maximum possible efficiency of the steam engine by considering it as a Carnot engine

$$\eta = 1 - \frac{T_L}{T_H} = 1 - \frac{300K}{523K} = 0.43$$

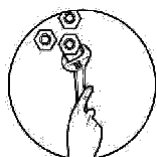
The steam engine can have maximum possible 43% of efficiency, implying this steam engine can convert 43% of input heat into useful work and remaining 57% is ejected as heat. In practice the efficiency is even less than 43%.



## What's More

Direction: Identify whether the following processes are run at constant volume and which are run at constant pressure.

- (a) an acid-base titration
- (b) decomposing  $\text{CaCO}_3$  by heating limestone in a crucible with a Bunsen burner
- (c) the reaction between zinc metal and an aqueous solution of  $\text{Cu}^{2+}$  ions to form copper metal and  $\text{Zn}^{2+}$  ions
- (d) measuring the calories in a 1-oz. serving of breakfast cereal by burning the cereal in a bomb calorimeter



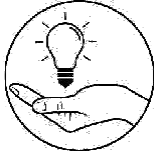
## What I Have Learned

### Activity 7.4 Solve It Up!

Directions: Solve the following heat engine work done related problems.

1. During a cyclic process, a heat engine absorbs 350 J of heat from a hot reservoir, does work and ejects an amount of heat 100 J into the surroundings (cold reservoir). Calculate the efficiency of the heat engine?
2. A steam engine boiler is maintained at  $250^\circ\text{C}$  and water is converted into steam. This steam is used to do work and heat is ejected to the surrounding air at temperature 300K. Calculate the maximum efficiency it can have?





## What I Can Do

Directions: From the lessons learned, briefly provide your answer & explanation on the following situations given in each number.

1. *In 1883, the volcano on Krakatau Island in the Pacific erupted violently in the largest explosion in Earth's recorded history, destroying much of the island in the process. Global temperature measurements indicate that this explosion reduced the average temperature of Earth by about 1 degree Celsius during the next two decades. Why?*



*Tip: You can search for science articles/news about the event.*



## Lesson

# 3

# The Second Law of Thermodynamics



## What's In

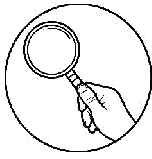
In the preceding lesson you were able to understand the first law of thermodynamics; Law of Conservation of Energy. You were also able to know the different thermodynamic processes, interpret PV diagrams, and do heat engine efficiency calculations. In this lesson, we will explore further the Second Law of Thermodynamics which is Entropy. The second law of thermodynamics is a powerful aid to help us understand why the world works as it does -- why hot pans cool down, why our bodies stay warm even in the cold, why gasoline makes engines run.



## What I Need to Know

After this lesson, you should be able to:

1. Describe reversible and irreversible processes.
2. Explain how entropy is a measure of disorder.
3. State the 2nd Law of Thermodynamics.
4. Calculate entropy changes for various processes e.g., isothermal process,
5. free expansion, constant pressure process, etc.



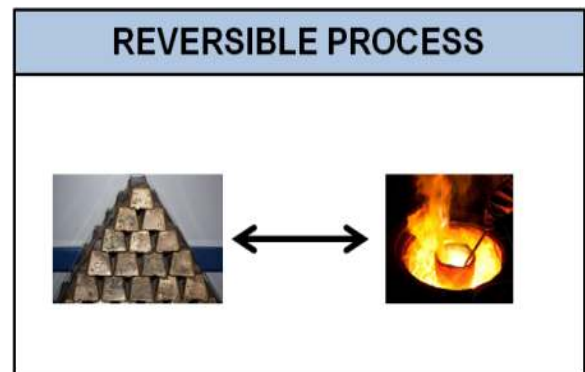
## What Is It

### **Reversible & Irreversible Processes**

A thermodynamic process is said to be **reversible** if the process can be turned back such that both the system and the surroundings return to their original states, with no other change anywhere else in the universe. As we know, in reality, no such processes as reversible processes can exist. Thus, the reversible processes can easily be defined as idealizations or models of real processes, on which the limits of the system or device are to be defined. They help us in incurring the maximum efficiency a system can provide in ideal working conditions and thus the target design that can be set.

#### **Examples of Reversible Process:**

- extension of springs
- slow adiabatic compression or expansion of gases
- electrolysis (with no resistance in the electrolyte)
- the frictionless motion of solids
- slow isothermal compression or expansion of gases



An **irreversible** process can be defined as a process in which the system and the surroundings do not return to their original condition once the process is initiated. Taking an example of an automobile engine, that has travelled a distance with the aid of fuel equal to an amount 'x'. During the process, the fuel burns to provide energy to the engine, converting itself into smoke and heat energy. We cannot retrieve the energy lost by the fuel and cannot get back the original form.

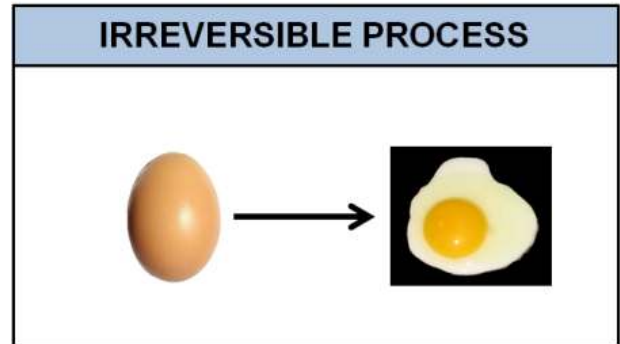
There are many factors due to which the irreversibility of a process occurs, namely:

1. The friction that converts the energy of the fuel to heat energy
2. The unrestrained expansion of the fluid which prevents from regaining the original form of the fuel Heat transfer through a finite temperature, the reverse of which is not possible as the forward process, in this case, is spontaneous
3. Mixing of two different substances which cannot be separated as the process of intermixing is again spontaneous in nature, the reverse of which is not feasible.

Thus, some processes are reversible while others are irreversible in nature depending upon their ability to return to their original state from their final state.

Examples of Irreversible Processes:

- Relative motion with friction
- Throttling
- Heat transfer
- Diffusion
- Electricity flow through a resistance.



The second law of thermodynamics deals with the direction taken by spontaneous processes. Many processes occur spontaneously in one direction only—that is, they are irreversible, under a given set of conditions.

Although irreversibility is seen in day-to-day life—a broken glass does not resume its original state, for instance—complete irreversibility is a statistical statement that cannot be seen during the lifetime of the universe. More precisely, an irreversible process is one that depends on path. If the process can go in only one direction, then the reverse path differs fundamentally, and the process cannot be reversible.

For example, heat involves the transfer of energy from higher to lower temperature. A cold object in contact with a hot one never gets colder, transferring heat to the hot object and making it hotter.

Furthermore, mechanical energy, such as kinetic energy, can be completely converted to thermal energy by friction, but the reverse is impossible. A hot stationary object never spontaneously cools off and starts moving.

**One-Way Processed in Nature:**

Examples of one-way processes in nature. (a) Heat transfer occurs spontaneously from hot to cold and not from cold to hot. (b) The brakes of this car convert its kinetic energy to heat transfer to the environment. The reverse process is impossible.

(c) The burst of gas let into this vacuum chamber quickly expands to uniformly fill every part of the chamber. The random motions of the gas molecules will never return them to the corner.

## Major players in developing the Second Law

- Nicolas Léonard Sadi Carnot was a French physicist, who is considered to be the "father of thermodynamics," for he is responsible for the origins of the Second Law of Thermodynamics, as well as various other concepts. The current form of the second law uses entropy rather than caloric, which is what Sadi Carnot used to describe the law. Caloric relates to heat and Sadi Carnot came to realize that some caloric is always lost in the motion cycle. Thus, the thermodynamic reversibility concept was proven wrong, proving that irreversibility is the result of every system involving work.
- Rudolf Clausius was a German physicist, and he developed the Clausius statement, which says "Heat generally cannot flow spontaneously from a material at a lower temperature to a material at a higher temperature."
- William Thompson, also known as Lord Kelvin, formulated the Kelvin statement, which states "It is impossible to convert heat completely in a cyclic process." This means that there is no way for one to convert all the energy of a system into work, without losing energy.
- Constantin Carathéodory, a Greek mathematician, created his own statement of the second law arguing that "In the neighborhood of any initial state, there are states which cannot be approached arbitrarily close through adiabatic changes of state."

## Entropy

Consider a system in two different conditions, for example 1kg of ice at 0 °C, which melts and turns into 1 kg of water at 0 °C. We associate with each condition a quantity called the entropy.

The **entropy** of any substance is a function of the condition of the substance. It is a physical property of the substance. For an ideal gas it is a function of its temperature and volume, and for a solid and liquid it is a function of its temperature and internal structure.

The entropy is independent of the history of the substance. The entropy of the 1 kg of water at 0 °C is the same if we obtained the water from ice, or if we cooled the water from room temperature down to 0 °C. When a small amount of heat  $\Delta Q$  is added to a substance at temperature  $T$ , without changing its temperature appreciably, the entropy of the substance changes by:

$$\Delta S = \Delta Q/T.$$

When heat is removed, the entropy decreases, when heat is added the entropy increases. Entropy has units of Joules per Kelvin.

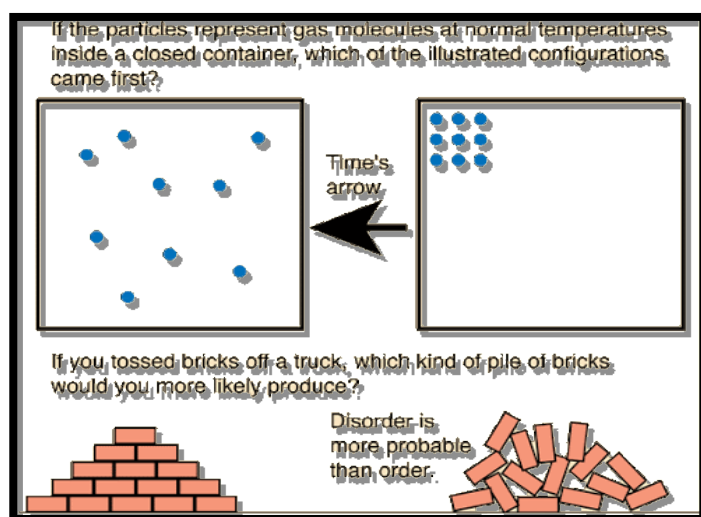
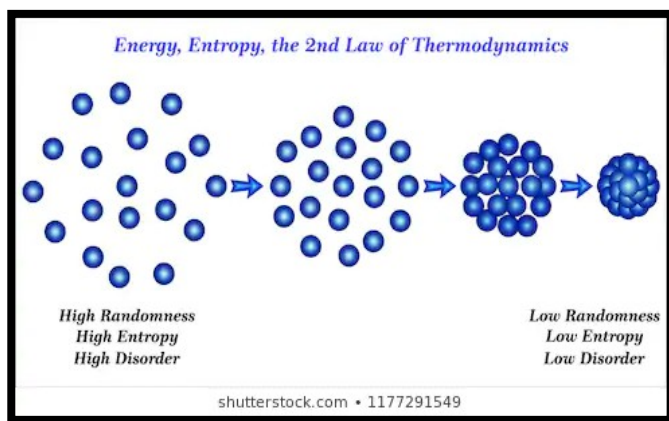
To calculate the change in entropy of a system for a finite process, when T changes appreciably, we find some reversible path that can take the system (infinitely slowly) from its initial to its final state and evaluate  $\Delta S$  for that path.

The actual path of the system from the initial to the final state may be very different and not reversible. But the change in entropy depends only on the initial and final state, not on the path.

We can define the entropy as the logarithm of the disorder times a constant of proportionality. When we change the entropy of a substance by an amount

$$\Delta S = \Delta Q/T$$

we change the **disorder** of the substance. Entropy always increases, because a high amount of disorder is, by definition, is more likely than a low amount of disorder.



### Sample Problem: Entropy Change for the Freezing of Water

1. An ice tray contains 500 g of water. Calculate the change in entropy of the water as it freezes completely and slowly at 0 °C.

Solution/Reasoning:

The water freezes at 0 °C = 273 K.  $\Delta S = \Delta Q/T$   
 $\Delta Q = -mL$ , m = mass of water, L = latent heat of fusion = 333000 J/kg.

Details of the calculation:  
 $\Delta S = -(0.5 \text{ kg}) (333000 \text{ J/kg}) / 273 \text{ K} = -610 \text{ J/K}.$

2. Suppose we have 1.50kg of water at a temperature of 0 degree Celsius. We put the water in a freezer, and enough heat is removed from the water to freeze it completely to ice at a temperature of 0 degree Celsius. How much does the entropy of the water ice system change during the freezing process?

Solution/Reasoning:

The melting of ice is an isothermal process, so we can use  $\Delta S = \Delta Q/T$ .

$$Q = mL_{\text{fusion}} = (1.50\text{kg})(334\text{kJ/Kg}) = 501\text{kJ}$$

Thus, the change in the entropy of the water-ice system is:

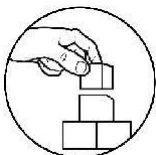
$$\Delta S = -501\text{kJ} / 273.15\text{K} = -1830\text{J/K}$$

### Entropy Change during Free Expansion of a Gas

3. Let us consider the free expansion of a gas. Initially 0.500mole of nitrogen gas is confined to a volume of 0.500 m<sup>3</sup>. When the barrier is removed, the gas expands to fill the new of 1.00m<sup>3</sup>. What is the change in entropy of the gas?

Solution:

We can use the equation below to calculate the number of microstates of the system before and after the expansion to calculate the change in entropy. In this system, the number of gas molecules



## What's More

### Activity 7.5: Calculation Time

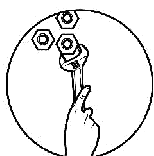
Direction: Solve for the following problem. Show your solutions.

1. The surface of the Sun is approximately at 5700 K, and the temperature of the Earth's surface is approximately 290 K. What entropy changes occur when 1000 J of thermal energy is transferred from the Sun to the Earth?



2. An ice tray contains 500 g of water. Calculate the change in entropy of the water as it freezes completely and slowly at 0 °C.
3. If you toss two dice, what is the total number of ways that you can obtain (a) a 12 and (b) a 7?

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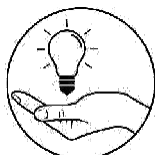
## What I Have Learned

### Activity 7.6: Create Your Own Problem

Make your own word problem about Entropy change. One (1) for Entropy Change for the Freezing of Water and another one (1) for Entropy Change during Free Expansion of a Gas. Show your illustrations and complete solutions to the problems. Be sure that the problems are not taken from the internet and that they are realistic.

### Activity 7.7: Give Me An Example

*What is an example of entropy from everyday life?*



## What I Can Do

### Activity 7.8 What is the second law of thermodynamics?

Directions: Read the excerpt and answer the following questions after reading.

Thermodynamics is the study of heat and energy. At its heart are laws that describe how energy moves around within a system, whether an atom, a hurricane or a black hole. The first law describes how energy cannot be created or destroyed, merely transformed from one kind to another. The second law, however, is probably better known and even more profound because it describes the limits of what the universe can do.

The second law can be expressed in several ways, the simplest being that heat will naturally flow from a hotter to a colder body. At its heart is a property of thermodynamic systems called entropy – in the equations above it is represented by "S" – in loose terms, a measure of the amount of disorder within a system. This can be represented in many ways, for example in the arrangement of the molecules – water molecules in an ice cube are more ordered than the same molecules after they have been heated into a gas. Whereas the water molecules were in a well-defined lattice in the ice cube, they float unpredictably in the gas. The entropy of the ice cube is, therefore, lower than that of the gas. Similarly, the entropy of a plate is higher when it is in pieces on the floor compared with when it is in one piece in the sink.

Heat engines work because heat naturally flows from hot to cold places. If there was no cold reservoir towards which it could move there would be no heat flow and the engine would not work. Because the cold reservoir is always above absolute zero, no heat engine can be 100% efficient.

The best-designed engines, therefore, heat up steam (or other gas) to the highest possible temperature then release the exhaust at the lowest possible temperature. The most modern steam engines can get to around 60% efficiency and diesel engines in cars can get to around 50% efficient. Petrol-based internal combustion engines are much more wasteful of their fuel's energy.

The inefficiencies are built into any system using energy and can be described thermodynamically. This wasted energy means that the overall disorder of the universe – its entropy – will increase over time but at some point, reach a maximum. At this moment in some unimaginably distant future, the energy in the universe will be evenly distributed and so, for all

Excerpt from [www.theguardian.com/science/2013/dec/01/what-is-the-second-law-of-thermodynamics](http://www.theguardian.com/science/2013/dec/01/what-is-the-second-law-of-thermodynamics)

### Questions:

1. What does the text tell about the second law of thermodynamics?

2. How is the second law expressed? Give examples.
3. What are heat engines? How do they work?
4. What are examples of entropy in real life?

### Summary:

- ✓ The properties of IDEAL gases are that: they are compressible, they are always in motion, colliding elastically within a container, they take the shape of the container and they behave inertly
- ✓ The IDEAL gases are compressible in the sense that because they have so much energy, it makes it easy to force them to move in an inward direction, thereby compressing them.
- ✓ The IDEAL gases are always in motion, colliding elastically with the container, meaning that they lose no energy after collisions and thus can keep moving.
- ✓ The IDEAL gases take the shape of the container as a result of their compressibility, and so the volume of the container can be assumed to be the overall volume of the gas(es) inside.
- ✓ The IDEAL gases behave inertly, meaning that they are assumed to not react with each other. This keeps things simple, so that their properties can be studied in the absence of complications where the gases could have been used up in gas reactions.
- ✓ The *first law of thermodynamics* states that the change in internal energy of a system equals the net heat transfer *into* the system minus the net work done *by* the system. In equation form, the first law of thermodynamics is  $\Delta U = Q - W$ .
- ✓ The first law of thermodynamics, also known as Law of Conservation of Energy, states that energy can neither be created nor destroyed; energy can only be transferred or changed from one form to another. For example, turning on a light would seem to produce energy; however, it is electrical energy that is converted.
- ✓ Heat transfer ( $Q$ ) and doing work ( $W$ ) are the two everyday means of bringing energy into or taking energy out of a system. The processes are quite different. Heat transfer, a less organized process, is driven by temperature differences. Work, a quite organized process, involves a macroscopic force exerted through a distance.
- ✓ The *internal energy*  $U$  of a system is the sum of the kinetic and potential energies of its atoms and molecules.
- ✓ Metabolism of living organisms, and photosynthesis of plants, are specialized types of heat transfer, doing work, and internal energy of systems.

- ✓ A *heat engine* is a cyclic device that takes heat  $Q_h$  in from a *hot reservoir*, converts *some* of it to work  $W$ , and rejects the rest of it  $Q_c$  to a *cold reservoir* so that at the end of a cycle it is in the same state (and has the same internal energy) with which it began. The net work done per cycle is (recall) the area inside the  $PV$  curve.
- ✓ The *Carnot Cycle* is the archetypical reversible cycle, and a Carnot Cycle-based heat engine is one that does not dissipate any energy internally and uses only reversible steps.

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- ✓ The Carnot efficiency is easy to compute (see text and lecture example). A Carnot Cycle consists of four steps:
  - a. Isothermal expansion (in contact with the heat reservoir)
  - b. Adiabatic expansion (after the heat reservoir is removed)
  - c. Isothermal compression (in contact with the cold reservoir)
  - d. Adiabatic compression (after the cold reservoir is removed)
- ✓ A thermodynamic process is said to be *reversible* if the process can be turned back such that both the system and the surroundings return to their original states, with no other change anywhere else in the universe.
- ✓ An irreversible process can be defined as a process in which the system and the surroundings do not return to their original condition once the process is initiated.
- ✓ The second law of thermodynamics deals with the direction taken by spontaneous processes. Many processes occur spontaneously in one direction only—that is, they are irreversible, under a given set of conditions.
- ✓ The second law of thermodynamics states that the total entropy of an isolated system can never decrease over time, and is constant if and only if all processes are reversible. Isolated systems spontaneously evolve towards thermodynamic equilibrium, the state with maximum entropy.
- ✓ Entropy  $S$  is a measure of disorder. The change in entropy of a system can be evaluated by integrating:

$$dS = \frac{dQ}{T}$$

- ✓ The entropy of the Universe never decreases. It either increases (for irreversible processes) or remains the same (for reversible processes).

**Assessment:**

**MULTIPLE CHOICE: Directions:** Read and understand each item and choose the letter of the correct answer. Write your answers on a separate sheet of paper.

1. Which one of the following statements is not consistent with the kinetic-molecular theory of gases?
  - a. Individual gas molecules are relatively far apart.
  - b. The actual volume of the gas molecules themselves is very small compared to the volume occupied by the gas at ordinary temperatures and pressures.
  - c. The average kinetic energies of different gases are different at the same temperature.
  - d. There is no net gain or loss of the total kinetic (translational) energy in collisions between gas molecules.
  - e. The theory explains most of the observed behavior of gases at ordinary temperatures and pressures.
2. Which statement best describe gases?
  - a. Gases consist of molecules that have enough spatial separation that the intermolecular bonding characteristics of liquids and solid is absent.
  - b. Gases are almost incompressible. In liquids molecules are close to each other.
  - c. Gases are characterized by structural rigidity and resistance to a force applied to the surface.
  - d. Gases are in the state in which matter adapts to the shape of its container but varies only slightly in volume

3. Which of the following term does not involve in ideal gas law?  
a. Pressure   b. Volume   c. Temperature   d. Time
4. Which mathematical formula summarizes the Ideal Gas Law?  
a.  $PV = Dt/r$    b.  $PV = ma$    c.  $PV = nRT$    d.  $PV = mgh$
5. A real gas most closely approaches the behavior of an ideal gas under conditions of:  
a. high P and low T   b. low P and high T   c. low P and T   d. STP
6. Which of the following statements best describe the first law of thermodynamics?  
a. Energy can be changed from one form to another, but it cannot be created or destroyed.  
b. Processes that involve the transfer or conversion of heat energy are irreversible.  
c. It is concerned with the limiting behavior of systems as the temperature approaches absolute zero.  
d. The total entropy of an isolated system can never decrease over time.
7. All the following statements define the second law of thermodynamics except,  
a. Processes that involve the transfer or conversion of heat energy are irreversible.  
b. The total entropy of an isolated system can never decrease over time.  
c. Energy can be changed from one form to another, but it cannot be created or destroyed.  
d. Heat is a form of energy, and thermodynamic processes are therefore subject to the principle of conservation of energy.

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8. Which one of the following definitions best describes the concept of work?  
a) the flow of energy from one object or substance to another due to a difference in temperature  
b) the flow of energy from one body to another through uniform molecular motion  
c) the force associated with molecular motion  
d) the random motion of molecules in a gas at low pressure
9. Consider the following thermodynamic properties.  
I) work done on a system  
II) heat absorbed  
III) entropy  
IV) enthalpy  
  
Which of these properties are state functions?  
a. I and II only   b. I and III only   c. I only   d. III and IV only
10. Which of the following statements will always apply when a reversible chemical reaction has attained equilibrium?  
a) All reactants will convert to products  
b) The reaction proceeds alternately in the forward and reverse directions  
c) The Gibbs free energy of the system reaches a minimum  
d) The forward reaction will dominate over the reverse reaction.

**Answer Key:**

### What I Know

(Pretest)

1. C
2. A
3. D
4. C
5. D
6. A
7. D
8. A
9. D
10. B

### What's More

Activity 7.2

1.  $T = 204.6\text{K}$
2.  $n = 1.26\text{moles}$
3.  $V = 2331\text{ L}$
4.  $P = 1.64\text{ atm}$

### What I Have Learned

Activity 7.3

1. D
2. C
3. D
4. B
5. A
6. D
7. A
8. B
9. C

10. D

### Activity 7.4

1.  $u = 0.71$
2.  $n = 0.43$

### What's More

Activity 7.5

1.  $S_{\text{sun}} = -0.175\text{J/K}$

### What's More

Activity 7.6 & Activity 7.7

Answers May Vary

Downloaded by aldrin alciso (aldrinalciso2@gmail.com)





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