

# Heat

- **Heat** is defined as the *transfer of energy* across the boundary of a system due to a **temperature difference** between the system and its surroundings
- The term *heat* will also be used to represent the *amount of energy transferred* by this method

# Changing Internal Energy

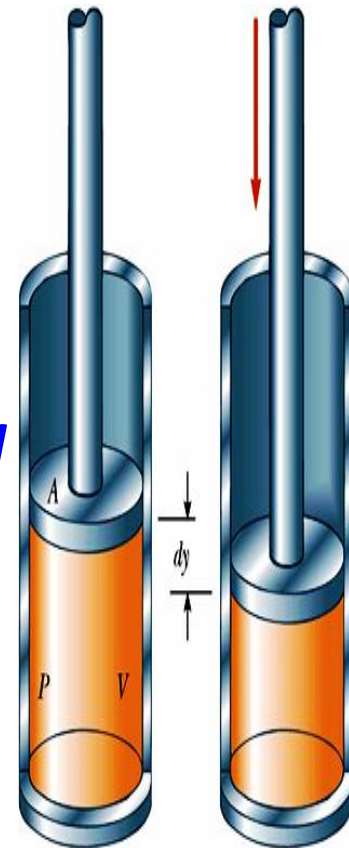
- Both *heat* ( $Q$ ) and *work* ( $W$ ) can change the *internal energy* ( $U$ ) of a system
- The *internal energy* can be changed even when *no energy is transferred by heat, but just by work*
- **Example:** *compressing gas with a piston*
  - Energy is transferred by work

# Units of Heat

- Historically, ***the calorie*** was **the unit used for heat** . *One calorie is the amount of energy transfer necessary to raise the temperature of 1 g of water from 14.5°C to 15.5°C*
- The “Calorie” used for food is actually **1 kilocalorie!!**
- In the **US Customary system**, the unit is a ***BTU (British Thermal Unit)*** . ***One BTU is the amount of energy transfer necessary to raise the temperature of 1 lb of water from 63°F to 64°F***
- The ***SI units*** for ***heat*** will be: **Joules**

# Work and Heat in Thermodynamic Processes

- **Work** can be done on a **deformable system**, such as a gas
- A force is applied to **slowly compress** the gas in a cylinder with a moveable piston
  - The **compression is slow** enough for all the system **to remain** essentially **in thermal equilibrium**
  - This is said to occur **quasi-statically**
- The piston is pushed downward by a force  **$F$**  through a displacement of  **$dr$** :



$$dW = \mathbf{F} \cdot d\mathbf{r} = -F\hat{\mathbf{j}} \cdot dy\hat{\mathbf{j}} = -Fdy = -PA dy$$

- ***Ady*** is the change in volume of the gas, ***dV***
- Therefore, the work done **ON** the gas is

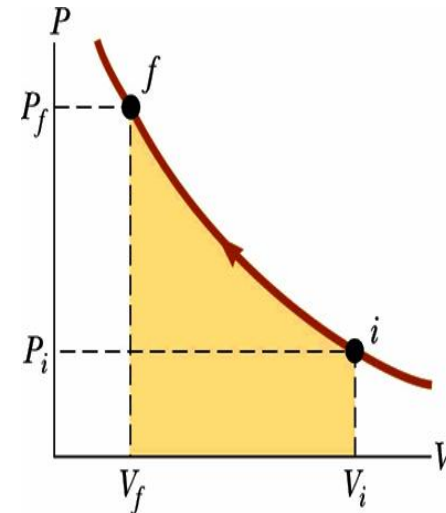
$$dW = -P dV$$

- Notes:
  - If the gas is ***compressed***, ***dV*** is ***negative*** and the ***work*** done on the gas is ***positive***
  - If the gas ***expands***, ***dV*** is ***positive*** and the ***work*** done on the gas is ***negative***
  - If the ***volume remains constant***, the ***work done is zero***
- The total work done is:

$$W = -\int_{V_i}^{V_f} P dV$$

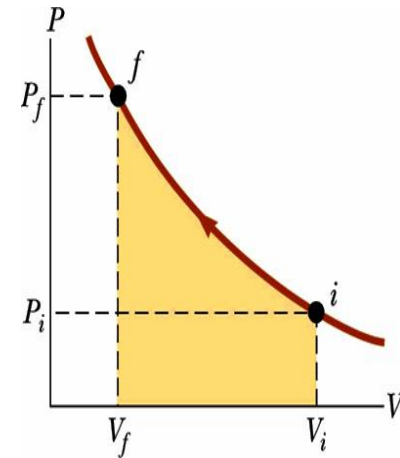
# PV Diagrams

- Used when the **pressure** and **volume** are known at each step of the process
- The state of the gas at each step can be plotted on a graph called a **PV diagram**
- This allows us to visualize the process through which the gas is progressing
- The curve is called the **path**



# PV Diagrams

- **The work** done on a gas in a quasi-static process that takes the gas from an **initial state** to a **final state** is the **negative of the area** under the curve on the **PV diagram**, evaluated between the initial and final states
- This is true whether or not the pressure stays constant
- The work done **does** depend on the path taken



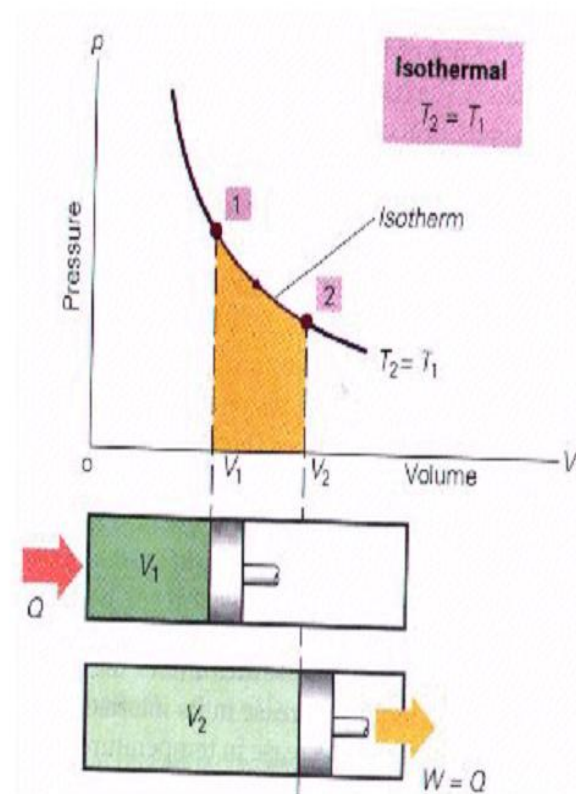
$$W = -\int_{V_i}^{V_f} P dV$$

# Work Done in Different Thermodynamic Processes for an Ideal Gas

- 1- Isothermal Process ( $T = \text{constant}$ ):

To keep the temperature constant both the pressure and volume change to compensate.

(Volume goes up,  
pressure goes down)





From the general formula of Work definition we have:

$$W = \int_{V_i}^{V_f} P \, dV \quad \text{----- (1)}$$

From the equation of state of an ideal gas we have:

$$Pv = RT \quad \text{or} \quad P = \frac{RT}{v} \quad \text{--- (2)}$$

Putting equation (2) in eq.(1) we get:

$$w = RT \ln \frac{v_2}{v_1} \quad \text{----- (3) for specific work}$$

Or

$$W = nRT \ln \frac{V_2}{V_1} \quad \text{----- (4) for normal work}$$

Equation (3) can be re-written in the following forms:

$$w = P_1 v_1 \ln \frac{P_1}{P_2} \text{ --- (5)}$$

or

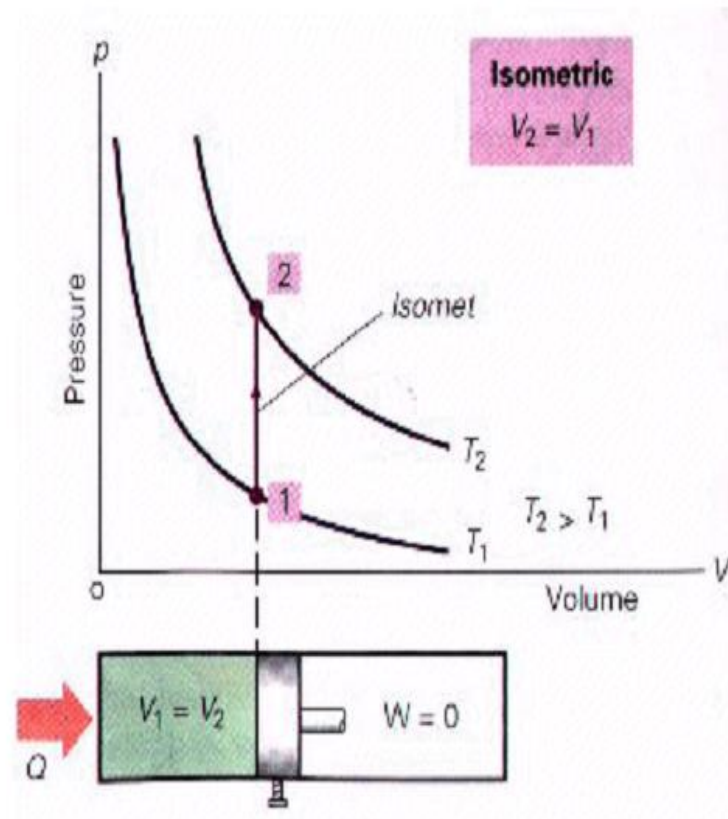
$$w = P_2 v_2 \ln \frac{P_1}{P_2} \text{ --- (6)}$$

## 2- Work in Isometric Process( $v=\text{constant}$ )

- From the general formula of Work definition we have:

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

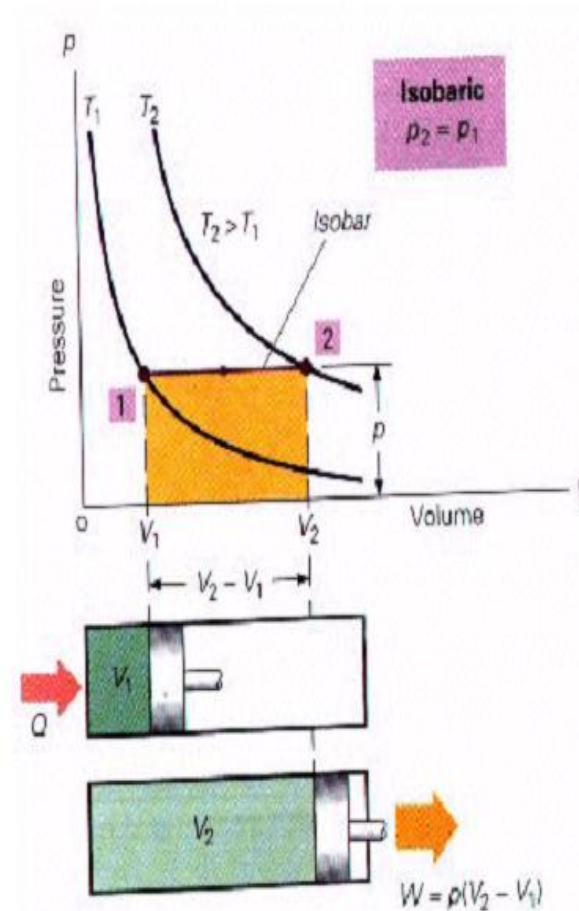
- Since  $v=\text{constant}$
- i.e  $dv=0$
- i.e : no work done in the isometric process



### 3- Work in Isobaric Process(P=constant)

- In this case the work can be found as :

$$W = \int_{V_1}^{V_2} P \, dv = P \int_{V_1}^{V_2} dv = P(v_2 - v_1)$$



**Example:** Calculate work done by expanding gas of 1 mole if initial pressure is 4000 Pa, initial volume is  $0.2 \text{ m}^3$ , and initial temperature is 96.2 K.

Assume a two processes: (1) *isobaric* expansion to  $0.3 \text{ m}^3$ ,  $T_f=144.3 \text{ K}$  (2) *isothermal* expansion to  $0.3 \text{ m}^3$ .

Given:

$$n = 1 \text{ mole}$$

$$T_i = 96.2 \text{ K}$$

$$T_f = 144.3 \text{ K}$$

$$V_i = 0.2 \text{ m}^3$$

$$V_f = 0.3 \text{ m}^3$$

$$P = \text{const}$$

Find:

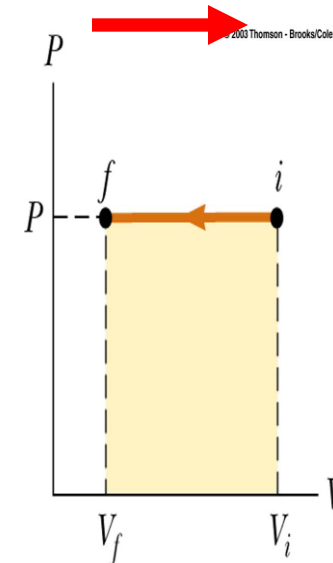
$$W = ?$$

1. Isobaric expansion:

$$W = P\Delta V = P(V_f - V_i) = 4000 \text{ Pa}(0.3 \text{ m}^3 - 0.2 \text{ m}^3) = 400 \text{ J}$$

Also:

$$\frac{T_f}{T_i} = \frac{\frac{P_f V_f}{nR}}{\frac{P_i V_i}{nR}} = \frac{V_f}{V_i} = \frac{0.3 \text{ m}^3}{0.2 \text{ m}^3} = 1.5$$



A 50% increase in temperature!

# Example:

Calculate work done by expanding gas of 1 mole if initial pressure is 4000 Pa, initial volume is  $0.2 \text{ m}^3$ , and initial temperature is 96.2 K.

Assume a two processes: (1) *isobaric* expansion to  $0.3 \text{ m}^3$ ,  $T_f=144.3 \text{ K}$  (2) *isothermal* expansion to  $0.3 \text{ m}^3$ .

Given:

$$n = 1 \text{ mole}$$

$$T_i = 96.2 \text{ K}$$

$$V_i = 0.2 \text{ m}^3$$

$$V_f = 0.3 \text{ m}^3$$

$$T = \text{const}$$

Find:

$$W = ?$$

2. Isothermal expansion:

$$\begin{aligned} W &= nRT \ln\left(\frac{V_f}{V_i}\right) = P_i V_i \ln\left(\frac{V_f}{V_i}\right) \\ &= (4000 \text{ Pa})(0.2 \text{ m}^3) \ln \frac{0.3 \text{ m}^3}{0.2 \text{ m}^3} = 324 \text{ J} \end{aligned}$$
