

Ideal Gas Law

- *About ideal gases*

very low pressure (or low density)

Observation/evidence

- (1) You can compress a gas, but not a solid or liquid.
- (2) Gases exert forces on the walls of a container.
- (3) Gases *keep* exerting forces on the walls of a container, even after a long time.

Feature

- (1) Molecules are far apart
- (2) Molecules are constantly moving around (translational motion)
- (3) Molecular collisions are elastic

Ideal Gas Law

- Variables:

- Number of moles (how many molecules)

- Temperature

- Volume

- Pressure

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

$$m = \frac{N}{N_A} M = nM \quad M \rightarrow \text{Molar Mass}$$

- PhET simulation - Gas Properties

<http://phet.colorado.edu/en/simulation/gas-properties>

Ideal Gas

Every quantity has to be in SI unit

How are these quantities related? A model for a gas...

$$PV = nRT_K$$

Number of moles

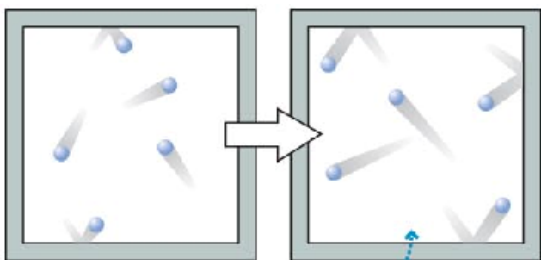
$$n = \frac{N}{N_A} \quad n = \frac{m}{\text{molar mass}}$$

Empirical discovery

Universal gas constant
Gas constant per mole

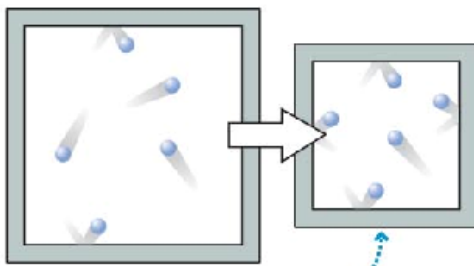
$$R = 8.31 \text{ J/mol} \cdot \text{K}$$

$$R = 0.082 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$



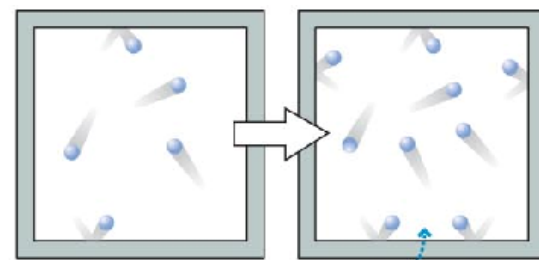
Increasing the temperature of the gas means the particles move at higher speeds. They hit the walls more often and with more force, so there is more pressure.

■ Increasing the temperature leads to an increase in pressure.



Decreasing the volume of the container means more frequent collisions with the walls of the container, and thus more pressure.

■ Decreasing the volume leads to an increase in pressure.



Increasing the number of particles in the container means more frequent collisions with the walls of the container, and thus more pressure.

■ Increasing the number of particles leads to an increase in pressure.

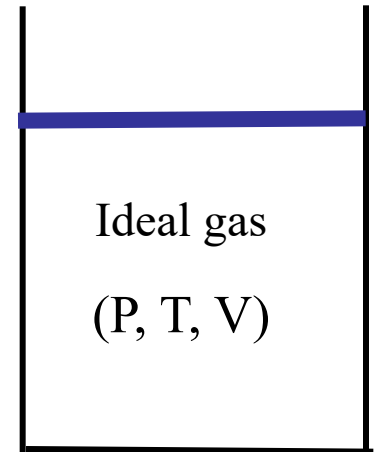
Ideal Gas in Sealed Containers

$$PV = nRT_K$$

For gases in a sealed container:

$$\frac{PV}{T} = nR = \text{constant}$$

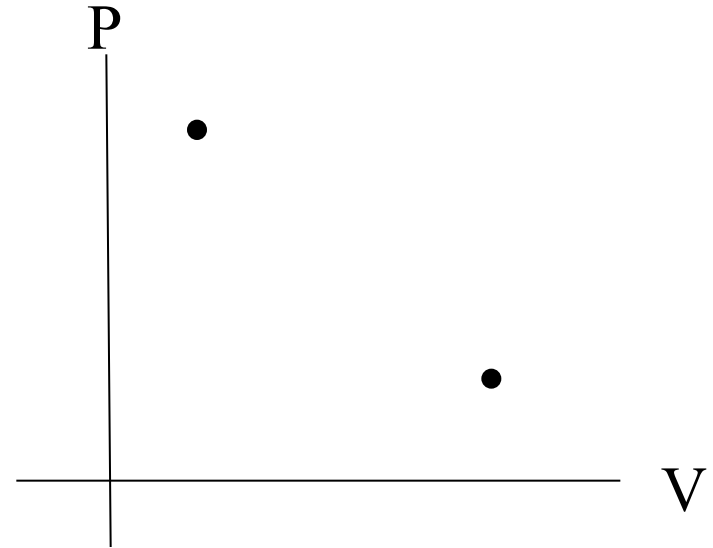
$$\frac{P_i V_i}{T_i} = \frac{P_f V_f}{T_f}$$



Ideal Gas Processes

- PV Diagram

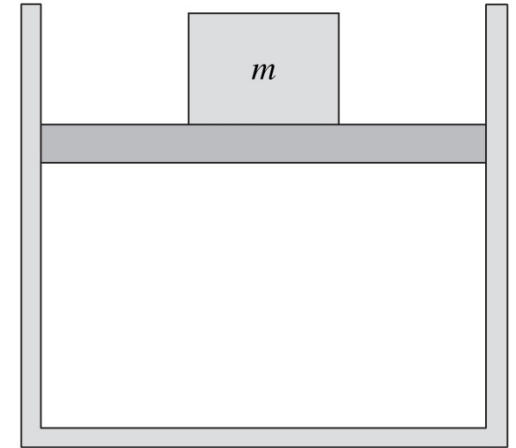
- Represent ideal gas processes.
- Each point on the graph represents a single, unique state of the gas (P , V , T).



- Ideal Gas Processes is the means by which that the ideal gas *changes from one state to another*
 - Constant-Volume Process (isochoric, isovolumetric)
 - Constant-Pressure Process (Isobaric)
 - Constant-Temperature Process (Isothermal)
 - “Insulated” -- Adiabatic ($Q = 0$) process

In the experimental arrangement shown, how would you decrease the volume without changing the pressure?

1. Can't be done, pressure is inversely proportional to volume.
2. Add more mass on top of the piston.
3. Remove some mass on top of the piston
4. Heat it.
5. Cool it.

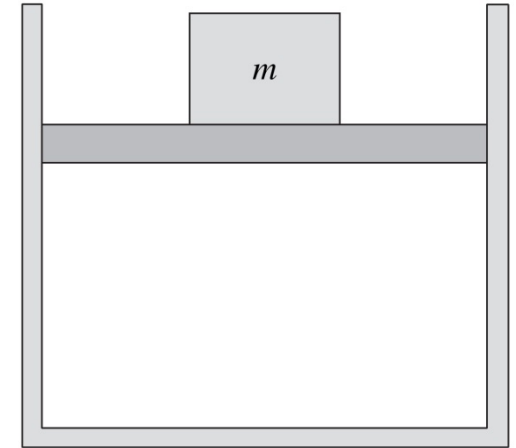


Constant Pressure Process
(isobaric)

PV diagram?

In the experimental arrangement shown, how would you decrease the volume without changing the temperature?

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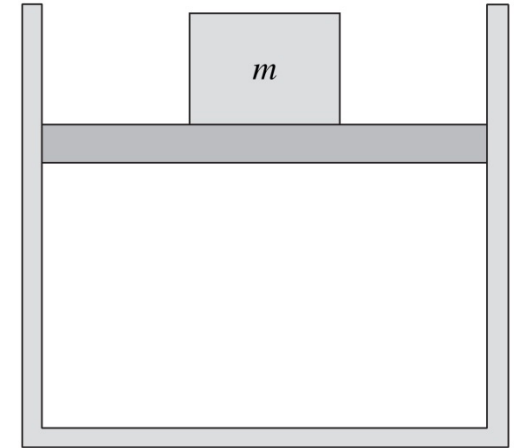


Constant Temperature Process
(isothermal)

PV diagram?

In the experimental arrangement shown, can you decrease the pressure without changing the volume? If so, how?

1. Can't be done, pressure is inversely proportional to volume.
2. Add more mass on top of the piston.
3. Remove some mass on top of the piston
4. Heat it.
5. Cool it.



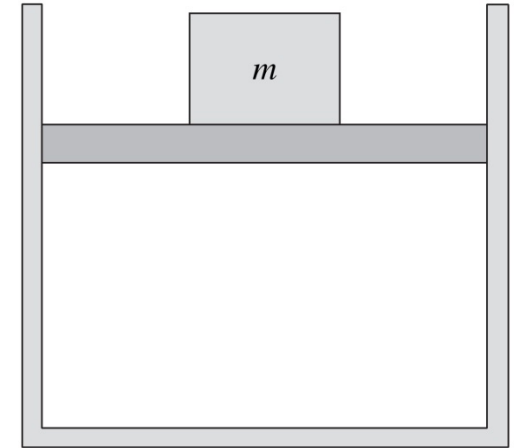
Constant Volume Process
(isochoric)

PV diagram?

In the experimental arrangement shown, how would you decrease the pressure without changing the temperature?

1. Can't be done, pressure is inversely proportional to volume.
2. Add more mass on top of the piston.
3. Remove some mass on top of the piston
4. Heat it.
5. Cool it.

PV diagram?

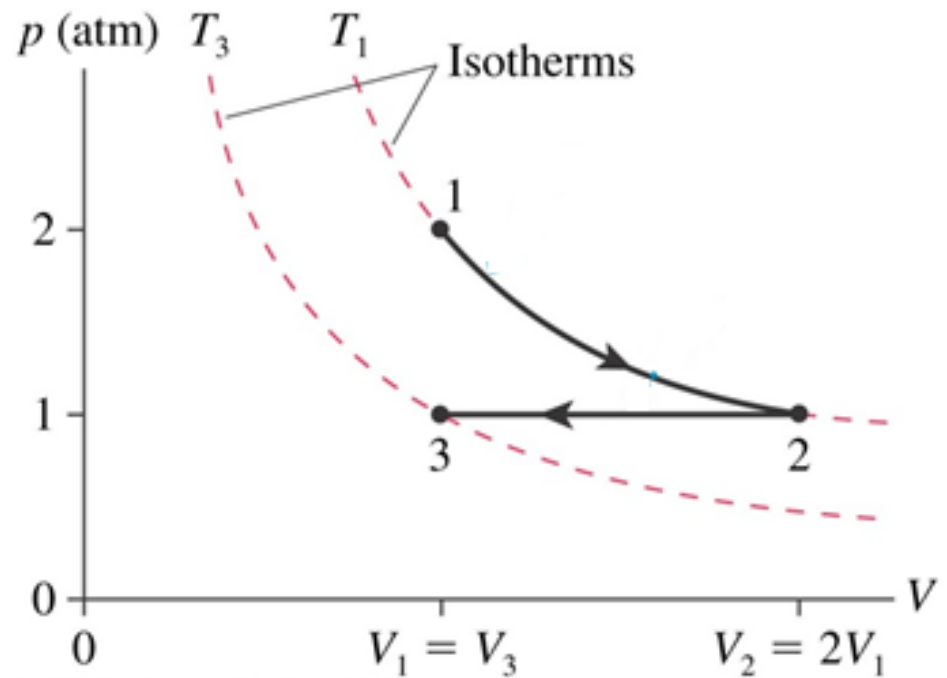


Constant Temperature Process
(isothermal)

By what factor does the absolute gas pressure change if the volume is tripled and the temperature is doubled?

1. 6
2. 1.5
3. $\frac{2}{3}$
4. $\frac{1}{6}$

A gas at 2.0 atm pressure and a temperature of 200°C is first expanded isothermally until its volume has doubled. It then undergoes an isobaric compression until it returns to its original volume. Find the final temperature and pressure.



Heat and Internal Energy

- Heat (Q)
 - 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.
- Internal Energy (U or E_{int})
 - Internal energy is all the energy of a system that is associated with its microscopic components: atoms and molecules when viewed from a reference frame at rest with respect to the center of mass of the system.
 - For any ideal gas, the internal energy (U) depends only on its temperature, not on its pressure or volume.

Internal Energy, Heat, and Work

- First Law of Thermodynamics $\Delta U = Q + W$
- Heat (Q) $Q = mc\Delta T$

Substance	Specific Heat (J/kg · °C)	Substance	Specific Heat (J/kg · °C)
<i>Elemental solids</i>		<i>Other solids</i>	
Aluminum	900	Brass	380
Beryllium	1 830	Glass	837
Cadmium	230	Ice (−5°C)	2 090
Copper	387	Marble	860
Germanium	322	Wood	1 700
Gold	129	<i>Liquids</i>	
Iron	448	Alcohol (ethyl)	2 400
Lead	128	Mercury	140
Silicon	703	Water (15°C)	4 186
Silver	234	<i>Gas</i>	
		Steam (100°C)	2 010

Note: To convert values to units of cal/g · °C, divide by 4 186.

Latent Heat

- Phase change
- Heat needed for a phase change

$$Q = \Delta mL$$

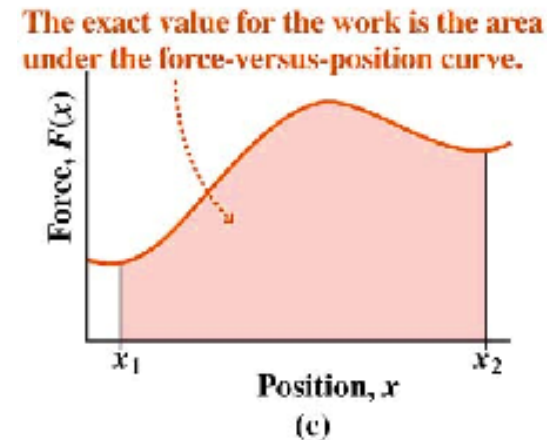
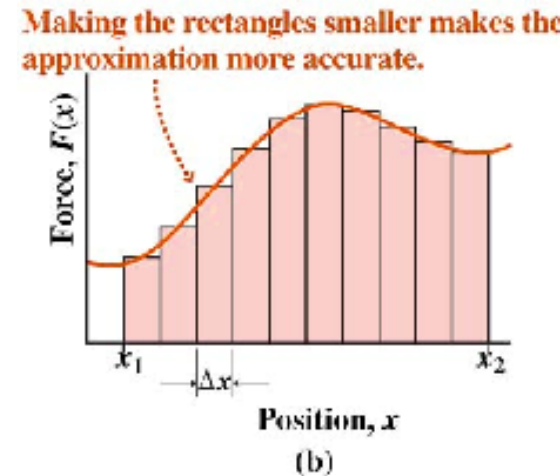
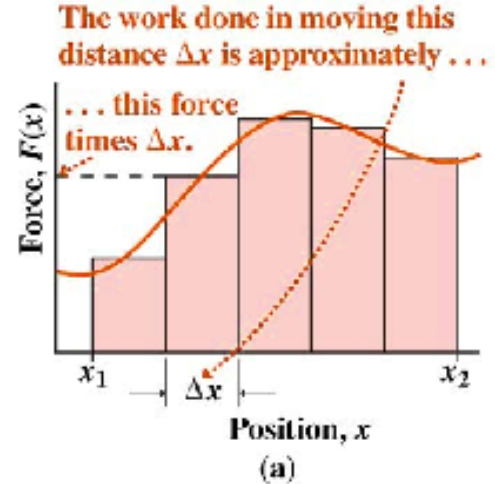
Substance	Melting Point (°C)	Latent Heat of Fusion (J/kg)	Boiling Point (°C)	Latent Heat of Vaporization (J/kg)
Helium	−269.65	5.23×10^3	−268.93	2.09×10^4
Oxygen	−218.79	1.38×10^4	−182.97	2.13×10^5
Nitrogen	−209.97	2.55×10^4	−195.81	2.01×10^5
Ethyl alcohol	−114	1.04×10^5	78	8.54×10^5
Water	0.00	3.33×10^5	100.00	2.26×10^6
Sulfur	119	3.81×10^4	444.60	3.26×10^5
Lead	327.3	2.45×10^4	1 750	8.70×10^5
Aluminum	660	3.97×10^5	2 450	1.14×10^7
Silver	960.80	8.82×10^4	2 193	2.33×10^6
Gold	1 063.00	6.44×10^4	2 660	1.58×10^6
Copper	1 083	1.34×10^5	1 187	5.06×10^6

Work Review

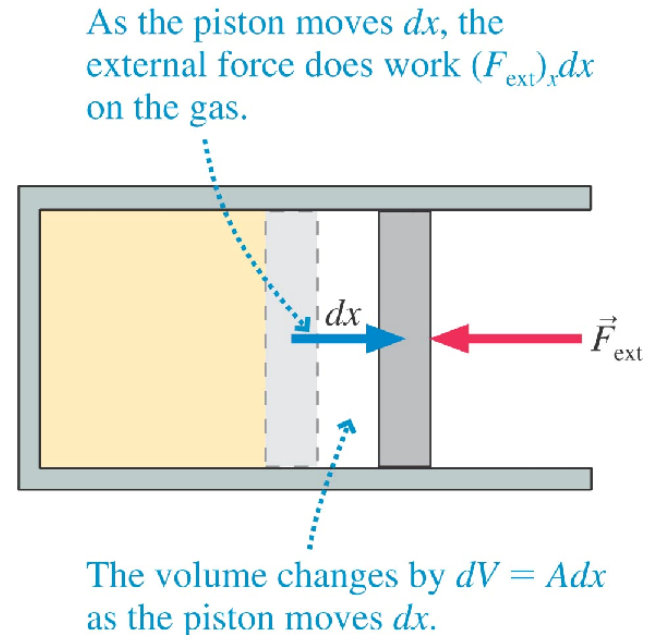
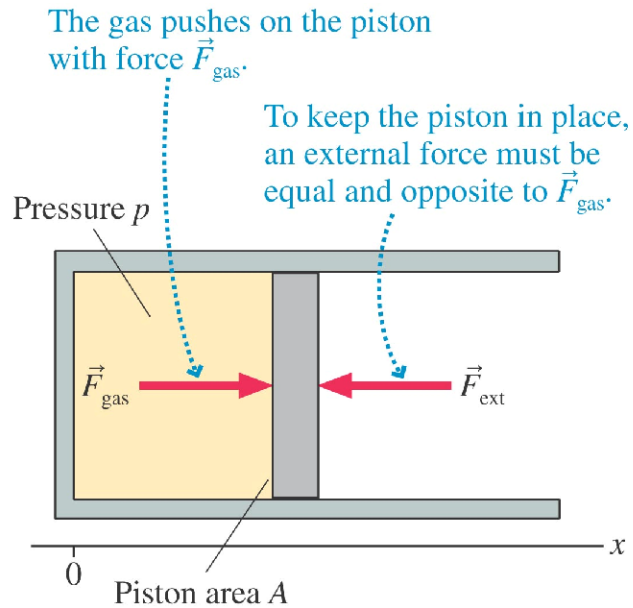
- Work: mechanical transfer of energy
- Work done by a varying force in one dimension

$$W = \int_{s_i}^{s_f} \vec{F}_s \cdot d\vec{s}$$

- SI unit: Joule (J)
- Work is a scalar quantity. It could be positive, negative or zero.
- Geometric interpretation: the area under the curve F_s between the limits.



Work in Ideal-Gas Processes



$$W_{\text{onGas}} = -\int_{V_i}^{V_f} P dV$$

$W_{\text{onGas}} > 0$ if the gas is compressed.

$W_{\text{onGas}} < 0$ if the gas is expanded.

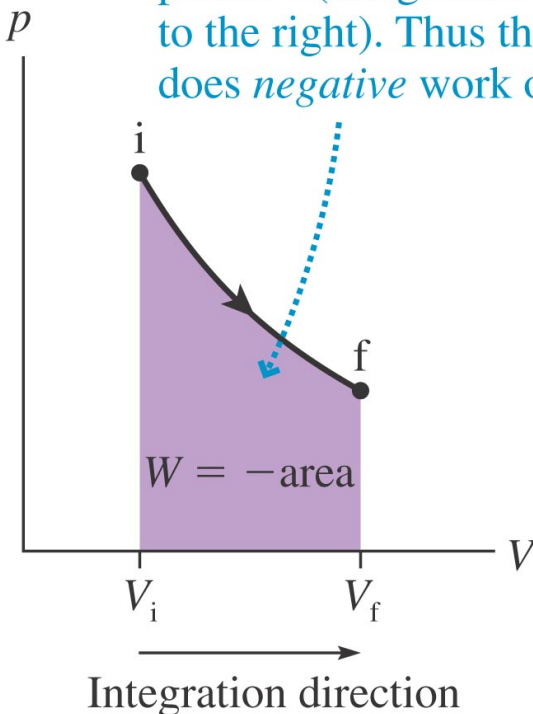
Finding Work from PV diagram

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

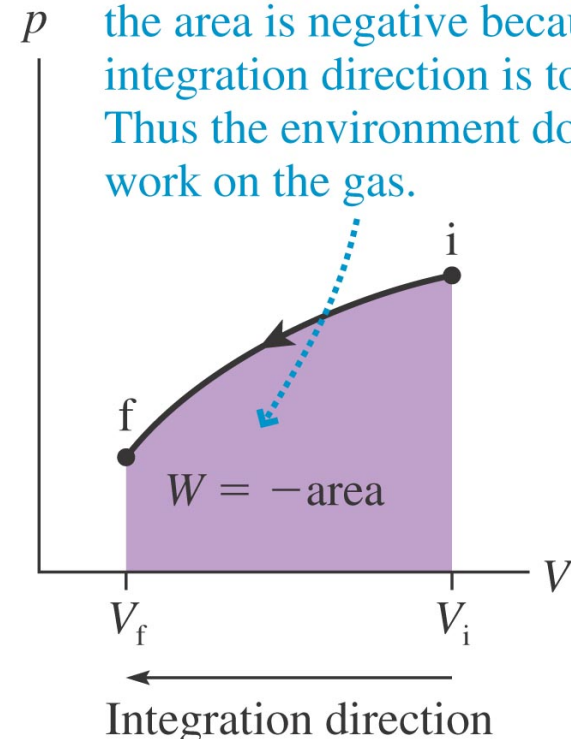
Geometric interpretation:

Work done on the gas is the negative area under the PV curve between the limits V_1 and V_2 .

(a) For an *expanding* gas ($V_f > V_i$), the area under the pV curve is positive (integration direction is to the right). Thus the environment does *negative* work on the gas.



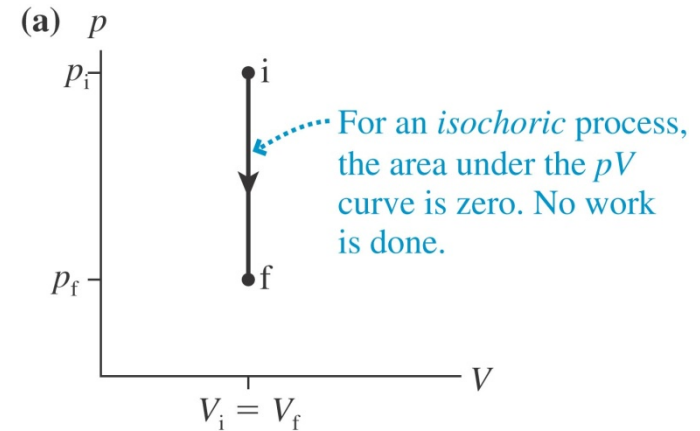
(b) For a *compressed* gas ($V_f < V_i$), the area is negative because the integration direction is to the left. Thus the environment does *positive* work on the gas.



Work in Isochoric & Isobaric Processes

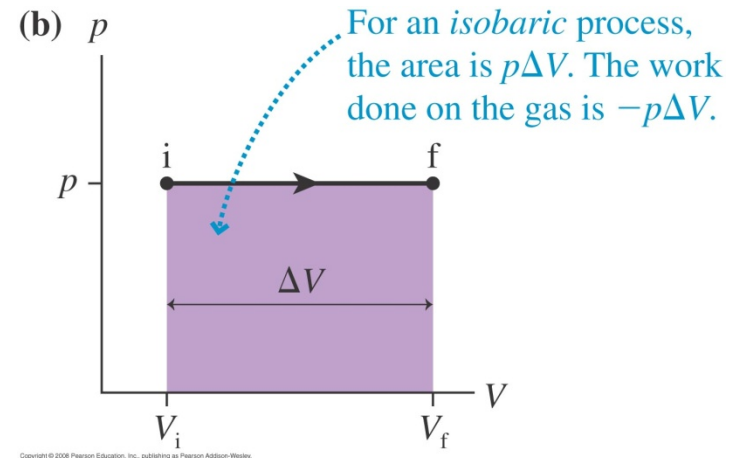
- Isochoric Process

$$W_{onGas} = -P\Delta V = 0$$



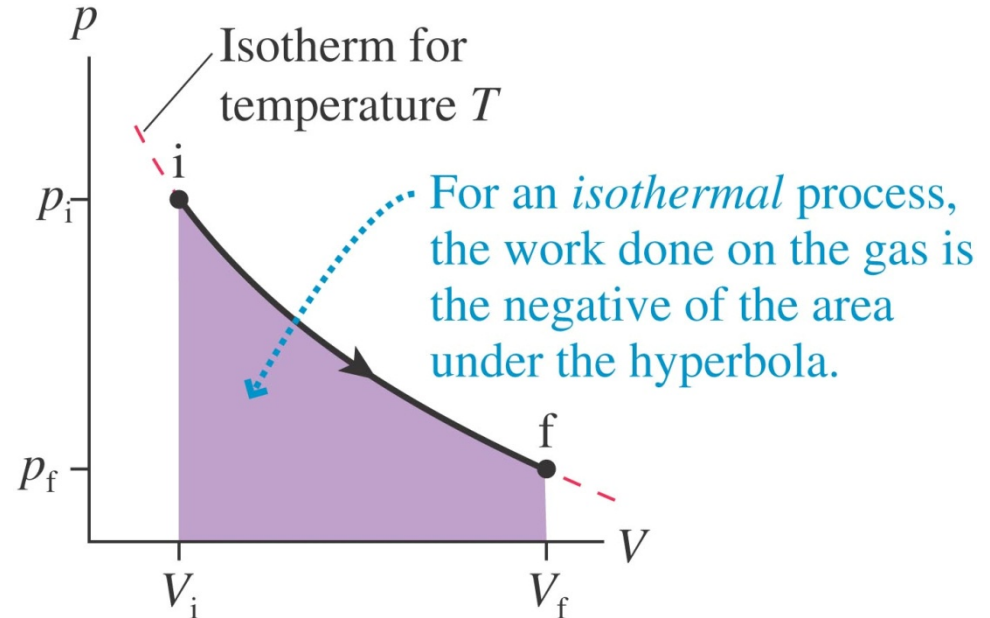
- Isobaric Process

$$W_{onGas} = -P\Delta V$$



Work in Isothermal Processes

$$W_{onGas} = -nRT \ln\left(\frac{V_f}{V_i}\right)$$



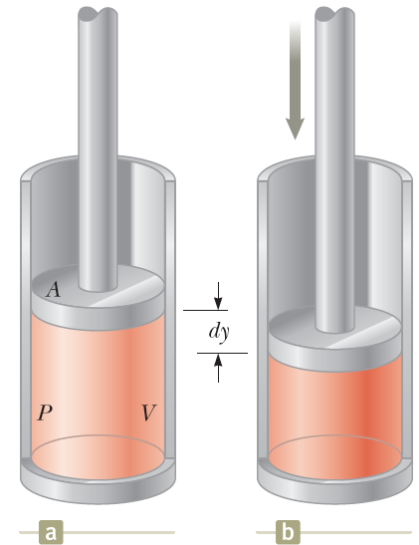
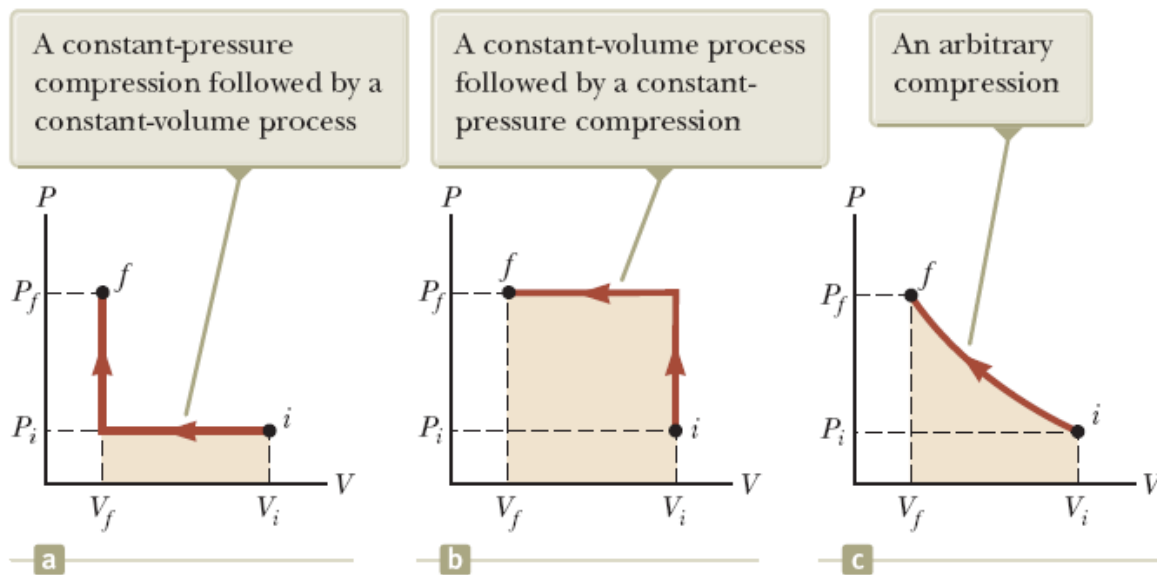
$$W = - \int_{V_i}^{V_f} P dV = - \int_{V_i}^{V_f} \frac{nRT}{V} dV \quad \left| \quad PV = nRT \right.$$

$$W = -nRT \int_{V_i}^{V_f} \frac{dV}{V} = -nRT \ln V \Big|_{V_i}^{V_f}$$

Work and Heat in Thermodynamics

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

Along which path the magnitude of the work done on the gas is the maximum?



- d. Same
- e. none of the above