EXAMPLE:

1. Given: The compressibility factor for a gas at 20°C is described by this equation for pressures up to 10 atm:

$Z = 1 - 2.024\times10^{-2}\,p$ with p in atm

The following properties of this gas are also known:

$C_V = 2.97 + 10.5\times10^{-3}\,T$ cal K^{-1} mol^{-1}

$C_p = 5.65 + 11.44\times10^{-3}\,T$ cal K^{-1} mol^{-1}

$(\partial U/\partial V)_T = 2\times10^{-3}$ cal L^{-1}

2.0 moles of this gas undergoes a reversible isothermal expansion at 20°C from 8.0 atm to 1.0 atm.

Calculate V_i, V_f, W, ΔU, q, ΔH.

EXAMPLE:

2. Given: One mole of a monatomic ideal gas is compressed adiabatically in a single stage with a constant opposing pressure equal to 10.0 atm. Initially the gas is at 27°C and 1.0 atm pressure. The final pressure is 10.0 atm.

Calculate the final temperature of the gas.

EXAMPLE:

3. Given: One mole of an ideal gas of $C_V = 5$ cal mol^{-1} K^{-1}, initially at 0°C and 1.0 atm pressure, is put through the following reversible cycle:

A: State 1 to state 2, heated at constant volume to twice the initial absolute temperature.

B: State 2 to state 3, expanded adiabatically until it is back to the initial temperature.

C: State 3 to state 1, compressed isothermally back to state 1.

Calculate q, W, ΔU, ΔH ***for steps A, B, and C.***

EXAMPLE:

1. Given: The compressibility factor for a gas at 20°C is described by this equation for pressures up to 10 atm:

$Z = 1 - 2.024\times10^{-2}\,p$ with p in atm

The following properties of this gas are also known:

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$(\partial \mathbf{U}/\partial V)_T = 2\times10^{-3}$ cal L^{-1}

2.0 moles of this gas undergoes a reversible isothermal expansion at 20°C from 8.0 atm to 1.0 atm.

Calculate V_i, V_f, W, ΔU, q, ΔH.

Draw a Picture:

Initial

$n = 2.0$ moles
$V = ?$
$p = 8.0$ atm
20°C

reversible
isothermal
$\rightarrow$
expansion

Final

$n = 2.0$ moles
$V = ?$
$p = 1$ atm
20°C

Properties of the gas:

$$Z = pV/nRT = 1 - 2.024\times10^{-2}\, p \qquad (p \text{ in atm})$$

for 20°C, up to 10 atm

$$C_V = 2.97 + 10.5\times10^{-3}\, T \ \text{cal K}^{-1}\,\text{mol}^{-1}$$

$$C_p = 5.65 + 11.44\times10^{-3}\, T \ \text{cal K}^{-1}\,\text{mol}^{-1}$$

$$(\partial \boldsymbol{U}/\partial V)_{\mathrm{T}} = 2\times10^{-3}\ \text{cal L}^{-1}$$

Question: Calculate V_i, V_f, W, $\Delta \boldsymbol{U}$, q, $\Delta \boldsymbol{H}$.

Principles and Definitions involved:

isothermal $dT = 0$

reversible $p_{\mathrm{op}} = p_{\mathrm{gas}}$

$\eth W = -\, p_{\mathrm{op}} dV$

The first law: $d\boldsymbol{U} = \eth q + \eth W$

$d\boldsymbol{U} = (\partial \boldsymbol{U}/\partial T)_V dT + (\partial \boldsymbol{U}/\partial V)_T dV$

$d\boldsymbol{H} = (\partial \boldsymbol{H}/\partial \mathrm{T})_p dT + (\partial \boldsymbol{H}/\partial p)_T dp$

$C_V = (\partial \boldsymbol{U}/\partial T)_V$ $\quad C_p = (\partial \boldsymbol{H}/\partial \mathrm{T})_p$

Solution:

Given the equation of state,

rearrange $pV/nRT = 1 - 2.024\times10^{-2}\, p$ to

$$V = nRT\, \{(1/p) - 2.024\times10^{-2}\}$$

can solve for V, given T and p

$V_i = \ldots$ L , $V_f = \ldots$ L

using $R = 8.20578\times10^{-2}$ L atm K^{-1} mol^{-1}

$\partial W = -p_{op}\mathrm{d}V$

reversible, ∴ $p_{op} = p_{gas}$

$W = -\int p_{gas}\,\mathrm{d}V$

$= -\int\{\mathrm{nRT}/(\mathrm{nRT}2.024\times10^{-2} + V)\}\,\mathrm{d}V$

$W = -2\mathrm{R}293\int \mathrm{d}V/(2\mathrm{R}\times293\times2.024\times10^{-2} + V)$

$= -2\mathrm{R}293\times\ln[\{2\mathrm{R}\times293\times2.024\times10^{-2} + V_2\}$

$/\{2\mathrm{R}\times293\times2.024\times10^{-2} + V_1\}]$

$\mathrm{R} = 0.082057\ \mathrm{L\ atm\ mol^{-1}\ K^{-1}}$

$W = \ldots\ldots$ L atm

$\Delta \boldsymbol{U}$: $\mathrm{d}\boldsymbol{U} = C_V\mathrm{d}T + (\partial \boldsymbol{U}/\partial V)_T\mathrm{d}V$

isothermal, ∴ $\mathrm{d}T = 0$ $\Delta \boldsymbol{U} = \int(\partial \boldsymbol{U}/\partial V)_T\,\mathrm{d}V$

Given, $(\partial \boldsymbol{U}/\partial V)_T = 2\times10^{-3}\ \mathrm{cal\ L^{-1}}$,

$\Delta \boldsymbol{U} = \int 2\times10^{-3}\,\mathrm{d}V$

$\Delta \boldsymbol{U} = 2\times10^{-3}\,(V_2 - V_1)$ cal

---

q: First law, $\Delta \boldsymbol{U} = q + W$

Change W to cal by multiplying by

{1.98718 cal/0.082057 L atm},

$q = \ldots\ldots$ cal

$\Delta \boldsymbol{H}$: $\quad \mathrm{d}\boldsymbol{H} = C_p \mathrm{d}T + (\partial \boldsymbol{H}/\partial p)_T \mathrm{d}p$

isothermal, $\therefore$ $\mathrm{d}T = 0$ $\quad \Delta \boldsymbol{H} = \int (\partial \boldsymbol{H}/\partial p)_T \,\mathrm{d}p$

$$(\partial \boldsymbol{H}/\partial p)_T = [p + (\partial \boldsymbol{U}/\partial V)_T](\partial V/\partial \mathrm{p})_T \quad +V$$

Given: $V = nRT\{ (1/p) - 2.024\times10^{-2} \}$,

which differentiates to

$$(\partial V/\partial p)_T = -nRTp^{-2}$$

Given, $(\partial \boldsymbol{U}/\partial V)_T = 2\times10^{-3}$ cal L^{-1}, convert to atm by

$\times\{0.082057 \text{ L atm} / 1.98718 \text{ cal}\}$

$\therefore$ $(\partial \boldsymbol{H}/\partial p)_T =$

$$[p + 2\times10^{-3}(.082057/1.98718)] \times -nRTp^{-2} \quad +V$$

$$= -nRT\{ p^{-1} + 2\times10^{-3}(.082057/1.98718)\, p^{-2}\} \; + \; nRT\,\{p^{-1} - 2.024\times10^{-2}\}$$

$$= -nRT\{ 2\times10^{-3}(.082057/1.98718)p^{-2} + 2.024\times10^{-2}\}$$

$$\Delta \boldsymbol{H} = \int (\partial \boldsymbol{H}/\partial p)_T \,\mathrm{d}p$$

$$= -nRT \int \{8.2586\times10^{-5}p^{-2} + 2.024\times10^{-2}\}\,\mathrm{d}p$$

$$= nRT\{8.2586\times10^{-5}[p^{-1}]^{1}_{8} + 2.024\times10^{-2}\,[1-8]\}$$

$$= nRT\{8.2586\times10^{-5}\times7/8 + 2.024\times10^{-2}\times7\}$$

$$= nRT\times2.024\times10^{-2}\times7$$

$= 2\times293\times1.987\times2.024\times10^{-2}\times7$ cal or

$= 2\times293\times8.3144\times2.024\times10^{-2}\times7$ joule

EXAMPLE:

2. Given: One mole of a monatomic ideal gas is compressed adiabatically in a single stage with a constant opposing pressure equal to 10.0 atm. Initially the gas is at 27°C and 1.0 atm pressure. The final pressure is 10.0 atm. ***Calculate the final temperature of the gas.***

Draw a Picture:

Initial

$n = 1$ mole
$V = ?$
$p = 1$ atm
27°C

monatomic
ideal gas

adiabatic
$p_{op} = 10$
→
compression

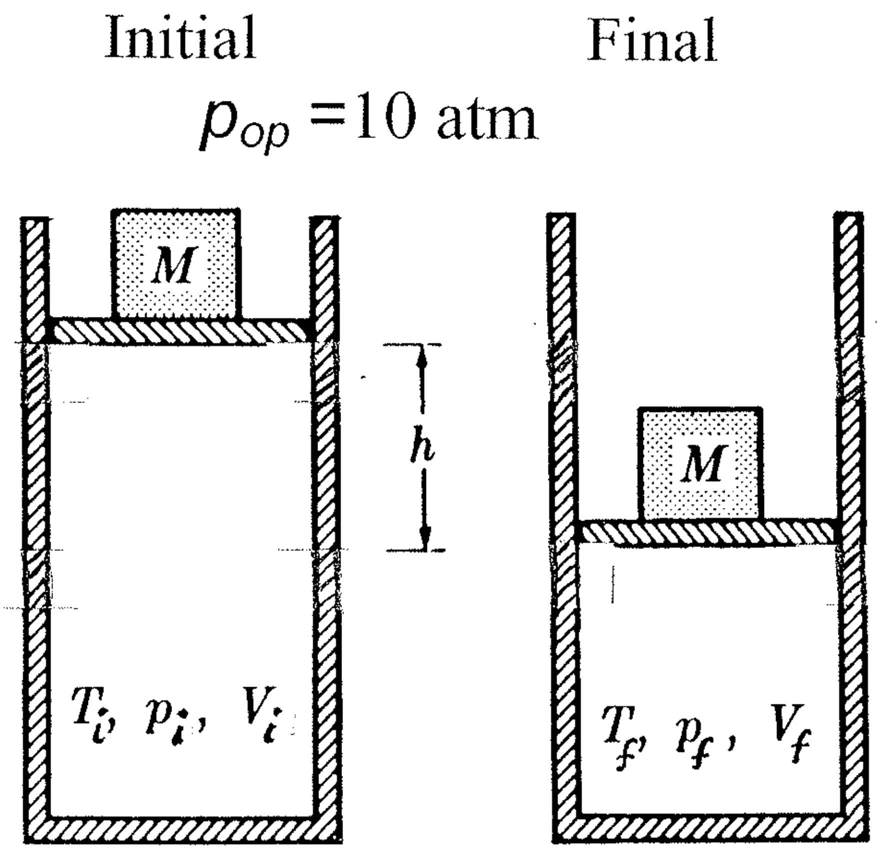
Final

$n = 1$ mole
$V = ?$
$p = 10$ atm
$T = ?$

Question:

$T_f = ?$

Draw a Picture:



adiabatic compression
monatomic ideal gas

n = 1 mole	n = 1 mole
p_i = 1 atm	p_f = 10 atm
T_i = 27+273.15	T_f =
V_i =?	V_f =?

Question: $T_f = ?$

Principles and Definitions involved:

adiabatic $\therefore\ q = 0$

$\eth W = -p_{\mathrm{op}}\mathrm{d}V$

The first law: $\mathrm{d}\boldsymbol{U} = \eth q + \eth W$

$\mathrm{d}\boldsymbol{U} = (\partial \boldsymbol{U}/\partial T)_V\,\mathrm{d}T + (\partial \boldsymbol{U}/\partial V)_T\,\mathrm{d}V$

$C_V = (\partial \boldsymbol{U}/\partial T)_V$

Ideal gas: $pV = nRT$ $\quad (\partial \boldsymbol{U}/\partial V)_T = 0$

monatomic: $C_V = (3/2)R$

Solution:

$pV=nRT$

using $R = 8.20578\times10^{-2}$ L atm K^{-1} mol^{-1}

$V_i = 1.0\ \mathrm{mol} \times R \times 300\ \mathrm{K} / 1.0\ \mathrm{atm}$

$= 300R$ L

$V_f = 1.0\ \mathrm{mol} \times R \times T_f\ \mathrm{K} / 10.\ \mathrm{atm}$

$= RT_f/10$ L

{ Can not get T_f from equation of state alone. Need to use the fact that the process is adiabatic. }

q: $q = 0$ adiabatic

W: $\delta W = -\, p_{op}\, dV$ $p_{op} = 10$ atm

$$W = -\int_{V_i}^{V_f} 10\, \mathrm{d}V = -10[V_f - V_i]$$

$$= -10\, [RT_f/10 - 300\, R\,]$$

$\Delta \boldsymbol{U}$: $d\boldsymbol{U} = C_V dT + (\partial \boldsymbol{U}/\partial V)_T dV$

for an ideal gas only, $(\partial \boldsymbol{U}/\partial V)_T = 0$

$\therefore\ \Delta \boldsymbol{U} = \int C_V\, \mathrm{d}T + 0$ $C_V = (3/2)R$ (monatomic)

$$\Delta \boldsymbol{U} = \int_{300}^{T_f} (3/2)R\, \mathrm{d}T = (3/2)R\, [T_f - 300]$$

First law: $\Delta \boldsymbol{U} = q + W$

$(3/2)R\, [T_f - 300] = 0 + -10R\, [T_f/10 - 300]$

Solve for T_f

EXAMPLE:

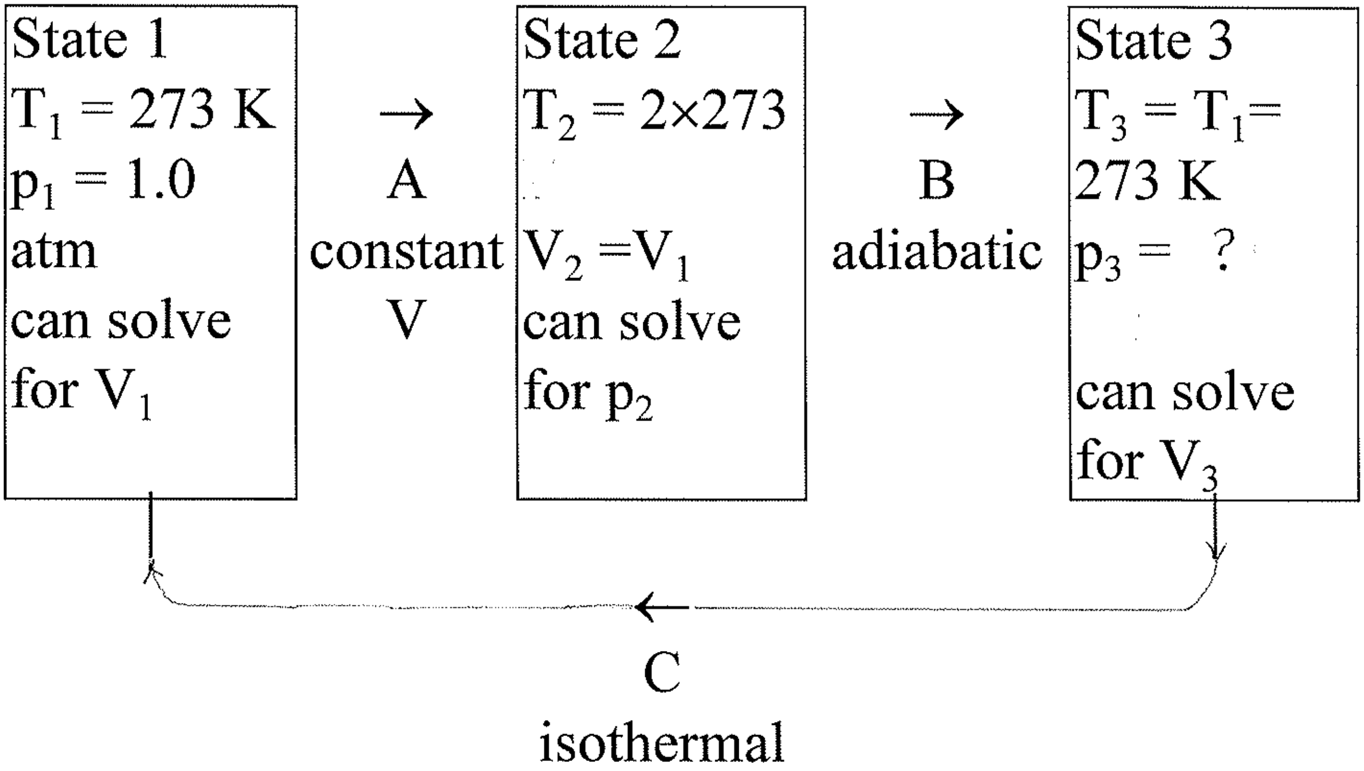
3. Given: One mole of an ideal gas of $C_V = 5$ cal $\text{mol}^{-1}\ \text{K}^{-1}$, initially at 0°C and 1.0 atm pressure, is put through the following reversible cycle:

A: State 1 to state 2, heated at constant volume to twice the initial absolute temperature.

B: State 2 to state 3, expanded adiabatically until it is back to the initial temperature.

C: State 3 to state 1, compressed isothermally back to state 1.

Calculate q , W , ΔU , ΔH for steps A, B, and C.



State 1
$T_1 = 273$ K
$p_1 = 1.0$ atm
can solve for V_1

→ A constant V

State 2
$T_2 = 2\times273$
$V_2 = V_1$
can solve for p_2

→ B adiabatic

State 3
$T_3 = T_1 =$ 273 K
$p_3 =$?
can solve for V_3

← C isothermal

Principles and Definitions involved:

$$\text{đ}W = -\, p_{\text{op}}\,\mathrm{d}V$$

The first law: $\mathrm{d}\boldsymbol{U} = \text{đ}q + \text{đ}W$

$\mathrm{d}\boldsymbol{U} = (\partial \boldsymbol{U}/\partial T)_V\,\mathrm{d}T + (\partial \boldsymbol{U}/\partial V)_T\,\mathrm{d}V$

$\mathrm{d}\boldsymbol{H} = (\partial \boldsymbol{H}/\partial \mathrm{T})_p\,\mathrm{d}T + (\partial \boldsymbol{H}/\partial p)_T\,\mathrm{d}p$

$C_V = (\partial \boldsymbol{U}/\partial T)_V \qquad C_p = (\partial \boldsymbol{H}/\partial \mathrm{T})_p$

Ideal gas: $pV = nRT \qquad (\partial \boldsymbol{U}/\partial V)_T = 0$

$C_p - C_V = R \qquad (\partial \boldsymbol{H}/\partial p)_T = 0$

reversible $p_{\text{op}} = p_{\text{gas}}$

Step A: constant volume $\mathrm{d}V = 0$

Step B: adiabatic $q = 0$

Step C: isothermal $\mathrm{d}T = 0$

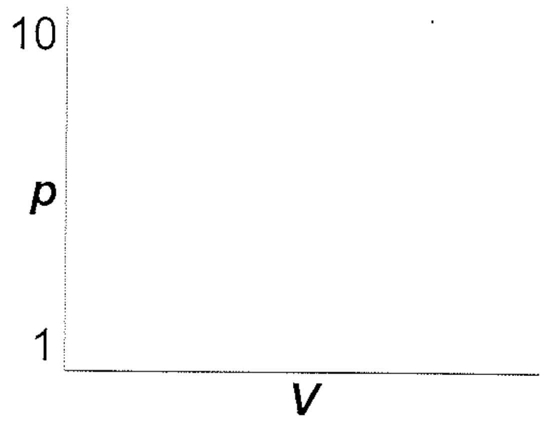
C_V = 5.0 cal mol^{-1} K^{-1} , ideal gas, C_p-C_V=R, ∴ C_p= 7 cal mol^{-1} K^{-1}

A reversible constant V, V_2=V_1	B reversible, adiabatic T_3=T_1	C reversible isothermal T_1
T_i =T_1=273 K T_f =T_2=2×273K p_1 = 1 atm, solve for V_1	T_i =T_2=2×273K T_f =T_3=273 K	T_i =T_3=273K T_f =T_1=273 K p_3 = ?, V_3= ?
q: $\Delta U= q+W$ take ΔU,W from below or else $C_V\equiv \text{đ}q_V/dT$ $q_V=\int C_V\,dT$	**q:** q = 0 (adiabatic) ∴ can solve for V_3 by $dU= \text{đ}q+\text{đ}W= 0+\text{đ}W$ $C_VdT = -p_{gas}dV$ $C_VdT/T = -\mathrm{R}dV/V$ $5\ln(T_3/T_2) = -\mathrm{R}\ln(V_3/V_1)$	**q:** $\Delta U= q+W$ take ΔU, W from below
W $\text{đ}W = -p_{op}dV$ given dV = 0, ∴ W = 0	**W:** $\Delta U= q+W$ $\Delta U= 0+W$ take ΔU from below	**W:** $\text{đ}W = -p_{op}dV$ $p_{op}= p_{gas}$ (reversible) $W =-\int(RT/V)dV$ $=-R273\ln(V_1/V_3)$
ΔU $dU = C_VdT$ $+(\partial U/\partial V)_TdV$ ideal gas, ∴ $(\partial U/\partial V)_T = 0$ or given dV=0, ∴ $\Delta U=\int C_V\,dT$	**ΔU** $dU = C_VdT$ $+(\partial U/\partial V)_TdV$ as on the left ...... $\Delta U=\int C_V\,dT$	**ΔU** $dU = C_VdT$ $+(\partial U/\partial V)_TdV$ ideal gas, ∴ $(\partial U/\partial V)_T = 0$ and given dT=0 ∴ ΔU= 0
ΔH $dH = C_pdT$ $+(\partial H/\partial p)_Tdp$ ideal gas, ∴ $(\partial H/\partial p)_T = 0$ ∴ $\Delta H=\int C_p\,dT$	**ΔH** $dH = C_pdT$ $+(\partial H/\partial p)_Tdp$ as on the left ...... $\Delta H=\int C_p\,dT$	**ΔH** $dH = C_pdT$ $+(\partial H/\partial p)_Tdp$ ideal gas, ∴ $(\partial H/\partial p)_T = 0$ given dT=0, ∴ ΔH=0

EXAMPLE:

Assume that air behaves as an ideal gas with $C_p = (7/2)R$.

(a) In one experiment 1.00 mole of “air molecules” is compressed from 1.00 atm to 10.0 atm at 25°C by the following reversible process: (1) heating at constant volume to the final pressure followed by (2) cooling at constant pressure to 25°C. Sketch these processes on a *pV* diagram.



Calculate ΔU, ΔH, q, and W, in kJ for each step in the process and for the overall process.

EXAMPLE:

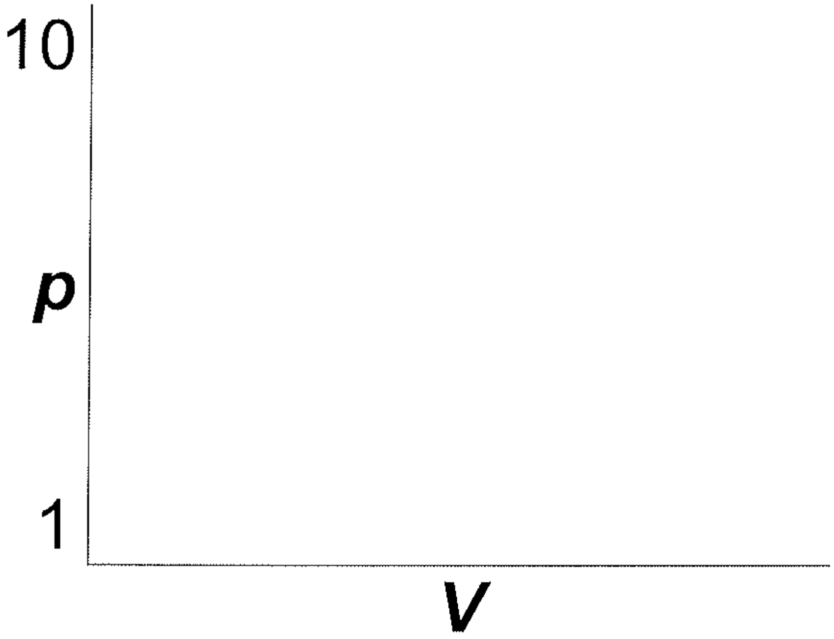
n moles of a gas obeying the eqn. of state $p(V-nb) = nRT$ *where* $b = 0.1\ \mathrm{L\ mol^{-1}}$ *and has* $(\partial U/\partial V)_T = 0$, $(\partial H/\partial p)_T = -b$ is subjected to an isothermal reversible expansion from an initial volume of 1.00 L to 24.89 L at 298 K.

Calculate the values of ΔU, ΔH, q, and W, in kJ in terms of *n*.

EXAMPLE:

Assume that air behaves as an ideal gas with $C_p = (7/2)R$.

(a) In one experiment 1.00 mole of “air molecules” is compressed from 1.00 atm to 10.0 atm at 25°C by the following reversible process: (1) heating at constant volume to the final pressure followed by (2) cooling at constant pressure to 25°C. Sketch these processes on a *pV* diagram.



Calculate ΔU, ΔH, q, and W, in kJ for each step in the process and for the overall process.

Principles and equations for this problem:

Ideal gas, for which: $\boldsymbol{pV} = n\boldsymbol{R}\boldsymbol{T}$ $(\partial \boldsymbol{U}/\partial V)_T = 0$

$(\partial \boldsymbol{H}/\partial p)_T = 0$, $C_p - C_V = R$

$C_p = (7/2)R$, given

ideal gas $\therefore$ $C_V = C_p - R = (5/2)R$

definition of heat capacity C: $\boldsymbol{q} = \int C \, \mathrm{d}\boldsymbol{T}$

First law $\Delta \boldsymbol{U} = \boldsymbol{q} + \boldsymbol{W}$

$\mathrm{d}W = -p_{op}\mathrm{d}V$

$\boldsymbol{U}(\boldsymbol{T},\boldsymbol{V})$: $\mathrm{d}\boldsymbol{U} = C_V \mathrm{d}T + (\partial \boldsymbol{U}/\partial V)_T \mathrm{d}V$

$\boldsymbol{H}(\boldsymbol{T},\mathrm{p}\boldsymbol{V})$: $\mathrm{d}\boldsymbol{H} = C_p \mathrm{d}T + (\partial \boldsymbol{H}/\partial p)_T \mathrm{d}p$

definition of ***reversible*** process: $p_{op} = p_{gas}$

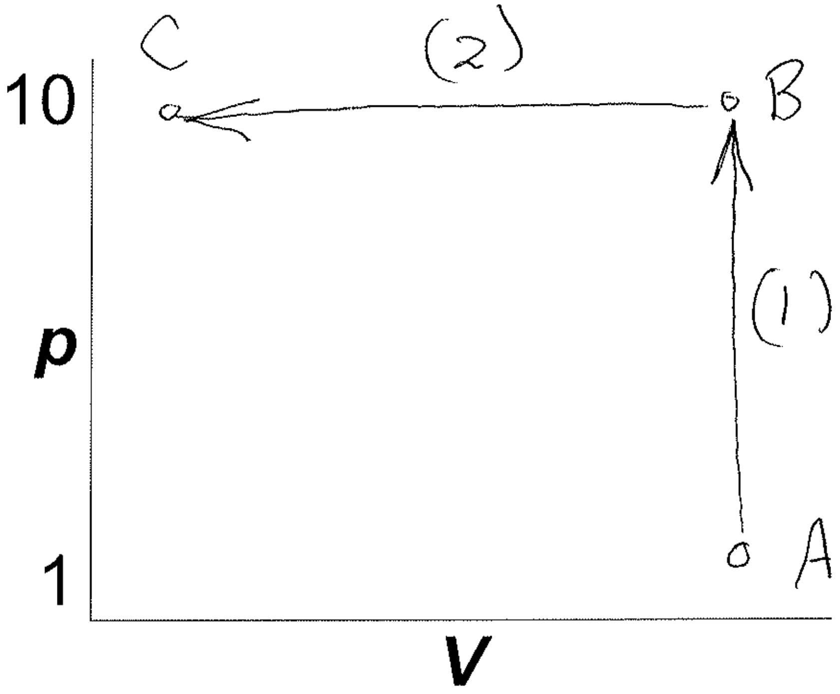
Solution:

step 1 A→B constant volume

step 2 B→C constant pressure

(state)$_A$	(state)$_B$	(state)$_C$
$p_A = 1$ given	$p_B = 10$ given	$p_C = 10$ given
$T_A = 298$ given	$T_B = 10 V_A/R$	$T_C = 298$ given
$V_A = R\,298/1$	$V_B = V_A$ given	$V_C = R\,298/10$

n=1, p in atm, T in K, V in L



step 1 dV= 0	step 2 dp=0	overall= 1+2
$\boldsymbol{q} = \boldsymbol{q}_V =$ $\int_{T_A}^{T_B} C_V \, dT =$ (5/2)8.314[2980-298] = 55.748 kJ	$\boldsymbol{q} = \boldsymbol{q}_p =$ $\int_{T_B}^{T_C} C_p \, dT =$ (7/2)8.314[298-2980] = -78.047 kJ	55.748 + (-78.047) = -22.3 kJ
d$\boldsymbol{W} = -\boldsymbol{p}_{op}dV$ dV= 0 $\therefore$ $\boldsymbol{W}$ = 0	$\boldsymbol{W} = \Delta \boldsymbol{U} - \boldsymbol{q}$ = -55.748 –(-78.047) = +22.3 kJ	0 + 22.3 kJ
d$\boldsymbol{U}$ = C_VdT + $(\partial \boldsymbol{U}/\partial V)_TdV$ dV=0 $\Delta \boldsymbol{U}$= $\int_{T_A}^{T_B} C_V \, dT$	d$\boldsymbol{U}$ = C_VdT + $(\partial \boldsymbol{U}/\partial V)_TdV$ $(\partial \boldsymbol{U}/\partial V)_T$ =0 $\Delta \boldsymbol{U}$= $\int_{T_B}^{T_C} C_V \, dT$ =(5/2)8.314[298-2980] = -55.748 kJ	$\Delta \boldsymbol{U}$=0 since $T_C = T_A$ and ideal
d$\boldsymbol{H}$ = C_pdT + $(\partial \boldsymbol{H}/\partial p)_Tdp$ $(\partial \boldsymbol{H}/\partial p)_T$ =0 $\Delta \boldsymbol{H}$= $\int_{T_A}^{T_B} C_p \, dT$ (7/2)8.314[2980-298] = 78.047 kJ	d$\boldsymbol{H}$ = C_pdT + $(\partial \boldsymbol{H}/\partial p)_Tdp$ dp=0 $\Delta \boldsymbol{H}$= $\int_{T_B}^{T_C} C_p \, dT$ (7/2)8.314[298-2980] = -78.047 kJ	$\Delta \boldsymbol{H}$=0 since $T_C = T_A$ and ideal

Note that d$\boldsymbol{W}$ = -$\boldsymbol{p}_{op}$dV *could also have been used for step 2, giving the same answer:*

since $\boldsymbol{p}_{op}$ *= constant= 10 atm in step 2,*

$$W = -10 \int_{V_B}^{V_C} dV$$

$$= -10 atm \int_{R298/1}^{R298/10} dV$$

$$= -10\left[\frac{R298}{10} - \frac{R298}{1}\right]$$

$$= 220.077 \text{ L atm} \times \frac{0.0083144 \text{ kJ}}{0.082057 \text{ L atm}}$$

= 22.3 kJ

EXAMPLE:

n moles of a gas obeying the eqn. of state $p(V-nb) = nRT$ where $b=0.1$ L mol^{-1} and has $(\partial \mathbf{U}/\partial V)_T=0$, $(\partial \mathbf{H}/\partial p)_T= - b$ is subjected to an isothermal reversible expansion from an initial volume of 1.00 L to 24.89 L at 298 K.

Calculate the values of $\Delta \mathbf{U}$, $\Delta \mathbf{H}$, $\mathbf{q}$, and $\mathbf{W}$, in kJ in terms of n.

Principles and equations for this problem:

First law $\Delta \boldsymbol{U} = \boldsymbol{q} + \boldsymbol{W}$

$\mathrm{d}W = -p_{op}\mathrm{d}V$

$\boldsymbol{U}(\boldsymbol{T},\boldsymbol{V})$: $\mathrm{d}\boldsymbol{U} = C_V\mathrm{d}T + (\partial \boldsymbol{U}/\partial V)_T\mathrm{d}V$

$\boldsymbol{H}(\boldsymbol{T},\mathrm{p})$: $\mathrm{d}\boldsymbol{H} = C_p\mathrm{d}T + (\partial \boldsymbol{H}/\partial p)_T\mathrm{d}p$

definition of ***reversible*** process: $p_{op} = p_{gas}$

definition of ***isothermal*** process: $T_{final} = T_{initial}$

Given:

eqn. of state: $p(V-nb) = nRT$ $b=0.1\ L\ mol^{-1}$

$(\partial \boldsymbol{U}/\partial V)_T=0$, $(\partial \boldsymbol{H}/\partial p)_T = -b$

Solution:

$(\text{state})_{\text{initial}}$	$(\text{state})_{\text{final}}$
$T_i = 298$ K given	$T_f = 298$ K given
$V_i = 1$ L given	$V_f = 24.8$ L given
$p_i = \dfrac{nR298}{(1.0-0.1n)}$	$p_f = \dfrac{nR298}{(24.8-0.1n)}$

$$\mathrm{d}\boldsymbol{W} = -\boldsymbol{p}_{op}\mathrm{d}V \qquad p_{op} = p_{gas}$$

$$\boldsymbol{W} = -\int \boldsymbol{p}_{gas}\mathrm{d}V$$

$$= -\int_{V_i}^{V_f} \frac{nR\ 298}{V - nb} dV \quad = -nR\,298 \ln\left(\frac{24.8-0.1n}{1.0-0.1n}\right)$$

$$\mathrm{d}\boldsymbol{U} = C_V\mathrm{d}T + (\partial \boldsymbol{U}/\partial V)_T\mathrm{d}V$$

$\mathrm{d}T = 0$, $(\partial \boldsymbol{U}/\partial V)_T = 0$ *given*, $\therefore\ \Delta \boldsymbol{U} = 0$

$$\boldsymbol{q} = \Delta \boldsymbol{U} - \boldsymbol{W} = -\boldsymbol{W} \quad = nR\,298 \ln\left(\frac{24.8-0.1n}{1.0-0.1n}\right)$$

$$\mathrm{d}\boldsymbol{H} = C_p\mathrm{d}T + (\partial \boldsymbol{H}/\partial p)_T\mathrm{d}p$$

$\mathrm{d}T = 0$, $(\partial \boldsymbol{H}/\partial p)_T = -b$, $\therefore \Delta \boldsymbol{H} = \int(\partial \boldsymbol{H}/\partial p)_T\mathrm{d}p =$

$$-b\int_{p_i}^{p_f} dp$$

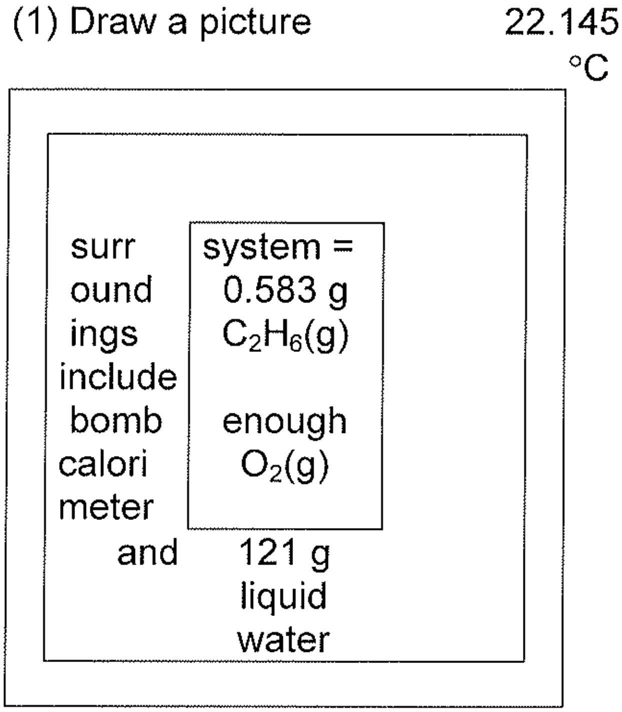
$$= -0.1\times n0.0083144\times 298\left[\frac{1}{24.8-0.1n} - \frac{1}{1.0-0.1n}\right]\text{kJ}$$

Chemistry 116

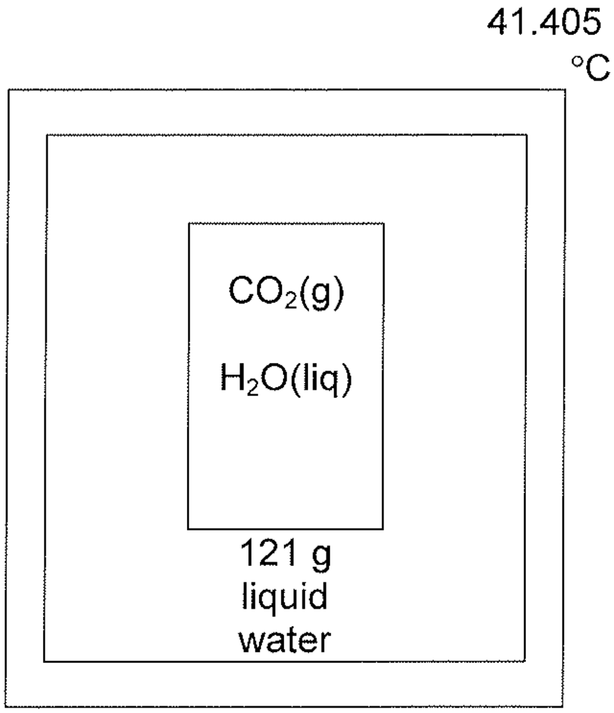
Quiz 1 on Chapter 7-8 Name ______________________

Problem: A 0.583 g sample of $C_2H_6(g)$ is burned in a bomb calorimeter containing 121 g of water as the heat exchange substance. The temperature rises from 22.145 to 41.405 °C. Taking the heat capacity of the calorimeter to be 1.0489 kJ/°C calculate q_V the molar heat of combustion of $C_2H_6(g)$ at constant volume, in kJ mol^{-1}. What is ΔE for the combustion of C_2H_6? Assuming that the gases behaved ideally, what is ΔH_c for C_2H_6? The specific heat capacity of liquid water is 4.184 J g^{-1}/°C. R = = 8.31451 J K^{-1} mol^{-1}

(1) Draw a picture



initial



final

(2) molar mass of C_2H_6 = ________g mol^{-1}

moles of $C_2H_6(g)$ = __________mol

(3) definition of heat capacity: $q = C\,\Delta T$

surroundings: q_{surr} =

__________kJ

(4) $\therefore$ $q_{system} = q_V =$ __________kJ

molar heat of combustion of C_2H_6(g) at constant volume, molar $q_V =$

______________kJ mol^{-1}.

(5) First law of thermodynamics leads to $\Delta E = q_V$ (general, when only pressure volume work is possible)

$\therefore \Delta E$ of combustion of C_2H_6(g) = _______________kJ mol^{-1}.

(6) Definition of enthalpy: $H = E + pV$

Ideal gas equation of state: $pV = nRT$

The balanced chemical equation:

(neglect the volume occupied by the liquid water resulting from the combustion.)

For 1 mole of C_2H_6 :

gases initial:____________________ gases final:_________________.

calculate ΔH here:

$\therefore \Delta H$ of combustion of C_2H_6(g)

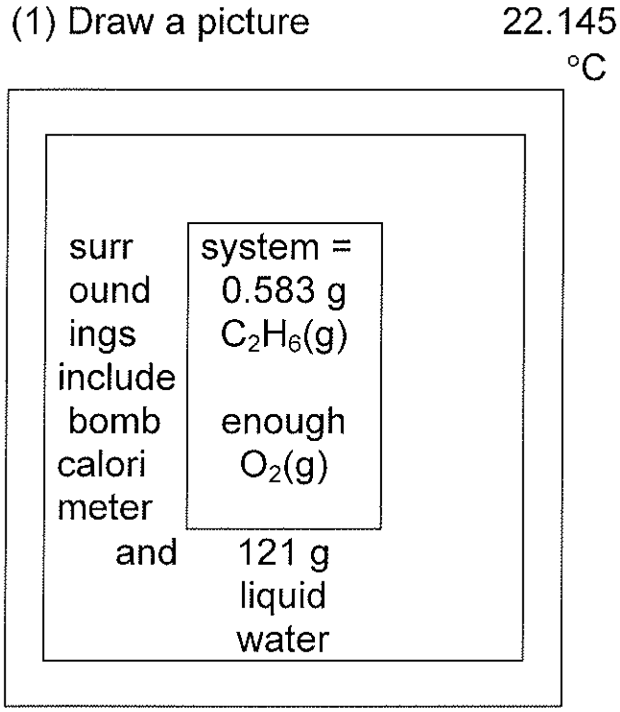
= _______________ kJ mol^{-1}

Chemistry 116 Perfecto Excellencio

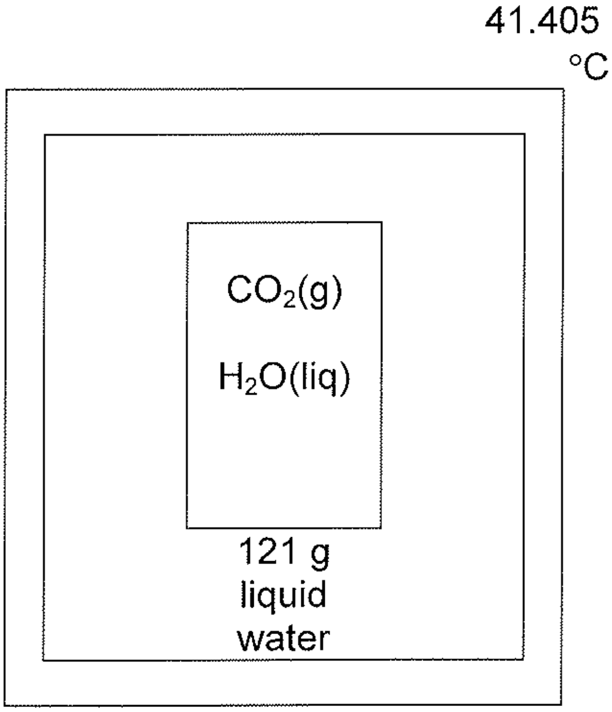
Quiz 1 on Chapter 7-8 Name

Problem: A 0.583 g sample of $C_2H_6(g)$ is burned in a bomb calorimeter containing 121 g of water as the heat exchange substance. The temperature rises from 22.145 to 41.405 °C. Taking the heat capacity of the calorimeter to be 1.0489 kJ/°C calculate q_V the molar heat of combustion of $C_2H_6(g)$ at constant volume, in kJ mol^{-1}. What is ΔE for the combustion of C_2H_6? Assuming that the gases behaved ideally, what is ΔH_c for C_2H_6 ? The specific heat capacity of liquid water is 4.184 J g^{-1}/°C. R = = 8.31451 J K^{-1} mol^{-1}

(1) Draw a picture



initial



final

(2) molar mass of $C_2H_6 = 2\times12+6\times1.008$ 30.048 g mol^{-1}

moles of $C_2H_6(g)$ = 0.583/30.048 g mol^{-1} = 0.0194 mol

(3) definition of heat capacity: $q = C\,\Delta T$

surroundings: q_{surr} = 1.0489 kJ/°C ×(41.405-22.145) °C +

121 g × 4.184 J g^{-1}/°C ×(41.405-22.145) °C × 10^{-3} kJ/J

= 20.202 kJ + 9.751 kJ =

29.953 kJ

Note that heat is discovered in the surroundings (the temperature of the calorimeter and the water increased!) This means that heat flowed from the system to the surroundings. The system LOST heat. Therefore $q_{system} < 0$.

(4) $\therefore\ q_{system} = q_V = -q_{surr} =$ $\underline{-29.953\text{ kJ}}$

molar heat of combustion of $C_2H_6(g)$ at constant volume, molar q_V = -29.953kJ/0.0194 mol =

$\underline{-1543.97}$ kJ mol^{-1}.

(5) First law of thermodynamics leads to $\Delta E = q_V$ (general, when only pressure volume work is possible)

$\therefore \Delta E$ of combustion of $C_2H_6(g)$ = $\underline{-1543.97}$ kJ mol^{-1}.

(6) Definition of enthalpy: $H = E + pV$

Ideal gas equation of state: $pV = nRT$

The balanced chemical equation:

$$C_2H_6(g) + (7/2)O_2 \rightarrow 2CO_2(g) + 3H_2O(liq)$$

(neglect the volume occupied by the liquid water resulting from the combustion.)

For 1 mole of C_2H_6 :

gases initial: $\underline{1\text{ mol } C_2H_6(g) + (7/2)\text{ mol } O_2(g)}$ gases final: $\underline{2CO_2(g)}$.

calculate ΔH here:

$$\begin{aligned}\Delta H = \Delta E + (pV)_{final} - (pV)_{initial} &= \Delta E + (nRT)_{final} - (nRT)_{initial}\\ &= -1543.97\text{ kJ mol}^{-1} + 2R\times314.56\text{K} - 4.5R\times295.30\text{K}\\ &= -1543.97\text{ kJ mol}^{-1}\\ &\quad + (2\times314.56-4.5\times295.30)\times8.31451\times10^{-3}\text{ kJ mol}^{-1}\\ &= -1543.97\text{ kJ mol}^{-1} - 699.73\times8.31451\times10^{-3}\text{ kJ mol}^{-1}\\ &= -1543.97\text{ kJ mol}^{-1} - 5.818\text{ kJ mol}^{-1}\end{aligned}$$

$\therefore \Delta H$ of combustion of $C_2H_6(g)$

= $\underline{-1549.79}$ kJ mol^{-1}