**Questions**

Explain the terms:-

1. System, surroundings, state of system, change of system.

2. What do you understand by the term internal energy and enthalpy of a system? How are they related?

3. Two litres of N2 at 0oC and 5 atm pressure are expanded isothermally against a constant pressure of 1 atm until the pressure of the gas is also 1 atm, assuming ideal behaviour.

What are the values of W, Δu, ΔH and q?

**4.**One mole of an ideal gas is allowed to expand reversibly against a confining pressure that is at all-time infinitesimally less than the gas pressure of 0.1 atm, the temperature being kept constant at 0oC. (a) How much work is done by the gas (b) What is the change in u and H. (c) How much heat is absorbed?

1)

**System;**A system is defined as collection of matter i.e part of the universe under study. The rest of the universe is being isolated.

(b) **Surrounding:**The rest of the universe not under study. A system could be isolated from its surrounding.

(c) **State of a system:**The state of a system is specified by determining all its properties such as pressure, volume etc.

d)change of state: a thermodynamics state of a system is defined by specifying values of a set of measurable properties sufficient to determine all other properties.

2) internal energy (U) is defined as the total [energy](https://www.thoughtco.com/energy-definition-and-examples-2698976) of a closed system.
Internal energy is the sum of potential energy of the system and the system's kinetic energy. The change in internal energy (ΔU) of a [reaction](https://www.thoughtco.com/chemical-reaction-definition-606755) is equal to the heat gained or lost enthalpy change in a reaction when the reaction is run at constant pressure.

Enthalpy is a thermodynamic property of a system. It is the sum of the internal energy added to the product of the pressure and volume of the system. It reflects the capacity to do non-mechanical work and the capacity to release heat.

The Relationship between ΔH and ΔE

If ΔH for a reaction is known, we can use the change in the enthalpy of the system to calculate its change in internal energy. When a reaction involves only solids, liquids, liquid solutions, or any combination of these, the volume does not change appreciably (ΔV = 0). Under these conditions, we can simplify Equation 18.11 to ΔH = ΔE. If gases are involved, however, ΔH and ΔE can differ significantly. We can calculate ΔE from the measured value of ΔH by using the right side of ΔH = ΔE + Δ(PV) = ΔE + PΔV together with the ideal gas law, PV = nRT. Recognizing that Δ(PV) = Δ(nRT), we can rewrite ΔH = ΔE + Δ(PV) = ΔE + PΔV as follows:

ΔH = ΔE + Δ(PV) = ΔE + Δ(nRT)

At constant temperature, Δ(nRT) = RTΔn, where Δn is the difference between the final and initial numbers of moles of gas. Thus

ΔE = ΔH − RTΔn

For reactions that result in a net production of gas, Δn > 0, so ΔE < ΔH. Conversely, endothermic reactions (ΔH > 0) that result in a net consumption of gas have Δn < 0 and ΔE > ΔH. The relationship between ΔH and ΔE for systems involving gases is illustrated in Example 4.

3)

Using ideal gas equation,

*PV*=*nRT*

For isothermal condition,

*PV*=Constant

∴*P*1​*V*1​=*P*2​*V*2​.....(1)

*P*1​=5*atm*

*P*2​=1*atm*

*V*1​=2*L*

*V*2​=

5×2=1×*V*⇒*V*=10*L*

∴*V*2​=10*L*

*W*=−*Pext*.​Δ*V*

Given that *Pext*.​=1*atm*

∴*W*=−1(10−2)=−8*L*−*atm*=−810.64*J*≈−810.10*J*

q = W = PV = nRT

as there is no change in heat

= 810.10 J

ΔU = q-w

810.10 - 810.10

=0

ΔH= qp

810.10 x 5

=4050.5 J

4) R= 0.08206

n=1 mole

T=00

P=0.1 atm

V=?

Using pv=nRT t is at o degrees

0.1 x ?=1 x 0.08206 x 0

?= 0.08206 / 0.1

v=0.8206

W=pv

= 0.1 x 0.8206

=0.08206

Δu = q-w

=0

ΔH = qp

q = w

=0.1 x 0.08206

=0.008206 J

H=u+pv

=0+0.1 x 0.08206

=0.008206