

## First Law Summary

- Statements of the first law:  
 $\Delta U = w + q$  (work plus heat)  
The internal energy is conserved.  
The internal energy is a state function.  
The internal energy change is not path-dependent.  
The internal energy change of a closed cycle is zero.  
The internal energy change of an isolated system is zero.
- Work is path dependent,  $w = -P\Delta V$   
Definitions: pressure always refers to external pressure.  
Sign convention: - work done by system, + work done on system.
  - Path 1: constant pressure path.  $w = -P_{\text{ext}}\Delta V$
  - Path 2: constant volume path.  $w = 0$ , because  $\Delta V = 0$ .
  - Path 3: isothermal path.  $\Delta U = 0$ .  $q = -w$ ,  $w = nRT \ln(V_f/V_i)$
  - Path 4: adiabatic path.  $q = 0$ ,  $\Delta U = w$
- The internal energy of a system increases as its temperature is raised.
  - In a constant volume process the constant of proportionality is called heat capacity at constant volume,  $C_v$ . (extensive, units  $\text{JK}^{-1}$ )
  - Molar heat capacity is intensive,  $C_v$ . (intensive, units  $\text{JK}^{-1}\text{mol}^{-1}$ )
  - Specific heat capacity is heat capacity per unit mass ( $\text{JK}^{-1}\text{g}^{-1}$ )
  - $C_v = (dU/dT)_v$ .  $dU = C_v dT$ .  $\Delta U = C_v \Delta T$ .
  - $C_v = 3/2nR$  for a monatomic ideal gas
  - $C_v = 5/2nR$  for a diatomic ideal gas
  - $C_v = 3nR$  for a polyatomic ideal gas
- The constant pressure energy is called the enthalpy
  - $\Delta H = \Delta U + P\Delta V$
  - $P\Delta V$  is the work term from moles of gas created or consumed
  - We can also write this as  $\Delta H = \Delta U + \Delta nRT$  where  $\Delta n$  is the change in moles of gas
  - For an ideal gas  $C_p = C_v + nR$