

# Chapter 6 : Thermochemistry

- **Internal Energy (E)** : Sum of potential and kinetic Energies

→ Change of energy in **System** = **opposite** Change Energy in **Surroundings**

$$\Delta E = E_{\text{final}} - E_{\text{initial}} = E_{\text{products}} - E_{\text{reactants}}$$

- Energy of System decreases → Energy released to surroundings
- ~ increases → ~ absorbed from ~

$$Q + W = \Delta E$$

heat → Q  
heat absorbed (+)  
heat released (-)

Work Done → W  
on system (+)  
by system (-)

(+) Energy absorbed  
(-) Energy released

$$\Delta E_{\text{universe}} = \Delta E_{\text{system}} + \Delta E_{\text{surroundings}} = 0$$

**Units:** Joule (SI system)

$$1 \text{ cal} = 4.184 \text{ J}$$

$$1 \text{ Btu} = 1055 \text{ J}$$

**Enthalpy:** Chemical change at Constant Pressure

## PV Work: Chemical Work

- Work done when the Volume of a system changes in the presence of an external Pressure

$$\Delta H = \Delta E + \Delta PV$$

- If P is Constant and the Volume Doesn't change a lot Then

$$\Delta PV = 0$$

And  $\Delta H = \Delta E$

Remember

$$-q_{\text{lost}} = q_{\text{gained}}$$

- Reactions That does not involve gases

- ~ ~ has a Constant moles of gas

## \* Exothermic and Endothermic Reactions

- heat is Given Out

$$\Delta H < 0$$

- heat is a product
- reactants has higher Energy

- heat is taken in

$$\Delta H > 0$$

- heat is a reactant
- products has higher Energy

## \* Calorimetry

$$q = C \times m \times \Delta T$$

specific heat capacity (C)

mass in g

$$T_{\text{final}} - T_{\text{initial}}$$

in Kelvin

(Quantity of heat required to change Temp of 1 gram of sub by 1 K)

# Stoichiometry of Thermochemical Equations

\* Equation that includes  $\Delta H_{rxn}$

- sign indicates Ex or Endo
- magnitude of  $\Delta H$  proportional to the amount of subst-

• How to solve Problems?

## \* Hess's Law

$$\Delta H_{\text{overall}} = \Delta H_1 + \Delta H_2 + \dots + \Delta H_n$$

• How to solve Problems?

→ find the amount of each reactant and product

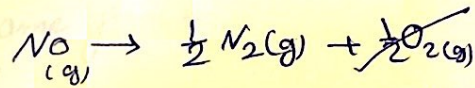
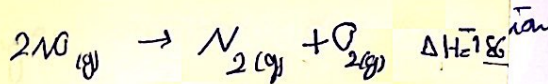
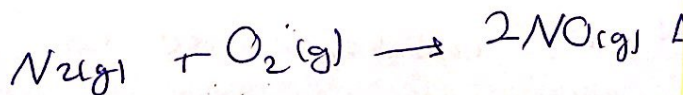
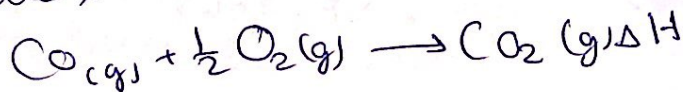
→ Change the sign of  $\Delta H$  when

→ Multiply moles and  $\Delta H$  by the

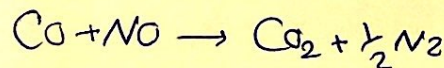
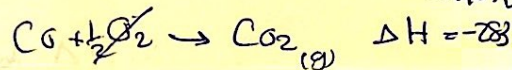
Ex



We have



$$\Delta H = -90.3$$



$$(-110.5 - 90.3) = -200.8$$

## Coffee Cup Calorimeter

→ This device measures the heat transferred at Constant Pressure ( $q_p$ )

→ If heat is transferred from a body to another then :-

$$\frac{q_{\text{body}(1)}}{q_{\text{body}(2)}} = 1 \Rightarrow \frac{C_1 \times m_1 \times \Delta T_1}{C_2 \times m_2 \times \Delta T_2} = 1$$

Law :-

$$-q_{\text{sub}_1} = q_{\text{sub}_2} + q_{\text{calorim}}$$

where  $q_{\text{sub}_1}$  is loss of heat and  $q_{\text{sub}_2}$  is Gaining heat

## A bomb calorimeter

→ This device measures the heat released at Constant Volume ( $q_v$ )

→ To determine  $\Delta H_{\text{rxn}}^\circ$  from  $\Delta H_f^\circ$  of reactants and products

$$\Delta H_{\text{rxn}}^\circ = \sum m \Delta H_f^\circ (\text{products}) - \sum n \Delta H_f^\circ (\text{reactants})$$

Standard Enthalpy of Reaction (Enthalpy for substance in the standard state)

Notes: \* Internal Energy is the same as  $\Delta E$

\* 28 has a Good idea

\* Standard state heat of formation:

$\Delta H_f^\circ$  :- formation of one compound in its standard state from elements in the standard state (1 atm / 1 M of solutes, stable form of liquids and solids)

$$\Delta E = \frac{3}{2} nRT \text{ for ideal gas}$$

\*  $\Delta H_f^\circ \text{O}_2 = 0$

