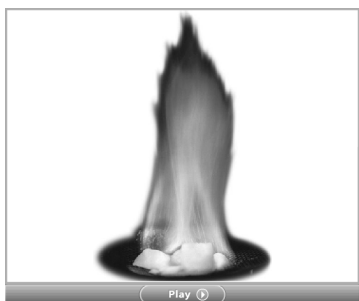


Chemical Thermodynamics

First Law of Thermodynamics

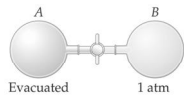
- The total energy of the universe is a constant.
- Energy can be
 - converted from one form to another
 - transferred from a system to the surroundings or vice versa.



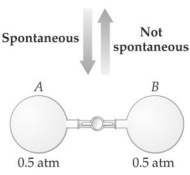
Spontaneous Processes

- Spontaneous processes proceed without any outside intervention

- The gas in vessel *B* will spontaneously diffuse into vessel *A*,



- but once the gas is in both vessels, it will *not* spontaneously

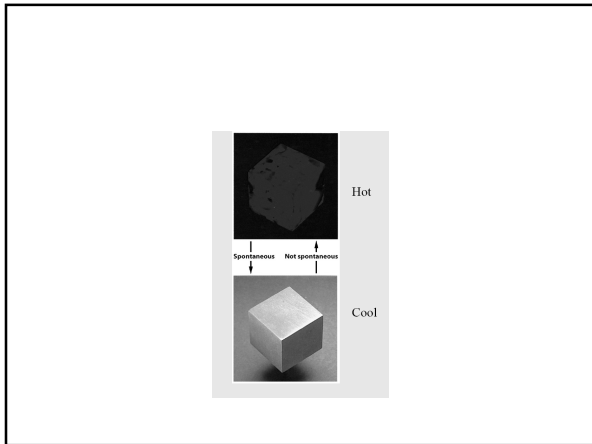


Spontaneous Processes



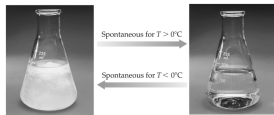
Spontaneous





Temperature can affect Spontaneous Processes

- Processes that are spontaneous at one temperature may be nonspontaneous at other temperatures.
- Above 0°C it is spontaneous for ice to melt.
- Below 0°C the reverse process is spontaneous.



What is a Joule (J)

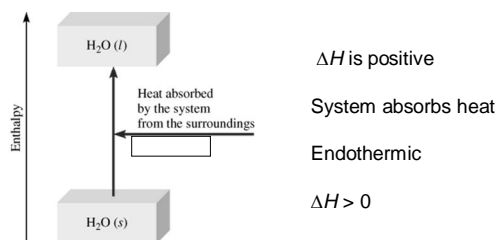
- The joule (symbol: J) is the SI unit of energy. It was named after James Prescott Joule for his work on the relationship between heat and mechanical work.
- One joule is the work done or energy required to exert a force of one Newton for a distance of one meter.

$$1\text{ J} = 1\text{ kg} \cdot \text{m}^2/\text{s}^2 \quad E=mc^2$$

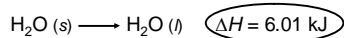
Calorie

- The calorie is the energy needed to increase the temperature of 1 gram of water by 1 °C. This is about 4.184 Joules.

Thermochemical Equations

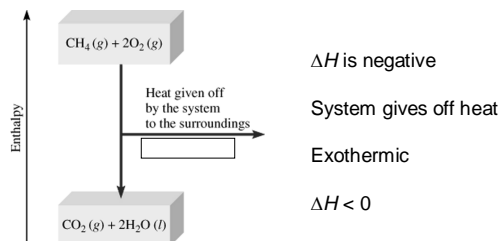


6.01 kJ are absorbed for every 1 mole of ice that melts at 0°C and 1 atm.

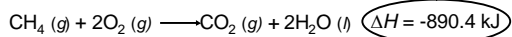


6.3

Thermochemical Equations



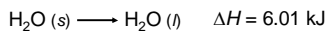
890.4 kJ are released for every 1 mole of methane that is combusted at 25°C and 1 atm.



6.3

Thermochemical Equations

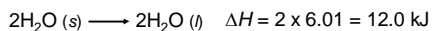
- The stoichiometric coefficients always refer to the number of moles of a substance



- If you reverse a reaction, the sign of ΔH changes



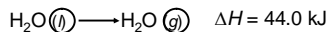
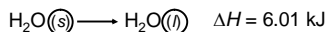
- If you multiply both sides of the equation by a factor n , then ΔH must change by the same factor n .



6.3

Thermochemical Equations

- The physical states of all reactants and products must be specified in thermochemical equations.



How much heat is evolved when 266 g of white phosphorus (P_4) burn in air?



$$266 \text{ g } \cancel{\text{P}_4} \times \frac{1 \text{ mol } \cancel{\text{P}_4}}{123.9 \text{ g } \cancel{\text{P}_4}} \times \frac{3013 \text{ kJ}}{1 \text{ mol } \cancel{\text{P}_4}} = 6470 \text{ kJ}$$

$$\Delta H = -6470 \text{ kJ}$$

6.3
