The Richard Stockton College of New Jersey
Chemistry Program, School of Natural Sciences and Mathematics
PO Box 195, Pomoma, NJ

CHEM 3410: Physical Chemistry I — Fall 2009

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Lecture 3: Heat capacity & working with the first law

References

1. Levine, *Physical Chemistry*, Sections 2.4–2.6, 2.8–2.9

Key Concepts

- We defined the *heat capacity*, C as the response of a system to heat flow:

$$C = \frac{\delta q}{dT}$$

  Since heat flow is path dependent, heat capacities must be defined for specific conditions, for example constant pressure (C_p) versus constant volume (C_v).

- For ideal gases:

$$C_p - C_v = R$$

  The constant pressure heat capacity is larger due to the fact that "extra" work is done in expanding the gas as it is heated under constant pressure. At constant volume, no PdV work is done, therefore all the heat goes into raising the temperature. For a monatomic ideal gas, $C_v = \frac{3}{2}R$ and for a diatomic ideal gas, $C_v = \frac{5}{2}R$.

- The amount of heat and work exchanged with the surroundings is dependent on the path from state 1 to state 2, but the change in internal energy must be the same since U is a state function.

Related Exercises in Levine

Exercises 2.9, 2.12,