

## 7.19: Isothermal Expansions of An Ideal Gas

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For an isothermal reversible expansion of an ideal gas, we have by definition that  $\Delta T = 0$ . Since the energy of an ideal gas depends only on the temperature, a constant temperature implies constant energy, so that  $\Delta E = 0 = q^{rev} + w^{rev}$ . Using the equation we find for  $w^{rev}$  in the previous section, we have

$$-q^{rev} = w^{rev} = -RT \ln \frac{V_2}{V_1}$$

(ideal gas, isothermal reversible expansion)

where  $V_1$  and  $V_2$  are the initial and final volumes, respectively. Since enthalpy is defined as  $H = E + PV$ , we have  $\Delta H = \Delta E + \Delta(PV) = \Delta E + \Delta(RT) = 0$ .

For the spontaneous isothermal expansion of an ideal gas from  $V_1$  to  $V_2$  against a constant applied pressure, we again have  $\Delta T = \Delta E = \Delta H = 0$ . These are state functions, and the amounts by which they change in this spontaneous process must be the same as those for the reversible process between the same two states. The heat and work exchanged in the spontaneous process are different, demonstrating that heat and work are not state functions. We have

$$-q = w = \int_{V_1}^{V_2} -P_{applied} dV = -P_{applied} (V_2 - V_1) = RT \left( \frac{P_{applied}}{P_1} - 1 \right)$$

(one mole ideal gas, isothermal free expansion,  $P_{applied} > 0$ )

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