

6. Let us consider the behavior of a 1.00-mol sample of hydrogen chloride gas.

- (a) (8 points) Using the ideal gas law, we would predict that 1.00 mol of HCl(g) at 8.22 atm and 506 K should occupy a volume of 5.05 L. However, careful experiments reveal that the gas actually occupies a volume of 5.00 L. Explain in detail why the actual volume is smaller than that predicted by the ideal gas law.

$PV = nRT$ neglects the attractive forces between molecules.

For many gases (like HCl), at moderate P 's > 1 atm, these forces make ~~the~~ the gases more compressible than $PV = nRT$ would predict.

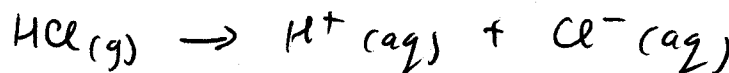
So HCl occupies slightly less volume than predicted. (The effect would be accounted for

by the "a" term of the more exact van der Waals equation.)

-2 or -1 incomplete logic -3 no discussion of a or b
-7 something -3 plausible but not true

- (b) (20 points) The 1.00-mol sample of HCl(g) is dissolved in a calorimeter containing 501 mL of water originally at 21.5°C. Assuming that there is no thermal energy transfer with the calorimeter or the surroundings, predict the temperature (in °C) of the water after the HCl has completely dissolved. (Note that $\Delta H_f^\circ(\text{Cl}^-, \text{aq}) = -167.16$ kJ/mol, and remember that $\Delta H_f^\circ(\text{H}^+, \text{aq})$ is defined to be zero.)

$$\Delta H_{\text{rxn}} + \Delta H_{\text{H}_2\text{O}} = 0$$



$$\Delta H_{\text{rxn}} = \Delta H_f^\circ(\text{H}^+, \text{aq}) + \Delta H_f^\circ(\text{Cl}^-, \text{aq}) - \Delta H_f^\circ(\text{HCl}, g)$$

dissolution rxn

$$= 0 - 167.16 \frac{\text{kJ}}{\text{mol}} - (-92.31 \frac{\text{kJ}}{\text{mol}})$$

$$\Delta H_{\text{rxn}} = \left(\frac{-74.85 \text{ kJ}}{\text{mol HCl}} \right) (1.00 \text{ mol HCl}) = -74.85 \text{ kJ} \left(\frac{10^3 \text{ J}}{\text{kJ}} \right)$$

$$= -7.485 \times 10^4 \text{ J}$$

$$\Delta H_{\text{H}_2\text{O}} = -\Delta H_{\text{rxn}} = 7.485 \times 10^4 \text{ J} = (501 \text{ mL}) \left(\frac{1.00 \text{ g}}{\text{mL}} \right) \left(4.184 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) \Delta T$$

$$\Delta T = 35.7^\circ\text{C}$$

$$\Rightarrow T_f = (21.5 + 35.7)^\circ\text{C} = \boxed{57.2^\circ\text{C}} \quad -1 \text{ math or s.f. error}$$