

CHEM 120
Fall 2009
Quiz 9

Name: _____

1a. (10 Points) The heat of solution (ΔH_{soln}) for potassium acetate dissolving in water is -15.3 kJ/mole. What will be the final temperature when 13.34 g of potassium acetate are dissolved in 525. g of water which is initially at 25.1 °C? Assume that the specific heat of the solution is 4.155 J·g⁻¹·K⁻¹ and that no heat is exchanged between the solution and the rest of the universe. You are given the molar mass of potassium acetate is 98.14 g/mole.

Start with $q_{reaction} = -q_{solution}$. Since the solution will only heat or cool, we will use $q = mC\Delta T$ for it, while for the reaction $q = n\Delta H$ is appropriate. Thus,

$$n\Delta H = -m_{solution} C_{solution} \Delta T_{solution}$$

$$\Delta T_{solution} = -\frac{n\Delta H}{m_{solution} C_{solution}}$$

Don't forget that $m_{solution}$ is the mass of everything in the solution (both the water and the reactants, which in this case is $KC_2H_3O_2$).

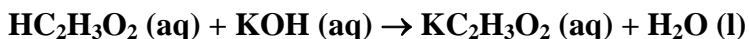
$$\Delta T_{solution} = -\frac{13.34 \text{ g} \left(\frac{1 \text{ mole}}{98.14 \text{ g}} \right) \left(\frac{-15.3 \text{ kJ}}{1 \text{ mole}} \right) \left(\frac{1000 \text{ J}}{1 \text{ kJ}} \right)}{(13.34 \text{ g} + 525. \text{ g}) (4.155 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1})} = +\frac{2.07_9 \times 10^3 \text{ J}}{(538.3_3 \text{ g}) (4.155 \text{ J} \cdot \text{g}^{-1} \cdot \text{K}^{-1})}$$

$$\Delta T_{solution} = +\frac{2.07_9 \times 10^3 \text{ J}}{2.23_6 \times 10^3 \text{ J} \cdot \text{K}^{-1}} = +0.929_7 \text{ K} = +0.930 \text{ }^\circ\text{C}$$

The final temperature is 25.1 °C + 0.930 °C = 26.0 °C.

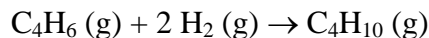
The final temperature of the solution is 26.0 °C.

b. (3 Points) Write the balanced chemical equation for the synthesis of potassium acetate from acetic acid.



Some people tried to use $\text{HC}_2\text{H}_3\text{O}_2 (\text{l}) + \text{K} (\text{s}) \rightarrow \text{KC}_2\text{H}_3\text{O}_2 (\text{s}) + \text{H}_2 (\text{g})$, or something similar. This would be an extremely violent and dangerous reaction and should not be attempted in real life.

2. (10 Points) The standard heats of combustion (ΔH_{comb}^0) for gaseous 1, 3-butadiene (C_4H_6), butane (C_4H_{10}), and hydrogen are -2540.2, -2877.6 and -285.8 kJ per mole of each substance, respectively. Use these data to calculate ΔH^0 for the hydrogenation of 1, 3-butadiene to butane (the reaction shown below).



You can't use $\Delta H_{rxn} = \sum \Delta H_f^0(\text{products}) - \sum \Delta H_f^0(\text{reactants})$, because the given ΔH are not ΔH_f^0 . This equation only works because all of the ΔH_f^0 are defined relative to the elemental state and thus have a common set of products/reactants. In general, this is not true and so we must write the reactions and use Hess's Law to calculate ΔH .

The reactions that we need and their ΔH_{comb}^0 are as follows.

Reaction	ΔH_{comb}^0 (kJ/mole)
$C_4H_6(g) + 11/2 O_2(g) \rightarrow 4 CO_2(g) + 3 H_2O(g)$	-2540.2
$C_4H_{10}(g) + 13/2 O_2(g) \rightarrow 4 CO_2(g) + 5 H_2O(g)$	-2877.6
$H_2(g) + 1/2 O_2(g) \rightarrow H_2O(g)$	-285.8

To obtain the reaction that we want, we need to reverse the second reaction and multiply the third by 2, as shown below.

Reaction	ΔH_{comb}^0
$C_4H_6(g) + 11/2 O_2(g) \rightarrow 4 CO_2(g) + 3 H_2O(g)$	(1 mole) (-2540.2 kJ/mole)
$4 CO_2(g) + 5 H_2O(g) \rightarrow C_4H_{10}(g) + 13/2 O_2(g)$	(1 mole) (+2877.6 kJ/mole)
$2 H_2(g) + 2/2 O_2(g) \rightarrow 2 H_2O(g)$	(2 mole) (-285.8 kJ/mole)
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$C_4H_6(g) + 2 H_2(g) \rightarrow C_4H_{10}(g)$	

And for the reaction as written

$$\Delta H_{rxn}^0 = -2540.2 \text{ kJ} + 2877.6 \text{ kJ} + -571.6 \text{ kJ} = -234.2 \text{ kJ}$$

The hydrogenation of 1, 3-butadiene to form butane has ΔH^0 of -234.2 kJ/mole.

3. (4 Points) We define ΔH_f^0 and ΔG_f^0 , but not ΔS_f^0 . Why?

The enthalpy and the Gibbs energy for a substance cannot be measured; only the change in these thermodynamic quantities can be. Therefore, we arbitrarily set H and G of the elements in their reference states to zero at any given temperature and pressure. We can then define ΔH_f^0 and ΔG_f^0 (formation of one mole of a substance from its elements in their reference states) and use them as if they were H^0 and G^0 for the substances. We don't need to do this for S because we can measure S directly (FYI from the temperature dependence of the constant pressure heat capacity) and the Third Law of Thermodynamics tells us where the zero of the entropy scale is. BTW, you can calculate ΔS^0 for the reactions that define ΔH_f^0 and ΔG_f^0 , just like one would for any other reaction, because you know S^0 for all of the reactants and products.