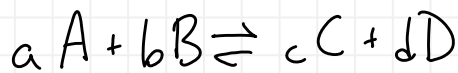


$$\begin{aligned}\Delta H^\circ &= \sum n \Delta H_f^\circ (\text{products}) - \sum n \Delta H_f^\circ (\text{reactants}) \\ &= -393.5 + 2(-241.8) - [-74.9 + 2(0)] \\ &= -802.2 \frac{\text{KJ}}{\text{units}}\end{aligned}$$

$$\begin{aligned}\Delta S^\circ &= \sum n S^\circ (\text{products}) - \sum n S^\circ (\text{reactants}) \\ &= 213.7 + 2(188.7) - [186.1 + 2(205.0)] \\ &= -5.0 \frac{\text{J/K}}{\text{units}} \Rightarrow \times \frac{1 \text{ KJ}}{1000 \text{ J}} = -0.0050 \frac{\text{KJ}}{\text{K}}\end{aligned}$$

$$\begin{aligned}\Delta G^\circ &= \Delta H^\circ - T \Delta S^\circ \\ &= -802.2 \text{ KJ} - (298 \text{ K})(-0.0050 \frac{\text{KJ}}{\text{K}}) \\ &= \underline{\underline{-800.7 \text{ KJ}}}\end{aligned}$$



$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

P = partial pressure
of gas, given
in atm

$$K_p = \frac{P_c^c P_d^d}{P_A^a P_B^b}$$

K = thermodynamic equilibrium constant

= K_c when all substances are present in
solution (measured in molarity)

= K_p when all substances are present in
gaseous phase (measured in atm)

$$\Delta G = W_{\max}$$

maximum amount of work that may be done by the system

Q = thermodynamic form of the reaction quotient

Finding ΔG under non standard conditions!

$$\Delta G = \Delta G^\circ + RT \ln Q$$

R = molar gas constant

$$8.315 \frac{\text{J}}{\text{mol}\cdot\text{K}}$$

absolute T

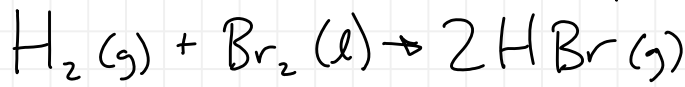
If at equilibrium, then $\Delta G = 0$

$$0 = \Delta G^\circ + RT \ln K$$

$$\Delta G^\circ = -RT \ln K$$

19.55

Find $\Delta G^\circ, K$ at $25^\circ\text{C} \Rightarrow$ tell whether spontaneous or not



$$0 \quad 0 \quad 2(-53.50)$$

$$\Delta G^\circ = 2(-53.50) - (0 + 0)$$

$$= -107.0 \text{ kJ} = \frac{107.0 \times 10^3 \text{ J}}$$

Spontaneous

$$R = 8.315 \frac{\text{J}}{\text{mol}\cdot\text{K}}$$

$$\frac{\Delta G^\circ}{-RT} = \frac{-RT \ln K}{-RT}$$

product-favored

large K

$$K = e^{-\frac{\Delta G^\circ}{-RT}} = e^{\frac{\Delta G^\circ}{-RT}} = K$$

$$K = e^{-\frac{-107.0 \times 10^3 \text{ J}}{8.315 \frac{\text{J}}{\text{mol}\cdot\text{K}} \cdot 298 \text{ K}}} = 5.7 \times 10^{18}$$

43.1822...

$$10^{-\text{pH}} = 10^{-11.52}$$

$$\Delta G = \Delta H - T\Delta S < 0 \quad (\text{spontaneous})$$

$$\Delta H - T\Delta S > 0 \quad (\text{non spontaneous})$$

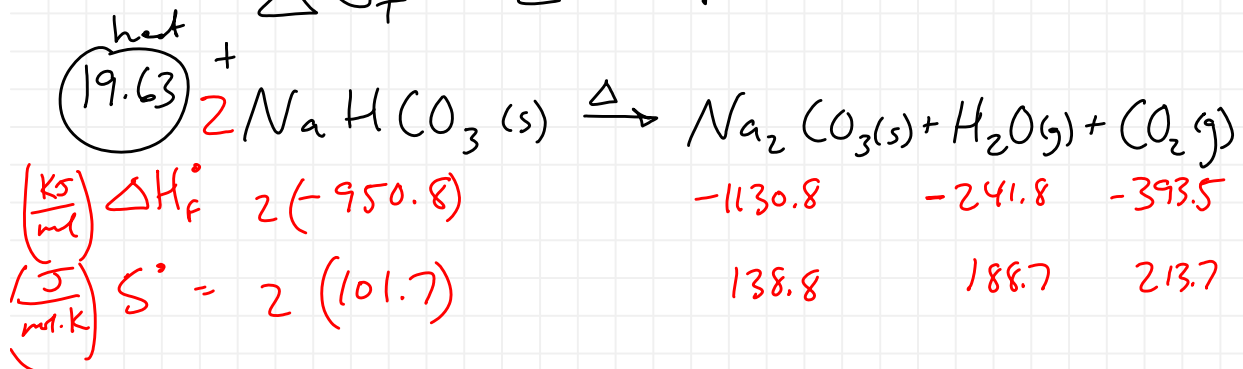
$$\Delta H - T\Delta S \approx 0 \quad (\text{equilibrium})$$

Sign of:

ΔH	ΔS	interpretation - Spontaneous or not
-	+	Spontaneous
+	-	non "
+	+	Spontaneous @ relatively large T
-	-	Spontaneous @ " low T

Approximation of "standard" free energy
(where every thing else is standard, except
the temperature)

$$\Delta G_T^\circ = \Delta H^\circ - T\Delta S^\circ$$



$$\Delta H^\circ = (-1130.8 + -241.8 + -393.5) - 2(-950.8)$$

$$= \boxed{135.5 \text{ KJ}}$$

$$\begin{aligned}\Delta S^\circ &= (138.8 + 188.7 + 213.7) - 2(101.7) \\ &= 337.8 \frac{\text{J}}{\text{K}} \times \frac{1 \text{ KJ}}{1000 \text{ J}} = 0.3378 \frac{\text{KJ}}{\text{K}}\end{aligned}$$

Estimate the T at which NaHCO_3 decomposes.

$$= 0 \rightarrow \Delta G_T^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$0 = \Delta H^\circ - T\Delta S^\circ$$

$$\frac{T\Delta S^\circ}{\Delta S^\circ} = \frac{\Delta H^\circ}{\Delta S^\circ} = \frac{135.5 \text{ KJ}}{0.3378 \text{ KJ/K}} = \boxed{401.1 \text{ K}}$$

$$- 273$$

$$\boxed{128^\circ\text{C}}$$

Ch. 20 - Electrochemistry

See redox handout:

