South Pasadena • AP Chemistry

Name KEY
Period ___ Date ___/___/___

20 • Entropy and Free Energy

STATION 1: ENTROPY CHANGE

For each of the following examples, decide whether the entropy is increasing (ΔS+) or decreasing (ΔS −). Is ΔS + or −?

ΔS is + The electrolytic decomposition of water. $2H_2O(l) \rightarrow 2\,H_2(g) + O_2(g)$

ΔS is − The freezing of water. $H_2O(l) \rightarrow H_2O(s)$

ΔS is + The reaction of sodium metal with water. $2Na(s) + 2H_2O(l) \rightarrow \underline{\underline{H_2(g)}} + 2Na^+(aq) + 2\,OH^-(aq)$

ΔS is + The boiling of water. $H_2O(l) \rightarrow H_2O\underline{(g)}$

ΔS is + The reaction of OF_2 and water. $\underbrace{OF_2(g) + H_2O(g)}_{\text{2 moles}} \rightarrow \underbrace{O_2(g) + 2HF(g)}_{\text{3 moles gas}}$

20 • Entropy and Free Energy

STATION 2: ΔH, ΔS, ΔG, GIBB'S FREE ENERGY

$C_2H_2(g) + 2\,H_2(g) \rightarrow C_2H_6(g)$

Substance	$S°$ (J/mol·K)	$\Delta H°_f$ (kJ/mol)
$C_2H_2(g)$	200.9	226.7
$H_2(g)$	130.7	0
$C_2H_6(g)$	229.6	-84.7

Calculate ΔS, ΔH, and ΔG for this reaction at 298 K.

$$\Delta S_{rxn} = \Delta S(C_2H_6) - (\Delta S(C_2H_2) + 2\,\Delta S(H_2))$$

$$= 229.6 - (200.9 + 2(130.7)) = \boxed{-232.7\ J/mol\cdot K}$$

$$\Delta H_{rxn} = \Delta H(C_2H_6) - \Delta H(C_2H_2) = -84.7 - 226.7 = \boxed{-311.4\ \frac{kJ}{mol}}$$

$$\Delta G = \Delta H - T\Delta S = -311.4\ \frac{kJ}{mol} - (298\ \cancel{K})\left(-232.7\ \frac{J}{mol\cdot \cancel{K}}\right)\left(\frac{1\ kJ}{1000\ J}\right)$$

$$= -311.4 - (-69.3) = \boxed{-242.1\ kJ/mol}$$

20 • Entropy and Free Energy

STATION 3: EQUILIBRIUM

Consider the boiling of liquid bromine: $Br_{2(l)} \leftrightharpoons Br_{2(g)}$

At ~~25°C~~ 298K, $\Delta H° = 30.84$ kJ/mol and $\Delta S° = 92.9$ J/mol·K for this reaction.
Calculate the value of ΔG°.

$$\Delta G = \Delta H - T\Delta S = 30.84\,\frac{kJ}{mol} - (298K)\left(92.9\,\frac{J}{mol\cdot K}\right)\left(\frac{1\,kJ}{1000\,J}\right)$$

$$= 30.84 - 27.68 = \boxed{3.16\ kJ/mol}$$

Assuming that ΔH and ΔS do not change at different temperatures, calculate the normal boiling point of liquid bromine.

EQUILIBRIUM; $\Delta G = 0$

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

$$\Delta H = T\Delta S$$

$$T = \frac{\Delta H}{\Delta S} = \frac{30.84\ kJ/mol}{92.9\ \cancel{J}/mol\cdot K \times \frac{1\,kJ}{1000\,\cancel{J}}} = 331.96986$$

$$= \boxed{332\ K}$$

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20 • Entropy and Free Energy

STATION 4: PREDICTING SPONTANEITY

Consider the reaction: $MgO(s) + SO_2(g) \rightarrow MgSO_3(s)$ + heat

What is the sign of ΔH for this reaction? __−__ Justify your answer.

Heat is needed to decompose $MgSO_{3(s)}$ into $MgO + SO_{2(g)}$
∴ Heat is released when they combine.

What is the sign of ΔS for this reaction? __−__ Justify your answer.

gas → solid less disorder.

This reaction will be:

ΔH $\ominus$ ΔS $\ominus$ spontaneous at low Temp

a) spontaneous at all temperatures
b) spontaneous at high temperatures
[✓ c)] spontaneous at low temperatures
d) non-spontaneous at all temperatures

STATION 5: K_{eq} & ΔG

Consider the reaction: $C_2H_5Cl(g) + Cl_2(g) \rightarrow C_2H_4Cl_2(g) + HCl(g)$

Standard Free Energies of Formation at 298 K

Substance	ΔG°_f $kJ \cdot mol^{-1}$
$C_2H_4Cl_2(g)$	-80.3
$C_2H_5Cl(g)$	-60.5
$HCl(g)$	-95.3
$Cl_2(g)$	0

Calculate the value of ΔG° for this reaction.

$$\Delta G_{Rxn} = \Delta G(C_2H_4Cl_2) + \Delta G(HCl) - \Delta G(C_2H_5Cl)$$

$$= -80.3\ \frac{kJ}{mol} + (-95.3) - (-60.5) = \boxed{-115.1\ kJ/mol}$$

Calculate the value of K_{eq} for the reaction at 298 K. [ΔG° = -RT *ln*K; R = 8.31 $J \cdot mol^{-1} \cdot K^{-1}$]

$$\Delta G^\circ = -R\ T\ \ln K$$

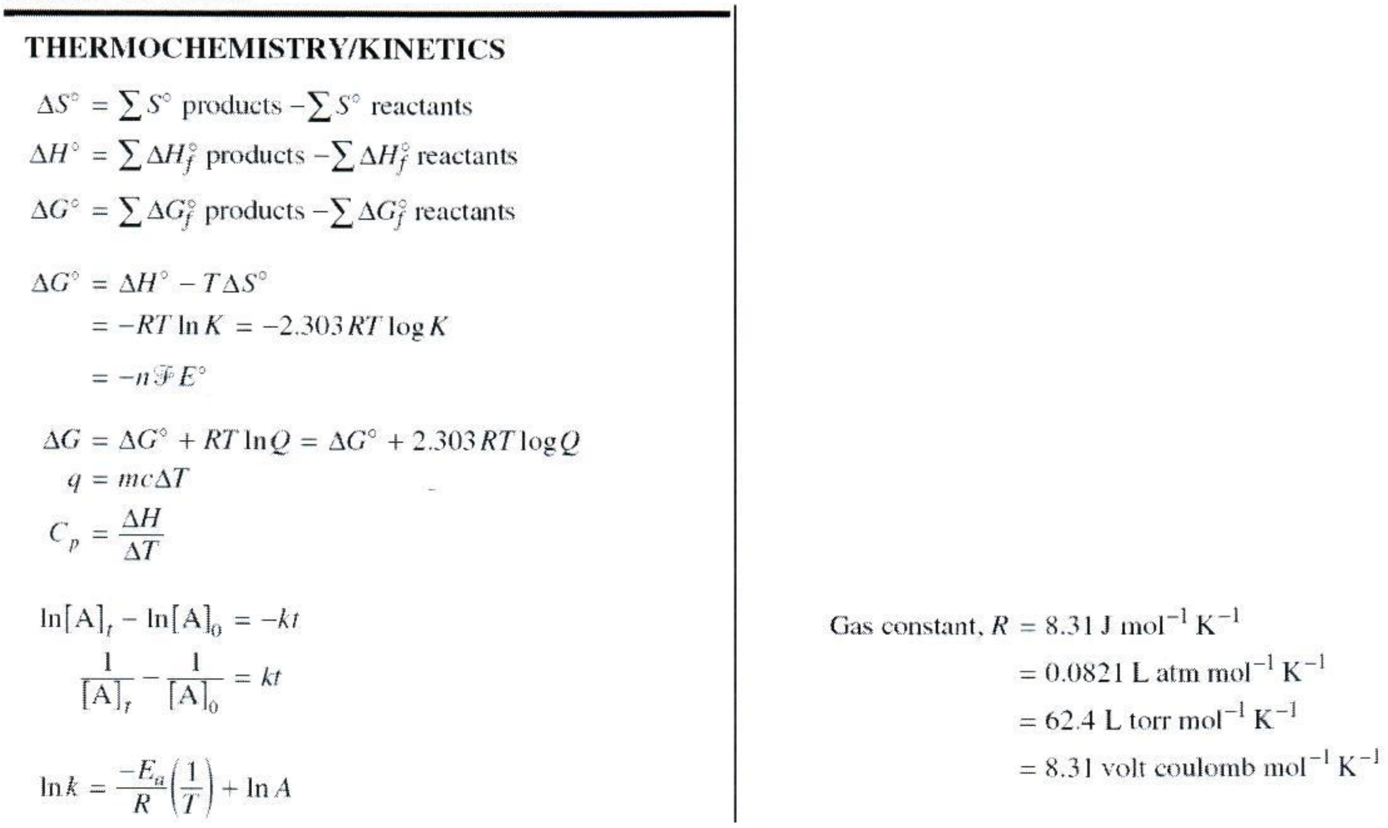
$$\left(\frac{1000\ J}{1\ kJ}\right)\left(-115.1\ \frac{kJ}{mol}\right) = -\left(8.31\ \frac{J}{mol \cdot K}\right)(298K) \ln K$$

$$\ln K = \frac{-115{,}100}{-2476.38} = 46.479$$

$$K = e^{46.479}$$

$$= \boxed{1.53 \times 10^{20}}$$

From the AP Exam:



THERMOCHEMISTRY/KINETICS

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

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

$$= -RT \ln K = -2.303\, RT \log K$$

$$= -n\mathcal{F}E^\circ$$

$$\Delta G = \Delta G^\circ + RT \ln Q = \Delta G^\circ + 2.303\, RT \log Q$$

$$q = mc\Delta T$$

$$C_p = \frac{\Delta H}{\Delta T}$$

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$\ln k = \frac{-E_a}{R}\left(\frac{1}{T}\right) + \ln A$$

Gas constant, $R = 8.31\ J\ mol^{-1}\ K^{-1}$

$= 0.0821\ L\ atm\ mol^{-1}\ K^{-1}$

$= 62.4\ L\ torr\ mol^{-1}\ K^{-1}$

$= 8.31\ volt\ coulomb\ mol^{-1}\ K^{-1}$