## Part A: Review: Conceptual Practice

1. All of the following processes lead to an increase in entropy EXCEPT?

a) decreasing the volume of a gas

b) melting a solid

c) chemicals reactions that increase the number of moles of gas

d) forming mixtures from pure substances

e) increasing the temperature of a gas.

2. If a chemical reaction is spontaneous yet endothermic, which of the following must be true?

a) $\Delta G > 0$; $\Delta S > 0$; $\Delta H > 0$
b) $\Delta G < 0$; $\Delta S > 0$; $\Delta H > 0$
c) $\Delta G > 0$; $\Delta S < 0$; $\Delta H > 0$
d) $\Delta G < 0$; $\Delta S < 0$; $\Delta H < 0$
e) $\Delta G > 0$; $\Delta S > 0$; $\Delta H < 0$

With your PAL team, discuss your answer and why the other options are incorrect.

3. If $\Delta G^o < 0$ for a reaction at all temperatures,
then $\Delta H^o$ is _____ and $\Delta S^o$ is_____.

a) negative, positive
b) positive, negative
c) negative, negative
d) positive, positive
e) positive, either positive or negative.

With your PAL team, discuss your answer and why the other options are incorrect.

4. For the following process, predict the algebraic sign of $\Delta G^o$, $\Delta H^o$ and $\Delta S^o$.

Dynamite is mostly nitroglycerin $C_3H_5N_3O_9$. Spontaneous explosive decomposition gives gaseous products $CO_2$ and $H_2O$ and a lot of heat is evolved.

$\Delta G^o$_____ $\Delta H^o$ _____ $\Delta S^o$ _____

5. A conceptual question. Is hydrogen peroxide stable at 298K? What about at higher temperatures or lower temperatures…is it stable?

$H_2O_2$ (g) → $H_2O$ (l) + ½ $O_2$ (g) $\Delta H° = -106$ kJ/mol and $\Delta S° = 58$ J/mol K

**Part B: Calculating the Change in Gibbs Free Energy for non-standard states**

6. Calculate $\Delta G^o$ for the following reaction at T = 298K ($K_b = 1.8 \times 10^{-5}$) .

$NH_3$ (aq) + $H_2O$ (l) → $NH_4^+$ (aq) + $OH^-$ (aq)

Use $\Delta G^o = -RT\ln K$

Now calculate $\Delta G$ at T = 298K when $[NH_3] = 0.10M$, $[OH^-] = 0.050M$, $[NH_4^+] = 0.10M$,
Use $\Delta G = \Delta G^o + RT\ln Q$

**Part C: Calculating the Equilibrium Constant**

7. Use the free energies of formation given below to calculate the equilibrium constant (K) for the following reaction at 298 K.

2 $HNO_3$ (aq) + NO(g) → 3 $NO_2$ (g) + $H_2O$(l) K = ?

| | | | | |
|---|---|---|---|---|
| $\Delta G°_f$ (kJ/mol) | -110.9 | 87.6 | 51.3 | -237.1 |

Use $\Delta Go = -RT\ln K$

8. Determine the equilibrium constant for the following reaction at 298 K.

$CH_2O(g)$ + 2 $H_2(g)$ → $CH_4(g)$ + $H_2O(g)$ $\Delta H° = -94.9$ kJ/mol; $\Delta S° = -224.2$ J/mol K

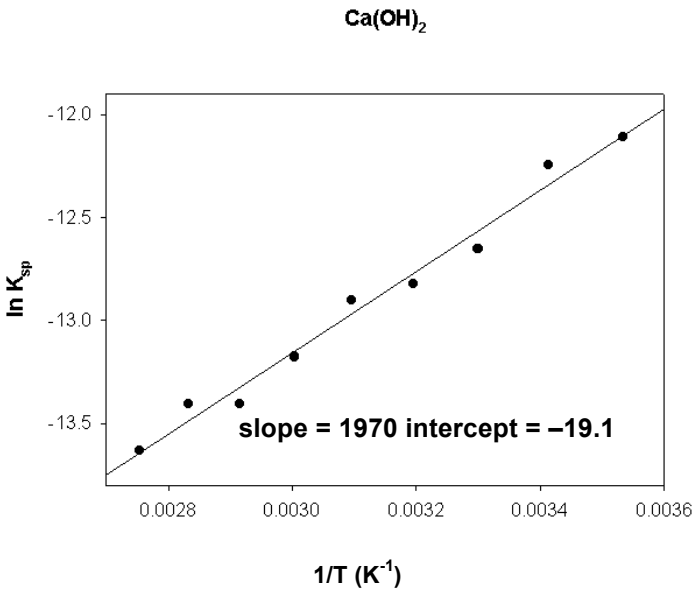
9. Equilibrium constants can be used to calculate $\Delta H^o$ and $\Delta S^o$.

Recall that $\Delta G^o = -RT\ln K_{eq}$. If we rearrange this to $\Delta H^o - T\Delta S^o = -RT\ln K_{eq}$ and after some algebra, we arrive at the following equation known as the van't Hoff equation.

$$\ln K_{eq} = -\frac{\Delta H^o}{RT} + \frac{\Delta S^o}{R}$$

Here is our example. The solubility of calcium hydroxide was measured at several temperatures, and the $K_{sp}$ was calculated. The data is below.

| T ($^o$C) | $K_{sp}$ | T (K) | 1/T ($K^{-1}$) | ln $K_{sp}$ |
|---|---|---|---|---|
| 10 | $5.5\times10^{-6}$ | 283 | 0.00353 | –12.11 |
| 20 | $4.8\times10^{-6}$ | 293 | 0.00341 | –12.25 |
| 30 | $3.2\times10^{-6}$ | 303 | 0.00330 | –12.65 |
| 40 | $2.7\times10^{-6}$ | 313 | 0.00319 | –12.82 |
| 50 | $2.5\times10^{-6}$ | 323 | 0.00310 | –12.90 |
| 60 | $1.9\times10^{-6}$ | 333 | 0.00300 | –13.17 |
| 70 | $1.5\times10^{-6}$ | 343 | 0.00292 | –13.41 |
| 80 | $1.5\times10^{-6}$ | 353 | 0.00283 | –13.41 |
| 90 | $1.2\times10^{-6}$ | 363 | 0.00275 | –13.63 |




1. On a separate piece of paper, determine the $\Delta H^o$ , $\Delta S^o$, and $\Delta G^o$ (this is an estimate since the temperatures above are not 298K).

2. Determine if it is spontaneous from $\Delta G^o$. Calculate the temperature at which the reaction will crossover from being spontaneous to nonspontaneous or vice versa.