SECTION 1

UNIT 11 – CHAPTER 6 STUDENT NOTES: THERMOCHEMISTRY

Calorimetry – Science of measuring heat

Terms

- **Heat Capacity = heat absorbed/change in temperature**

EX 1: H_2O | C_2H_5OH

Add 10 kJ of heat | Add 10 kJ of heat

10 – 15°C | 10 – 22°C

Heat Capacity = $\frac{10\text{ kJ}}{|10-15|} = \boxed{2\ \frac{\text{kJ}}{^\circ C}}$ | Heat Capacity = $\frac{10}{|10-22|} = \boxed{.83\ \frac{\text{kJ}}{^\circ C}}$

Specific Heat Capacity = heat capacity/gram $H_2O = 4.18$ J/g·°C

Molar Heat Capacity = heat capacity/mole $H_2O = 75.2$ J/mol·°C

ENTHALPY CHANGE: $\Delta H = \frac{q}{\text{MOLES CONSUMED}}$ $H_2O = 18.0$ cal/mol·°C

Heat capacities of common substances

H_2O = 4.18 J/g·°C	Aluminum = .897 J/g·°C
Iron = .412 J/g·°C	Carbon dioxide = .839 J/g·°C
Steel = .466 J/g·°C	Lead = .129 J/g·°C

Terms

- **Work** – Work = Force X Distance; Pressure X Δ Volume
- **Heat** – $q = m \cdot C_p \cdot \Delta t$ (Joules)
- **Internal Energy** – Sum of work and heat $E = q + w$ (-) work done by a gas – expansion
  (+) work done to a gas - compression

Coffee-cup calorimetry – uses a simple polystyrene cup to measure heat exchange

EX 2: When 1.00 grams of calcium chloride is added to 50.0 grams of water in a coffee-cup calorimeter, it dissolves and the temperature rises from 25.00°C to 28.51°C. Assuming that all the heat given off by the reaction is transferred to the water, calculate the heat for this reaction. (Heat of solution)

$q = m \cdot C \cdot \Delta t$

$q = (50\text{ g})(4.18)(3.51)$

$\boxed{q = 733.6\text{ J}}$

Example problems

1) Consider the following reaction: $Ba(NO_3)_{2(aq)} + Na_2SO_{4(aq)} \leftrightarrow 2\ NaNO_{3(aq)} + BaSO_{4(s)}$

Calculate the enthalpy change if 1.00 liter of each reactant at a concentration of 1.00 *M* is reacted to form the precipitate and the temperature of the solution goes from 25.0°C to 28.1°C an the density of the solution is 1.3 g/mL.

$C_p = 4.18$

$1L + 1L = 2L$ TOTAL $\frac{1000\,mL}{1L} \cdot \frac{1.3g}{1mL} = 2600g$

$q = mC_p\Delta t$

$q = (2600g)(4.18)(3.10°C) = 33{,}643.6J$ OR $\boxed{33.64KJ}$

1 MOL $Ba(NO_3)_2$

1 MOL Na_2SO_4

$\Delta H = \frac{q}{\text{MOLES USED}}$

EACH REACTANT USED ENTIRE 1.0 MOLES

$\Delta H = \frac{33.64KJ}{1\ \text{MOL USED}} = \boxed{-33.64KJ}$

2) The specific heat capacity of aluminum is 0.900 J/g ·°C.

a) Calculate the energy needed to raise the temperature of an 850-gram block of aluminum from 22.8°C to 94.6°C. $q = mC\Delta t \rightarrow (850)(.900)(71.8) = \boxed{54{,}927J \text{ or } 54.93KJ}$

b) Calculate the molar heat capacity of aluminum.

$M_{HC} = \frac{.900}{1/27} = \boxed{\frac{24.3J}{MOL \cdot °C}}$

3) In a coffee-cup calorimeter, 50.0 mL of 0.100 *M* $AgNO_3$ and 50.0 mL of 0.100 *M* HCl are mixed to yield the following reaction: $Ag^+_{(aq)} + Cl^-_{(aq)} \leftrightarrow AgCl_{(s)}$

The two solutions were initially at 22.6°C and the final temperature is 23.4°C. Calculate the heat that accompanies this reaction. Assume that the combined solution has a mass of 100.0 grams and has a specific heat capacity of 4.17 J/g ·°C. Calculate the enthalpy change for the reaction in kJ/mol.

$q = mC\Delta t$

$(100g)(4.17)(.80°C) = 333.6J$ or $\boxed{.3336KJ}$

$Ag^+ + Cl^- \rightarrow AgCl$

$\frac{.1\,MOL}{1L} \cdot .05L = .005$ MOL of $Ag^+ + Cl^-$

BOTH USE ALL MOLES

$\Delta H = \frac{.3336\,KJ}{.005\,MOL}$

$\boxed{\Delta H = -66.72 KJ/MOL}$

Thermochemical equation – a chemical equation that shows the enthalpy relation between products and reactants

EX 3: Consider the following thermochemical equation:

4 MOL

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)} \qquad \Delta H = -508\ kJ$$

a) How much heat is evolved when 4 moles of H_2 are ignited?

$\frac{4\,MOL\,H_2}{} \cdot \frac{508KJ}{2\,MOL\,H_2} = \boxed{1{,}016KJ \text{ EMITTED}}$

b) How much heat is evolved when 20.0 grams of H_20 is produced?

$\frac{20.0g\,H_2O}{} \cdot \frac{1\,MOL\,H_2O}{18g\,H_2O} \cdot \frac{508KJ}{2\,MOL\,H_2O} = \boxed{282.2KJ}$

c) How much heat is evolved when 10.0 grams of H_2 and 20.0 grams of O_2 react?

$\frac{10.0g\,H_2}{} \cdot \frac{1\,MOL\,H_2}{2g\,H_2} \cdot \frac{508KJ}{2\,MOL\,H_2} = 1{,}270\ KJ$

$\frac{20.0g\,O_2}{} \cdot \frac{1\,MOL\,O_2}{32g\,O_2} \cdot \frac{508KJ}{1\,MOL\,O_2} = \boxed{317.5KJ}$

$\boxed{\text{END OF SECTION 1}}$

BEGIN SECTION 2

Hess's Law – The change in enthalpy is the same whether the reaction takes place in one step or in a series of steps.

EX 4: Calculate the heat change for the conversion of graphite to diamond using the following information:

$H_{combustion\ (graphite)}$ = -394 kJ/mole

$C_{GRAPHITE} + O_2 \rightarrow CO_2 + H_2O \quad \Delta H = -394KJ$

$H_{combustion\ (diamond)}$ = -396 kJ/mole

$C_{DIAMOND} + O_2 \rightarrow CO_2 + H_2O \quad \Delta H = -396KJ$

$C_{graphite} \longleftrightarrow C_{diamond}$

THEN REVERSE TO CREATE DIAMOND PRODUCT

$H_2O + CO_2 \rightarrow O_2 + C_{DIAMOND} \quad \Delta H = +396KJ$

NET RXN: GRAPHITE $\longleftrightarrow$ DIAMOND

$-394 + 396 = \boxed{+2\frac{KJ}{MOL}}$ (ENDOTHERMIC)

EX 5: Diborane (B_2H_6) is a highly reactive boron hydride that was once considered as a possible rocket fuel for the U.S. space program. Calculate the $H_{reaction}$ for the synthesis of diborane.

$$2\,B_{(s)} + 3\,H_{2(g)} \longleftrightarrow B_2H_{6(s)}$$

a) $2\,B_{(s)} + 3/2\,O_{2(g)} \longleftrightarrow B_2O_{3(s)}$ $\quad \Delta H = -1273$ kJ

b) $B_2O_{6(s)} + 3\,O_{2(g)} \longleftrightarrow B_2O_{3(s)} + 3\,H_2O_{(g)}$ $\quad \Delta H = -2035$ kJ

c) $H_{2(g)} + ½\,O_{2(g)} \longleftrightarrow H_2O_{(l)}$ $\quad \Delta H = -286$ kJ

d) $H_2O_{(l)} \longleftrightarrow H_2O_{(g)}$ $\quad \Delta H = +44$ kJ

a) $2B + 3/2\,O_2 \longleftrightarrow B_2O_3 \quad \Delta H = -1273KJ$

REVERSED FROM ABOVE TO MAKE B_2O_6 PRODUCT! → b) $B_2O_3 + 3H_2O \longleftrightarrow B_2H_6 + 3O_2 \quad \Delta H = +2035KJ$

X3 TO MATCH H_2O IN b) → c) $3H_2 + 3/2\,O_2 \longleftrightarrow 3H_2O \quad \Delta H = 3(-286KJ) = -858KJ$

→ d) $3\,H_2O \longleftrightarrow 3\,H_2O \quad \Delta H = 3(44KJ) = +132KJ$

$2B + 3H_2 \longleftrightarrow B_2H_6 \quad \boxed{\Delta H = +36KJ \text{ (ENDOTHERMIC)}}$

EX 6: Given the following data:

FLIP + X2 TO MATCH H_2O BELOW → $H_{2(g)} + ½\,O_{2(g)} \rightarrow H_2O_{(l)}$ $\quad$ H = -258.8 kJ → $2(+258.8KJ) = +517.6KJ$

FLIP + X2 TO MATCH N_2O_5 PRODUCT → $N_2O_{5(g)} + H_2O_{(l)} \rightarrow 2\,HNO_{3(l)}$ $\quad$ H = -76.6 kJ → $2(+76.6KJ) = +153.2KJ$

X4 TO MATCH ENDING N_2 → $½\,N_{2(g)} + 3/2\,O_{2(g)} + ½\,H_{2(g)} \rightarrow HNO_{3(l)}$ $\quad$ H = -174.1 kJ → $4(-174.1KJ) = -696.4KJ$

$\boxed{\Delta H = -25.6KJ}$

Calculate the enthalpy change for the reaction

$$2\,N_{2(g)} + 5\,O_{2(g)} \rightarrow 2\,N_2O_{5(g)}$$

Standard Enthalpies of Formation – of a compound is equal to the enthalpy change when one mole of the compound is formed at a constant pressure of 1.0 atm and a fixed temperature, 25°C, from the elements in their states at that pressure and temperature.

- Most enthalpies of formation are negative numbers, meaning that the compound forms from its elements & is exothermic
- Elements in their standard states have a standard enthalpy of formation equal to zero.

In standard enthalpies of formation calculations, keep in mind:

1) When a reaction is reversed, the magnitude of all Heats remains the same, but its sign changes.

2) When the balanced equation for a reaction is multiplied by an integer, the value of H for that reaction must be multiplied by the same integer.

3) Elements in their standard states are not included in the Heat calculated, that is standard heat for an element in its standard state is **zero**.

$\Delta H_f^\circ\ NH_3 = -46.11\,kJ$
$\Delta H_f^\circ\ O_2 = \emptyset$ (ELEMENTAL STATE)
$\Delta H_f^\circ\ NO_2 = 33.18\,kJ$
$\Delta H_f^\circ\ H_2O = -285.83\,kJ$

EX 7: Calculate the enthalpy change for the following reaction:

① COFFEE CUP CALORIMETRY
② HESS'S LAW
③ ΔH_f IN FORMATION

$4\ NH_{3(g)} + 7\ O_{2(g)} \rightarrow 4\ NO_{2(g)} + 6\ H_2O_{(l)}$ $\Delta H = ?$

$\Delta H_f^0 = \sum H_{fp} - \sum H_{fr}$

$[4(34\,kJ)+6(-286\,kJ)] - [4(-46\,kJ)+7(\emptyset)]$

$\boxed{\Delta H_f = -1397.82\,kJ}$

EX 8: How much heat is released by 13.3 grams of Al in the equation:

$2\ Al_{(s)} + Fe_2O_{3(s)} \rightarrow Al_2O_{3(s)} + 2\ Fe_{(s)}$ $\Delta H = -850\ kJ$

$\Delta H_f^\circ = [1(-1676\,kJ) + 2(\emptyset kJ)] - [2(\emptyset kJ) + (-826\,kJ)]$

$\Delta H_f^\circ = -850\,kJ$ (IF NOT GIVEN!)

$\frac{13.3\,g\ Al}{} \times \frac{1\,mol\ Al}{27.0\,g\ Al} \times \frac{-850\,kJ}{2\,mol\ Al} = \boxed{-209.4\,kJ}$

END SECTION 2

THERMOCHEMISTRY – Chapter 17

UNIT 12 SECTION 1

Chemical reactions are driven by two factors:

Energy Factor – Many spontaneous reactions proceeded with a decrease of energy

- Almost all exothermic reactions are spontaneous
- Most phase changes are endothermic, yet spontaneous

  $H_2O_{(s)} \rightarrow H_2O_{(l)}$ $\Delta H = +6.0\ kJ$
- Some reactions, though not spontaneous become spontaneous at higher temperatures.

  $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$ $\Delta H = +178.3\ kJ$

  (at 1100 K, this reaction becomes spontaneous)

Randomness Factor – Nature tends to move spontaneously from a state of lower probability to higher probability.

- Nature tends to move spontaneously from a more ordered to more random state.
- Entropy (*S*) is the measure of the randomness factor.

$S = S^0_{final} - S^0_{initial}$ **-S = nonspontaneous; +S = spontaneous**

Factors that influence entropy

- A liquid has a higher entropy than a solid from which it is formed.
- A gas has a higher entropy than the liquid from which it is formed.
- Increasing the temperature of a substance increases its entropy (especially in phase changes).

Predict whether the *S* is positive or negative for each of the following processes.

1) taking dry ice from the freezer where the temperature is -80°C and allowing it to warm to room temperature + (SOLID → GAS = MOST ORGANIZED TO MORE RANDOM)

2) dissolving bromine in hexane + (MOST ORGANIZED TO MORE RANDOM)

3) condensing gaseous bromine to liquid bromine — (GOING FROM MORE RANDOM TO LESS RANDOM)

Calculating ΔS^0 for a reaction

$$\boxed{\Delta S^0 = \sum S^0_{products} - \sum S^0_{reactants}}$$

Calculate the ΔS for each of the following reactions:

$CaCO_3 = 92.9$ J/mol·K
$CaO = 39.8$ J/mol·K
$CO_2 = 213.7$ J/mol·K

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$ *↑ GAS MOLES (+) ENTROPY

$\Delta S^0 = (39.8 + 213.6)(-92.9) = +160.5$ SPONTANEOUS

A reaction that results in an increase in the number of moles of a gas is accompanied by an increase in entropy; if the gas molecules decrease, entropy is a negative number

Calculate ΔS^0 for $Ca(OH)_2 \rightarrow Ca^{2+}{}_{(aq)} + 2\,OH^-{}_{(aq)}$

$Ca(OH)_2 = 83.4$
$Ca^{+2} = 74.8$
$OH^- = 34.8$

a) dissolving one mole of calcium hydroxide in water

$\Delta S^0 = [74.8 + (2 \times 34.8)] - 83.4$

$\Delta S^0 = 61$

b) the combustion of one mole of methane to form carbon dioxide and liquid water

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

$\Delta S^0 = [213.6 + 139.8] - [186.2 + 2(205.0)] = \boxed{-242.8 \text{ *LESS GAS FORMED}}$

NONSPONTANEOUS

Free Energy (ΔG)

Two quantities affect reaction spontaneity: enthalpy (ΔH) and entropy (ΔS).

Gibbs Free Energy (ΔG) – represents that portion of the total energy change that is available (i.e., free) to do useful work.

(+) energy must be supplied

(-) energy is released for work

Gibbs-Helmholtz Equation: **$\Delta G = \Delta H - T \cdot \Delta S$**

TEMP MUST BE KELVIN!
WHEN FINDING ΔG, ΔS MUST BE CONVERTED TO KJ

Sign determines spontaneity: ΔG (-) reaction is spontaneous

ΔG (+) reaction will not take place

$\Delta G = 0$ system is at equilibrium

Two factors tend to make ΔG negative:

1) A negative value of ΔH

2) A positive value of ΔS (many physical changes, the entropy increase is the major or driving force i.e., melting ice)

Standard Free Energy Change

Standard conditions – gases at 1 atmosphere and solutions 1.0 *M* and 25°C

For the reaction: $CaSO_{4(s)} \rightarrow Ca^{2+}{}_{(aq)} + SO_4^{2-}{}_{(aq)}$

Calculate a) ΔH^0 b) ΔS^0 c) ΔG^0 at 25°C

$CaSO_4$ ΔH = -1433, ΔS = 107, ΔG = -1320
Ca^{2+} ΔH = -543, ΔS = -55, ΔG = -553
SO_4^{-2} ΔH = -909.3, ΔS = 20, ΔG = -745

ΔH: [-543+(-909.3)] - [-1433] = -19.3 KJ
ΔS: [-55 + 20] - [107] = -142 J
ΔG = ΔH - TΔS in KJ → (-19.3) - (298 · .142 KJ) = $\boxed{-61.6}$

Calculation of ΔG^0 at other temperatures than standard

To a good degree of approximation, the temperature variation for ΔH^0 and ΔS^0 can be neglected. This means that to apply the Gibbs-Helmholtz equation to temperatures other than 25°C, you need only change the value of the temperature.

Calculate the ΔG^0 at 230°C for the reaction of one mole of Fe_2O_3 with hydrogen. The products are iron metal and water vapor. $Fe_2O_3 + 3H_2 \rightarrow 2Fe + 3H_2O_{(g)}$

Fe_2O_3 ΔH = -826, ΔS = 90, ΔG = -740
H_2 ΔH = ∅ (ELEMENTAL FORM), ΔS = 131, ΔG = ∅
Fe ΔH = ∅, ΔS = 27, ΔG = ∅
$H_2O_{(g)}$ ΔH = -242, ΔS = 189, ΔG = -229

ΔH = [(2×Fe)+(3×H₂O)] - [(1×Fe₂O₃)+3(H₂)] → [(2×∅)+(3×-242)] - [(1×-826)+(3×∅)]
ΔH = 100 KJ
ΔS = [(2×27)+(3×189)] - [(1×90)+(3×131)]
ΔS = 138 J
ΔG = 100 - (503K · .138KJ) = $\boxed{30.6 \therefore \text{NOT SPONTANEOUS}}$

Calculate the ΔG^0 at 335 K for each reaction below and tell if it is spontaneous or nonspontaneous.

a) $CO_{(g)} + 2\ H_{2(g)} \rightarrow CH_3OH_{(l)}$

$CH_3OH_{(l)}$ ΔH = -239, ΔG = -166, ΔS = 127
$CO_{(g)}$ ΔH = -110.5, ΔG = -137, ΔS = 198
$H_{2(g)}$ ΔH = ∅, ΔS = 131, ΔG = ∅

b) $N_{2(g)} + O_{2(g)} \rightarrow 2\ NO_{(g)}$

$NO_{(g)}$ ΔH = 90, ΔG = 87, ΔS = 211
N_2 ΔH = ∅, ΔG = ∅, ΔS = 192
O_2 ΔH = ∅, ΔG = ∅, ΔS = 205

a) ΔH_f^0 = [-239] - [(1× -110.5)+(2×∅)] = -128.5 KJ
ΔS_f^0 = [127] - [(1×198)+(2×131)] = -333 J
ΔG_{335K} = -128 - (335K · -.333 KJ) = $\boxed{-16.4 \therefore \text{SPONTANEOUS}}$

b) ΔH_f^0 = [2(90)] - [(1×∅)+(1×∅)] = 180 KJ
ΔS_f^0 = [(2×211)] - [(1×192)+(1×205)] = 25 J
ΔG_{335K} = 180 - (335 × .025 KJ) = $\boxed{171.6 \therefore \text{NONSPONTANEOUS}}$

BEGIN SECTION 2

Effect of temperature on spontaneity

	ΔH^0	ΔS^0	$\Delta G^0 = \Delta H^0 - T \cdot \Delta S^0$	
I	(-)	(+)	always (-)	Spontaneous at all temps reverse rxn always nonspontaneous
II	(+)	(-)	always (+)	Nonspontaneous at all temperatures
III	(+)	(+)	(+) at low T, (-) at high T	***If ΔH^0 and ΔS^0 have opposite signs, it is impossible to reverse the direction of spontaneity by a change in temperature. Both terms ΔH^0 and $T \cdot \Delta S^0$ reinforce each other**
IV	(-)	(-)	(-) at low T, (+) at high T	

At what temperature does ΔG^0 become zero for the reaction:

$\Delta H = -826$, $\Delta S = 90$; $\Delta H = \emptyset$, $\Delta S = 131$; $\Delta H = \emptyset$, $\Delta S = 27$; $\Delta H = -242$, $\Delta S = 189$

$$Fe_2O_{3(s)} + 3\,H_{2(g)} \rightarrow 2\,Fe_{(s)} + 3\,H_2O_{(g)}$$

$\Delta H = [(2\times0)+(3\times-242)] - [(1\times-826)+(3\times\emptyset)]$

$\Delta H = 100\,kJ$

$\Delta S = [(2\times27)+(3\times189)] - [(1\times90)+(3\times131)]$

621 − 483

$\Delta S = 138\,J$

$100\,kJ = T \cdot (.138\,kJ)$

$\boxed{T = 724.6\,K}$

SETTING $\Delta G^0 = \emptyset$ GIVES US: $\emptyset = \Delta H - (T \cdot \Delta S)$ or $\Delta H^0 = T \cdot \Delta S^0$

* $\Delta H^0 = 98.8\,kJ$

* $\Delta S^0 = .1415\,kJ$

$98.8\,kJ = T \cdot (.1415) \rightarrow \boxed{T = 698\,K}$

Temperature for spontaneity: **$T = \Delta H^0 / \Delta S^0$**

At what temperature does the following reaction occur: $H_2O_{(l)} \rightarrow H_2O_{(g)}$

Effect of pressure and temperature

$G = \Delta G^0 + RT\ln Q$

T = degrees Kelvin

$R = 8.31 \times 10^{-3}$ kJ/K

Q rules

1) gases enter as their partial pressures in atmospheres
2) aqueous solutions enter as their molar concentrations
3) pure liquids and solids do not appear

$\Delta H\ Zn^{2+} = -153.39$

$\Delta S\ Zn^{2+} = -112.1$

$\Delta G = \emptyset$, $\Delta G = \emptyset$, $\Delta G = -147.1$, $\Delta G = \emptyset$

Consider the reaction: $Zn_{(s)} + 2\,H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$ @ 25°C

Calculate

a) G^0 $\Delta G^0 = [(1\times-147.1)+(1\times0.0)] - [(1\times0.0)+(1\times0.0)] = \boxed{-147.1\,kJ/mol}$

b) G when P_{H2} = 750 mm Hg, $[Zn^{2+}] = 0.10\,M$, $[H^+] = 1.0 \times 10^{-4}\,M$

$G = \Delta G^0 + RT\ln Q$

$= -147.1 + \left[(.00831)(298K)\left(\ln\frac{[Zn^{2+}][P_{H_2}]}{[H^+]^2}\right)\right]$

$= -147.1 + \left[(.00831)(298K)\left(\ln\frac{(.10M)\left(750mmHg\,\frac{1\,Atm}{760mmHg}\right)}{(1.0\times10^{-4})^2}\right)\right]$

$= -147 + 39.9\,kJ$

$\boxed{= -107.2\,kJ}$

$PbCl_2 \rightarrow Pb^{2+} + 2Cl^-$

$\Delta G = -314.1$ $\Delta G = -24.4$ $\Delta G = -131.2$

Show by calculation whether dissolving lead (II) chloride is spontaneous when:

a) $[Pb^{2+}] = 1.0\ M$, $[Cl^-] = 2.0$ M

b) $[Pb^{2+}] = 1.0 \times 10^{-5}\ M$, $[Cl^-] = 2.0 \times 10^{-5} M$

a) $\Delta G^0 = [-24.4 + (2 \times -131.2)] - [-314.1] = +27.3\text{KJ}$

$\Delta G = +27.3\text{KJ} - [(.00831)(298)(\ln(1.0)^1 \cdot (2.0)^2)$ $\ln 4$

$\boxed{\Delta G = +30.71\text{KJ}}$ *NONSPONTANEOUS *NO SOLUBILITY

b) $\Delta G^0 = +27.3\text{KJ} - [(.00831)(298)(\ln(1.0\times10^{-5})^1 \cdot (2.0\times10^{-5})^2)$ $\ln(4\times10^{-15})$

$\boxed{\Delta G = -54.8\text{ KJ}}$ *SPONTANEOUS

Free energy change and the equilibrium constant, *K*

For spontaneity: ΔG^0 must be negative (-)

K must be greater than 1

ΔE^0 must be (+)

<u>ΔG^0 and *K* are related</u>: $\Delta G^0 = -RT \ln K$ (must be at standard conditions)

8.314 298

Using ΔG^0_f tables in appendix 1, calculate the solubility product constant, K_{sp}, for $PbCl_2$ at 25^0C.

$PbCl_{2(s)} \leftrightarrow Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$

$\Delta G^0_{RXN} = \sum \Delta G^0_{PRODUCTS} - \sum \Delta G^0_{REACTANTS}$

$\Delta G^0_{RXN} = [\Delta G^0_f(Pb^{2+}) + 2\Delta G^0_f(Cl^-)] - [\Delta G^0_f(PbCl_2)]$

$\Delta G^0_{RXN} = [(-24.4) + (2 \times -131.2)] - [-314.2]$

$\Delta G^0_{RXN} = 27.4\text{ KJ} = 27{,}400\text{ J/MOL}$

$27{,}400 = (-8.314)(298) \ln K$

$-11.06 = \ln K \rightarrow \boxed{e^x}^{-11.06}$

$\boxed{K = 1.6\times10^{-5}}$

END SECTION 2