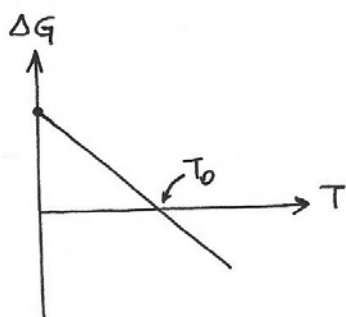


LIZARDI BHI	2008-09	Topics: Thermochemistry Chemical kinetics Chemical equilibrium	Marks:
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1st exam			
2008 - XI - 26			
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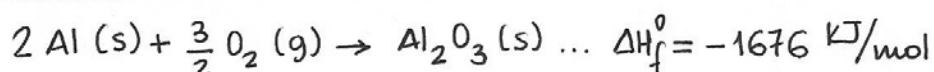
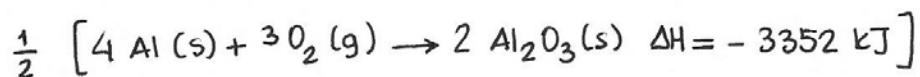
- ①
- a) The reaction is endothermic
- b) The entropy-change is positive; one of the products is a gas
- c) If  $\Delta H > 0$  and  $\Delta S > 0$ , the graphic  $\Delta G$ -T for the function  $\Delta G = \Delta H - T\Delta S$  has this shape:



the system is spontaneous between  $T_0$  and  $\infty$ , where

$$T_0 = \frac{\Delta H}{\Delta S}$$

- d) The formation enthalpy-change of aluminum oxide is



- e) • The volume of  $\text{O}_2$  formed at STP

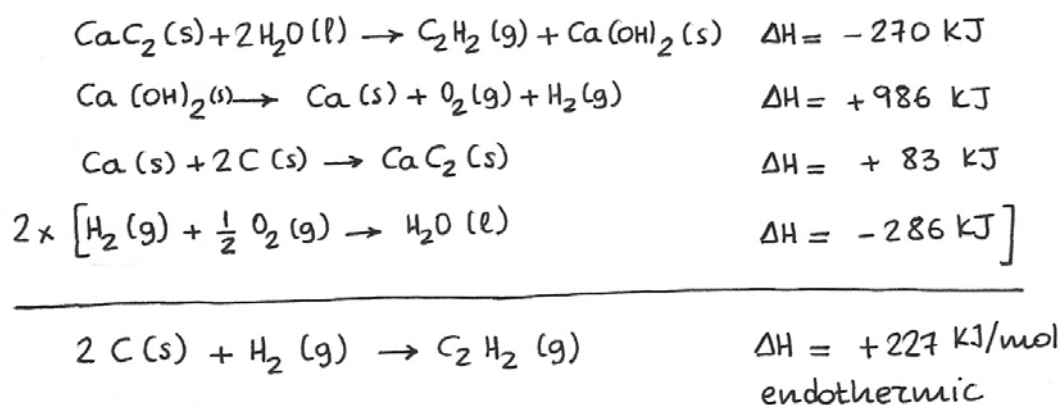
$$x = 30 \text{ g Al} \times \frac{3 \text{ mol O}_2}{(4 \times 27) \text{ g Al}} \times \frac{22.4 \text{ L}}{1 \text{ mol O}_2} = 18.67 \text{ L O}_2$$

- The price of the energy supplied:

$$x = 30 \text{ g Al} \times \frac{3.352 \times 10^6 \text{ J}}{108 \text{ g Al}} \times \frac{0.107 \text{ €}}{3.6 \times 10^6 \text{ J}} = 0.028 \text{ €}$$

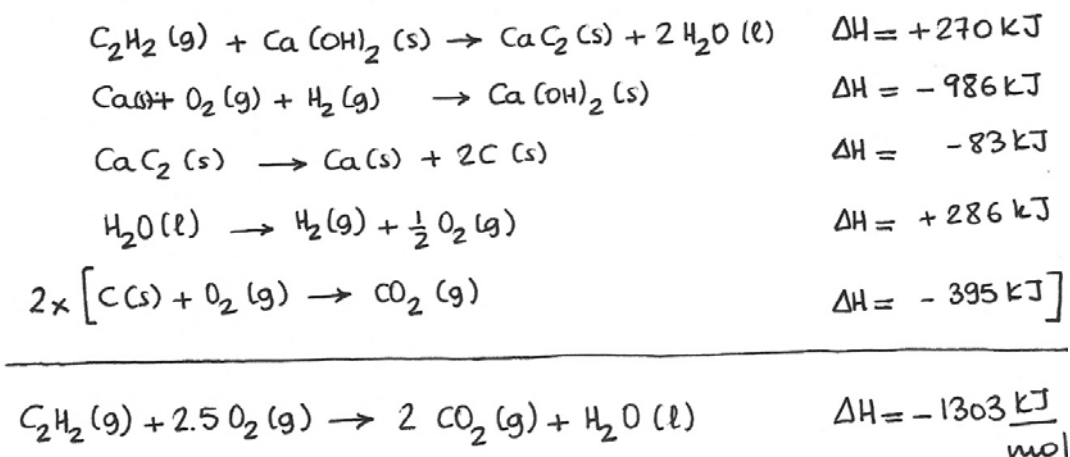
② a) The formation enthalpy-change of acetylene:  $(2\text{C}(s) + \text{H}_2(g) \rightarrow \text{C}_2\text{H}_2(g))$

The reactions we need to combine are:



⑥ b) The combustion enthalpy-change of acetylene:  $(\text{C}_2\text{H}_2(g) + 2.5\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + \text{H}_2\text{O}(l))$

The reactions we need to combine are:



③ We start by building up the table of # of moles

	$N_2O_4(g)$	$\rightleftharpoons$	$2 NO_2(g)$	$n_{total}$
$n_{initial}$	1.3		—	1.3
$n_{change}$	-x		2x	x
$n_{equilib.}$	1.3-x		2x	1.3+x

$n_{initial}$  can be calculated:

$$n = 120 \text{ g } N_2O_4 \times \frac{1 \text{ mol}}{92 \text{ g}} = 1.3 \text{ mol } N_2O_4$$

$$M(N_2O_4) = (2 \times 14) + (4 \times 16) = 92 \text{ g/mol}$$

If we apply the ideal gas law, the number of moles at equilibrium can be known:

$$PV = nRT \rightarrow n = \frac{PV}{RT} = \frac{33.8 \text{ atm} \times 2 \text{ L}}{0.082 \frac{\text{atm} \cdot \text{L}}{\text{K} \cdot \text{mol}} \times 373 \text{ K}}$$

$$n = 2.21 \text{ moles of gas}$$

From here, we get the value of x:

$$1.3 + x = 2.21 \rightarrow x = 2.21 - 1.3 = 0.91 \text{ mol}$$

Ⓐ The dissociation percent ( $\alpha$ )

$$\alpha = \frac{x}{n_{initial}} = \frac{0.91}{1.3} = 0.7 \rightarrow 70\% \text{ of initial } N_2O_4 \text{ has dissociated}$$

Ⓑ Mole fractions

The number of moles at equilibrium are

$$n(N_2O_4) = 1.3 - x = 0.39 \text{ moles}$$

$$n(NO_2) = 2x = 1.82 \text{ moles}$$

$$n_{total} = 1.3 + x = 2.21 \text{ moles}$$

The mole fractions are:

$$x(N_2O_4) = \frac{0.39 \text{ moles}}{2.21 \text{ moles}} = 0.176$$

$$x(\text{NO}_2) = \frac{1.82 \text{ moles}}{2.21 \text{ moles}} = 0.824$$

© Partial pressures

$$P_{\text{N}_2\text{O}_4} = x_{\text{N}_2\text{O}_4} \times P_T = 0.176 \times 33.8 \text{ atm} = 5.95 \text{ atm}$$

$$P_{\text{NO}_2} = x_{\text{NO}_2} \times P_T = 0.824 \times 33.8 \text{ atm} = 27.85 \text{ atm}$$

④ The concentrations at equilibrium are:

$$[\text{N}_2\text{O}_4] = \frac{0.39 \text{ moles}}{2 \text{ L}} = 0.195 \text{ M}$$

$$[\text{NO}_2] = \frac{1.82 \text{ moles}}{2 \text{ L}} = 0.91 \text{ M}$$

⑤ The equilibrium constants

$$K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = \frac{0.91^2}{0.195} = 4.25$$

$$K_p = \frac{P_{\text{NO}_2}^2}{P_{\text{N}_2\text{O}_4}} = \frac{27.85^2}{5.95} = 130.36$$

Ⓕ Disturbance	→	How the system shifts
a. $\text{N}_2\text{O}_4 \uparrow$		$\text{N}_2\text{O}_4 \downarrow$ $\rightleftharpoons$ shifts to products
b. $\text{NO}_2 \downarrow$		$\text{NO}_2 \uparrow$ $\rightleftharpoons$
c. $P \downarrow$		$P \uparrow$ $n \uparrow$ $\rightleftharpoons$
d. $V \downarrow$ $P \uparrow$		$P \downarrow$ $n \downarrow$ $\leftleftharpoons$ shifts to reactants
e. $T \uparrow$ heat $\uparrow$		heat $\downarrow$ $\rightleftharpoons$ shifts to products
f. It has no effect on composition		

- ④  
 (a) The table of # of moles and concentrations

	$B(g)$	$+ C(g)$	$\rightleftharpoons$	$2A(g)$
n initial	0.5	0.3		-
n change	$-x$	$-x$		$2x$
n equil.	$0.5-x$	$0.3-x$		$2x = 0.2 \text{ mol}$
[ ] equil.	$\frac{0.5-x}{V}$	$\frac{0.3-x}{V}$		$\frac{2x}{V}$

From the value of  $x$ , we can conclude that the moles at equilibrium are:

$$n(B) = 0.5 - x = 0.4 \text{ mol}$$

$$n(C) = 0.3 - x = 0.2 \text{ mol}$$

$$n(A) = 2x = 0.2 \text{ mol}$$

The equilibrium constants are:

$$K_c = \frac{[A]^2}{[B][C]} = \frac{(0.2/V)^2}{(0.4/V)(0.2/V)} = 0.5$$

$$K_p = K_c \cdot (RT)^{\Delta n} \xrightarrow{\Delta n=0} K_p = K_c = 0.5$$

- (b) After equilibrium is disturbed

	$B(g)$	$+ C(g)$	$\rightleftharpoons$	$2A(g)$
n initial	0.4	0.2		1.2
n change	$x$	$x$		$-2x$
n equil.	$0.4+x$	$0.2+x$		$1.2-2x$
[ ] equil.	$\frac{0.4+x}{V}$	$\frac{0.2+x}{V}$		$\frac{1.2-2x}{V}$

$$K_c = 0.5 = \frac{(1.2-2x)^2}{(0.4+x)(0.2+x)} = \frac{4x^2 - 4.8x + 1.44}{x^2 + 0.6x + 0.08} \rightarrow$$

$$\rightarrow 0.5x^2 + 0.3x + 0.04 = 4x^2 - 4.8x + 1.44 \rightarrow$$

$$\rightarrow 3.5x^2 - 5.1x + 1.4 = 0$$

$$x = \frac{5.1 \pm \sqrt{26.01 - 19.6}}{7} \rightarrow x = 0.367$$

The composition at equilibrium will be:

$$n(B) = 0.4 + x = 0.767 \text{ mol}; \quad n(C) = 0.2 + x = 0.567;$$

$$n(A) = 1.2 - 2x = 0.466 \text{ mol}$$