Form
A|B
C|D
CHEMISTRY 102
Exam III

Detailed Key Exam 3 Sp24

Spring 2024
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1. Consider the following reaction:

I   2.0M   2.0M   0
C   −4x   −x   +2x
     $4 NO_2(g) + O_2(g) \rightleftharpoons 2 N_2O_5(g)$ $K = 6.2 \times 10^{-6}$
E   2.0−4x 2.0−x   2x

$= \dfrac{[N_2O_5]^2}{[NO_2]^4[O_2]} = \dfrac{(2x)^2}{(2.0−4x)^2(2.0−x)}$
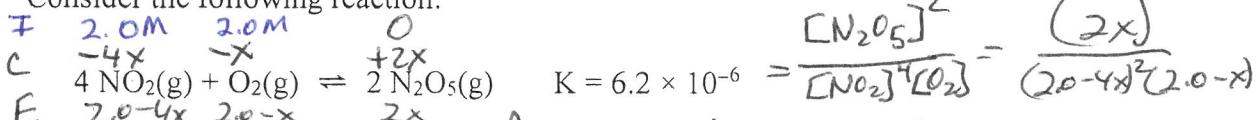
Assume $2.0−4x \approx 2.0$ and $2.0−x \approx 2.0$.

If 4.0 moles of NO_2 and 4.0 mol of O_2 are placed in a 2.0 L container, what is the equilibrium concentration of N_2O_5?

$6.2 \times 10^{-6} = \dfrac{4x^2}{(2.0)^4(2.0)} = \dfrac{x^2}{8}$

a) 1.4×10^{-2} M
b) 4.0×10^{-2} M
c) 8.0×10^{-2} M
d) 7.0×10^{-3} M
e) 2.5×10^{-3} M

Solving: $x = 7.04 \times 10^{-3}$ mol/L (Assumptions good.)

$[N_2O_5]_e = 2x = 2(7.04 \times 10^{-3})$ ⟨1.4×10^{-2} M⟩

Consider the following data for the next two questions:
Break problem into 5 steps.
Specific heat capacity of ice = 2.03 J/°C•g
Specific heat capacity of water = 4.18 J/°C•g
Specific heat capacity of steam = 2.02 J/°C•g
ΔH_{fusion} = 6.02 kJ/mol
$\Delta H_{vaporization}$ = 40.7 kJ/mol

$q_1 = \dfrac{2.03 J}{g \cdot °C} \times 18.02 g \times 30.0 °C = 1.097 \times 10^3 J$

$q_2 = 1.00 \text{ mol} \left(\dfrac{6020 J}{mol}\right) = 6020 J$

$q_3 = \dfrac{4.18 J}{g \cdot °C} \times 18.02 g \times 100.0 °C = 7.532 \times 10^3 J$

$q_4 = 1.00 \text{ mol}\left(\dfrac{40,700 J}{mol}\right) = 40,700 J$; $q_5 = \dfrac{2.02 J \times 18.02 \times 40.0 °C}{g \cdot °C} = 1.456 \times 10^3 J$

2. Consider a 1.00 mole sample of ice at −30.0°C, which is heated to steam at 140.0°C. Calculate q for the entire process.

$q_{TOTAL} = q_1 + q_2 + q_3 + q_4 + q_5 = 56,800 J$ ⟨56.8 kJ⟩

a) 123 kJ
b) 1070 kJ
c) 6.75 kJ
d) 234 kJ
e) 56.8 kJ

|heat loss by element| = |heat gain by ice-water mixture|

heat gain = 109.5g ice $\left(\dfrac{1 mol H_2O}{18.02g}\right)\left(\dfrac{6.02 kJ}{mol}\right) = 36.58$ kJ

3. A 500.0 g sample of an element at 195°C is dropped into an ice-water mixture. 109.5 g of ice melts and an ice-water mixture remains. Calculate the specific heat capacity of the element.

heat loss by element = S × mass × ΔT, make this a positive value by keeping ΔT positive.

$36.58 \times 10^3 J = S \times 500.0 \times (195−0)$

a) 2.67 J/°C•g
b) 0.375 J/°C•g
c) 0.789 J/°C•g
d) 1.62 J/°C•g

⟨S = 0.375 J/g•°C⟩ ; $T_f = 0$ since a ice/water mixture is present at end.

4. 1st law: energy of universe is constant ($\Delta E_{univ} = 0$).

Which of the following mathematical relationships is **false**?

$\Delta E_{univ} = 0 = \Delta E_{sys} + \Delta E_{surr}$, so $\Delta E_{sys} = −\Delta E_{surr}$.

T a) $\Delta E_{system} = −\Delta E_{surroundings}$

T b) At constant pressure, q = ΔE + PΔV.

$\Delta H = \Delta E + \Delta(PV)$, at constant P, $\Delta H = q$
and $\Delta(PV) = P\Delta V$, so $q = \Delta E + P\Delta V$.

F c) $\Delta H_{system} − \Delta H_{universe} = \Delta H_{surroundings}$

T d) At constant volume, ΔE = q. $\Delta E = q + w$, at constant V, $w = 0$. So $\Delta E = q$.

T e) For an overall cyclic process, $w_{overall} = −q_{overall}$.   For cyclic process, $\Delta E_{overall} = 0$

$\Delta E_{overall} = 0 = q_{overall} + w_{overall}$, so $w_{overall} = −q_{overall}$.

→ $\Delta H_{univ} = \Delta H_{sys} + \Delta H_{surr}$, $\Delta H_{univ} − \Delta H_{sys} = \Delta H_{surr}$

5/1
16/12

5. Consider the following reaction:

$$I \quad 4.0$$
$$C \quad -2x \qquad +2x \qquad +x$$
$$2\,NO_2(g) \rightleftharpoons 2\,NO(g) + O_2(g)$$
$$E \quad 4.0-2x \qquad 2x \qquad x$$
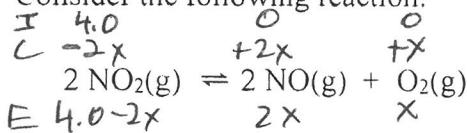
$$K = ? \frac{[NO]^2[O]}{[NO_2]^2} = \frac{(1.0)^2(0.50)}{(3.0)^2} = 0.056$$

4.0 moles of NO_2 are placed in a 1.0 L container and are allowed to react to reach equilibrium. At equilibrium the concentration of $NO(g)$ is 1.0 M. Calculate the value of the equilibrium constant, K, for this reaction.

From problem, $[NO]_e = 1.0 = 2x$, $x = 0.50\ m$.

a) 0.056   b) 0.11   c) 0.17   d) 25   e) 590

So $[NO_2] = 4.0 - 2(0.50) = 3.0\ m$ and $[O_2] = 0.50$.

6/2
17/13

6. The K_{sp} value for BaF_2 is 5.0×10^{-7}. When 10.0 mL of 0.010 M NaF is mixed with 10.0 mL of 0.010 M $Ba(NO_3)_2$, will a precipitate form? Equal volumes of 2 different solutions were added together, so initial concentrations are halved.

a) No, because $Q = 1.0 \times 10^{-6}$ and since it is greater than K_{sp}, no precipitate will form.
b) Yes, because $Q = 1.0 \times 10^{-6}$ and since it is greater than K_{sp}, a precipitate will form.
c) No, because $Q = 1.3 \times 10^{-7}$ and since it is less than K_{sp}, no precipitate will form.
d) Yes, because $Q = 1.3 \times 10^{-7}$ and since it is less than K_{sp}, a precipitate will form.

$BaF_2 \rightleftharpoons Ba^{2+} + 2F^-$ $Q = [Ba^{2+}]_0[F^-]^2 = (0.0050)(0.0050)^2 = 1.25 \times 10^{-7}$
Here, $Q < K (=5.0 \times 10^{-7})$, so ion concentrations are too small. No precip. will form.

7/13
18/14
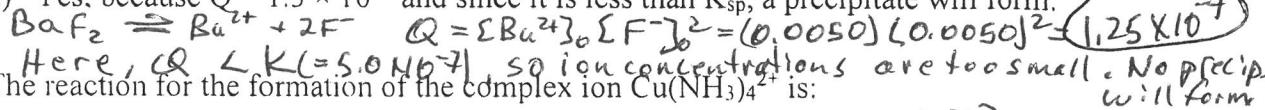
7. The reaction for the formation of the complex ion $Cu(NH_3)_4^{2+}$ is:

$$Cu^{2+}(aq) + 4\,NH_3(aq) \rightleftharpoons Cu(NH_3)_4^{2+}(aq)$$
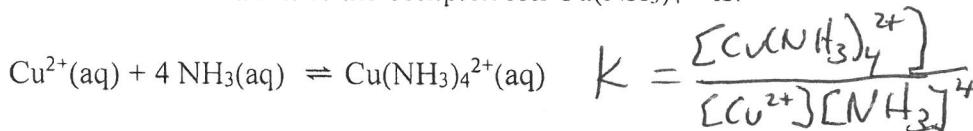
$$K = \frac{[Cu(NH_3)_4^{2+}]}{[Cu^{2+}][NH_3]^4}$$

Which of the following is the correct expression for the equilibrium constant, K, for this reaction?

a) $K = \dfrac{[Cu(NH_3)_4^{2+}]}{[Cu^{2+}][NH_3]}$ b) $K = \dfrac{[Cu(NH_3)_4^{2+}]^4}{[Cu^{2+}][NH_3]^4}$ c) $K = \dfrac{1}{[Cu^{2+}][NH_3]}$

d) $K = \dfrac{1}{[Cu^{2+}][NH_3]^4}$ e) $K = \dfrac{[Cu(NH_3)_4^{2+}]}{[Cu^{2+}][NH_3]^4}$

Energy added to bomb from CH_4 combustion $= 6.79\,g\,CH_4 \left(\dfrac{1\,mol\,CH_4}{16.04\,g}\right)\left(\dfrac{802\,kJ}{mol\,CH_4}\right) = 339.5\,kJ$

8/4
19/15

8. The heat capacity of a bomb calorimeter was determined by combusting 6.79 g of CH_4 (energy of combustion of $CH_4 = -802$ kJ/mol CH_4); the temperature of the bomb calorimeter increased by 10.8°C. In a second experiment with the same bomb calorimeter, combustion of a 12.6 g sample of ethyne, C_2H_2, produced a temperature increase of 16.9°C. Calculate the energy of combustion of C_2H_2 in kJ/mol. The molar mass of CH_4 is 16.04 g/mol and the molar mass of C_2H_2 is 26.04 g/mol.

Heat capacity of bomb $= 339.5\,kJ/10.8°C = 31.44\,kJ/°C$

a) -31.4 kJ/mol   b) -1.10×10^3 kJ/mol   c) 1.47×10^3 kJ/mol

Heat added to bomb from C_2H_2 combustion $= 16.9°C \left(\dfrac{31.44\,kJ}{°C}\right) = 531.3\,kJ$

d) -532 kJ/mol   e) 340. kJ/mol

$$\Delta E_{comb} = \frac{-531.3\,kJ}{12.6\,g\,C_2H_2} \cdot \frac{26.04\,g\,C_2H_2}{mol\,C_2H_2} = -1.10 \times 10^3\,kJ/mol\,C_2H_2$$

Note: temperature of the bomb increased, so reaction is exothermic.

Form
A|B
C|D

9|16
——
24|5

9. Consider the following reaction which has an enthalpy change of ΔH_T:

$$A \rightarrow X \qquad \Delta H = \Delta H_T$$

This reaction can be broken down into a series of steps as shown in the following diagram:
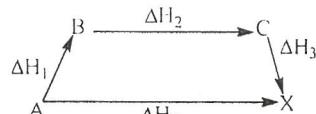
From Hess's law $= \Delta H_T = \Delta H_1 + \Delta H_2 + \Delta H_3$, *so*

Which of the following relationships **must** be **true** for this reaction?

$\Delta H_T - \Delta H_1 - \Delta H_2 - \Delta H_3 = 0$

a) $\Delta H_T - \Delta H_1 - \Delta H_2 - \Delta H_3 = 0$ ⟵ (circled)

b) $\Delta H_2 - (\Delta H_3 + \Delta H_1) = \Delta H_T$

c) $\Delta H_T + \Delta H_1 + \Delta H_2 + \Delta H_3 = 0$

d) $\Delta H_T + \Delta H_2 = \Delta H_1 + \Delta H_3$

e) $\Delta H_3 - (\Delta H_1 + \Delta H_2) = 0$

10|17
——
25|6

10. Consider the following reaction at some temperature:

I 3.00 atm 3.00 atm   0
C   -x   -x   +2x

$$Cl_2(g) + F_2(g) \rightleftharpoons 2\, ClF(g) \qquad K_p = 16 = \frac{(P_{ClF})^2}{P_{Cl_2} \times P_{F_2}} = \frac{(2x)^2}{(3-x)^2} = \left(\frac{2x}{3-x}\right)^2$$

E 3-x   3-x   2x
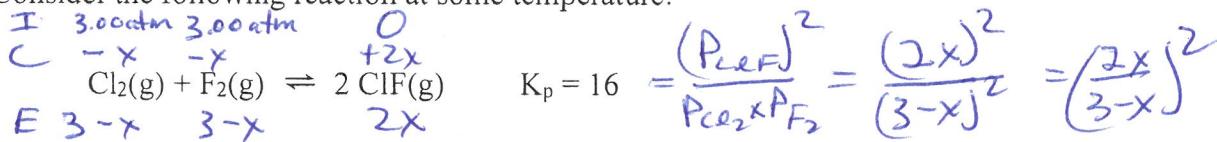
If initially 3.00 atm of F_2 and 3.00 atm of Cl_2 are reacted in a rigid container, calculate the equilibrium partial pressure of $Cl_2(g)$.

Taking the square root of both sides: $4 = \frac{2x}{3-x}$

a) 0.25 atm   b) 0.50 atm   c) 1.0 atm (circled)

$12 - 4x = 2x$, $12 = 6x$, $x = 2.0$ atm

d) 1.5 atm   e) 2.0 atm

$P_{Cl_2} = 3.00 - x = 3.00 - 2.0 = \boxed{1.0\ atm}$

$Q = [Ca^{2+}]_o [F^-]_o^2 = 0.10(0.10)^2$
$Q = 1.0 \times 10^{-3}$ which is greater
than K, so rxn shifts
left to reach equilibrium.
Note the initial concentrations
are halved
since volume
doubled.

11|18
——
26|7

11. A 50.0 mL sample of 0.200 M $Ca(NO_3)_2$ is mixed with 50.00 mL of 0.200 M NaF. When the system has come to equilibrium, what are the equilibrium concentrations of Ca^{2+} and F^-? K_{sp} for $CaF_2 = 4.0 \times 10^{-11}$.

Let rxn go to completion then solve an equilibrium problem

	$[Ca^{2+}]_e$	$[F^-]_e$
a)	5.0×10^{-2} M	1.4×10^{-5} M
b)	1.0×10^{-1} M	1.0×10^{-5} M
c)	2.2×10^{-4} M	4.3×10^{-4} M
d)	1.0×10^{-1} M	2.0×10^{-5} M
e)	5.0×10^{-2} M	2.8×10^{-5} M

(d and e circled)

$CaF_2(s) \rightleftharpoons Ca^{2+} + 2F^- \quad K_{sp} = 4.0 \times 10^{-11}$

I	-	0.10	0.10
C	+0.05	-0.05	-0.10 (F⁻ limiting)
A	-	0.05	0
I	-	0.05	0
C	-s	+s	+2s
E	-	0.05+s	2s

$K_{sp} = 4.0 \times 10^{-11} = (0.05+s)(2s)^2$

$4.0 \times 10^{-11} \approx 0.05(4s^2)$, solving: $s = 1.41 \times 10^{-5}$ mol/L (Assumption good.)

At equil: $[Ca^{2+}] = 0.05 + 1.41 \times 10^{-5} \approx \boxed{0.050\ M}$

$[F^-] = 2(1.41 \times 10^{-5}) = \boxed{2.8 \times 10^{-5}\ M}$

Handwritten at top: $AlF_3(s) \rightleftharpoons Al^{3+} + 3F^-$ $K_{sp} = [Al^{3+}][F^-]^3$

	$AlF_3(s)$	Al^{3+}	$3F^-$
I		0	0
C	$-S$	$+S$	$+3S$
E		S	$3S$
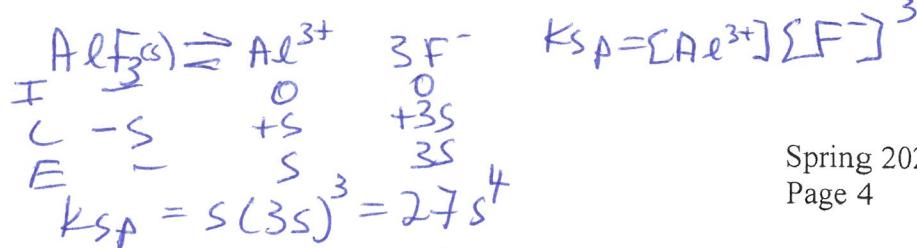
$K_{sp} = S(3S)^3 = 27S^4$

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12. Aluminum fluoride has a solubility of 6.66×10^{-2} mol/L. Calculate the K_{sp} value for $AlF_3(s)$.

Handwritten: From problem, $S = 6.66 \times 10^{-2}$.
$K_{sp} = 27(6.66 \times 10^{-2})^4 = 5.31 \times 10^{-4}$

a) 4.44×10^{-3}

b) 1.97×10^{-5}

c) 5.90×10^{-5}

(d) 5.31×10^{-4}

e) 5.90×10^{-6}

Handwritten box: For problem below, if we break all bonds and form all bonds, $\Delta H = 0$. Or we can break only the bonds that we have to in order to convert reactants to products. ΔH still $= 0$.

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13. Using the bond energies below, estimate the enthalpy change in kJ for the reaction:
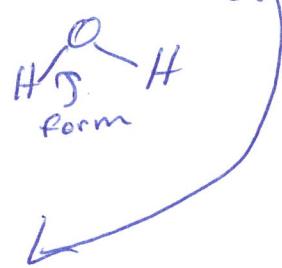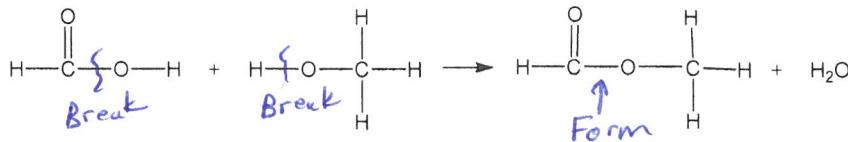
Bond Energy (kJ/mol)	
C–H	413
C–O	358
C=O	745
O–H	467
C–C	347
H–H	432

Handwritten:
Break	Form
C–O	C–O
O–H	O–H

Since we break the exact same bonds that are formed, $\Delta H = 0$ from bond energies.

a) $\Delta H = -2654$ kJ   b) $\Delta H = -567$ kJ   c) $\Delta H = 567$

(d) $\Delta H = 0$ kJ   e) $\Delta H = 1258$ kJ

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14. An equilibrium mixture for the reaction:

$$2\, NH_3(g) \rightleftharpoons N_2(g) + 3\, H_2(g) \qquad K = ? \quad \frac{[N_2][H_2]^3}{[NH_3]^2}$$

contains 9.0 mol N_2, 6.0 mol H_2, and 6.0 mol NH_3, all in a 3.0 L container at a certain temperature. Calculate the value of K for this reaction at this temperature.

a) 2.0   b) 4.0   (c) 6.0

d) 9.0   e) 12

Handwritten:
$$K = \frac{\left(\frac{9.0\,mol\,N_2}{3.0\,L}\right)^1 \left(\frac{6.0\,mol\,H_2}{3.0\,L}\right)^3}{\left(\frac{6.0\,mol\,NH_3}{3.0\,L}\right)^2} = \frac{3(2)^3}{2^2} = 6.0$$

Left margin handwritten: 12|5 13|16 ; 13|6 14|17 ; 14|7 15|18

Form
A|B
C|D

(handwritten top) $I \quad Ag_2CO_3(s) \rightleftharpoons 2Ag^+ + CO_3^{2-} \quad k_{sp} = 8.2 \times 10^{-12} = [Ag^+]^2[CO_3^{2-}]$

$C \quad -S \qquad 0 \qquad 0.50m$

$\qquad +2S \qquad +S$

$E \qquad - \qquad 2S \qquad 0.50+S$
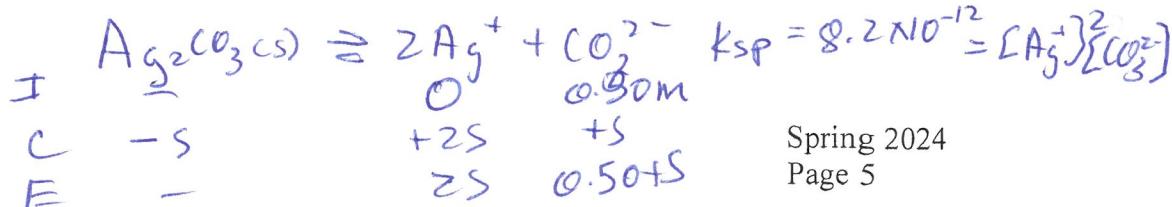
15. Calculate the molar solubility of $Ag_2CO_3(s)$ in an aqueous solution that is 0.50 M in K_2CO_3. K_{sp} for $Ag_2CO_3 = 8.2 \times 10^{-12}$

(handwritten) $k_{sp} = 8.2 \times 10^{-12} = (2S)^2(0.50+S), \quad 8.2 \times 10^{-12} = 4S^2(0.50)$

a) 2.0×10^{-6} mol/L   b) 2.6×10^{-4} mol/L   c) 1.3×10^{-4} mol/L

(handwritten) Solving: $(S = 2.0 \times 10^{-6} \text{ mol/L})$ Assumption good.

d) 2.8×10^{-6} mol/L   e) 3.4×10^{-3} mol/L

(handwritten) The bond energy equation for HCl is: $HCl \rightarrow H + Cl \quad \Delta H = HCl$ bond energy

16. Given the following set of gas phase reactions:

$$H_2 + Cl_2 \rightarrow 2\ HCl \qquad \Delta H° = -184 \text{ kJ}$$
$$H_2 \rightarrow 2\ H \qquad \Delta H° = 432 \text{ kJ}$$
$$Cl_2 \rightarrow 2\ Cl \qquad \Delta H° = 240 \text{ kJ}$$

(handwritten)
$HCl \rightarrow \tfrac{1}{2}H_2 + \tfrac{1}{2}Cl_2 \quad \Delta H_1 = 184/2$
$\tfrac{1}{2}H_2 \rightarrow H \quad \Delta H_2 = 432/2$
$\tfrac{1}{2}Cl_2 \rightarrow Cl \quad \Delta H_3 = 240/2$
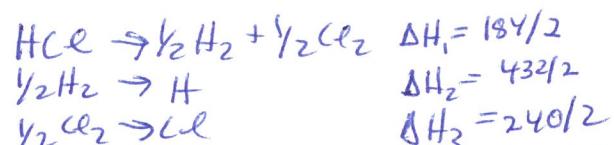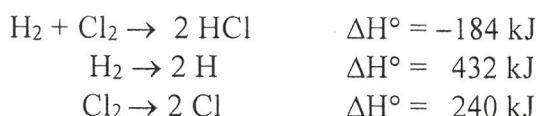
calculate the bond energy of the H–Cl bond.

(handwritten) $HCl \rightarrow H + Cl \qquad \Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3$

a) 248 kJ   b) 428 kJ   c) 864 kJ

(handwritten) $\Delta H = 92 + 216 + 120 = (428 \text{ kJ}) = HCl$ bond energy

d) 92 kJ   e) 184 kJ

(handwritten) $\Delta E_{overall} = \Delta E_I + \Delta E_{II}$
$\Delta E_I = \Delta E$ from Step I
$\Delta E_{II} = \Delta E$ from Step II

17. Determine ΔE for a system which undergoes both of the following two processes:

(handwritten) $\Delta E_I = q_I + w_I = 30J + 15J = 45J$

I. 15 J of work is done on the system, while 30 J of heat is absorbed by it; then:

(handwritten) $\Delta E_{II} = q_{II} + w_{II} = -20J - 25J = -45J$

II. 20 J of heat is given off by the system, while 25 J of work is done by the system.

(handwritten) $\Delta E_{overall} = 45J - 45J = 0$

a) 30 J   b) –30 J   c) 40 J

d) 45 J   e) 0

18. Consider the reaction:

(handwritten) $Q = [N_2O]_0[H_2O]_0^2 = 2.0(2.0)^2 = 8.0$

$$NH_4NO_3(s) \rightleftharpoons N_2O(g) + 2\ H_2O(g) \qquad K = 4.85$$
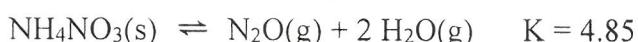
(handwritten) $Q > K (= 4.85)$, so rxn shifts left to reach equilibrium.

200. g of $NH_4NO_3(s)$ are placed in a reaction vessel containing $H_2O(g)$ at a concentration of 2.0 M and $N_2O(g)$ at a concentration of 2.0 M. After equilibrium is reached, will the mass of $NH_4NO_3(s)$ increase, decrease or remain unchanged?

(handwritten) So products get smaller, and reactants $[NH_4NO_3(s)]$ get bigger.

a) The mass of $NH_4NO_3(s)$ will decrease (< 200. g).

b) The mass of $NH_4NO_3(s)$ will increase (> 200. g).

c) The mass of $NH_4NO_3(s)$ will remain unchanged (= 200. g).

(handwritten) It doesn't matter that the reactant is a solid. When a rxn shifts left to get to equilibrium, all reactants get bigger, whether they are a solid or a liquid or a solute or a gas.

When container volume is increased, the reaction shifts to the side with more moles of gas. Here the rxn shifts left.

19 / 8
5 / 1

19. Consider the following endothermic reaction at equilibrium:

→ heat is a reactant

heat + $\quad SiHCl_3(g) + 3\,H_2O(g) \rightleftharpoons SiH(OH)_3(s) + 3\,HCl(g)$

4 mol gas $\qquad$ *3 mol g*
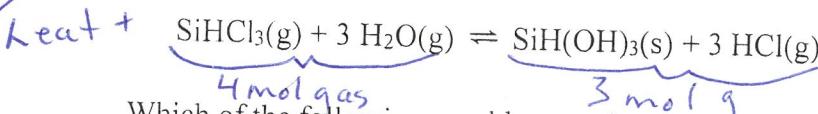
Which of the following would cause the reaction to shift toward reactants (shift left) to reestablish equilibrium?

a) Adding $H_2O(g)$. *—rxn shifts right when reactant gas is added*
b) Increasing the temperature. *—rxn shifts right when heat (a reactant) is added.*
c) Adding $Ar(g)$ (assume a constant volume container). *No effect since no concentrations change.*
yes d) Increasing the volume of the reaction container.
e) Adding $SiH(OH)_3(s)$. *—No effect; how much solid is present has no effect on the equilibrium.*

20 / 9
6 / 2

20. How many of the following five quantities is/are state functions?

State functions have values that are path independent. $\Delta P, \Delta H, \Delta T$ are all path independent, so they are state functions.

I. work $= w$ $\qquad$ II. enthalpy $= \Delta H$ $\quad$ III. highway distance from Peoria to Champaign

IV. pressure $= P$ $\quad$ V. temperature $= T$

Path functions are ones which depend on the path one takes to go from the initial state to the final state.

a) 1 $\qquad$ b) 2 $\qquad$ c) 3 $\qquad$ d) 4 $\qquad$ e) 5 (All are state functions.)

work (w), heat (q) are the main path functions in thermo. Here, highway distance is also a path function.

21 / 10
7 / 3

21. Consider the reaction:

$$Ca_3P_2(s) + 6\,H_2O(l) \rightarrow 3\,Ca(OH)_2(s) + 2\,PH_3(g) \qquad \Delta H = -721.70\,kJ = q$$
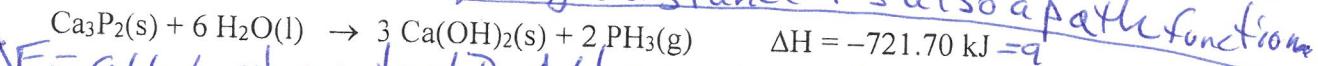
$\Delta E = q + w$, at constant P, $\Delta H = q$, $w = -P\Delta V = -RT\Delta n$

If this reaction were carried out at 298 K in a beaker open to a constant-pressure atmosphere, which of the following would be **false**?

$w = -RT\Delta n = -8.3145\,J (298k)(2-0) = -4960\,J = -4.96\,kJ$
$\frac{}{K\cdot mol}$

no a) $\Delta E = -716.74\,kJ$ $\qquad$ b) $w = -4.96\,kJ$ $\qquad$ c) $q_p = -721.70\,kJ$

$\Delta E = q + w = -721.70 - 4.96 = -726.66$

d) The reaction is exothermic. $\qquad$ e) Heat is released.

ok $\qquad\qquad\qquad\qquad\qquad$ *ok*

Reaction is exothermic, so final temp of calorimeter will be greater than 25.00°C.

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8 / 4

22. The following reaction was studied in a coffee cup calorimeter:

mol HCl $= 0.1L\left(\frac{0.5mol}{L}\right) = 0.0500$ mol HCl $\leftarrow$ HCl is limiting so use HCl info to calculate heat released.

$$2\,HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + 2\,H_2O(l) \qquad \Delta H = -118\,kJ$$

mol $Ba(OH)_2 = 0.3L\left(\frac{0.50mol}{L}\right) = 0.15$ mol $Ba(OH)_2$

In the coffee cup calorimeter, 100.00 mL of 0.500 M HCl was mixed with 300.0 mL of 0.500 M $Ba(OH)_2$. Assuming a solution density of 1.00 g/mL, a solution heat capacity of 4.18 J/°C·g, an initial solution temperature of 25.00°C, and no heat loss to the coffee cups, calculate the final temperature of the reaction mixture.

Total mass surroundings = 100g + 300g = 400g

heat released = 0.0500 mol HCl $\left(\frac{-118 kJ}{2 mol HCl}\right) = -2.95\,kJ$

a) 23.24°C $\qquad$ b) 25.92°C $\qquad$ c) 2.89°C $\qquad$ d) 27.89°C $\qquad$ e) 26.76°C

|heat gain by calorimeter| = |heat loss by rxn| = 2950 J

$2950\,J = S \times mass \times \Delta T$, $2950 = \frac{4.18 J}{g\cdot °C} \times 400g \times \Delta T$, $\Delta T = 1.76°C$

$T_f = 25.00 + 1.76 = 26.76°C$

A|B
C|D

$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ $\Delta H = -1411.1 kJ$
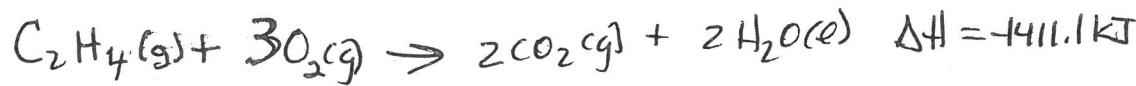
$\Delta H_{rxn} = \Sigma \Delta H^\circ_{f,prod} - \Sigma \Delta H^\circ_{f,react}$

$-1411.1 kJ = [2(-3935) + 2(-285.9)] - [x + 3(0)]$ where $x = \Delta H^\circ_{f,C_2H_4}$

23. When ethene gas, C_2H_4, is combusted at 298 K, $CO_2(g)$ and $H_2O(l)$ are produced. The enthalpy of combustion of ethene at 298 K is –1411.1 kJ/mol ethene. Given the following standard enthalpies of formation, calculate the standard enthalpy of formation for ethene.

Solving: $x = \Delta H^\circ_{f,C_2H_4} = 52.3 kJ/mol$
Note that $\Delta H^\circ_{f,O_2} = 0$ as is the case for all elements in their standard state.

ΔH_f°

CO$_2$(g)   –393.5 kJ/mol

H$_2$O(l)   –285.9 kJ/mol

a) 52.3 kJ/mol   b) 731.7 kJ/mol   c) –2769.9 kJ/mol

d) –126.2 kJ/mol   e) –296.0 kJ/mol

	$La(IO_3)_3(s)$ $\rightleftharpoons$	La^{3+}	$+3IO_3^-$	$Ksp=?$
I		0	0.10M	
C	–s	+s	+3s	
E		s	0.10+3s	
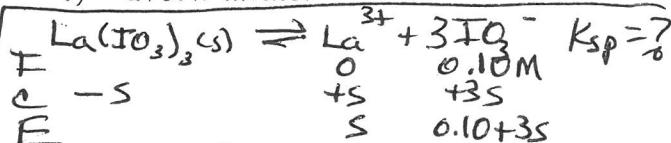
24. The solubility of La(IO$_3$)$_3$ in a 0.10 M KIO$_3$ solution is 1.0×10^{-7} mol/L. Calculate the K$_{sp}$ value for La(IO$_3$)$_3$.

$Ksp = [La^{3+}][IO_3^-]^3 = s(0.10+3s)^3$; from problem, $s = 1.0 \times 10^{-7}$ mol/L.

a) 1.0×10^{-8} b) 2.7×10^{-9} c) 1.0×10^{-10}

$Ksp = 1.0 \times 10^{-7}[0.10 + 3(1.0 \times 10^{-7})]^3 = 1.0 \times 10^{-10} = Ksp$

d) 2.7×10^{-27} e) 6.2×10^{-8}

25. Consider the following five reactions:

I. P$_4$(s) + 6 Cl$_2$(g) → 4 PCl$_3$(g)

II. P$_4$(s) + 5 O$_2$(g) → P$_4$O$_{10}$(g)

III. PCl$_3$(g) + ½ O$_2$(g) → Cl$_3$PO(l)

IV. PCl$_3$(g) + Cl$_2$(g) → PCl$_5$(g)

V. P$_4$O$_{10}$(s) + 6 PCl$_5$(g) → 10 Cl$_3$PO(l)

Bond energies give a good estimate for ΔH_{rxn} when only gases are present in the reaction. Only rxn IV contains only gases, so it is the correct answer. The other reactions contain solids or liquids which have intermolecular forces which are not accounted for when using bond energies.

For which reaction should bond energies give the best estimation for $\Delta H_{reaction}$?

a) I   b) II   c) III   d) IV   e) V

Let s = Molar solubility in mol/L; least soluble has smallest s value.

26. Which of the following compounds has the smallest molar solubility (mol/L) in water?

a) Al(OH)$_3$ K$_{sp}$ = 2 × 10^{-32} = 27s^4, s = 5.2 × 10^{-9} mol/L

b) CdS   K$_{sp}$ = 1 × 10^{-28} = s^2, s = 1.0 × 10^{-14} mol/L

c) PbSO$_4$ K$_{sp}$ = 1 × 10^{-8} = s^2, s = 1.0 × 10^{-4} mol/L

d) Sn(OH)$_2$ K$_{sp}$ = 3 × 10^{-27} = 4s^3, s = 9.1 × 10^{-10} mol/L

e) MgCO$_3$ K$_{sp}$ = 7 × 10^{-6} = s^2, s = 2.6 × 10^{-3} mol/L

These relationships come from setting up the ICE table for each salt.

The ΔH_f° reaction always ⊛ the reaction of producing 1 mol of the compound of interest [here $H_2O(l)$] from the elements that make up the compound all in their standard states.

27 | 27
27 | 27

27. The standard enthalpy of formation for liquid water is –286 kJ/mol. For which of the following reactions does $\Delta H_{reaction}$ = –286 kJ?

Answer d in correct.

a) $H_2O(l) \rightarrow 2\ H(g) + O(g)$

b) $2\ H(g) + O(g) \rightarrow H_2O(l)$ → *H and O not standard states of $H_2 + O_2$.*

c) $H_2O(l) \rightarrow H_2(g) + \frac{1}{2}\ O_2(g)$ → *reverse of ΔH_f° reaction for $H_2O(l)$*

d) $H_2(g) + \frac{1}{2}\ O_2(g) \rightarrow H_2O(l)$

e) $H_2O(g) \rightarrow H_2O(l)$
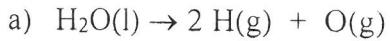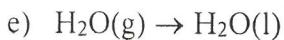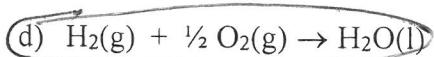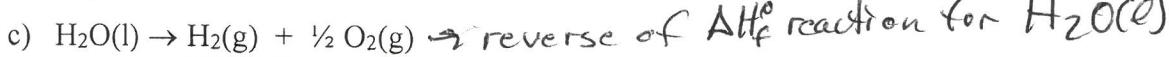
28 | 28
28 | 28

28. Consider when 1.0 mole of A(g) and 1.0 mole of B(g) are added to a 2.0 L container and the following balanced reaction occurs:

We know nothing about the K value for the reaction. K could be a big or a very small number.

$$I\ \ 0.50M\ \ 0.50$$
$$\ \ -x\ \ -2x\ \ +2x\ \ +3x$$
$$A(g) + 2\ B(g) \rightleftharpoons 2\ C(g) + 3\ D(g)$$
$$E\ \ 0.50-x\ \ 0.50-2x\ \ 2x\ \ 3x$$

Which of the following statements **must** be **true** regarding this reaction once equilibrium has been reached?

maybe a) The value of the equilibrium constant for this reaction must be greater than 1 (K > 1).

maybe b) The value of the equilibrium constant for this reaction must be greater than 1 (K < 1).

F c) At equilibrium, the rate of the reverse reaction must be greater than the rate of the forward reaction. *Rates are equal at equilibrium.*

maybe d) At equilibrium, the concentration of B(g) must be equal to the concentration of C(g) ([B]$_e$ = [C]$_e$). *0.50–2x may equal 2x, but it doesn't have to*

T e) At equilibrium, the concentration of B(g) must be smaller than the concentration of A(g) ([B]$_e$ < [A]$_e$). *$[A]_e = 0.50-x$, $[B]_e = 0.50-2x$*

[B] must be smaller since we are subtracting a larger value (2x) from 0.50 to calculate concentration.

29. My chemistry 102 exam should be graded with the answer sheet associated with:

a) Form A b) Form B c) Form C d) Form D e) Form E