Form Detailed Key Exam 3 Fall 22

A|B
C|D

1|12
24|5

1. Initially, 1 mol of SO_3 is placed into a 2.0 L container at a constant temperature. Some SO_3 reacts to form SO_2 and O_2 by the following equilibrium reaction:

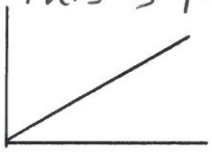
I 0.50M 0 0
C −2x +2x +x

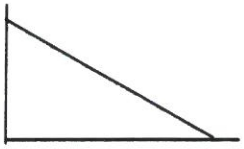
$$2\,SO_3(g) \rightleftharpoons 2\,SO_2(g) + O_2(g)$$

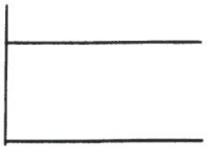
E 0.50−2x 2x x

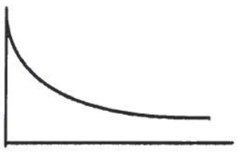
Which of the following graphs most accurately represents the SO_3 concentration versus time plot for this equilibrium reaction? [SO_3] is the y-axis, time is the x-axis.

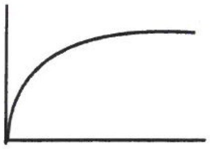
SO_3 starts at 0.50M, then decreases to 0.50−2x at equilibrium. this is plot d.

a) 

b) 

c) 

[✓ d)] 

e) 

2|13
25|6

2. Some A(g) is reacted by the following the reaction at a constant pressure of 1.0 atm:

ΔE = q+w; we don't know how much A reacted, so q may not be

$$A(g) \rightarrow 2\,B(g) \qquad \Delta H = -100.\ kJ$$

−100kJ if 1 mol A wasn't reacted. Let's determine q to see how much A reacted.

As a result of this reaction, the volume of the reaction container increases by 50.0 L and the internal energy changes by −30.0 kJ. How many moles of A were consumed in the reaction?

$w = -P\Delta V = -1.0\,atm(50.0\,L) = -50.\ L\cdot atm\left(\frac{101.3\,J}{L\cdot atm}\right) = -5065\,J$

$\Delta E = -30.0\,kJ = q + (-5.065\,kJ), \quad q = -24.94\,kJ$

a) 1.25 mol b) 0.80 mol c) 0.35 mol [✓ d) 0.25 mol] e) 0.20 mol

$A_{reacted} = \frac{-24.94\,kJ \;|\; 1\,mol\,A}{\;|\; -100\,kJ} = 0.25$ mol A reacted to release −24.94 kJ of heat

3|14
26|7

3. A system undergoes a process consisting of the following two steps:

$\Delta E_{overall} = \Delta E_1 + \Delta E_2$, $\Delta E = q+w$

Step 1: The system absorbs 40 J of heat while 60 J of work is done on it.
Step 2: The system absorbs 60 J of heat while performing 40 J of work.

$\Delta E_1 = 40\,J + 60\,J = 100\,J$

Calculate ΔE for the overall process.

$\Delta E_2 = 60\,J - 40\,J = 20\,J$

a) 200 J [✓ b) 120 J] c) 0 J d) −120 J e) −200 J

$\Delta E_{overall} = 100 + 20 = \boxed{120\,J}$

Form
A | B
C | D

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

	$CaF_2(s)$	Ca^{2+}	$2F^-$
F		0.500M	0.500M
C		−0.250	−0.500
A		0.250	0
I		0.250	0
C	−s	+s	+2s
E		0.250+s	2s

F^- is limiting

when equal volumes of 2 different solutions are mixed, concentration are halved since volume doubled

Consider the following information for the next two questions. When 50.0 mL of 1.00 *M* $Ca(NO_3)_2$ is mixed with 50.0 mL of 1.00 *M* KF, a precipitate of $CaF_2(s)$ forms. K_{sp} for CaF_2 = 4.0×10^{-11}. $4.0 \times 10^{-11} = (0.250+s)(2s)^2 \approx 0.250(4s^2)$

solving: $s = 6.3 \times 10^{-6}$ mol/L (Assumption great)

4 | 11
20 | 23

4. Calculate the equilibrium concentration of Ca^{2+} in solution ($[Ca^{2+}]$ = ?).

$[Ca^{2+}] = 0.250 + 6.3 \times 10^{-6} = \boxed{0.250M}$

a) 1.00 *M*   b) 0.75 *M*   c) 0.50 *M*   [✓ d) 0.25 *M*]   e) 0.050 *M*

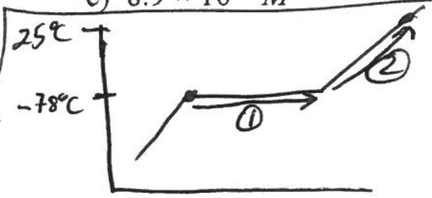
5 | 12
21 | 24

5. Calculate the equilibrium concentration of F^- in solution ($[F^-]$ = ?).

$[F^-] = 2s = 2(6.3 \times 10^{-6}) = \boxed{1.3 \times 10^{-5} M}$

a) 4.5×10^{-6} *M*   [✓ b) 1.3×10^{-5} *M*]   c) 8.9×10^{-6} *M*

d) 6.3×10^{-6} *M*   e) 2.0×10^{-11} *M*



$q_{Tot} = q_1 + q_2$

6 | 3
22 | 25

6. At 1.00 atm and −78.0°C, carbon dioxide converts directly from the solid phase to the gaseous phase without going through the liquid phase; this is called sublimation. The sublimation process and energy change for $CO_2(s)$ at −78.0°C is:

Step 1: $CO_2(s, -78°C) \rightarrow CO_2(g, -78°C)$, $q_1 = 25.3 kJ/mol (88.0 g)\left(\frac{1 mol CO_2}{44.01 g}\right) = 50.6 kJ$

$CO_2(s) \rightarrow CO_2(g)$ $\Delta H_{sublimation}$ = 25.3 kJ/mol

Step 2: $CO_2(g, -78°C) \rightarrow CO_2(g, 25°C)$, $q_2 = \frac{0.844 J}{g\,°C}(88.0 g)[25-(-78)] = 7.65 \times 10^3 J$

Calculate the energy required to convert 88.0 g of $CO_2(s)$ at −78.0°C to $CO_2(g)$ at 25°C. The specific heat capacity of $CO_2(g)$ is 0.844 J/°C•g.

$q_{TOTAL} = q_1 + q_2 = 50.6 kJ + 7.65 kJ = 58.25 = \boxed{58.3 kJ}$

a) 3.99 kJ   b) 3.99×10^3 kJ   [✓ c) 58.3 kJ]   d) 7.70×10^3 kJ   e) 54.5 kJ

7 | 4
23 | 26

7. Consider the following equilibrium constant vs. temperature data for some reaction:

K	Temp
2.54×10^4	109°C
5.04×10^2	225°C
6.33×10^1	303°C
2.25×10^{-1}	412°C
3.03×10^{-3}	539°C

From data, as T↑, K decreases. This is indicative of an exothermic rxn.

$A \rightarrow B + heat$

As heat (a product) is added, rxn shifts left producing more reactants. End result is that K decreases as T increases for an exothermic rxn.

Which of the following statements (a-c) is/are **true**?

[✓ a) The reaction is exothermic.]
b) The reaction is endothermic.
c) As temperature increases, more products are produced in order to get to equilibrium.
d) Statements a and c are both true.
e) Statements b and c are both true.

Form

A | B
C | D

8 | 23
16 | 8

8. Which of the following statements (a-d) is **true**?

F a) If the reaction quotient Q is greater than K, the reaction must shift to the ~~right~~ left to establish equilibrium.

F b) The value of K can never equal the value of K_p. When $\Delta n = 0$, $K = K_p$.

F c) At equilibrium, the rate of the forward reaction and the rate of the reverse reaction both equal ~~zero~~ each other, but is not zero.

F d) A reaction contains mostly ~~products~~ reactants at equilibrium when K << 1. $K = \frac{[Prod]^x}{[React]^y}$

[✓ e)] None of the above statements (a-d) are true.

9 | 24
17 | 9

9. Consider the following reaction at some constant temperature:

$$\begin{array}{llll} I & 1.0\,M & 0 & 0 \\ C & -2x & +2x & x \\ & 2\,SeO_3(g) & \rightleftharpoons 2\,SeO_2(g) & +\ O_2(g) \\ E & 1.0-2x & 2x & x \end{array}$$

$K = 4.0 \times 10^{-12} = \frac{(2x)^2(x)}{(1.0-2x)^2} \approx \frac{4x^3}{(1.0)^2}$

1.0 mol of SeO_3 is placed in a 1.0 L container, which then reacts to reach equilibrium according to the above reaction. At equilibrium, what is the concentration of $SeO_2(g)$?

$x = [(4.0\times10^{-12})/4]^{1/3} = 1.0\times10^{-4}$ (assumption good)

a) 4.0×10^{-12} *M* b) 2.0×10^{-6} *M* ~~c) 1.0×10^{-6} *M*~~

$[SeO_2] = 2x = 2(1.0\times10^{-4}) = \boxed{2.0\times10^{-4}\,M}$

d) 1.0×10^{-12} *M* [✓ e)] 2.0×10^{-4} *M*

10 | 25
18 | 10

10. Consider the following solubilities of cerium iodate ($K_{sp} = 3.5 \times 10^{-10}$):

I. Solubility of $Ce(IO_3)_3(s)$ in water

II. Solubility of $Ce(IO_3)_3(s)$ in 0.10 *M* $Ce(NO_3)_3(aq)$

III. Solubility of $Ce(IO_3)_3(s)$ in 0.10 *M* $KIO_3(aq)$

these are common ions, they have the smallest s.

$Ce(IO_3)_3(s) \rightleftharpoons Ce^{3+} + 3IO_3^-$
E $0.10+s$ $3s$
$K_{sp} = (0.10+s)(3s)^3 \approx 0.10(27s^3)$
solving: $s = 5.1\times10^{-4}$ mol/L

$Ce(IO_3)_3(s) \rightleftharpoons Ce^{3+} + 3IO_3^-$
E s $0.10+3s$

Which of the above solubilities (I-III) is **smallest**?

$K_{sp} = 3.5\times10^{-10} = s(0.10+3s)^3 \approx s(0.10)^3$

a) I ~~b) II~~ ~~[✓ c) III]~~ ~~d) All~~ of the solublities (I-III) are the same.

solving: $\boxed{s = 3.5\times10^{-7}\text{ mol/L}}$ Solubility is smallest for III.

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11 | 26
19 | 11

11. Consider the following data:

$|\text{heat gain by ice}| = \frac{2.03\,J}{g\cdot°C} \times 18.0\,g \times 5.0°C + 1\,mol\left(\frac{6020\,J}{mol}\right) + \frac{4.18\,J}{g\cdot°C} \times 18\,g \times (T_f - 0)$

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

keep all quantities positive ✓

$|\text{heat loss by hot water}| = 4.18\,J/g\cdot°C \times 18.0\,g \times (90.0 - T_f)$

One mole of ice (18.0 g) at –5.0°C is added to one mole of water at 90.0°C. Calculate the final temperature of the mixture assuming no heat loss to the surroundings.

$|\text{heat loss of hot water}| = |\text{heat gain by ice}|$

[✓ a)] 3.8°C b) 5.9°C c) 12.6°C d) 31.4°C e) 68.5°C

$182.7\,J + 6020\,J + 75.24\,T_f = 6771.6 - 75.24\,T_f$

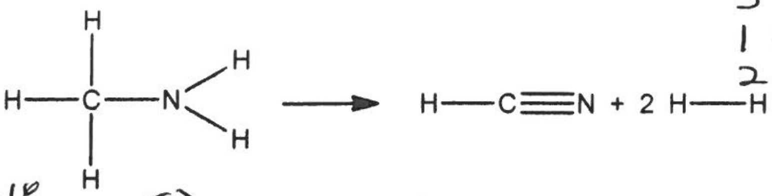
$150.48\,T_f = 568.9,\quad T_f = \frac{568.9}{150.48} = \boxed{3.8°C}$

Form
A | B
C | D

12|5
13|20

12. Hydrogen cyanide can also be produced by the reaction:

Break	Form
3 C–H	1 C–H
1 C–N	1 C≡N
2 N–H	2 H–H



$\Delta H° = 158$ kJ

$\Delta H^\circ_{rxn} = \Sigma D_{broken} - \Sigma D_{formed}$

Given the following bond energies, calculate the bond energy of the H–H bond.

	bond energy (kJ/mol)
H–C	413
H–N	391
C–N	305
C≡N	891

$158\,kJ = [3(413) + 1(305) + 2(391)] - [1(413) + 1(891) + 2x]$

Solving for $x = D_{H-H} = \boxed{432\,kJ/mol}$

a) 316 kJ b) 207 kJ [✓ c) 432 kJ] d) 158 kJ e) 734 kJ

13|6
14|21

13. The solubility of lead phosphate in 1.00 L of 0.010 *M* $Pb(NO_3)_3$ is 5.0×10^{-25} mol/L. Calculate the K_{sp} value for $Pb_3(PO_4)_2(s)$.

$Pb_3(PO_4)_2(s) \rightleftharpoons 3\,Pb^{2+} + 2\,PO_4^{3-}$

I 0.010 0

E 0.010+3s 2s

a) 1.3×10^{-48} b) 1.3×10^{-25} c) 1.0×10^{-48}

$K_{sp} = [Pb^{2+}]^3[PO_4^{3-}]^2 = (0.010+3s)^3(2s)^2$, From problem, $s = 5.0 \times 10^{-25}$

d) 1.1×10^{-268} [✓ e) 1.0×10^{-54}]

$K_{sp} = [0.010 + 3(5.0 \times 10^{-25})]^3[2(5.0 \times 10^{-25})]^2 = \boxed{1.0 \times 10^{-54} = K_{sp}}$

14|7
15|22

14. Hydrogen cyanide can be produced by the following balanced reaction:

$2\,NH_3(g) + 3\,O_2(g) + 2\,CH_4(g) \rightarrow 2\,HCN(g) + 6\,H_2O(g)$ $\Delta H° = ?$

$\Delta H^\circ_{rxn} = \Sigma \Delta H^\circ_{f,prod} - \Sigma \Delta H^\circ_{f,react}$

Given the following standard enthalpies of formation:

	$\Delta H_f°$ (kJ/mol)
$NH_3(g)$	–46
$CH_4(g)$	–75
$HCN(g)$	135
$H_2O(g)$	–242

$\Delta H° = [2(135) + 6(-242)] - [2(-46) + 3(0) + 2(-75)]$

$\Delta H° = -1182\,kJ - (-242\,kJ) = \boxed{-940\,kJ}$

calculate ΔH° for this reaction.

a) 505 kJ b) –505 kJ [✓ c) –940. kJ] d) –1424 kJ e) 940. kJ

Form

$Ag_2CrO_4(s) \rightleftharpoons 2Ag^+ + CrO_4^{2-}$ $K_{sp} = 9.0 \times 10^{-12}$

A|B
C|D

In all solutions, $[CrO_4^{2-}] = 1.0 \times 10^{-4}$ M. Let's calculate $[Ag^+]$ where $Q = K_{sp}$.

15|19
9|12

15. Consider the following five solutions:

$9.0 \times 10^{-12} = [Ag^+]^2 (1.0 \times 10^{-4})$, $[Ag^+] = 3.0 \times 10^{-4}$ mol/L

I. A solution which is 1.0×10^{-8} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
II. A solution which is 1.0×10^{-7} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
III. A solution which is 1.0×10^{-6} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
IV. A solution which is 1.0×10^{-4} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.
[✓ V.] A solution which is 1.0×10^{-3} *M* $AgNO_3$ and 1.0×10^{-4} *M* K_2CrO_4.

When $[Ag^+] > 3.0 \times 10^{-4}$ mol/L, then $Q > K_{sp}$ and a precipitate forms.

In how many of these five solutions will a precipitate of $Ag_2CrO_4(s)$ form? K_{sp} for $Ag_2CrO_4 = 9.0 \times 10^{-12}$.

Only solution V has $[Ag^+] > 3.0 \times 10^{-4}$ mol/L.

[✓ a) 1] b) 2 c) 3 d) 4

In all other solutions, $Q < K_{sp}$, so no precipitate forms.

e) 5; a precipitate will form in all five of the solutions (I-V).

16|20
10|13

16. Consider the following reaction for which K = 4.0 at some temperature.

$$2\,NF_3(g) \rightleftharpoons N_2(g) + 3\,F_2(g) \qquad K = 4.0 = \frac{[N_2][F_2]^3}{[NF_3]^2}$$

Plug in these concentrations into K expression. If K = 4.0, then they are equilibrium concentrations. When K ≠ 4.0, then the concentrations are not at equilibrium.

Which of the following sets of concentrations does **not** represent an equilibrium mixture for this reaction at this temperature?

a) $[NF_3] = 1.0$ *M*, $[N_2] = 4.0$ *M*, $[F_2] = 1.0$ *M* K=4.0
b) $[NF_3] = 2.0$ *M*, $[N_2] = 2.0$ *M*, $[F_2] = 2.0$ *M* K=4.0
c) $[NF_3] = 1.0$ *M*, $[N_2] = 0.50$ *M*, $[F_2] = 2.0$ *M* K=4.0
d) $[NF_3] = 3.0$ *M*, $[N_2] = 4.5$ *M*, $[F_2] = 2.0$ *M* K=4.0
[✓ e)] $[NF_3] = 2.0$ *M*, $[N_2] = 8.0$ *M*, $[F_2] = 1.0$ *M* gives K = 2.0; not at equilibrium

NO

$$K = \frac{(8.0)(1.0)^3}{(2.0)^2} = 2.0$$

17|21
11|14

17. Consider the generic reaction A(g) → B(g). If this reaction is run in a constant volume container, then the heat released or gained when 1 mol of A(g) is reacted is equal to:

$\Delta E = q_v$

a) ΔH [✓ b) ΔE] c) ΔV d) ΔT e) w

18|22
12|15

18. Which of the following processes (a-c) is/are endothermic?

~~a)~~ Natural gas is burned in a furnace. exothermic (heat is a product)
[✓ b)] When NH_4NO_3 dissolves in water, the solution gets colder. endothermic
~~c)~~ Two oxygen atoms react to form the O_2 molecule. exothermic (heat is a product)
~~d)~~ Both processes b and c are endothermic.
~~e)~~ All of the processes (a-c) are endothermic

$2O(g) \rightarrow O_2(g) + \text{heat}$; when O_2 bond forms, energy is released.

Form
A|B
C|D

$|\overline{Br}-\overline{Br}|$ $|\overline{F}-\overline{F}|$ $|\overline{I}-\overline{I}|$

$|\overline{Cl}-\overline{Cl}|$ $|N \equiv N|$

19|8
5|16

19. Which of the following diatomic molecules has the largest bond energy?

a) Br_2, 14e⁻
b) F_2, 14e⁻
[✓ c)] N_2, 10e⁻
d) Cl_2, 14e⁻
e) I_2, 14e⁻

the triple bond in N_2 will be a much stronger bond than the single bonds in the halogens.

20|9
6|17

20. Using the following reactions and ΔH values, calculate ΔH_{rxn} for:

$ClF + \frac{1}{2}O_2 \rightarrow \frac{1}{2}Cl_2O + \frac{1}{2}F_2O$ $\Delta H_1 = \frac{1}{2}(167.4)$
$F_2 + \frac{1}{2}O_2 \rightarrow F_2O$ $\Delta H_2 = \frac{1}{2}(-43.4)$
$\frac{1}{2}Cl_2O + \frac{3}{2}F_2O \rightarrow ClF_3 + O_2$ $\Delta H_3 = -\frac{1}{2}(341.4)$
$ClF + F_2 \rightarrow ClF_3$
$\Delta H_{TOT} = \Delta H_1 + \Delta H_2 + \Delta H_3$

$ClF(g) + F_2(g) \rightarrow ClF_3(g)$ $\Delta H_{rxn} = ?$

$2\,ClF(g) + O_2(g) \rightarrow Cl_2O(g) + F_2O(g)$ ΔH = 167.4 kJ/mol

$2\,ClF_3(g) + 2\,O_2(g) \rightarrow Cl_2O(g) + 3\,F_2O(g)$ ΔH = 341.4 kJ/mol

$2\,F_2(g) + O_2(g) \rightarrow 2\,F_2O(g)$ ΔH = –43.4 kJ/mol

$\Delta H_{TOT} = \frac{1}{2}(167.4) + \frac{1}{2}(-43.4) - \frac{1}{2}(341.4) = \boxed{-108.7\ kJ}$

a) –217.5 kJ/mol b) –130.2 kJ/mol c) 217.5 kJ.mol

[✓ d)] –108.7 kJ/mol e) 108.7 kJ/mol

21|10
7|18

21. Consider the following reaction at some temperature:

Solids are not included in K expressions.

$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ $K = [CO_2] = \frac{0.50\,mol}{2.0\,L} = 0.25$

An equilibrium mixture contains 2.0 mol $CaCO_3(s)$, 4.0 mol $CaO(s)$, and 0.50 mol $CO_2(g)$ in a 2.0 L container. If at this same temperature another equilibrium mixture contains 6.0 mol $CaCO_3(s)$ and 4.0 mol $CaO(s)$ in a 2.0 L container, determine the equilibrium concentration of $CO_2(g)$ present in this mixture.

The endpoint (equilibrium) is reached when $[CO_2] = 0.25\,M$.

[✓ a)] 0.25 *M* b) 0.33 *M* c) 0.50 *M* d) 0.75 *M* e) 1.50 *M*

the relative amounts of the solids do not matter.

22|11
8|19

22. The gas arsine, AsH_3, decomposes by the following reaction:

$$\begin{array}{lccc} I & [AsH_3]_0 & - & 0 \\ C & -2x & - & +3x \\ & 2\,AsH_3(g) & \rightleftharpoons\ 2\,As(s)\ + & 3\,H_2(g) \\ E & [AsH_3]_0 - 2x & & 3x \end{array}$$

In an experiment, pure $AsH_3(g)$ of unknown initial concentration is placed into a 1.0 L container. After equilibrium is reached, 3.0 moles of $H_2(g)$ is produced and 6.0 moles of $AsH_3(g)$ is also present. Determine the initial concentration of $AsH_3(g)$ before the reaction took place ($[AsH_3]_{initial} = ?$).

From problem, $[H_2]_e = \frac{3.0\,mol}{1.0\,L} = 3.0\,M = 3x$, $x = 1.0\,M$

a) 9.0 *M* [✓ b)] 8.0 *M* c) 3.0 *M* d) 10.5 *M* e) 15.0 *M*

$[AsH_3]_e = \frac{6.0\,mol}{L} = [AsH_3]_0 - 2(1.0)$ $\boxed{[AsH_3]_0 = 8.0\,M}$

Form
A|B
C|D

$Q = \frac{(2.0)^2}{2.0(2.0)} = 1.0$, $Q > K_{sp}$, so rxn shifts left to reach equilibrium.

23|15
1|1

23. Consider the following reaction at some constant temperature:

$$\begin{array}{lccc} I & 2.0\,M & 2.0\,M & 2.0\,M \\ C & +x & +x & -2x \\ & I_2(g) & + H_2(g) \rightleftharpoons & 2\,HI(g) \\ E & 2.0+x & 2.0+x & 2.0-2x \end{array} \qquad K = 0.75$$

Initially, 2.0 mol of I_2, 2.0 mol of H_2 and 2.0 mol of HI are mixed together in a 1.0 L rigid container and are allowed to react to reach equilibrium. Which of the following statements is **true** once this reaction reaches equilibrium?

$2.0+x > 2.0-2x$

T [✓ a)] At equilibrium, $[I_2] > [HI]$.
F b) At equilibrium, $[I_2] =$ ~~2.0 *M*.~~ $2.0+x$
F c) At equilibrium, $[I_2] > [H_2]$. Both are equal to $2.0+x$.
F d) At equilibrium, $[H_2] < 2.0$ *M*. $[H_2] = 2.0+x > 2.0$
F e) At equilibrium, $[HI] = 2.5$ *M*. $[HI]$ must be less than 2.0

24|16
2|2

24. How many of the following four statements (I-IV) about enthalpy is/are **true**?

T I. Enthalpy is a state function.
T II. Enthalpy is an extensive property.
T III. Hess's law can be applied to determine the enthalpy change for a reaction.
T IV. The standard enthalpy of formation for an element in its standard state is zero.

a) 0 (None are true.)
b) 1
c) 2
d) 3
[✓ e)] 4 [All of the statements (I-IV) are true.]

25|17
3|3

25. Consider the following salts:

$AgOH(s)$, $K_{sp} = 1.0 \times 10^{-8}$
$Cu(OH)_2(s)$, $K_{sp} = 4.0 \times 10^{-18}$
$Cr(OH)_3(s)$, $K_{sp} = 2.7 \times 10^{-19}$

$AgOH \rightleftharpoons Ag^+ + OH^-$
E $\quad s \quad s$ $\qquad K_{sp} = s^2$

$Cu(OH)_2 \rightleftharpoons Cu^{2+} + 2OH^-$
E $\quad s \quad 2s$ $\qquad K_{sp} = 4s^3$

$Cr(OH)_3 \rightleftharpoons Cr^{3+} + 3OH^-$
E $\quad s \quad 3s$ $\qquad K_{sp} = 27s^4$

Place these salts in order of **increasing** molar solubility (from least soluble to most soluble). AgOH: $s = \sqrt{1.0 \times 10^{-8}} = 1.0 \times 10^{-4}$ mol/L

a) $AgOH < Cu(OH)_2 < Cr(OH)_3$
b) $Cu(OH)_2 < AgOH < Cr(OH)_3$
$10^{-6} \quad 10^{-5} \quad 10^{-4}$
[✓ c)] $Cu(OH)_2 < Cr(OH)_3 < AgOH$
d) $Cr(OH)_3 < Cu(OH)_2 < AgOH$
e) $Cr(OH)_3 < AgOH < Cu(OH)_2$

$Cu(OH)_2$: $s = \left(\frac{4.0 \times 10^{-18}}{4.0}\right)^{1/3} = 1.0 \times 10^{-6}$ mol/L

$Cr(OH)_3$: $s = \left(\frac{2.7 \times 10^{-19}}{27}\right)^{\frac{1}{4}} = 1.0 \times 10^{-5}$ mol/L

26|18
4|4

26. Consider a theoretical salt which is composed of M^{4+} and X^- ions. If the solubility of the MX_4 salt is 1.00×10^{-2} mol/L, calculate the K_{sp} value for $MX_4(s)$.

$K_{sp} = [M^{4+}][X^-]^4 = s(4s)^4 = 256s^5$

a) 1.08×10^{-10}
[✓ b)] 2.56×10^{-8}
c) 2.70×10^{-7}
d) 4.00×10^{-6}
e) 1.00×10^{-8}

$$\begin{array}{lccc} & MX_4(s) \rightleftharpoons & M^{4+} + & 4X^- \\ I & & 0 & 0 \\ C & -s & +s & +4s \\ E & & s & 4s \end{array}$$

$K_{sp} = 256(1.0 \times 10^{-2})^5 = \boxed{2.56 \times 10^{-8}}$

Since the temperature of the surroundings increased, reaction is exothermic. Only answers b and d are possible.

Form
A | B
C | D

27 | 27
27 | 27

27. A 10.0 g-sample of a salt is dissolved in 100.0 g of water at an initial temperature of 25.0°C. After the salt dissolves, the final temperature of the solution increases to 37.2°C. Assume the heat capacity of the solution is 4.184 J/°C•g and assume no heat is lost to the calorimeter. Given the following enthalpies of dissolution and molar masses, which salt (a-e) was dissolved?

|heat loss by dissolving salt| = |heat gain by solution|, mass of solution = 110.0 g

Salt	ΔH (kJ/mol)	Molar Mass (g/mol)
a) NaCl	3.00	58.45
b) NaI	–7.53	149.9
c) KBr	19.9	119.0
[✓ d)] LiBr	–48.8	86.84
e) KCl	17.2	74.55

$\text{heat gain} = \frac{4.184\,J}{g\,^\circ C} \times 110.0\,g\,(37.2-25.0)$

$\text{heat gain} = 5615\,J$

If NaI is salt: $\Delta H = \frac{-5615\,J}{10.0\,g\,NaI\left(\frac{1\,mol}{149.9\,g}\right)} = -8.42\times10^4\,J$ salt is not NaI.

If salt is LiBr: $\Delta H = \frac{-5615\,J}{10.0\,g\,LiBr\left(\frac{1\,mol\,LiBr}{86.84\,g}\right)} = -4.88\times10^4\,J = \boxed{-48.8\,kJ/mol}$ salt is LiBr.

28 | 28
28 | 28

28. Consider the following reaction at equilibrium:

$$\underbrace{N_2(g) + 3\,H_2(g)}_{4\text{ mol gas}} \rightleftharpoons \underbrace{2\,NH_3(g)}_{2\text{ mol gas}} + \text{heat} \quad \Delta H = -92\text{ kJ}$$

Which of the following changes will shift the equilibrium to the right (to products)?

1. increasing the temperature
[✓ 2.] decreasing the temperature remove heat (a product), rxn shifts right.
3. increasing the volume of the reaction container
[✓ 4.] decreasing the volume of the reaction container, decrease volume, rxn shifts right to side with fewer moles of gas.
[✓ 5.] removing some NH_3 - remove product, rxn shifts right.
6. adding some NH_3

a) 1, 4, 6 b) 2, 3, 5 c) 1, 3, 6 [✓ d) 2, 4, 5] e) 1, 3, 5

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8\,kJ = \Delta H^\circ_{f,H_2O(l)}$

29 | 29
29 | 29

29. The standard enthalpy of formation for $H_2O(l)$ is –285.8 kJ/mol. Calculate the change in internal energy for the following process at 298 K and 1 atm:

This reaction is the reverse, so $\Delta H = q = 285.8\,kJ$.

$H_2O(l) \rightarrow H_2(g) + 1/2\,O_2(g) \qquad \Delta E^\circ = ?$

$w = -P\Delta V = -RT\Delta n = -\frac{8.3145\,J}{K\cdot mol} \times 298\,K \times 1.5 = -3717\,J = -3.72\,kJ$

[✓ a)] 282.1 kJ b) 289.5 kJ c) 285.8 kJ d) –282.1 kJ e) –289.5 kJ

$\Delta E = q + w = 285.8\,kJ - 3.72\,kJ = \boxed{282.1\,kJ}$

30 | 30
30 | 30

30. My answers for this 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