Form A/B C/D

Detail Key Exam 3 Fall 2023

CHEMISTRY 102 Exam III

Fall 2023

1/5 1/9/23

1. At some temperature, the gas arsine (AsH_3) decomposes by the following reaction:

I 3.0 mol/1.0 L

$$2\ AsH_3(g) \rightleftharpoons 2\ As(s) + 3\ H_2(g) \qquad K = 3.0 \times 10^{-6} = \frac{[H_2]^3}{[AsH_3]^2}$$

E 3.0 − 2x — 3x

; solids like As are not included in K expressions.

If 3.0 mol of AsH_3 are initially placed in a 1.0 L container, calculate the equilibrium H_2 concentration ($[H_2]_{equilibrium} = ?$). $3.0\times10^{-6} = \frac{(3x)^3}{(3.0-2x)^2} \approx \frac{27x^3}{(3.0)^2} = 3x^3$

a) $1.4 \times 10^{-2}\ M$ b) $0.14\ M$ c) $2.8 \times 10^{-3}\ M$

Solving: $x = 1.0 \times 10^{-2}$ mol/L (assumption good)

d) $1.7 \times 10^{-6}\ M$ (e) $3.0 \times 10^{-2}\ M$

$[H_2]_e = 3x = 3(0.010) = 0.030$ mol/L

2/6 2024

2. Consider the following generic reaction:

When rate forward is greater than rate reverse, products are being produced faster than reactants. This occurs when the reaction shifts right to reach equilibrium.

$$2\ A_2(g) + B(g) \rightleftharpoons BA_4(g)$$

At a specific set of conditions, it is determined that the rate of the forward reaction is greater (faster) than the rate of the reverse reaction. Which of the following statements is **false** regarding what happens as this reaction proceeds to equilibrium?

T a) The rate of the forward reaction **must** decrease as the reaction proceeds to equilibrium. At equilibrium, rates are equal, so the faster rate must decrease to reach equilibrium.

T b) When this reaction reaches equilibrium, the value of K_p will **not** equal the value of K for this reaction ($K_p \neq K$). $\Delta n = 1 - 3 = -2$; since $\Delta n \neq 0$, $K \neq K_p$.

T c) The concentration of $BA_4(g)$ **must** increase as the reaction proceeds to equilibrium. When rxn shifts right to equilibrium, [products] increase

F (d) The value of the equilibrium constant for this reaction **must** be greater than one ($K > 1$). It could be greater than 1, or less than 1, or equal to 1. From the info in the problem, can't tell anything about the value of K.

3/7 2/25

3. Consider the following exothermic reaction:

Heat is a product

$$\underbrace{CH_4(g) + 2\ O_2(g)}_{3\ mol\ gas} \rightleftharpoons \underbrace{CO_2(g) + 2\ H_2O(l)}_{1\ mol\ gas} \quad \Delta H = -891\ kJ$$

than

How many of the following five statements (I-V) are **true**?

F I. If CH_4 is added at some constant temperature, the reaction shifts right and K decreases. Rxn shifts right, but value of K is unchanged since temperature didn't change.

F II. If CO_2 is added at some constant temperature, the reaction shifts left and K decreases. Rxn shifts left, but K value is unchanged.

(T) III. If the temperature is increased, the reaction shifts left and K decreases. Since heat is a product, adding heat shifts rxn left. As a result, K decreases.

(T) IV. If the pressure is decreased by increasing the volume at some constant temperature, the reaction shifts left. Rxn shifts to side with more moles of gas present.

F V. If, in a rigid container at constant temperature, the pressure is increased by adding argon gas, the reaction shifts right. – since $[CH_4]$, $[O_2]$, $[CO_2]$, and K do not change when Ar is added, there is no effect when Ar is added.

a) 1 (b) 2 c) 3 d) 4 e) 5 (All statements are true.)

4 | 13
10 | 19

4. Consider the following reaction at some temperature:

$SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$ $K = 4.0$

I: 2.00 mol / 2.00 L; 2.00 mol / 2.00 L; 0; 0
C: $-x$; $-x$; $+x$; $+x$
E: $2.00 - x$; $2.00 - x$; x; x

$K = \frac{x^2}{(2.00-x)^2}$ ← perfect square

If 2.00 mol of $SO_2(g)$ and 2.00 mol of $NO_2(g)$ are reacted in a 2.00 L rigid container, calculate the equilibrium concentration of $NO_2(g)$.

Taking square root: $2.0 = \frac{x}{2.00-x}$, solving: $x = 2/3 = 0.6667$ mol/L

a) 0.33 *M* b) 1.41 *M* c) 1.00 *M* d) 0.50 *M* e) 0.67 *M*

$[NO_2]_e = 1.00 - x = 1.00 - 0.6667 = 0.33$ mol/L

5 | 14
11 | 20

5. Consider the following reaction which takes place in a 4.0 L rigid container at 800. K:

$N_2(g) + 3\,F_2(g) \rightleftharpoons 2\,NF_3(g)$ $K_p = ?$

$K_p = \frac{P_{NF_3}^2}{P_{N_2} \times P_{F_2}^3} = \frac{(0.48)^2}{0.021(0.063)^3} = 4.4 \times 10^4$

An equilibrium mixture contains the following partial pressures: $P_{N_2} = 0.021$ atm, $P_{F_2} = 0.063$ atm, $P_{NF_3} = 0.48$ atm. Calculate K_p for this reaction at 800. K.

a) 1.9×10^5 b) 280 c) 4.4×10^4 d) 170 e) 360

6 | 15
12 | 21

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

T I. In general, when the bond strengths of the products are stronger than the bond strengths of the reactants, an exothermic reaction results.

T II. In general, bond energies give a good estimate of ΔH for a reaction when all reactants and products are in the gas phase. In gas phase reactions, amount of intermolecular forces are minimal.

T III. In an endothermic reaction, the products have a higher potential energy than the reactants. Added heat goes to increase the potential energy as reactants are converted to products.

T IV. When determining ΔH for a reaction using bond energies, the strengths of the intermolecular forces are not taken into consideration.

a) 0 (none) b) 1 c) 2 d) 3 e) 4 (All statements are true.)

7 | 16
13 | 22

7. Consider the following data:

$P_4(s) + 6\,Cl_2(g) \rightarrow 4\,PCl_3(g)$ $\Delta H° = -1226$ kJ

$P_4(s) + 5\,O_2(g) \rightarrow P_4O_{10}(s)$ $\Delta H° = -2967$ kJ

$PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$ $\Delta H° = -84$ kJ

$PCl_3(g) + 1/2\,O_2(g) \rightarrow Cl_3PO(g)$ $\Delta H° = -286$ kJ

$P_4O_{10} \rightarrow P_4 + 5O_2$ $\Delta H_1 = 2967$

$6\,PCl_5 \rightarrow 6\,PCl_3 + 6Cl_2$ $\Delta H_2 = 6(84)$

$10\,PCl_3 + 5O_2 \rightarrow 10\,Cl_3PO$ $\Delta H_3 = 10(-286)$

$P_4 + 6\,Cl_2 \rightarrow 4\,PCl_3$ $\Delta H_4 = -1226$

$P_4O_{10} + 6PCl_5 \rightarrow 10\,Cl_3PO$

$\Delta H = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4$

Calculate ΔH° for the reaction:

$P_4O_{10}(s) + 6\,PCl_5(g) \rightarrow 10\,Cl_3PO(g)$ $\Delta H° = ?$

$\Delta H = -615$ kJ

a) −111 kJ b) −615 kJ c) −2680 kJ d) −7555 kJ e) 111 kJ

This is a cyclic process where the initial state equals the final state. When this occurs, all state functions like ΔH and ΔE equal 0. This is not the case for path functions like q and w.

8 | 21 | 1 | 14

8. Suppose you add 45 J of heat to a system and let it do 10 J of expansion work, then return the system to its initial state by cooling and compression. Which of the following statements **must** be **true** for the overall process?

F a) $\Delta H < \Delta E$ (both = 0)

F b) The quantity of work done in the compression step must exactly equal the quantity of work done by the system in the expansion step. q + w depend on the path

F c) $\Delta H = \cancel{70.\text{ J}} = 0$

F d) In the compression step, q = –45 J. q depends on path; doesn't have to the opposite of q in step 1.

T e) The change in the internal energy for this overall process is zero. $\Delta E = 0 = \Delta H$

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$ $\Delta H = -2222$ kJ

9 | 22 | 2 | 15

9. The standard enthalpy of combustion of propane, $C_3H_8(g)$, is –2222 kJ/mol. The ΔH_f° values for $CO_2(g)$ and $H_2O(l)$ are –394 kJ/mol and –286 kJ/mol, respectively. Calculate ΔH_f° for propane.

$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f,prod} - \sum \Delta H^\circ_{f,react}$; let $x = \Delta H^\circ_{f,C_3H_8(g)}$

a) 2326 kJ/mol b) 76 kJ/mol c) –104 kJ/mol

$-2222\text{ kJ} = [3(-394) + 4(-286)] - [x + 5(0)]$

d) –76 kJ/mol e) –2326 kJ/mol

Solving: $x = \Delta H^\circ_{f,C_3H_8(g)} = -104$ kJ/mol

10 | 23 | 3 | 16

10. Calculate the molar solubility of solid $Pb_3(PO_4)_2$ in a 0.10 *M* $Pb(NO_3)_2$ solution. K_{sp} for $Pb_3(PO_4)_2 = 1.0 \times 10^{-54}$.

Let x = solubility in mol/L

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

E: — 0.10 + 3x 2x (assume 0.10 + 3x ≈ 0.10)

a) 6.2×10^{-17} mol/L b) 6.2×10^{-12} mol/L c) 2.5×10^{-52} mol/L

$K_{sp} = 1.0 \times 10^{-54} = (0.10)^3(2x)^2$, solving: $x = 1.6 \times 10^{-24}$ = solubility in mol/L

d) 1.6×10^{-27} mol/L e) 1.6×10^{-26} mol/L

Assumption great!

11 | 24 | 4 | 17

11. Consider the following information:

N_2 bond energy = 941 kJ/mol

F_2 bond energy = 154 kJ/mol

$N_2(g) + 3F_2(g) \rightarrow 2NF_3(g)$ $\Delta H = -206$ kJ

Use this information to calculate the N–F bond energy.

Break: $N \equiv N$, 3 F–F Form: 6 N–F

Let x = N–F bond energy

$\Delta H = -206 = [941 + 3(154)] - (6x)$

Solving: x = N–F bond energy = 268 kJ/mol

a) 233 kJ/mol b) 268 kJ/mol c) 317 kJ/mol

d) 66 kJ/mol e) 434 kJ/mol

12 | 25 | 5 | 18

12. Calcium chloride ($CaCl_2$) is a common de-icing agent used in the winters to prevent ice formation on roads. The solubility of $CaCl_2$ is 74.5 g of $CaCl_2$ per 200.0 mL of solution. Calculate the K_{sp} value for solid calcium chloride. The molar mass of calcium chloride is 110.98 g/mol.

$CaCl_2(s) \rightleftharpoons Ca^{2+} + 2Cl^-$ $K_{sp} = [Ca^{2+}][Cl^-]^2$

E: — S 2S Let S = solubility in mol/L

correct answer

a) 151 b) 2.06×10^8 c) 3.36 d) 451 e) 45.1

$K_{sp} = S(2S)^2 = 4S^3$, $S = \frac{74.5\text{ g } CaCl_2 \left(\frac{1\text{ mol } CaCl_2}{110.98\text{ g}}\right)}{0.2000\text{ L}} = 3.36$ mol/L

$K_{sp} = 4(3.36)^3 = 151$

13 | 1 | 22 | 10

13. A 1.00 L flask was initially filled with 3.00 mol N_2 and 3.00 mol Br_2, which then reacts by the following reaction at some constant temperature:

I 3.00 mol / 1.00 L, 3.00 mol / 1.00 L, 0

$$N_2(g) + 3\,Br_2(g) \rightleftharpoons 2\,NBr_3(g) \qquad K = ? \frac{[NBr_3]^2}{[N_2][Br_2]^3}$$

E 3.00 − x, 3.00 − 3x, 2x

At equilibrium, 2.75 mol of N_2 remains. Calculate the value of K for the above reaction.

$[N_2]_e = 2.75\,mol/1.00\,L = 3.00 - x$, $x = 0.25\,M$

a) 0.43 b) 2.6×10^{-2} c) 8.0×10^{-3} d) 1.8×10^{-5} e) 6.5×10^{-2}

$[Br_2]_e = 3.00 - 3(0.25) = 2.25\,M$ $NBr_3 = 2(0.25) = 0.50\,M$

$K = \frac{(0.50)^2}{2.75(2.25)^3} = 8.0 \times 10^{-3}$

14 | 2 | 23 | 11

14. Consider the following reaction at some constant temperature:

I 0, 0, $[TeO_3]_0$

$$2\,TeO_2(g) + O_2(g) \rightleftharpoons 2\,TeO_3(g) \qquad K = 4.0$$

E 2x, x, $[TeO_3] - 2x$

Some TeO_3 is added to an otherwise empty 4.0 L container. After equilibrium is reached, 16.0 mol of TeO_3 and 8.0 mol of TeO_2 are present. How many **moles** of O_2 are present at equilibrium? $2x = [TeO_2]_e = \frac{8.0\,mol}{4.0\,L}$, $x = 1.0\,mol/L$

a) 8.0 mol b) 2.0 mol c) 1.0 mol d) 4.0 mol e) 16.0 mol

$[O_2] = x = 1.0\,mol/L$, $4.0\,L \times \frac{1.0\,mol\,O_2}{L} = 4.0\,mol\,O_2$

15 | 3 | 24 | 12

15. Consider the following four (I-IV) solutions:

when Q > Ksp a precip will form. Note that the initial concentrations have been halved since volume of solution doubled.

no I. 50.0 mL of 2×10^{-5} *M* $Cd(NO_3)_2$ is added to 50.0 mL of 2×10^{-5} *M* KOH; K_{sp} for $Cd(OH)_2(s) = 1 \times 10^{-14}$ $Q = [Cd^{2+}]_0[OH^-]_0^2 = 1\times10^{-5}(1\times10^{-5})^2 = 1.\times10^{-15} < K_{sp}$

$Cd(OH)_2(s)$ does not form.

no II. 50.0 mL of 2×10^{-5} *M* $Mn(NO_3)_2$ is added to 50.0 mL of 2×10^{-5} *M* KOH: K_{sp} for $Mn(OH)_2(s) = 1 \times 10^{-13}$ $Q = [Mn^{2+}]_0[OH^-]_0^2 = 1\times10^{-15}$

Q < Ksp so no ppt forms.

no III. 50.0 mL of 2×10^{-5} *M* $Mg(NO_3)_2$ is added to 50.0 mL of 2×10^{-5} *M* KOH; K_{sp} for $Mg(OH)_2(s) = 1 \times 10^{-9}$ $Q = [Mg^{2+}]_0[OH^-]^2 = 1\times10^{-15}$

Q < Ksp so no precipitate forms.

yes IV. 50.0 mL of 2×10^{-5} *M* $Zn(NO_3)_2$ is added to 50.0 mL of 2×10^{-5} *M* KOH; K_{sp} for $Zn(OH)_2(s) = 1 \times 10^{-16}$ $Q = [Zn^{2+}]_0[OH^-]^2 = 1\times10^{-15}$

Here Q > Ksp, so rxn shifts left and a precipitate forms.

In how many of the above four solutions (I-IV) will a precipitate form?

$Zn(OH)_2(s) \rightleftharpoons Zn^{2+} + 2OH^- \quad K_{sp} = [Zn^{2+}][OH^-]^2$

a) 0 (none) b) 1 c) 2 d) 3

e) 4 (A precipitate will form in all four of the solutions.)

$q = s \times mass \times \Delta T$, $\Delta T = \frac{q}{s \times mass}$

16 | 4 | 25 | 13

16. Two metals of equal mass with different heat capacities are subjected to the same amount of heat. Which metal undergoes the **smallest** change in temperature?

For both metals, q and mass are the same. The metal with the smallest ΔT has the largest heat capacity (s).

a) The metal with the larger heat capacity.

b) The metal with the smaller heat capacity.

c) Because they have the same mass, both undergo the same change in temperature.

d) We need to know the initial temperature of the metals.

e) The identity of the metals must be known.

$M_3Y_4(s) \rightleftharpoons 3M^{4+} + 4Y^{3-}$ $K_{sp} = [M^{4+}]^3[Y^{3-}]^4$

		$3M^{4+}$	$4Y^{3-}$
I		0	0
C	$-s$	$+3s$	$+4s$
E	—	$3s$	$4s$

17 | 8 | 14 | 5

17. Consider a theoretical ionic compound formed from M^{4+} and Y^{3-} ions. Which of the following mathematical statements correctly relates K_{sp} to the molar solubility for the ionic compound formed from M^{4+} and Y^{3-} ions? Note: s = molar solubility.

$K_{sp} = (3s)^3[4s]^4 = 6912s^7$

a) $K_{sp} = 7s^{12}$ b) ~~$K_{sp} = 256s^5$~~ c) $K_{sp} = 6912s^7$

d) ~~$K_{sp} = 9775s^{12}$~~ e) ~~$K_{sp} = 12s^7$~~

18 | 9 | 15 | 6

18. Which of the following statements correctly describes the signs of q and w for the following process at P = 1 atm and T = 298 K?

Heat is released when a gas condenses. So this is an exothermic process where q is negative.

$H_2O(g) \rightarrow H_2O(l)$ + heat

For this process, $\Delta n = 0 - 1 = -1$. The moles of gas decrease, so we have a compression. For compressions, the surroundings does work on the system (w is positive).

a) q and w are negative.
b) q is positive, w is negative.
c) q is negative, w is positive.
d) q and w are both positive.
e) q and w are both zero.

19 | 10 | 16 | 7

19. An ideal gas absorbs 100 J of heat and is simultaneously compressed by a constant external pressure of 2.00 atm from an initial volume of 10.0 L to 6.00 L. What is the change in internal energy?

$\Delta E = q + w$, $q = 100\,J$, $w = -P\Delta V = -2.00\,atm(6.00 - 10.0\,L) = 8.0\,L\cdot atm$

$8.0\,L\cdot atm\left(\frac{101.3\,J}{1\,L\cdot atm}\right) = 810\,J$

a) –910 J b) –710 J c) 710 J

$\Delta E = q + w = 100\,J + 810\,J = 910\,J$

d) 810 J e) 910 J

20 | 11 | 17 | 8

20. Solid calcium carbonate, $CaCO_3$, decomposes to form solid calcium oxide, CaO, and gaseous carbon dioxide, CO_2. What is the correct equilibrium expression for this reaction?

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ $K = [CO_2]$

solids are not included in K expressions

a) $K = \frac{[CO_2][CaO]}{[CaCO_3]}$ b) $K = \frac{1}{[CO_2]}$ c) $K = \frac{[CaCO_3]}{[CO_2][CaO]}$

d) $K = [CO_2][CaO]$ e) $K = [CO_2]$

21 | 12 | 18 | 9

21. An initial solution contains 0.10 *M* Pb^{2+}, 0.10 *M* Cu^+, 0.10 *M* Ag^+, and 0.10 *M* La^{3+}. Into the initial solution, KCl(aq) is added dropwise until a precipitate forms. Which of the following precipitates forms first as KCl(aq) is added dropwise?

A precipitate forms when $Q > K_{sp}$. For each salt, calculate $[Cl^-]$ where $Q = K_{sp}$.

a) $PbCl_2(s)$, $K_{sp} = 1 \times 10^{-5}$
$Q = 1\times10^{-5} = (0.10)[Cl^-]^2$
$[Cl^-] = 1.0\times10^{-2}$ mol/L

b) $CuCl(s)$, $K_{sp} = 1 \times 10^{-7}$
$Q = 1\times10^{-7} = 0.10[Cl^-]$
$[Cl^-] = 1\times10^{-6}$ mol/L

c) $AgCl(s)$, $K_{sp} = 1 \times 10^{-10}$
$Q = 1\times10^{-10} = 0.10[Cl^-]$
$[Cl^-] = 1\times10^{-9}$ mol/L

d) $LaCl_3(s)$, $K_{sp} = 1 \times 10^{-22}$
$Q = 1\times10^{-22} = 0.10[Cl^-]^3$
$[Cl^-] = 1\times10^{-7}$ mol/L

For all of these salts, AgCl requires the smallest amount of Cl^- to form a precip. Therefore, AgCl(s) will form first as Cl^- is added slowly.

The enthalpy of formation equation that we must solve for is:

$C_{(gr)} + 2S(s) \rightarrow CS_2(l)$ $\Delta H_{rxn} = \Delta H^\circ_{f,CS_2}$

22 | 17 | 6 | 1

22. Which expression correctly gives the value for the standard enthalpy of formation for liquid carbon disulfide (CS_2) using the following data:

Use Hess's law to solve.

$C(\text{graphite}) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ_{rxn} = -394$ kJ

$S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H^\circ_{rxn} = -296$ kJ

$CS_2(l) + 3\,O_2(g) \rightarrow CO_2(g) + 2\,SO_2(g)$ $\Delta H^\circ_{rxn} = -1072$ kJ

$\Delta H^\circ_{f,\,CS_2} = ?$

$C_{(gr)} + O_{2(g)} \rightarrow CO_{2(g)}$ $\Delta H_1 = -394$

$2S(s) + 2O_{2(g)} \rightarrow 2SO_{2(g)}$ $\Delta H_2 = 2(-296)$

$CO_{2(g)} + 2SO_{2(g)} \rightarrow CS_{2(l)} + 3O_2$ $\Delta H_3 = -(-1072)$

a) [−1072 + 394 + 296] kJ

b) [1072 − 394 − 2(296)] kJ

$C_{(g)} + 2S(s) \rightarrow CS_2(l)$ $\Delta H = 1072 - 394 - 2(296)$

answer b

c) [2(1072) − 394 − 296] kJ

d) [−1072 + 394 + 2(296)] kJ

e) [1072 − 394 − 296] kJ

$BiI_3(s) \rightleftharpoons Bi^{3+} + 3I^-$ $K_{sp} = [Bi^{3+}][I^-]^3$

I: 0 0

E: s 3s

23 | 18 | 7 | 2

23. The concentration of I− in a solution saturated with $BiI_3(s)$ is 3.9×10^{-5} mol/L. Calculate the K_{sp} value for $BiI_3(s)$.

from ICE table: $[I^-]_e = 3s = 3.9 \times 10^{-5}$, $s = 1.3 \times 10^{-5}$ mol/L

a) 6.2×10^{-17} b) 5.1×10^{-10} c) 2.3×10^{-18} d) 7.7×10^{-19} e) 5.9×10^{-14}

$K_{sp} = s(3s)^3 = 27s^4$, $K_{sp} = 27(1.3 \times 10^{-5})^4 = 7.7 \times 10^{-19}$

24 | 9 | 8 | 3

24. Consider the reaction $2\,N_2O_5(g) \rightarrow 4\,NO_2(g) + O_2(g)$ and the following data:

Substance	ΔH°_f
$N_2O_5(g)$	11.3 kJ/mol
$NO_2(g)$	33.2 kJ/mol
$O_2(g)$	?

$\Delta H^\circ = \sum \Delta H^\circ_{f,\text{prod}} - \sum \Delta H^\circ_{f,\text{react}}$

$\Delta H = [4(33.2) + 0] - [2(11.3)] = 110.2$ kJ

At constant P, $\Delta H = q$.

$\Delta E = q + w = 110.2 + w$

$w = -P\Delta V = -RT\Delta n = -8.3145\,J/K\cdot mol\,(298\,K)(5-2) = -7.43 \times 10^3\,J = -7.43$ kJ

Calculate ΔE°, the internal energy change, for this reaction at 1 atm and 25°C.

$\Delta E = 110.2\text{ kJ} + (-7.43\text{ kJ}) = 102.8$ kJ

a) 102.8 kJ b) −7.4 kJ c) 117.6 kJ d) 7.4 kJ e) 110.2 kJ

Since temp. of calorimeter decreased as salt dissolved, this is endothermic ($\Delta H > 0$).

25 | 20 | 9 | 4

25. A coffee-cup calorimeter contains 60.00 g of water at 22.0°C. A 4.25 g sample of NH_4NO_3 is added to the water in the calorimeter. After the NH_4NO_3 has dissolved, the temperature of the water is 16.9°C. Calculate the enthalpy change for the dissolution of ammonium nitrate in units of kJ/mol. Assume no heat loss to the calorimeter and assume the solution has a heat capacity of 4.18 J/°C·g. The molar mass of H_2O is 18.02 g/mol and the molar mass of NH_4NO_3 is 80.05 g/mol. I will keep all quantities positive.

heat loss to surroundings = $4.18\,J/g\cdot°C\,(64.25\,g)(22.0-16.9) = 1370$ J

a) 26 kJ/mol b) 1.4×10^3 kJ/mol c) 150 kJ/mol

heat gain by dissolution of 4.25 g NH_4NO_3 = 1370 J

d) -1.4×10^3 kJ/mol e) −24 kJ/mol

$$\Delta H = \frac{1370\,J}{4.25\,g\,NH_4NO_3\left(\frac{1\,mol\,NH_4NO_3}{80.05\,g}\right)} = 2.58 \times 10^4\,J/mol = 26\,kJ/mol$$

Form A | B / C | D

CHEMISTRY 102
Exam III

Breaking the intermolecular forces in 18 g of $H_2O(l)$ at 100°C to $H_2O(g)$ at 100°C takes the most energy (40,700 J - see below for calc); converting

Fall 2023

Consider the following H_2O data for the next two questions:

Specific heat capacity of ice = 2.03 J/°C•g; ΔH_{fusion} = 6.02 kJ/mol;

Specific heat capacity of water = 4.18 J/°C•g; $\Delta H_{vaporization}$ = 40.7 kJ/mol;

Specific heat capacity of steam = 2.02 J/°C•g

26 | 26 / 26 | 26

26. Consider the process of heating 18.0 g of ice from −100.0°C to steam at 200.0°C. Which part of this heating process requires the **largest** amount of energy?

a) Heating 18.0 g of ice at −100.0°C to ice at 0.0°C. $q = 2.03 J/g\cdot°C(18.0g)(100°C) = 3.7\times10^3 J$

b) Converting 18.0 g of ice at 0.0°C to water at 0.0°C. $q = 1mol\left(\frac{6020J}{mol}\right) = 6020J$

c) Heating 18.0 g of water at 0.0°C to water at 100.0°C. $q = 4.18J/g\cdot°C(18.0g)(100°C) = 7.5\times10^3 J$

d) Converting 18.0 g of water at 100.0°C to steam at 100.0°C. $q = 1mol\left(\frac{40.7kJ}{mol}\right) = 40.7kJ$

e) Heating 18.0 g of steam at 100.0°C to steam at 200.0°C. $q = 2.02J/g\cdot°C(18.0g)(100°C) = 3.6\times10^3 J$

Heat gain of ice = Heat loss of hot water = $4.18J/g\cdot°C(36.0g)(95.0 - T_f)$

I will keep all quantities positive.

27 | 27 / 27 | 27

27. A coffee cup calorimeter is filled with 36.0 g of water initially at 95.0°C. A 36.0 g-sample of ice at −5.0°C is then added to the calorimeter contents. Calculate the final temperature of the mixture assuming no heat loss to the surroundings or to the calorimeter.

heat gain $= 2.03J/g\cdot°C(36.0g)(5°C) + 2mol\left(\frac{6020J}{mol}\right) + 4.18J/g\cdot°C(36.0g)(T_f - 0)$

a) 50.0°C b) 2.3°C c) 92.7°C d) 6.3°C e) 0.0°C

set the 2 expressions equal to each other and solve for $T_f = 6.3°C$.

28 | 28 / 28 | 28

28. When 1.00 L of 2.40 *M* $AgNO_3$ is added to 1.00 L of 2.00 *M* K_3PO_4, a precipitate of Ag_3PO_4 forms (K_{sp} for $Ag_3PO_4 = 1.8 \times 10^{-18}$). Calculate the equilibrium Ag^+ concentration in the resulting solution ($[Ag^+]_e$ = ?).

Note: initial concentrations of ions are halved since volume doubled.

$Ag_3(PO_4)(s) \rightleftharpoons 3Ag^+ + PO_4^{3-}$

I	2.00 M	-1.20	1.00 M
C	-1.20	0	-0.40
A	0	+3s	0.60
I			0.60 M
C	-s	+3s	+s
E		3s	0.60+s

a) 4.8×10^{-7} *M* b) 1.0×10^{-18} *M* c) 6.0×10^{-17} *M*

d) 1.4×10^{-6} *M* e) 1.20 *M*

29 | 29 / 29 | 29

29. When 1.00 L of 2.40 *M* $AgNO_3$ is added to 1.00 L of 2.00 *M* K_3PO_4, a precipitate of Ag_3PO_4 forms (K_{sp} for $Ag_3PO_4 = 1.8 \times 10^{-18}$). Calculate the equilibrium PO_4^{3-} concentration in the resulting solution ($[PO_4^{3-}]_e$ = ?).

$1.8\times10^{-18} = (3s)^3(0.60+s) \approx 27s^3(0.60)$, $s = 4.8\times10^{-7} mol/L$ Assumption good.

a) 0.60 *M* b) 0.80 *M* c) 1.00 *M*

d) 0.40 *M* e) 1.20 *M*

$[Ag^+]_e = 3s = 3(4.8\times10^{-7}) = 1.4\times10^{-6} mol/L$

$[PO_4^{3-}]_e = 0.60 - 4.8\times10^{-7} = 0.60 M$

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