

1/5
19/23

1.

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

$$K = 3.0 \times 10^{-6} = \frac{[\text{H}_2]^3}{[\text{AsH}_3]^2}$$

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 ($[\text{H}_2]_{\text{equilibrium}} = ?$).

- a) $1.4 \times 10^{-2} \text{ M}$ b) 0.14 M c) $2.8 \times 10^{-3} \text{ M}$
 d) $1.7 \times 10^{-2} \text{ M}$ e) $3.0 \times 10^{-2} \text{ M}$

$$[\text{H}_2]_e = 3x = 3(0.010) = 0.030 \text{ M}$$

2/6
20/24

2.

Consider the following generic reaction:
 $2 \text{A}_2(g) + \text{B}(g) \rightleftharpoons \text{BA}_4(g)$
 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.

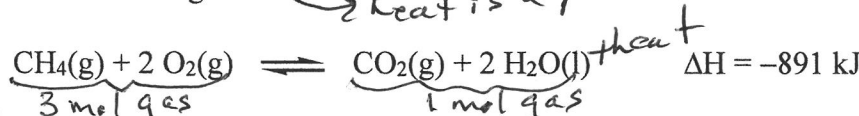
At a specific set of conditions, it is determined that the rate of the forward reaction is greater (faster) than the rate of the reverse reaction. Which of the following statements is false regarding what happens as this reaction proceeds to equilibrium?

- 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.
 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 $\Delta n \neq 0$, $K \neq K_p$.
 c) The concentration of $\text{BA}_4(g)$ **must** increase as the reaction proceeds to equilibrium. When rxn shifts right to equilibrium, products increase.
 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
21/25

3.

Consider the following exothermic reaction:

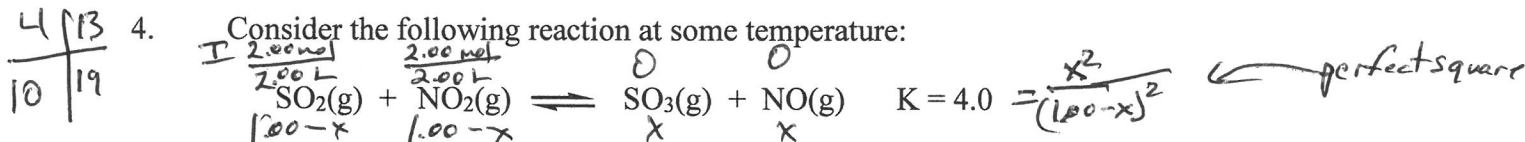


How many of the following five statements (I-V) are true?

- 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.
 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.
 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.
 IV. If the pressure is decreased by increasing the volume at some constant result, K decreases. temperature, the reaction shifts left. Rxn shifts to side with more moles of gas present.
 V. If, in a rigid container at constant temperature, the pressure is increased by adding argon gas, the reaction shifts right. - Since $[\text{CH}_4]$, $[\text{CO}_2]$, and $[\text{H}_2\text{O}]$ are not affected, there is no effect.

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

K do not change when Ar is added, there is no effect when Ar is added.



If 2.00 mol of $\text{SO}_2(\text{g})$ and 2.00 mol of $\text{NO}_2(\text{g})$ are reacted in a 2.00 L rigid container, calculate the equilibrium concentration of $\text{NO}_2(\text{g})$.

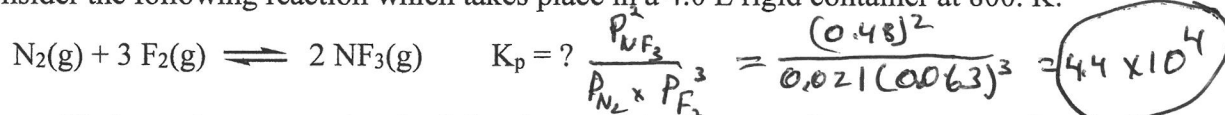
Taking square root: $2.0 = \frac{x}{1.00-x}$, solving: $x = 2/3 = 0.6667 \text{ mol/L}$

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

$[\text{NO}_2]_e = 1.00 - x = 1.00 - 0.6667 = 0.33 \text{ mol/L}$

5/14
11/20

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



An equilibrium mixture contains the following partial pressures: $P_{\text{N}_2} = 0.021 \text{ atm}$, $P_{\text{F}_2} = 0.063 \text{ atm}$, $P_{\text{NF}_3} = 0.48 \text{ atm}$. 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

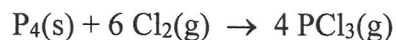
6. How many of the following four statements (I-IV) is/are **true**?

- I. In general, when the bond strengths of the products are stronger than the bond strengths of the reactants, an exothermic reaction results.
- II. In general, bond energies give a good estimate of ΔH for a reaction when all reactants and products are in the gas phase. *Handwritten note: For gas phase reactions, amount of intermolecular forces are minimal.*
- III. In an endothermic reaction, the products have a higher potential energy than the reactants. *Handwritten note: Added heat goes to increase the potential energy as reactants are converted to products.*
- IV. When determining ΔH for a reaction using bond energies, the strengths of the intermolecular forces are not taken into consideration.

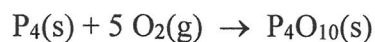
- a) 0 (none) b) 1 c) 2 d) 3 e) 4 (All statements are true.)

7/16
13/22

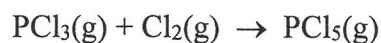
7. Consider the following data:



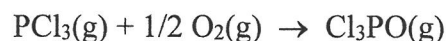
$$\Delta H^\circ = -1226 \text{ kJ}$$



$$\Delta H^\circ = -2967 \text{ kJ}$$

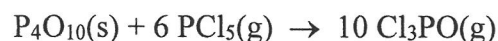


$$\Delta H^\circ = -84 \text{ kJ}$$



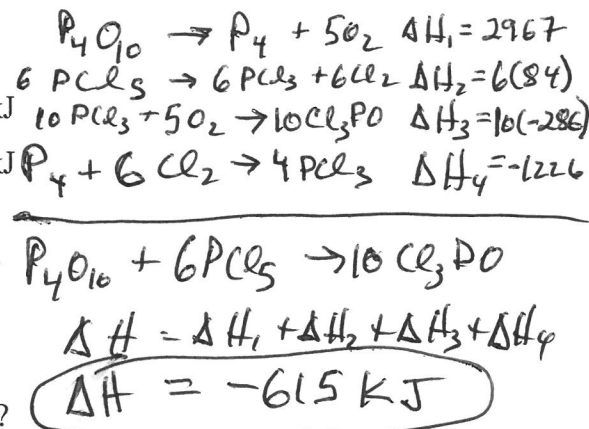
$$\Delta H^\circ = -286 \text{ kJ}$$

Calculate ΔH° for the reaction:



$$\Delta H^\circ = ?$$

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



Form

A/B
C/DCHEMISTRY 102
Exam III

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. Page 3

this is not the case for path functions like q and w .
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?

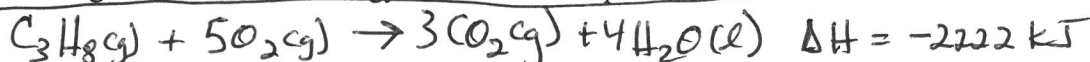
F a) $\Delta H < \Delta E$ (both = 0)

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 and w depend on the path

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

F d) In the compression step, $q = -45 \text{ J}$. q depends on path; doesn't have to be the opposite of q in step 1.

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



9. The standard enthalpy of combustion of propane, $\text{C}_3\text{H}_8\text{(g)}$, is -2222 kJ/mol . The ΔH_f° values for $\text{CO}_2\text{(g)}$ and $\text{H}_2\text{O(l)}$ are -394 kJ/mol and -286 kJ/mol , respectively. Calculate ΔH_f° for propane.

$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_{f,\text{prod}}^\circ - \sum \Delta H_{f,\text{react}}^\circ; \text{ let } x = \Delta H_{f,\text{C}_3\text{H}_8\text{(g)}}^\circ$$

a) 2326 kJ/mol

b) 76 kJ/mol

☒ c) -104 kJ/mol

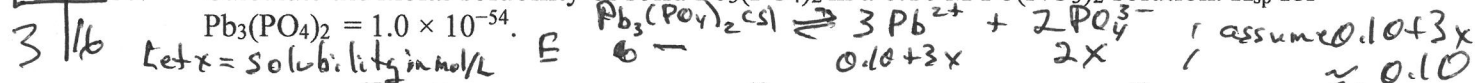
$$-2222 \text{ kJ} = [3(-394) + 4(-286)] - [x + 5(0)]$$

d) -76 kJ/mol

e) -2326 kJ/mol

$$\text{Solving: } x = \Delta H_{f,\text{C}_3\text{H}_8\text{(g)}}^\circ = -104 \text{ kJ/mol}$$

10. Calculate the molar solubility of solid $\text{Pb}_3(\text{PO}_4)_2$ in a $0.10 \text{ M Pb}(\text{NO}_3)_2$ solution. K_{sp} for $\text{Pb}_3(\text{PO}_4)_2 = 1.0 \times 10^{-54}$.



a) $6.2 \times 10^{-17} \text{ mol/L}$

b) $6.2 \times 10^{-12} \text{ mol/L}$

☒ c) $2.5 \times 10^{-52} \text{ mol/L}$

$$K_{sp} = 1.0 \times 10^{-54} = (0.10)^3 (2x)^2$$

d) $1.6 \times 10^{-27} \text{ mol/L}$

☒ e) $1.6 \times 10^{-26} \text{ mol/L}$

$$\text{Solving: } x = 1.6 \times 10^{-26} = \text{solubility in mol/L}$$

Assumption great!

Consider the following information:

N_2 bond energy = 941 kJ/mol

F_2 bond energy = 154 kJ/mol



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

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

$$\text{Solving: } x = \text{N-F bond energy} = 268 \text{ kJ/mol}$$

a) 233 kJ/mol

☒ b) 268 kJ/mol

c) 317 kJ/mol

d) 66 kJ/mol

e) 434 kJ/mol

Break
 $\text{N} \equiv \text{N}$

3 F-F

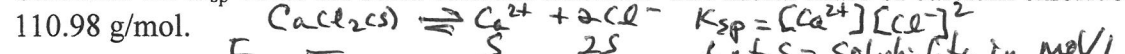
Let $x = \text{N-F bond energy}$

Form
 6N-F

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 .

12/25
5/18

Correct answer



Let $S = \text{solubility in mol/L}$

$$K_{sp} = S(2S)^2 = 4S^3$$

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

☒ a) 151

b) 2.06×10^8

☒ c) 3.36

d) 451

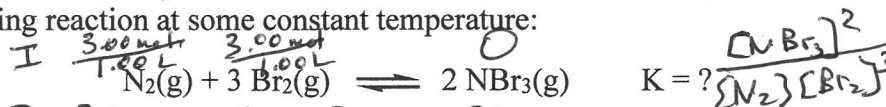
e) 45.1

$$S = \frac{74.5 \text{ g CaCl}_2}{110.98 \text{ g}} \cdot \frac{1 \text{ mol CaCl}_2}{110.98 \text{ g}} = 3.36 \text{ mol/L}$$

$$0.2000 \text{ L}$$

13/11
22/10

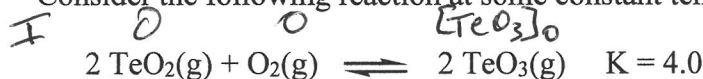
A 1.00 L flask was initially filled with 3.00 mol N₂ and 3.00 mol Br₂, which then reacts by the following reaction at some constant temperature:



At equilibrium, 2.75 mol of N₂ remains. Calculate the value of K for the above reaction.

E $3.00 - x \quad 3.00 - 3x \quad 2x$
 $[\text{N}_2]_e = 2.75 \text{ mol}/1.00 \text{ L} = 3.00 - x, x = 0.25 \text{ M}$
 $[\text{Br}_2]_e = 3.00 - 3(0.25) = 2.25 \text{ M}$
 $[\text{NBr}_3]_e = 2(0.25) = 0.50 \text{ M}$
 $K = \frac{(0.50)^2}{2.75(2.25)^3} = 8.0 \times 10^{-3}$

Consider the following reaction at some constant temperature.



E $2x \quad x \quad [\text{TeO}_3] - 2x$
 Some TeO₃ is added to an otherwise empty 4.0 L container. After equilibrium is reached, 16.0 mol of TeO₃ and 8.0 mol of TeO₂ are present. How many moles of O₂ are present at equilibrium? $2x = [\text{TeO}_2]_e = \frac{8.0 \text{ mol}}{4.0 \text{ L}}, x = 1.0 \text{ mol/L}$

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

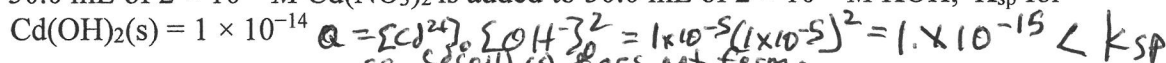
$[\text{O}_2] = x = 1.0 \text{ mol/L}, 4.0 \text{ L} \times \frac{1.0 \text{ mol O}_2}{\text{L}} = 4.0 \text{ mol O}_2$

15/3
24/12

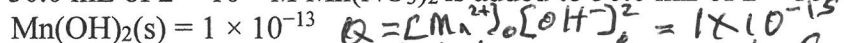
Consider the following four (I-IV) solutions:

when $Q > K_{sp}$, a precipitate will form. Note that the initial concentrations have been halved since volume of solution doubled.

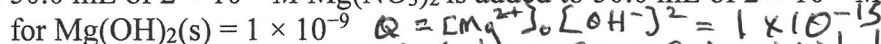
I. 50.0 mL of $2 \times 10^{-5} \text{ M Cd}(\text{NO}_3)_2$ is added to 50.0 mL of $2 \times 10^{-5} \text{ M KOH}$; K_{sp} for



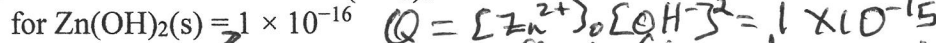
II. 50.0 mL of $2 \times 10^{-5} \text{ M Mn}(\text{NO}_3)_2$ is added to 50.0 mL of $2 \times 10^{-5} \text{ M KOH}$; K_{sp} for



III. 50.0 mL of $2 \times 10^{-5} \text{ M Mg}(\text{NO}_3)_2$ is added to 50.0 mL of $2 \times 10^{-5} \text{ M KOH}$; K_{sp} for

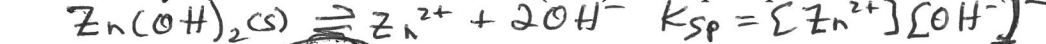


IV. 50.0 mL of $2 \times 10^{-5} \text{ M Zn}(\text{NO}_3)_2$ is added to 50.0 mL of $2 \times 10^{-5} \text{ M KOH}$; K_{sp} for



Here $Q > K_{sp}$, so rxn shifts left and a precipitate forms.

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



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

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

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

a) The metal with the larger heat capacity. Smallest ΔT has the largest heat capacity (C).

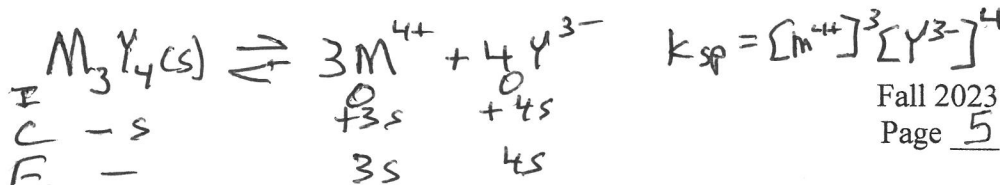
b) The metal with the smaller heat capacity.

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.

16/4
25/13



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.



- 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.

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).

19/10
16/7

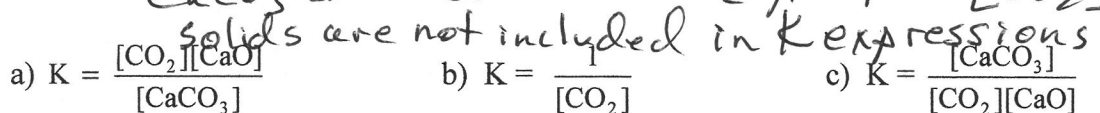
19. An ideal gas absorbs 100 J of heat and is simultaneously compressed by a constant external pressure of 2.00 atm from an initial volume of 10.0 L to 6.00 L. What is the change in internal energy?

- a) -910 J b) -710 J c) 710 J
d) 810 J e) 910 J

$$\Delta E = q + w = 100 \text{ J} + 810 \text{ J} = 910 \text{ J}$$

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?



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

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

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?

$$Q = 1 \times 10^{-5} = (0.10)[Cl^-]^2$$

$$[Cl^-] = 1.0 \times 10^{-2} \text{ mol/L}$$

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

$$Q = 1 \times 10^{-7} = 0.10[Cl^-]$$

b) $CuCl(s), K_{sp} = 1 \times 10^{-7}$
 $[Cl^-] = 1 \times 10^{-6} \text{ mol/L}$
d) $LaCl_3(s), K_{sp} = 1 \times 10^{-22}$
 $Q = 1 \times 10^{-22} = 0.10[Cl^-]^3$
 $[Cl^-] = 1 \times 10^{-7} \text{ mol/L}$

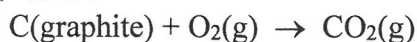
For all of these salts, $AgCl$ requires the smallest amount of Cl^- to form a precip. therefore, $AgCl(s)$ will form first as Cl^- is added slowly.

21/12
18/9

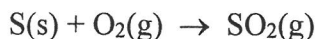
22/17
6/1

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

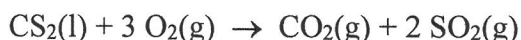
Use Hess's law to solve.



$$\Delta H_{rxn}^\circ = -394 \text{ kJ}$$

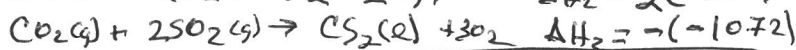
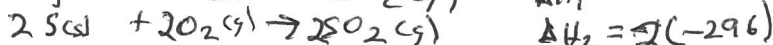
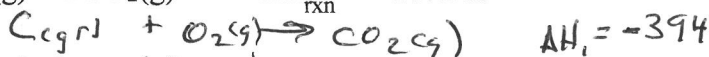


$$\Delta H_{rxn}^\circ = -296 \text{ kJ}$$



$$\Delta H_{rxn}^\circ = -1072 \text{ kJ}$$

$$\Delta H_{f, CS_2}^\circ = ?$$

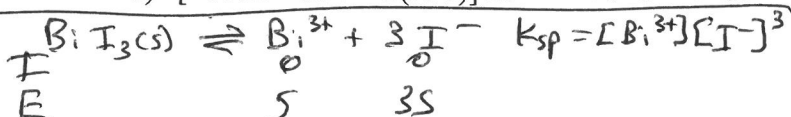


a) $[-1072 + 394 + 296] \text{ kJ}$

c) $[2(1072) - 394 - 296] \text{ kJ}$

e) $[1072 - 394 - 296] \text{ kJ}$

b) $[1072 - 394 - 2(296)] \text{ kJ}$
 $C_{(gr)} + 2 S(s) \rightarrow CS_2(l) \quad \Delta H = 1072 - 394 - 2(296)$
 d) $[-1072 + 394 + 2(296)] \text{ kJ}$ answer b

23/18
7/2

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

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

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

24/19
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

 $O_2(g)$

?

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

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

$$\Delta H = 9.$$

$$\Delta E = 9 + w = 110.2 + w$$

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

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

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

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

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 $4.18 \text{ J/}^\circ\text{C}\cdot\text{g}$. 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.

$$\text{heat loss to surroundings} = 4.18 \text{ J/g}\cdot^\circ\text{C} (64.25 \text{ g})(22.0 - 16.9) = 1370 \text{ J}$$

a) 26 kJ/mol

b) $1.4 \times 10^3 \text{ kJ/mol}$

c) 150 kJ/mol

heat gain by dissolution of $4.25 \text{ g } NH_4NO_3 = 1370 \text{ J}$

d) $-1.4 \times 10^3 \text{ kJ/mol}$

e) -24 kJ/mol

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

Breaking the intermolecular forces in 18 g of $H_2O(l)$ at $100^\circ C$ to $H_2O(g)$ at $100^\circ C$ takes the most energy (40,700 J - see below for calc) converting 18 g of $H_2O(l)$ at $100^\circ C$ to $H_2O(g)$ at $100^\circ C$ takes the most energy (40,700 J - see below for calc) Page 1

Consider the following H_2O data for the next two questions:

Specific heat capacity of ice = $2.03 \text{ J/}^\circ\text{C}\cdot\text{g}$;

$\Delta H_{\text{fusion}} = 6.02 \text{ kJ/mol}$;

Specific heat capacity of water = $4.18 \text{ J/}^\circ\text{C}\cdot\text{g}$;

$\Delta H_{\text{vaporization}} = 40.7 \text{ kJ/mol}$;

Specific heat capacity of steam = $2.02 \text{ J/}^\circ\text{C}\cdot\text{g}$

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 . $q = 2.03 \text{ J/g}\cdot^\circ\text{C} (18.0 \text{ g}) (100^\circ\text{C}) = 3.7 \times 10^3 \text{ J}$
 b) Converting 18.0 g of ice at 0.0°C to water at 0.0°C . $q = 1 \text{ mol} \left(\frac{6020 \text{ J}}{\text{mol}} \right) = 6020 \text{ J}$
 c) Heating 18.0 g of water at 0.0°C to water at 100.0°C . $q = 4.18 \text{ J/g}\cdot^\circ\text{C} (18.0 \text{ g}) (100^\circ\text{C}) = 7.5 \times 10^3 \text{ J}$
 d) Converting 18.0 g of water at 100.0°C to steam at 100.0°C . $q = 1 \text{ mol} \left(\frac{40.7 \text{ kJ}}{\text{mol}} \right) = 40.7 \text{ kJ}$
 e) Heating 18.0 g of steam at 100.0°C to steam at 200.0°C . $q = 2.02 \text{ J/g}\cdot^\circ\text{C} (18.0 \text{ g}) (100^\circ\text{C}) = 3.6 \times 10^3 \text{ J}$

Heat gain of ice = heat loss of hot water = $4.18 \text{ J/g}\cdot^\circ\text{C} (36.0 \text{ g}) (95.0 - T_f) = 3.6 \times 10^3 \text{ J}$

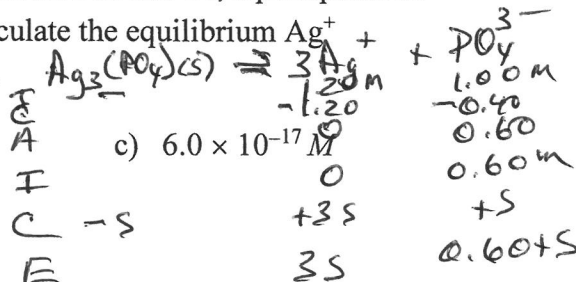
27. A coffee cup calorimeter is filled with 36.0 g of water initially at 95.0°C . A 36.0 g sample of ice at -5.0°C is then added to the calorimeter contents. Calculate the final temperature of the mixture assuming no heat loss to the surroundings or to the calorimeter.

- 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 $T_f = 6.3^\circ\text{C}$.

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 $\text{Ag}_3\text{PO}_4 = 1.8 \times 10^{-18}$). Calculate the equilibrium Ag^+ concentration in the resulting solution ($[\text{Ag}^+]_e = ?$).

- a) $4.8 \times 10^{-7} \text{ M}$ b) $1.0 \times 10^{-18} \text{ M}$ c) $6.0 \times 10^{-17} \text{ M}$ d) $1.4 \times 10^{-6} \text{ M}$ e) 1.20 M



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 $\text{Ag}_3\text{PO}_4 = 1.8 \times 10^{-18}$). Calculate the equilibrium PO_4^{3-} concentration in the resulting solution ($[\text{PO}_4^{3-}]_e = ?$).

- a) 0.60 M b) 0.80 M c) 1.00 M d) 0.40 M e) 1.20 M

$1.8 \times 10^{-18} = (3s)^3 (0.60 + s) \approx 27s^3 (0.60)$, $s = 4.8 \times 10^{-7} \text{ mol/L}$
 Assumption good.
 $[\text{Ag}^+]_e = 3s = 3(4.8 \times 10^{-7}) = 1.4 \times 10^{-6} \text{ mol/L}$
 $[\text{PO}_4^{3-}]_e = 0.60 - 4.8 \times 10^{-7} = 0.60 \text{ M}$

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