

Sample Chemistry Question (Ch. 15, 16, 17) - CH 223

Questions for Chapters Seventeen Part II, Nineteen and Twenty:

1. What is the concentration of F^- in a saturated solution of BaF_2 if $K_{sp} = 1.7 \times 10^{-6}$?
 - a. $7.5 \times 10^{-3} M$
 - b. $8.2 \times 10^{-4} M$
 - c. $1.5 \times 10^{-2} M$
 - d. $4.3 \times 10^{-7} M$
 - e. $1.5 \times 10^{-6} M$
2. For $BaSO_4$, $K_{sp} = 1.1 \times 10^{-10}$. What is the molar solubility of $BaSO_4$ in a solution which is 0.018 M in Na_2SO_4 ?
 - a. 0.018 M
 - b. $7.8 \times 10^{-5} M$
 - c. $1.1 \times 10^{-5} M$
 - d. $6.1 \times 10^{-9} M$
 - e. $1.1 \times 10^{-10} M$
3. In which of the following reactions do you expect to have the largest increase in entropy?
 - a. $I_{2(s)} \rightarrow I_{2(g)}$
 - b. $2 IF_{(g)} \rightarrow I_{2(g)} + F_{2(g)}$
 - c. $Mn_{(s)} + O_{2(g)} \rightarrow MnO_{2(s)}$
 - d. $Hg_{(l)} + S_{(s)} \rightarrow HgS_{(s)}$
 - e. $CuSO_{4(s)} + 5 H_2O_{(l)} \rightarrow CuSO_4 \cdot 5H_2O_{(s)}$
4. For a particular reaction the equilibrium constant is 1.50×10^{-2} at $370^\circ C$. ΔH° is $+16.0 kJ$. What is ΔS° for the reaction?
 - a. $-18.8 J/K$
 - b. $+18.8 J/K$
 - c. $-10.0 J/K$
 - d. $+10.0 J/K$
 - e. None of the above
5. How many electrons are transferred in the following reaction:
 $2 ClO_3^- + 12 H^+ + 10 I^- \rightarrow 5 I_2 + Cl_2 + 6 H_2O$
 - a. 12
 - b. 5
 - c. 2
 - d. 30
 - e. 10

6. If a current of 6.0 amps is passed through a solution of $\text{Ag}^+_{(\text{aq})}$ for 1.5 hours, how many grams of silver are produced?

- a. 0.60 g
- b. 36 g
- c. 0.34 g
- d. 3.0 g
- e. 1.0 g

Here are the answers to the previous questions:

1. What is the concentration of F^- in a saturated solution of BaF_2 if $K_{\text{sp}} = 1.7 \times 10^{-6}$?

- a. 7.5×10^{-3} M
- b. 8.2×10^{-4} M
- c. 1.5×10^{-2} M
- d. 4.3×10^{-7} M
- e. 1.5×10^{-6} M

Answer: Recall that the K_{sp} expression can be written as $K_{\text{sp}} = [\text{Ba}^{2+}][\text{F}^-]^2$ for BaF_2 . One mole of BaF_2 creates two moles of F^- and one mole of Ba^{2+} per mol of BaF_2 dissolved. If x represents the amount of Ba^{2+} dissolved, then $2x$ represents the F^- , and we can re-write the K_{sp} expression as $K_{\text{sp}} = 1.7 \times 10^{-6} = (x)(2x)^2 = (x)(4x^2) = 4x^3$. x represents the solubility of the BaF_2 , and here $x = (1.7 \times 10^{-6}/4)^{1/3} = 7.5 \times 10^{-3}$ M. The concentration of $[\text{F}^-]$ will be twice as much as the solubility (since there are two F^- ions per BaF_2 molecule); hence, the concentration of F^- in a saturated BaF_2 solution will be $2(7.5 \times 10^{-3} \text{ M}) = \mathbf{1.5 \times 10^{-2} \text{ M}}$, answer c.

2. For BaSO_4 , $K_{\text{sp}} = 1.1 \times 10^{-10}$. What is the molar solubility of BaSO_4 in a solution which is 0.018 M in Na_2SO_4 ?

- a. 0.018 M
- b. 7.8×10^{-5} M
- c. 1.1×10^{-5} M
- d. 6.1×10^{-9} M
- e. 1.1×10^{-10} M

Answer: If we have a 0.018 M Na_2SO_4 solution, we have, therefore, a solution with $[\text{SO}_4^{2-}] = 0.018$ M and $[\text{Na}^+] = 0.036$ M.

K_{sp} for $\text{BaSO}_4 = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$, and normally we would write $K_{\text{sp}} = x \cdot x$ since both ions dissociate in a 1:1 ratio. x here is the solubility of the BaSO_4 .

Here, however, we have a common ion present - namely sulfate. We can re-write the equation for K_{sp} as following:

$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = (x)(x+0.018)$ since the sulfate has contributions from both the BaSO_4 and the Na_2SO_4 .

Normally x is much smaller than 0.018, and $(x+0.018)$ will likely be approximately equal to 0.018 using significant figures. We can re-write the equation as:

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = (x)(x+0.018) = (x)(0.018)$$

Solving for x gives the solubility: $x = K_{sp}/0.018 = 1.1 \times 10^{-10}/0.018 = 6.1 \times 10^{-9}$

Note that x is much less than 0.018, making the assumption valid. Answer = 6.1×10^{-9} , answer **d**.

3. In which of the following reactions do you expect to have the largest increase in entropy?

- a. $\text{I}_{\text{s(s)}} \rightarrow \text{I}_{\text{2(g)}}$
- b. $2 \text{IF}_{\text{(g)}} \rightarrow \text{I}_{\text{2(g)}} + \text{F}_{\text{2(g)}}$
- c. $\text{Mn}_{\text{(s)}} + \text{O}_{\text{2(g)}} \rightarrow \text{MnO}_{\text{2(s)}}$
- d. $\text{Hg}_{\text{(l)}} + \text{S}_{\text{(s)}} \rightarrow \text{HgS}_{\text{(s)}}$
- e. $\text{CuSO}_{\text{4(s)}} + 5 \text{H}_2\text{O}_{\text{(l)}} \rightarrow \text{CuSO}_{\text{4}} \cdot 5\text{H}_2\text{O}_{\text{(s)}}$

Answer: Entropy will increase upon an increase in disorder. This can occur through several methods, including solids going to liquids and/or liquids going to gases. Another disorder enhancement comes about when a large molecule splits into many particles.

Option a shows a solid going to a gas. This would increase disorder; hence, this would increase the entropy.

Option b shows two gas molecules rearranging to make two more gas molecules. There is no disorder here except possibly in the fact that a mixed molecule (IF) is dissociating to "pure" elements (I_2 and F_2), which might lead to a decrease in entropy (less randomness in the elements than in the molecules.) Hence, this option will probably lead to a decrease in entropy instead of an increase.

Option c has a solid and gas combining to make a solid. Anytime gases become solids there is usually a decrease in entropy, not an increase. Also, two molecules are changing into one product: this is less disorder, and a lowering of entropy.

Option d is similar to option c except that a liquid is being forced to become a solid. This will also lead to a decrease in entropy. Also, two molecules are changing into one product: this is less disorder, and a lowering of entropy.

Option e is similar to option d in that a liquid is being turned into a solid. Note also how six reactant molecules are converting to one product molecule: this is a more ordered product state, which defeats entropy. Entropy will diminish here.

The only option which should increase the entropy is option **a**, the correct **answer** for this problem.

4. For a particular reaction the equilibrium constant is 1.50×10^{-2} at 370°C . ΔH° is $+16.0 \text{ kJ}$. What is ΔS° for the reaction?

- a. -18.8 J/K
- b. $+18.8 \text{ J/K}$
- c. -10.0 J/K
- d. $+10.0 \text{ J/K}$
- e. None of the above

Answer: To solve this problem we need two equations: $\Delta G = \Delta H - T\Delta S$, and $\Delta G = -RT \ln K$. Combining the equations leads to:

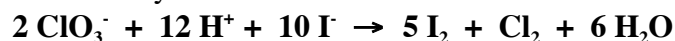
$$\Delta G = \Delta H - T\Delta S = -RT \ln K, \text{ or}$$

$$\Delta S = (RT \ln K + \Delta H)/T$$

Converting °C to K and kJ to J gives:

$$\Delta S = ((8.314 * 643 \text{ K}) \ln (1.50 * 10^{-2}) + (16.0 * 10^3 \text{ J}))/643 \text{ K} = \mathbf{-10.0 \text{ J/K}}$$
, answer **c**.

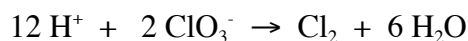
5. How many electrons are transferred in the following reaction:



- a. 12
- b. 5
- c. 2
- d. 30
- e. 10

Answer: To solve this equation, we need to break the oxidizing and reducing portions into half reactions. Cl in chlorate changes its oxidation number upon going to Cl_2 ; I also changes its oxidation number upon going to I_2 .

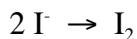
For the Cl:



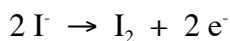
gives a balanced reaction for mass, but not for charge. The product side has no charge, the reactant side has a net +10 charge; hence, add 10 electrons to the reactant side to balance charge. The result:



For the I side,



This equation is balanced for mass, but not charge: the reactant side has a net -2 charge, and the product side has no charge. Add 2 electrons to balance the equation for both mass and charge to get:



We can now clearly see that the I is being oxidized (losing electrons) and the ClO_3^- is being reduced (gaining electrons.)

To balance the reaction, multiple the I equation by five to cancel the electrons. This results in ten net electrons being transferred from the I to the ClO_3^- , meaning that **10 electrons are being transferred**, answer **e**.

6. If a current of 6.0 amps is passed through a solution of $\text{Ag}^+_{(\text{aq})}$ for 1.5 hours, how many grams of silver are produced?

- a. 0.60 g
- b. 36 g
- c. 0.34 g
- d. 3.0 g
- e. 1.0 g

Answer: Remember that an amp equals a Coulomb per second. In addition, the Faraday (96485 C/mol e^-) will be helpful, and each Ag^+ will need one electron to become $\text{Ag}_{(\text{s})}$. Unit analysis will help to solve this problem:

$$6.0 \text{ amps} = 6.0 \text{ C/s} * (60 \text{ s/minute}) * (60 \text{ minutes/hour}) * 1.5 \text{ hours} = 32400 \text{ C delivered to the } \text{Ag}^+.$$

To find out the quantity of silver produced, use the Faraday constant and the atomic mass of silver.

$$32400 \text{ C} * (\text{mol } e^- / 96485 \text{ C}) * (\text{mol } \text{Ag}^+ / \text{mol } e^-) * (108 \text{ g Ag} / \text{mol Ag}) = \mathbf{36 \text{ g}} \text{ of silver, answer } \mathbf{b}.$$