This print-out should have 30 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

Mlib 05 3027

001 10.0 points

At a certain elevated temperature and pressure, diamond and graphite are in equilibrium. When graphite changes to diamond under these conditions

1. the change in standard molar Gibbs free energy is zero.

2. the change in molar Gibbs free energy is a minimum.

3. the molar Gibbs free energy for diamond is zero and so is that of graphite.

4. the change in molar Gibbs free energy is a maximum.

5. the change in molar Gibbs free energy is zero. **correct**

Explanation:

At equilibrium, ΔG for the process is zero.

Msci 14 0212

002 10.0 points

Consider a series of chloride salts (MCl_2). As the charge-to-size ratio of M^{2+} (decreases, increases) the hydration energy of M^{2+} (decreases, increases, does not change) in magnitude and the crystal lattice energy of MCl_2 (decreases, increases, may increase or decrease) in magnitude.

1. increases; decreases; decreases

2. decreases; increases; increases

3. increases; does not change; may increase or decrease

4. increases; increases; decreases

5. increases; increases; increases **correct**

6. decreases; decreases; increases

7. increases; decreases; increases

Explanation:

As charge-to-size ratio increases, hydration energy and crystal latice energy also increase.

Sparks vp 010

003 10.0 points

Consider two closed containers. Container X is a 2 L container that contains 0.5 L of acetone. Container Y is a 3 L container that contains 1.8 L of acetone. Both containers and contents are at 28°C. Which of the following is true?

1. You would need information about the shape of the containers to be able to answer this question.

2. The vapor pressure in container X is greater.

3. The vapor pressure in container Y is greater.

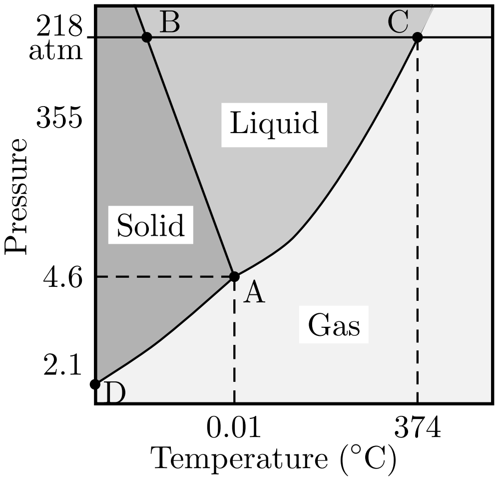
4. The vapor pressures in both containers are equal. **correct**

Explanation:

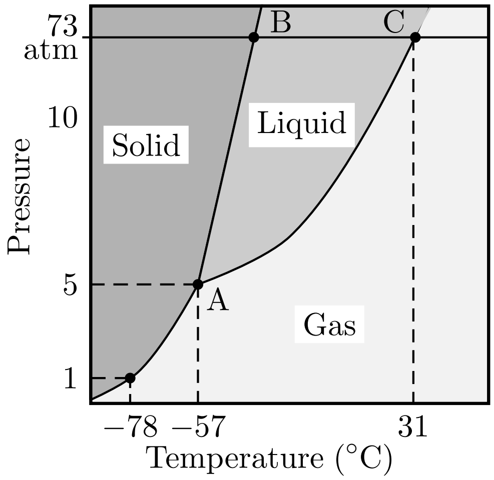
Msci 13 1304

004 10.0 points

Consider the phase diagram for water *(not to scale)*



and for carbon dioxide *(not to scale)*



Which of the following statements is NOT true?

1. Liquid water is more dense than ice.

2. Carbon dioxide cannot exist as a liquid at temperatures below −57°C.

3. Water cannot exist as a liquid at −5°C. **correct**

4. Water cannot exist as a liquid at pressures below 4.6 torr.

5. We could cause gaseous carbon dioxide to solidify at −78°C by increasing the pressure to greater than 1 atm.

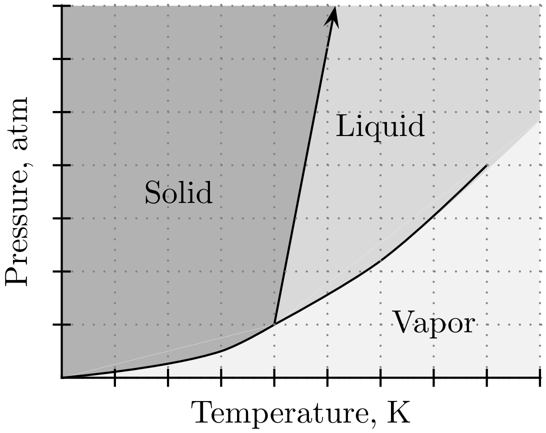
Explanation:

Starting in the solid phase of water and increasing the pressure (*i.e.*, increasing density) it becomes liquid and vice versa. The liquid section for carbon dioxide is to the right of −57°. By increasing pressure water can remain liquid at temperatures well below its standard freezing point. The liquid section for water is completely above 4.6 torr. Any point for carbon dioxide that is below −78° and above 1 atm is in the solid section.

ChemPrin3e T08 35

005 10.0 points

The phase diagram for CO_2 is given below.



The triple point is at 5.1 atm and 217 K. What happens if carbon dioxide at −50°C and 25 atm is suddenly brought to 1 atm?

1. The solid vaporizes. **correct**

2. The solid remains stable.

3. The liquid and solid are in equilibrium.

4. The solid and vapor are in equilibrium.

5. The solid melts.

Explanation:

Sparks phase change calc 001

006 10.0 points

How much energy is released when 150 g water at 52°C freezes and forms ice with a temperature of −14°C? The specific heat of water in the liquid state is 4.18 J/g°C, in the solid state is 2.09 J/g°C, and in the gaseous state is 2.03 J/g°C. The heat of fusion is 334 J/g and the heat of vaporization is 2260 J/g.

1. 93 kJ

2. 45 kJ

3. 22 kJ

4. 102 kJ

5. 37 kJ

6. 87 kJ **correct**

Explanation:

Msci 14 0707

007 10.0 points

The solubility of a gas in water increases with

1. increase of pressure or decrease of temperature. **correct**

2. decrease of pressure or decrease of temperature.

3. the effect of temperature and pressure depend on the identity of the gas.

4. decrease of pressure or increase of temperature.

5. increase of pressure or increase of temperature.

Explanation:

Henry's Law states that as the pressure of the gas above a solution surface increases, the concentration of the gas increases. In other words, it becomes more soluble. Conversely, solubility decreases with pressure.

Solubility of a gas also increases when temperature is decreased. Gas dissolving in water is exothermic, therefore according to LeChatelier's principle, if you add more heat (increase temperature), the gas is going to bubble out (be less soluble). Inversely, if you decrease temperature, the gas is going to be more soluble.

Mlib 04 4055

008 10.0 points

Which of the following alcohols would be the least miscible with water?

1. pentanol ($CH_3CH_2CH_2CH_2CH_2OH$)

2. hexanol ($CH_3CH_2CH_2CH_2CH_2CH_2OH$) **correct**

3. propanol ($CH_3CH_2CH_2OH$)

4. methanol (CH_3OH)

5. ethanol (CH_3CH_2OH)

Explanation:

The polar OH group is miscible with water but as the nonpolar hydrocarbon chain lengthens, solubility decreases.

Msci 14 0904

009 10.0 points

The vapor pressure of pure CH_2Cl_2 (molecular weight = 85 g/mol) is 133 torr at 0°C and the vapor pressure of pure CH_2Br_2 (molecular weight 174 g/mol) is 11 torr at the same temperature. What is the total vapor pressure of a solution prepared from equal masses of these two substances?

1. vapor pressure = 93 torr **correct**

2. vapor pressure = 124 torr

3. vapor pressure = 72 torr

4. vapor pressure = 89 torr

5. vapor pressure = 144 torr

6. vapor pressure = 105 torr

Explanation:

For CH_2Cl_2,

$P^0 = 133$ torr MW = 85 g/mol

For CH_2Br_2,

$P^0 = 11$ torr MW = 174 g/mol

This is a combination of Raoult's Law and Dalton's Law of Partial Pressures. The answer does not depend on what the masses are, as long as they are equal. You can choose any mass you like, but to speed up calculations, it is convenient to choose the mass the same as one of the molecular weights given, so that the number of moles for one of the components is exactly ONE.

So, for argument's sake, choose 85 g to be the mass of each of the components. That way you have:

$$(85\text{ g }CH_2Cl_2)\left(\frac{1\text{ mol }CH_2Cl_2}{85\text{ g }CH_2Cl_2}\right) = 1.0\text{ mol }CH_2Cl_2$$

Now calculate the moles of the other component.

$$(85\text{ g }CH_2Br_2)\left(\frac{1\text{ mol }CH_2Br_2}{174\text{ g }CH_2Br_2}\right) = 0.49\text{ mol }CH_2Br_2$$

Once you have the two values for moles you can calculate the mole fraction of each component.

$$n_{\text{total}} = 1.0 + 0.49 = 1.49\text{ mol}$$

$$X_{CH_2Cl_2} = \frac{1.0\text{ mol}}{1.49\text{ mol}} = 0.67$$
$$X_{CH_2Br_2} = \frac{0.49\text{ mol}}{1.49\text{ mol}} = 0.33$$

Then use those values in Raoult's Law to get the vapor pressure for each component. Raoult's Law states that:

$$P_A = X_A\, P_A^0$$

$$P_{CH_2Cl_2} = (0.67)(133\text{ torr}) = 89\text{ torr}$$
$$P_{CH_2Br_2} = (0.33)(11\text{ torr}) = 3.6\text{ torr}$$

Add the two together to get the total vapor pressure (Dalton's Law).

$$\begin{aligned}P_{\text{total}} &= P_A + P_B + \cdots\\ &= 89\text{ torr} + 3.6\text{ torr} = 93\text{ torr}\end{aligned}$$

You might want to check to see that you get the same answer no matter what value you assume as the equal masses of the two components. As an additional challenge, can you solve this problem WITHOUT assuming a definite number of grams, by setting the mass of each component equal to an algebraic variable?

Msci 13 0915

010 10.0 points

The heat of vaporization of water is 9.73 kcal/mol. At what pressure (in torr) would pure water boil at 92°C?

1. 760 torr

2. 570 torr **correct**

3. 428 torr

4. 1140 torr

5. 285 torr

Explanation:

Here we use the Clausius-Clapeyron equation:

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{\text{vap}}}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

Don't forget to convert from °C to K:

$$\text{K} = {}^\circ\text{C} + 273$$

We also need to convert the unit of ΔH_{vap} into cal/mol, so it will match the unit of R.

Here you should remember that the normal boiling point of water occurs at 760 torr (1 atm) at 100°C.

$T_1 = 100^\circ\text{C} = 373\text{ K}$

$T_2 = 92^\circ\text{C} = 365\text{ K}$

$R = 1.987\text{ cal/mol}\cdot\text{K}$

$$\begin{aligned}\Delta H_{\text{vap}} &= 9.73\text{ kcal/mol}\left(\frac{1000\text{ cal}}{1\text{ kcal}}\right)\\ &= 9730\text{ cal/mol}\end{aligned}$$

Substituting these values into the Clausius-Clapeyron equation and solving for P_2, we have

$$\begin{aligned}\ln\left(\frac{P_2}{760\text{ torr}}\right) &= \frac{\frac{9730\text{ cal}}{\text{mol}}}{\frac{1.987\text{ cal}}{\text{mol}\cdot\text{K}}} \\ &\quad\times\left(\frac{1}{373\text{ K}}-\frac{1}{365\text{ K}}\right) \\ &= -0.287752 \\ \frac{P_2}{760\text{ torr}} &= e^{-0.287752} \\ P_2 &= (760\text{ torr})e^{-0.287752} \\ P_2 &= 569.96\text{ torr}\end{aligned}$$

ChemPrin3e T09 66

011 10.0 points

For the decomposition of ammonia to nitrogen and hydrogen, the equilibrium constant is 1.47×10^{-6} at 298 K. Calculate the temperature at which $K = 1.00$. For this reaction, $\Delta H^\circ = 92.38\text{ kJ}\cdot\text{mol}^{-1}$.

1. 466 K **correct**

2. 353 K

3. 193 K

4. 219 K

5. 492 K

Explanation:

ChemPrin3e T08 72

012 10.0 points

An animal cell assumes its normal volume when it is placed in a solution with a total solute molarity of 0.3 M. If the cell is placed in a solution with a total solute molarity of 0.1 M,

1. no movement of water takes place.

2. water enters the cell, causing expansion. **correct**

3. water leaves the cell, causing contraction.

4. the escaping tendency of water in the cell increases.

Explanation:

Msci 14 1111B

013 10.0 points

What is the boiling point elevation of a solution of Na_2SO_4 (142.1 g/mol) made by dissolving 5.00 g of Na_2SO_4 in 250 grams of water? Note that $K_b = 0.512°\text{C}/m$. Assume 100 percent dissociation.

1. 0.072°C

2. 0.108°C

3. 0.216°C **correct**

4. 0.018°C

5. 0.141°C

6. 0.363°C

Explanation:

When Na_2SO_4 dissolves it dissociates into two Na^+ ions and one SO_4^{2-} ion, which is a tripling of the stated molality. Use triple the stated molality in the formula.

Msci 17 0203

014 10.0 points

Consider the reaction

$$2\,HgO(s) \rightleftharpoons 2\,Hg(\ell) + O_2(g)\,.$$

What is the form of the equilibrium constant K_c for the reaction?

1. None of the other answers is correct.

2. $K_c = \dfrac{[O_2]}{[HgO]^2}$

3. $K_c = \dfrac{[Hg]^2\,[O_2]}{[HgO]^2}$

4. $K_c = [Hg]^2\,[O_2]$

5. $K_c = [O_2]$ **correct**

Explanation:

Solids and liquids are not included in the K expression.

Concept DeltaG and K W

015 10.0 points

If $\Delta G^\circ_{\rm rxn}$ is positive, then the forward reaction is (spontaneous / nonspontaneous) and K is (less / greater) than one.

1. spontaneous, less

2. None of these; ΔG is not directly related to K.

3. nonspontaneous; greater

4. nonspontaneous; less **correct**

5. spontaneous, greater

Explanation:

A positive $\Delta G^\circ_{\rm rxn}$ (standard reaction free energy) denotes an endothermic reation, which is nonspontaneous. Also, $\Delta G^\circ_{\rm rxn} = -RT\ \ln K$, so a positive $\Delta G^\circ_{\rm rxn}$ would result in a K that is between the values of zero and one.

Msci 17 0801

016 10.0 points

Given that CO_2(g) reacts with C(s) via the reaction

$$\mathrm{C(s) + CO_2(g) \rightleftharpoons 2\,CO(g)}$$

and $K_{\rm p} = 1.90$ atm, what is the equilibrium partial pressure of CO_2 if 1.00 atm of CO_2 is placed in a vessel with PURE SOLID CARBON? (Note: There was no CO initially.)

1. 0.55 atm

2. 0.51 atm **correct**

3. 0.60 atm

4. 0.43 atm

5. 0.85 atm

Explanation:

$K_{\rm p} = 1.9$ atm

	C(s) +	CO_2(g) $\rightleftharpoons$	2 CO(g)
ini, atm		1	0
Δ, atm		$-x$	$2\,x$
eq, atm		$1-x$	$2\,x$

$$\begin{aligned} K_{\rm p} &= \frac{P^2_{\rm CO}}{P_{\rm CO_2}} = 1.9 \\ \frac{(2x)^2}{1-x} &= 1.9 \\ 4x^2 + 1.9x - 1.9 &= 0 \end{aligned}$$

Therefore

$$x = \frac{-1.9 \pm \sqrt{(1.9)^2 + 16(1.9)}}{8} = 0.491476$$

$$P_{\rm CO_2} = 1 - x = 0.508524 \text{ atm}$$

Msci 17 0501

017 10.0 points

The equilibrium constant for thermal dissociation of F_2

$$\mathrm{F_2(g) \rightleftharpoons 2\,F\,(g)}$$

is 0.300. If initially 1.00 mol F_2 is placed in a 1.00 L container, which of the following is the correct number of moles of F_2 that have dissociated at equilibrium?

1. 0.130 mol

2. 0.956 mol

3. 0.176 mol

4. 0.548 mol

5. 0.474 mol

6. 0.239 mol **correct**

7. 0.213 mol

8. 0.418 mol

Explanation:

$K_c = 0.300$ $[F_2]_{ini} = 1$ M

	$F_2(g) \rightleftharpoons$	$2\,F(g)$
Initial, M	1	0
Change, M	$-x$	$+2\,x$
Equilibrium, M	$1-x$	$2\,x$

$$K_c = \frac{[F]^2}{[F_2]}$$

$$0.3 = \frac{(2x)^2}{1-x} = \frac{4x^2}{1-x}$$

Using the quadratic equation, $x = 0.239$. If you substitute x back into the equilibrium concentration $(1-x)$ of F_2, you get 0.761. That means that there are still 0.761 moles of F_2 at equilibrium; therefore, only 0.239 dissociated.

Msci 17 0614

018 10.0 points

A 10.0 L vessel contains 0.0015 mole CO_2 and 0.10 mole CO. If a small amount of carbon is added to this vessel and the temperature is raised to 1000°C

$$CO_2(g) + C(s) \rightleftharpoons 2\,CO(g)\,,$$

will more CO form? The value of K_c for this reaction is 1.17 at 1000°C. Assume that the volume of the gas in the vessel is 10.0 L.

1. Yes, the rate of the forward reaction will increase to produce more CO. **correct**

2. No, the rate of the reverse reaction will increase to produce more CO_2.

3. Unable to determine this from the data provided.

Explanation:

$$[CO] = \frac{0.1\text{ mol}}{10\text{ L}} \qquad [CO_2] = \frac{0.0015\text{ mol}}{10\text{ L}}$$

Carbon, being a solid, has no effect on equilibrium.

$$[Q] = \frac{[CO]^2}{[CO_2]} = \frac{\left(\frac{0.1}{10.0}\text{ M}\right)^2}{\left(\frac{0.0015}{10.0}\text{ M}\right)} = 0.666667 < K_c = 1.17$$

Therefore equilibrium will shift to the right.

Rxn Anal 09 75

019 10.0 points

Which part(s) of the reaction

$$2\,O_3(g) \rightleftharpoons 3\,O_2(g)$$

will be favored by an increase in the total pressure (resulting in compression)?

1. Neither is favored.

2. Unable to determine

3. products

4. reactants **correct**

Explanation:

By Le Chatelier's principle, if the forward reaction leads to a net increase in the number of moles of gas, then applying pressure will shift the reaction toward the reactants in order to remove the stress applied by increasing the pressure.

Msci 17 1203

020 10.0 points

Given the reaction

$$2\,ICl(s) \rightleftharpoons I_2(s) + Cl_2(g)$$

and the thermodynamic data

Species	ΔH_f kJ/mol	S^0 J/mol·K
ICl(s)	17.78	242.4
$I_2(s)$	0.0	116.1
$Cl_2(g)$	0.0	223.0

calculate K_p at 100°C.

1. 0.57

2. 7.562

3. 0.023

4. 0.75

5. 0.0023 **correct**

Explanation:
$T = 100°\text{C} + 273\text{ K} = 373\text{ K}$

$$\begin{aligned}\Delta H^0 &= 0 + 0 + 0 - (2)(17.78)\\ &= -35.56\text{ kJ/mol}\end{aligned}$$

$$\begin{aligned}\Delta S^0 &= 116.1 + 223.0 - (2)(242.4)\\ &= -145.7\text{ J/K}\end{aligned}$$

$$\begin{aligned}\Delta G^0 &= \Delta H^0 - T\,\Delta S^0\\ &= -35.56 - (373)(-0.1457)\\ &= 18.7861\text{ kJ/mol}\end{aligned}$$

$$\begin{aligned}\Delta G^0 &= -RT\,\ln K\\ K &= e^{-\Delta G^0/(RT)}\\ &= \exp\left[\frac{-18786.1\text{ J/mol}}{(8.314\text{ J/mol}\cdot\text{K})\,(373\text{ K})}\right]\\ &= 0.0023\end{aligned}$$

Mlib 07 0165

021 10.0 points

Which of the following expressions correctly describes the relationship between $[H_3O^+]$ and $[OH^-]$ in any aqueous solution at 25°C?

1. $[H_3O^+][OH^-] = 10^{-14}$ **correct**

2. $\dfrac{[OH^-]}{[H_3O^+]} = 10^{-14}$

3. $\dfrac{[H_3O^+]}{[OH^-]} = 10^{-14}$

4. $[H_3O^+] - [OH^-] = 14$

5. $[H_3O^+][OH^-] = 14$

Explanation:
At 25°C,

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

DAL 02 0303

022 10.0 points

Which of the following statements is true with respect to the autodissociation of water when sipping a glass of ice water?

I. pH = pOH = 7
II. pH < 7
III. pH = pOH
IV. pH > 7

1. I and III only

2. IV only

3. II only

4. III and IV only **correct**

Explanation:
Since water autodissociates, $[H^+] = [OH^-]$. For any given temperature, the pH of pure water is defined as neutral and pH = pOH. At 25°C neutral pH = 7. K_a will be smaller than 1×10^{-7} at 0°C because water autodissociates less than 25°C. pH will therefore be greater than 7 at 0°C.

Msci 18 0453

023 10.0 points

A typical fresh egg white will have a pH of 7.80. This corresponds to

1. $[H_3O^+]$ of 1.6×10^{-8} M; $[OH^-]$ of 6.3×10^{-7} M **correct** .

2. $[H_3O^+]$ of 8.5×10^{-7} M; $[OH^-]$ of 5.5×10^{-7} M .

3. $[H_3O^+]$ of 8.0×10^{-7} M; $[OH^-]$ of 1.3×10^{-8} M .

4. $[H_3O^+]$ of 7.0×10^{-8} M; $[OH^-]$ of 1.4×10^{-7} M .

5. $[H_3O^+]$ of 3.0×10^{-8} M; $[OH^-]$ of 3.3×10^{-7} M .

Explanation:

Acid Strength 10 23

024 10.0 points

Arrange the acids

I) hydrogen selenate ion ($HSeO_4^-$), $pK_a = 1.92$;

II) phosphorous acid (H_3PO_3), $pK_{a1} = 2.00$;

III) phosphoric acid (H_3PO_4), $pK_a = 2.12$;

IV) selenous acid (H_2SeO_3), $pK_a = 2.46$;

in *decreasing* order of strengths.

1. III, IV, II, I

2. IV, III, II, I

3. II, IV, III, I

4. None of these

5. I, III, II, IV

6. I, II, IV, III

7. I, II, III, IV **correct**

8. IV, II, III, I

9. Cannot be determined

10. I, III, IV, II

Explanation:

The stronger the acid, the higher the K_a value and the lower the pK_a value:

$$pK_a = -\log(K_a)$$
$$K_a = 10^{-pK_a}$$

I. For the hydrogen selenate ion,

$$K_a = 10^{-1.92} = 0.0120226$$

II. For phosphorous acid,

$$K_a = 10^{-2.00} = 0.01$$

III. For phosphoric acid,

$$K_a = 10^{-2.12} = 0.00758578$$

IV. For selenous acid,

$$K_a = 10^{-2.46} = 0.00346737$$

$$HSeO_4^- > H_3PO_3 > H_3PO_4 > H_2SeO_3$$

Msci 18 0387

025 10.0 points

The pH of a solution of hydrochloric acid is 2.80. What is the molarity of the acid?

1. 6.3×10^{-2} M

2. 6.3×10^{-3} M

3. 4.2×10^{-3} M

4. 4.2×10^{-2} M

5. 1.6×10^{-3} M **correct**

Explanation:

pH = 2.80

$$[H^+] = 10^{-pH} = 10^{-2.80} = 1.6 \times 10^{-3} \text{ M}$$

Since HCl is a strong acid, it is completely dissociated, so its concentration here is also 1.6×10^{-3} M.

Msci 18 0402

026 10.0 points

What is the H^+ ion concentration in a 0.50 mol/L solution of a weak base that has an ionization constant (K_b) of 2.0×10^{-8}?

1. 1.0×10^{-8} mol/L

2. 2.0×10^{-10} mol/L

3. 1.0×10^{-10} mol/L **correct**

4. 8.0×10^{-16} mol/L

5. 1.0×10^{-4} mol/L

Explanation:

Msci 18 0408

027 10.0 points

0.50 moles of HCN are added to a liter of water. What is the pH? (K_a of HCN is 4.0×10^{-10})

1. 9.40

2. 4.69

3. 4.85 **correct**

4. 4.35

5. 5.35

Explanation:

HCN is *not* a strong acid so $[H^+]$ will *not* be 0.5 M. To figure it out, we must look at the K_a.

$$HCN \longrightarrow H^+ + CN^-$$

Initial	0.5	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.5 - x$ (but x is negligible)	x	x

x is negligible compared to 0.5 in this situation because the K_a is so small (which means the reaction isn't going to go very much. We leave in the other two x's in because they are not negligible compared to zero:

$$K_a = \frac{[CN^-][H^+]}{[HCN]}$$

$$4 \times 10^{-10} = \frac{x^2}{0.5}$$

$$x = 1.4 \times 10^{-5} = [H^+]$$

$$pH = -\log\left(1.4 \times 10^{-5}\right) = 4.85$$

Brodbelt 04 05

028 10.0 points

Everyone should recognize that

__?__ is a strong acid,
__?__ is a weak acid,
__?__ is a strong base, and
__?__ is a salt.

1. HCl; HCN; $Cu(OH)_2$; LiCl

2. HNO_3; HCN; KOH; CO

3. HNO_3; HCN; KOH; LiCl **correct**

4. CH_3COOH; HF; KOH; KBr

5. HCl; HNO_3; NaOH; LiCl

Explanation:

Of the answer choices, hydrochloric acid (HCl) and nitric acid (HNO_3) are strong acids because they dissociate completely in water. Acetic acid (CH_3COOH) and HCN are weak acids due to their partial dissociation. Sodium hydroxide (NaOH) potassium hydroxide (KOH) are strong bases, dissociating completely to yield hydroxide (OH^-). Lithium chloride (LiCl) and potassium bromide (KBr) are salts, compounds composed of a cation (metal) other than H^+ and an anion (non-metal) other than OH^- or O^{2-}.

DAL 0301 10

029 10.0 points

Each of the following samples was placed in 1 liter of water.

I) 0.6 mol NaOH
II) 0.7 mol KCl
III) 0.5 mol Na_2NO_3
IV) 1 mol of sugar

Rank the solutions that are made in terms of increasing order of boiling point elevation. (Remember your solubility rules.)

1. II, IV, I, III

2. IV, I, II, III **correct**

3. III, I, II, IV

4. III, I, IV, II

5. IV, II, I, III

Explanation:

Colligative properties of a solution depend on the number of solute particles in solution, not the type.

I) will have 1.2 mol of particles because NaOH completely dissociates in water.

II) will have 1.4 mol of particles because KCl completely dissociates in water.

III) will have 1.5 mol of particles because Na_2CO_3 completely dissociates in water.

IV) will have 1 mol of sugar in solution because it is soluble but does not dissociate.

The higher the concentration, the greater the boiling point elevation, so

$$\mathrm{IV} > \mathrm{I} > \mathrm{II} > \mathrm{III}\,.$$

Msci 17 0641

030 10.0 points

Consider an aqueous solution with these equilibria present:

$$\mathrm{AgBr(s)} \rightleftharpoons \mathrm{Ag^+(aq)} + \mathrm{Br^-(aq)}$$

$$\mathrm{Ag^+(aq)} + 2\,\mathrm{S_2O_3^{2-}(aq)} \rightleftharpoons \mathrm{Ag(S_2O_3)_2^{3-}(aq)}$$

$$\mathrm{S_2O_3^{2-}(aq)} + \mathrm{H_3O^+(aq)} \rightleftharpoons \mathrm{HS_2O_3^-(aq)} + \mathrm{H_2O}(\ell)$$

Which suggested change INCREASES the solubility of AgBr(s) in the solution?

1. add the soluble salt KBr

2. add the soluble salt $Na_2S_2O_3$ **correct**

3. add concentrated HCl(aq)

4. None of these increases the solubility.

5. add more AgBr(s)

Explanation:

More AgBr will dissolve if Ag^+ and/or Br^- is removed from solution. Adding AgBr(s) has no effect on the equilibrium. Added KBr increases $[Br^-]$. Concentrated HCl can precipitate AgCl but since it is more soluble than AgBr this will not affect the solubility of AgBr. Added $S_2O_3^{2-}$ (from $Na_2S_2O_3$) removes Ag^+ by precipitating $Ag_2S_2O_3$.