

## JANUARY 2000 PROVINCIAL EXAMINATION Multiple Choice Explanations

Note: To navigate to a question go to “Insert” menu above and choose “Bookmark” or press “Control + Shift + F5”. When the “Bookmark” box appears, first click the “location” button then either double click on the question or select the question and press “Go To”. It will take you to the question you want. To get rid of the box, click the “X” on the top right corner or click the “Close” box. When you want to go to another question, press “Control + Shift + F5” again to get the “Bookmark” box.

Another note: The questions ARE a little fuzzy. It is not a problem with your eyes!

1. Which of the following is most likely to have the **greatest** reaction rate at room temperature?
- A.  $2\text{H}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{H}_2\text{O}_{(l)}$
- B.  $2\text{Ag}^+_{(aq)} + \text{CrO}_4^{2-}_{(aq)} \rightarrow \text{Ag}_2\text{CrO}_{4(s)}$
- C.  $\text{Pb}_{(s)} + 2\text{HCl}_{(aq)} \rightarrow \text{PbCl}_{2(aq)} + \text{H}_{2(g)}$
- D.  $\text{CH}_{4(g)} + 2\text{O}_{2(g)} \rightarrow \text{CO}_{2(g)} + 2\text{H}_2\text{O}_{(g)}$

The answer is **B**



Explanation:

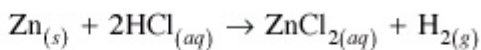
Reaction “A” looks simple but the diatomic bonds in  $\text{H}_2$  &  $\text{I}_2$  must be broken before the reaction can take place. This means there is a higher energy requirement (activation energy), so the reaction is fairly slow at room temp. The “at room temp” is important because this reaction could be very fast at a high temp.

Reaction “B” is the answer because there are no bonds broken, the  $\text{Ag}^+$  and the  $\text{CrO}_4^{2-}$  simply join together (precipitation reactions are always fast!)

Reaction “C” is slow because of the solid (it is heterogeneous). Reactions can take place only on the surface of the lead.

Reaction “D” is slow because covalent C-H bonds in  $\text{CH}_4$  must be broken. This requires more energy.

2. Consider the following reaction involving 1.0 g of powdered zinc:



| Trial | Temperature (°C) | Concentration of HCl |
|-------|------------------|----------------------|
| 1     | 40               | 3.0                  |
| 2     | 20               | 3.0                  |
| 3     | 40               | 6.0                  |

The rates, in order of fastest to slowest, are

- A. 1, 2, 3
- B. 2, 1, 3
- C. 3, 1, 2
- D. 3, 2, 1

The answer is **C**



Explanation:

The higher the Temp. and the higher the [HCl], the faster the rate, so Rx. 3 is fastest (high temp. and high [HCl]), followed by Rx. 1 (high temp. and low [HCl]) and slowest is Rx. 2 (low temp. and low [HCl])

3. Activation energy can be described as the

- A. energy of motion.
- B. energy of the activated complex.
- C. energy difference between the reactants and the products.
- D. energy difference between the reactants and the activated complex.

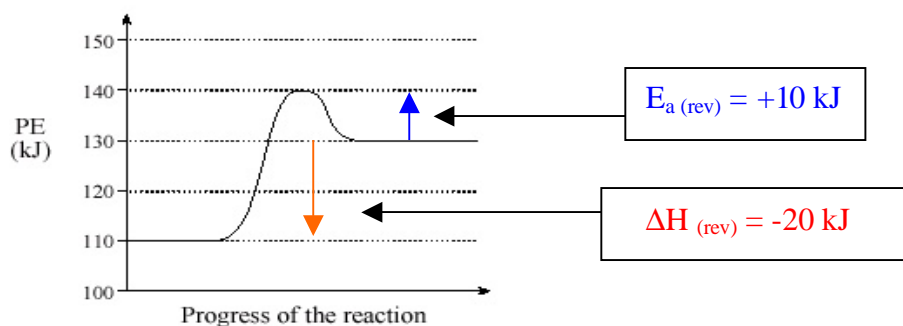
The answer is **D**



Explanation:

“D” is the **definition** of Activation Energy. “A” is just Kinetic Energy, “B” is the actual energy of the AC, not an energy difference on the PE Graph, “C” of course is the  $\Delta H$ .

4. Consider the following potential energy diagram for a reversible reaction:



Which of the following describes the system above?

|    | Reaction | Activation Energy (kJ) | $\Delta H$ (kJ) |
|----|----------|------------------------|-----------------|
| A. | reverse  | 10                     | -20             |
| B. | reverse  | 10                     | -30             |
| C. | forward  | 30                     | +10             |
| D. | forward  | 20                     | +30             |

The answer is **A**



**Explanation:**

The  $E_A$  for the reverse reaction is shown by the blue arrow placed on the diagram above and the  $\Delta H$  for the reverse reaction is shown by the red arrow. It is -20 because it's going **down** from 130 to 110.

The  $E_A$  for the *forward* reaction would be  $140 - 110 = 30$  and the  $\Delta H$  for the *forward* reaction would be  $130 - 110 = +20$ . Look at the graph and make sure you know how these are obtained!

5. Increasing the temperature of a reaction increases the reaction rate by

|      |  |
|------|--|
| I.   | increasing frequency of collisions           |
| II.  | increasing the kinetic energy of collision   |
| III. | decreasing the potential energy of collision |

- A. I only.
- B. I and II only.
- C. II and III only.
- D. I, II and III.

The answer is **B**



Explanation:

Increasing the temperature increases the average kinetic energy of the molecules so there are more collisions with a high kinetic energy (II) and if they are moving faster, there is also a greater probability (chance) of collisions (I). The potential energy barrier for a successful collision is called the activation energy, which **cannot** be altered by changing the temp. It can only be altered by using a suitable catalyst.

6. What effect does a catalyst have on a reaction?

- A. It changes the  $\Delta H$  of a reaction.
- B. It increases the kinetic energy of the reactants.
- C. It decreases the potential energy of the products.
- D. It provides a reaction mechanism with a lower activation energy.

The answer is **D**



Explanation:

A catalyst can change only the activation energy (the energy difference between the reactants or products and the *activated complex*). It cannot change the energy difference between the reactants and products ( $\Delta H$ ) or the potential energy of the reactants or products. Looking at “B”, only a change in *temperature* can change the kinetic energy.

7. Consider the following equilibrium:



Equal moles of  $\text{N}_2$  and  $\text{O}_2$  are added, under certain conditions, to a closed container. Which of the following describes the changes in the reverse reaction which occur as the system proceeds toward equilibrium?

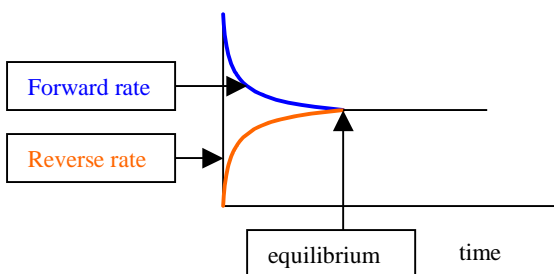
|    | Rate of Reverse Reaction | $[\text{NO}_2]$ |
|----|--------------------------|-----------------|
| A. | increases                | increases       |
| B. | decreases                | increases       |
| C. | increases                | decreases       |
| D. | decreases                | decreases       |

The answer is **A**



Explanation:

Since  $\text{N}_2$  &  $\text{O}_2$  are the only things added, the concentrations of these are both high. Since these are the reactants, the rate of the forward reaction starts out fast. Since there is no product to begin with, the rate of the reverse reaction is initially zero. As time goes on and the system approaches equilibrium,  $\text{N}_2$  &  $\text{O}_2$  are used up, their concentrations decrease and the rate of the forward reaction decreases. In the meantime, and as  $\text{N}_2$  &  $\text{O}_2$  were reacting,  $\text{NO}_2$  was being formed, so the  $[\text{NO}_2]$  would be increasing and the chance that one  $\text{NO}_2$  molecule could collide with another is increasing, therefore the rate of the reverse reaction would be increasing as the system approaches equilibrium. Remember the graph of forward and reverse reaction rate vs. time:



8. A chemical equilibrium is described as “dynamic” because
- maximum randomness has been achieved.
  - the pressure and temperature do not change.
  - both reactants and products continue to form.
  - the concentrations of chemical species remain constant.

The answer is **C**



Explanation:

“Dynamic” means things are continuing to happen. In other words the forward and reverse reactions continue to occur. Remember “A” is not true because maximum randomness (entropy) is NOT always achieved at equilibrium. “B” and “D” are both characteristics of equilibrium but these do not explain why it’s called *dynamic*.

9. Which of the following reactions results in an entropy increase?

- $2\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{CO}_{(g)}$
- $\text{N}_{2(g)} + 2\text{H}_{2(g)} \rightarrow \text{N}_2\text{H}_{4(l)}$
- $2\text{SO}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{SO}_{3(g)}$
- $\text{Ag}^+_{(aq)} + \text{Cl}^-_{(aq)} \rightarrow \text{AgCl}_{(s)}$

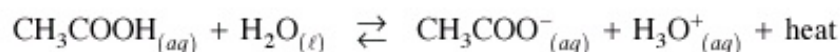
The answer is **A**



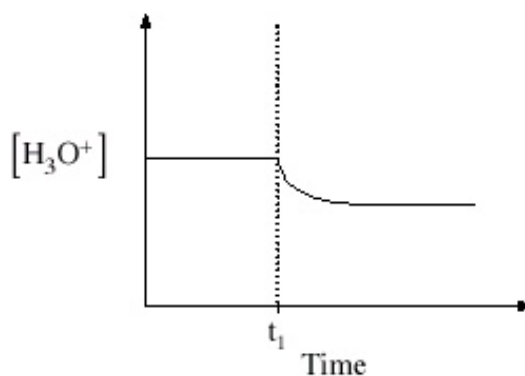
Explanation:

In reaction “A”, there is 1 mole of gas on the left side and 2 moles of gas on the right, therefore entropy is increasing as the reaction proceeds. In “B”, entropy is decreasing because gases are forming a liquid, in “C”, 3 moles of gas are forming 2 moles of gas and in “D”, (aq) is going to solid.

10. Consider the following equilibrium:



A stress was applied at time  $t_1$  and the data was plotted on the following graph:



The stress that was imposed at time  $t_1$  is the result of

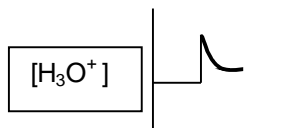
- A. the addition of HCl.
- B. decreasing the temperature.
- C. the addition of  $\text{NaCH}_3\text{COO}$ .
- D. increasing the volume of the container.

The answer is **C**



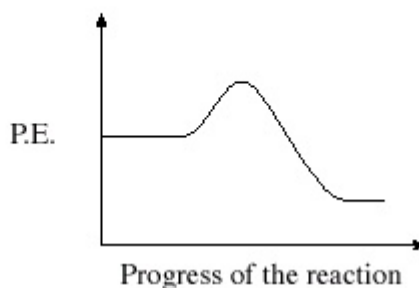
Explanation:

The fact that the  $[\text{H}_3\text{O}^+]$  gradually decreases means that the equilibrium is shifting to the left side. “A” is not correct because if HCl is added, the  $[\text{H}_3\text{O}^+]$  would go up and then partially down:



If the temperature was decreased as in “B” the equilibrium would shift to the right and  $[\text{H}_3\text{O}^+]$  would gradually increase. Answer “C” is correct because adding  $\text{NaCH}_3\text{COO}$  would produce  $\text{CH}_3\text{COO}^-$  which would shift the equilibrium to the left. Increasing the volume would decrease the pressure. This would have no effect on this equilibrium because there are no gases present.

11. Consider the following potential energy diagram for an equilibrium system:



When the temperature of the system is increased, the equilibrium shifts to the

- A. left and the  $K_{eq}$  increases.
- B. left and the  $K_{eq}$  decreases.
- C. right and the  $K_{eq}$  increases.
- D. right and the  $K_{eq}$  decreases.

The answer is **B**



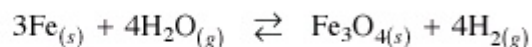
Explanation:

The graph shows that this is an **exothermic** reaction, so the heat term is on the *right*: e.g.)  $A + B \rightleftharpoons C + \text{heat}$ . When the temperature is increased, the equilibrium shifts to the left. The  $K_{eq}$  expression is:

$$K_{eq} = \frac{[C]}{[A][B]}$$

so as the equilibrium shifts to the left, the  $[C]$  decreases and the  $[A][B]$  increases, therefore the value of the ratio ( $K_{eq}$ ) decreases. Remember **TEMPERATURE IS THE ONLY THING THAT CHANGES THE VALUE OF  $K_{eq}$ !**

12. What is the  $K_{eq}$  expression for the following equilibrium?



- A.  $K_{eq} = [\text{H}_2]^4$   
 B.  $K_{eq} = \frac{[\text{H}_2]}{[\text{H}_2\text{O}]}$   
 C.  $K_{eq} = \frac{[\text{H}_2]^4}{[\text{H}_2\text{O}]^4}$   
 D.  $K_{eq} = \frac{[\text{Fe}_3\text{O}_4][\text{H}_2]^4}{[\text{Fe}]^3[\text{H}_2\text{O}]^4}$

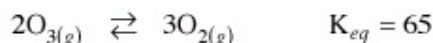
The answer is **C**



Explanation:

The  $\text{Fe}_3\text{O}_4$  and the Fe are left out of the  $K_{eq}$  expression because they are solids and their concentrations are constant. Remember the  $K_{eq}$  expressions only contain gases and aqueous species. Also remember that *coefficients* in the equation become *exponents* in the  $K_{eq}$  expression.

13. Consider the following equilibrium:



Initially, 0.10 mole of  $\text{O}_3$  and 0.10 mole of  $\text{O}_2$  are placed in a 1.0L container. Which of the following describes the changes in concentrations as the reaction proceeds toward equilibrium?

|    | $[\text{O}_3]$ | $[\text{O}_2]$ |
|----|----------------|----------------|
| A. | decreases      | decreases      |
| B. | decreases      | increases      |
| C. | increases      | decreases      |
| D. | increases      | increases      |

The answer is **B**



Explanation:

This is a Trial  $K_{eq}$  type of problem. The  $K_{eq}$  expression is:  
And the volume of the container is 1.0 L, so substituting the initial concentrations of  $\text{O}_2$  and  $\text{O}_3$  into the  $K_{eq}$  expression, we get:

$$K_{eq} = \frac{[\text{O}_2]^3}{[\text{O}_3]^2}$$

$$\text{Trial } K_{eq} = \frac{(0.10)^3}{(0.10)^2} = 0.10 \quad \text{The Trial } K_{eq}(0.10) < \text{Actual } K_{eq} (65) \text{ so the}$$

reaction will shift to the Right side as equilibrium is approached. In this process of shifting to the right  $[\text{O}_3]$  will decrease and  $[\text{O}_2]$  will increase.

14. Which of the following does **not** define solubility?

- A. the concentration of solute in a saturated solution
- B. the moles of solute dissolved in a given volume of solution
- C. the maximum mass of solute that can dissolve in a given volume of solution
- D. the minimum moles of solute needed to produce one litre of a saturated solution

The answer is **B**



Explanation:

Remember a saturated solution is one in which all the solute that can be, is dissolved in a given volume. Solubility can be defined as how much solute is needed to form a saturated solution. "A", "C" and "D" all state this. "B" is the one that does not define solubility because it does not mention saturation or solubility equilibrium. It could describe any solution.

15. The ion concentrations in 0.25M  $\text{Al}_2(\text{SO}_4)_3$  are

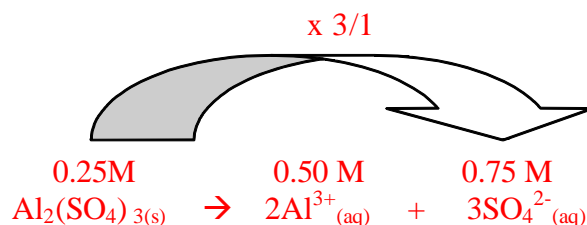
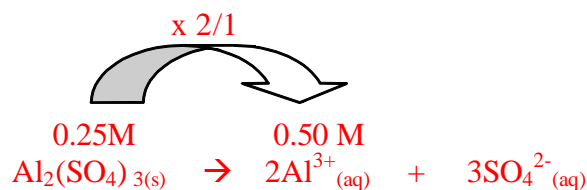
|    | $[\text{Al}^{3+}]$ | $[\text{SO}_4^{2-}]$ |
|----|--------------------|----------------------|
| A. | 0.25 M             | 0.25 M               |
| B. | 0.50 M             | 0.75 M               |
| C. | 0.75 M             | 0.50 M               |
| D. | 0.10 M             | 0.15 M               |

The answer is **B**



Explanation:

Looking up  $\text{Al}_2(\text{SO}_4)_3$  on the solubility table tells us that it is *soluble*. The question does not mention anything about the solution being saturated, so this is simply an “individual ion concentration” problem. The  $[\text{Al}^{3+}]$  and  $[\text{SO}_4^{2-}]$  can be obtained by using the balanced *dissociation* equation and the *coefficient ratios* as follows:



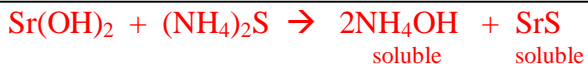
16. Which of the following will **not** produce a precipitate when equal volumes of 0.20 M solutions are combined?
- A. KOH and CaCl<sub>2</sub>
- B. Zn(NO<sub>3</sub>)<sub>2</sub> and K<sub>3</sub>PO<sub>4</sub>
- C. Sr(OH)<sub>2</sub> and (NH<sub>4</sub>)<sub>2</sub>S
- D. Na<sub>2</sub>SO<sub>4</sub> and Pb(NO<sub>3</sub>)<sub>2</sub>

The answer is **C**



Explanation:

For this one, we need to look at the Solubility Table. It's important here to realize that the compounds that could form precipitates are not the ones listed here, but the *products* that would result by mixing each pair. You could use "ion boxes" to determine the possible products in each case and use the solubility table to find which products form precipitates and which ones don't.



For this reaction (C) both products are soluble so no precipitate will be produced.



17. What is observed when H<sub>2</sub>SO<sub>4</sub> is added to a saturated solution of CaSO<sub>4</sub>?
- A. the pH increases
- B. the [Ca<sup>2+</sup>] increases
- C. bubbles of H<sub>2</sub> are given off
- D. additional CaSO<sub>4</sub> precipitates

The answer is **D**



Explanation:

The equilibrium equation describing a saturated solution of CaSO<sub>4</sub> is:



Adding the H<sub>2</sub>SO<sub>4</sub> will increase the [SO<sub>4</sub><sup>2-</sup>].

This will cause the equilibrium to shift to the left, which will produce more solid CaSO<sub>4</sub>.

18. The solubility of  $\text{CdS} = 2.8 \times 10^{-14}$ . The value of  $K_{sp}$  is

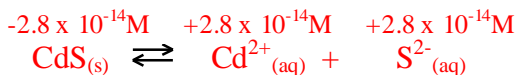
- A.  $7.8 \times 10^{-28}$
- B.  $2.8 \times 10^{-14}$
- C.  $5.6 \times 10^{-14}$
- D.  $1.7 \times 10^{-7}$

The answer is **A**



Explanation:

Remember solubility and  $K_{sp}$  are NOT the same thing! The solubility equilibrium equation for  $\text{CdS}$  is:  $\text{CdS}_{(s)} \rightleftharpoons \text{Cd}^{2+}_{(aq)} + \text{S}^{2-}_{(aq)}$   
 If we call the solubility “ $2.8 \times 10^{-14}\text{M}$ ”, we can show the changes in concentration as the reaction reaches equilibrium:



The  $K_{sp}$  expression is:

$K_{sp} = [\text{Cd}^{2+}] [\text{S}^{2-}]$  substituting the concentrations:

$$K_{sp} = (2.8 \times 10^{-14})^2 = 7.8 \times 10^{-28}$$

19. How many moles of solute are dissolved in 200.0 mL of a saturated solution of FeS?

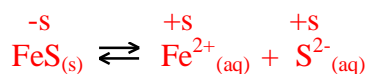
- A.  $1.2 \times 10^{-19}$
- B.  $6.0 \times 10^{-19}$
- C.  $1.5 \times 10^{-10}$
- D.  $7.7 \times 10^{-10}$

The answer is C



Explanation:

First of all we must find the molar solubility of FeS. We do this by writing the equation for the solubility equilibrium of FeS and using the  $K_{sp}$  expression. Since we don't know the molar solubility, we call it "s":



$K_{sp} = [\text{Fe}^{2+}] [\text{S}^{2-}]$  and substituting "s" for the concentrations:

$$K_{sp} = s^2$$

Looking up the  $K_{sp}$  for FeS on the  $K_{sp}$  table, it is  $6.0 \times 10^{-19}$ .

$$\text{So the molar solubility, } s = \sqrt{K_{sp}} = \sqrt{6.0 \times 10^{-19}} = 7.75 \times 10^{-10} \text{ M}$$

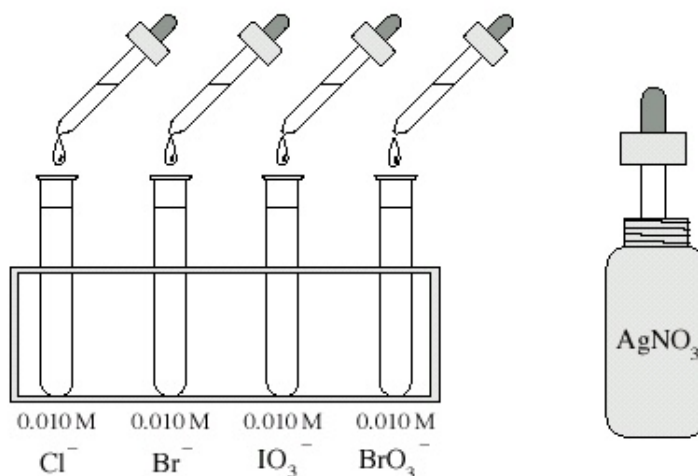
Remember:

$$\frac{\text{moles}}{\text{M} \cdot \text{L}}$$

We can now use: moles = M x L to calculate the moles:

$$\text{Moles} = 7.75 \times 10^{-10} \text{ M} \times 0.2000 \text{ L} = 1.5 \times 10^{-10} \text{ moles}$$

20. Consider the following 10.0 mL solutions:



Equal moles of AgNO<sub>3</sub> are added to each solution. It is observed that a precipitate forms in all but one solution. Which solution does **not** form a precipitate?

- A. Cl<sup>-</sup>
- B. Br<sup>-</sup>
- C. IO<sub>3</sub><sup>-</sup>
- D. BrO<sub>3</sub><sup>-</sup>

The answer is **D**



Explanation:

If you look on the solubility table you will notice that IO<sub>3</sub><sup>-</sup> and BrO<sub>3</sub><sup>-</sup> are not on the table. The next step would be to look at the K<sub>sp</sub> table. Comparing the K<sub>sp</sub>'s of the possible precipitates that could form:

| Possible precipitate | K <sub>sp</sub>         |
|----------------------|-------------------------|
| AgCl                 | 1.8 x 10 <sup>-10</sup> |
| AgBr                 | 5.4 x 10 <sup>-13</sup> |
| AgIO <sub>3</sub>    | 3.2 x 10 <sup>-8</sup>  |
| AgBrO <sub>3</sub>   | 5.3 x 10 <sup>-5</sup>  |

We see that AgBrO<sub>3</sub> has the highest K<sub>sp</sub>. Since these compounds are all the same type (AB compounds), the higher the K<sub>sp</sub>, the higher the solubility. So AgBrO<sub>3</sub> has the highest solubility of the four. Therefore if only one of these does not form a precipitate, the AgBrO<sub>3</sub> must be the one or BrO<sub>3</sub><sup>-</sup> is the ion that does not precipitate with Ag<sup>+</sup>.

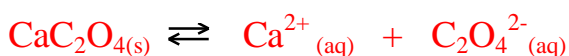
21. Which of the following could dissolve a precipitate of  $\text{CaC}_2\text{O}_4$  in a saturated solution of  $\text{CaC}_2\text{O}_4$  ?
- A.  $\text{NaOH}$
  - B.  $\text{CaC}_2\text{O}_4$
  - C.  $\text{H}_2\text{C}_2\text{O}_4$
  - D.  $\text{Ca}(\text{NO}_3)_2$

The answer is **A**



Explanation:

When you are asked about compounds dissolving precipitates in a saturated solution, the first thing to do is write the solubility equilibrium equation for the compound of LOW solubility. In this case, that is  $\text{CaC}_2\text{O}_4$  :



First, why B, C and D don't work!  $\text{CaC}_2\text{O}_4$  and  $\text{H}_2\text{C}_2\text{O}_4$  both release the  $\text{C}_2\text{O}_4^{2-}$  ion to the solution. This increases the  $[\text{C}_2\text{O}_4^{2-}]$  in the equilibrium above, causing a shift to the *left*, forming more solid precipitate, rather than dissolving it. The  $\text{Ca}^{2+}$  in  $\text{Ca}(\text{NO}_3)_2$  would increase the  $[\text{Ca}^{2+}]$  in the equilibrium above also causing a shift to the *left*, again forming more solid precipitate, rather than dissolving it. Looking at the solubility table,  $\text{OH}^-$  forms a compound of low solubility with  $\text{Ca}^{2+}$ . So when  $\text{NaOH}$  is added to saturated  $\text{CaC}_2\text{O}_4$ , the  $\text{OH}^-$  precipitates the  $\text{Ca}^{2+}$ , forming  $\text{Ca}(\text{OH})_{2(s)}$ . This decreases the  $[\text{Ca}^{2+}]$  in the equilibrium above, causing it to shift to the *right*, thus dissolving the solid  $\text{CaC}_2\text{O}_4$ .

22. Which of the following is a general property of bases?
- A. taste sour
  - B. turn litmus red
  - C. conduct electric current in solution
  - D. concentration of  $\text{H}_3\text{O}^+$  is greater than concentration of  $\text{OH}^-$

The answer is **C**



Explanation:

Answers A, B and D are all true of *acids* ONLY. Both acids and bases conduct electric current in solution, so this would also be a general property of a base.

23. Water will act as an acid with which of the following?

|      |                         |
|------|-------------------------|
| I.   | $\text{H}_2\text{CO}_3$ |
| II.  | $\text{HCO}_3^-$        |
| III. | $\text{CO}_3^{2-}$      |

- A. I only.
- B. III only.
- C. I and II only.
- D. II and III only.

The answer is **D**



**Explanation:**

Be careful how you read this. They are NOT asking which ones will act as an acid! They are asking with which ones will *water* act as an acid. In other words, which ones will act as *bases* in water (If water is the Bronsted acid, the other species would be the Bronsted base.)  $\text{H}_2\text{CO}_3$  is a weak acid on the acid table, so it will be an acid.  $\text{CO}_3^{2-}$  is a weak base because it is found on the right side of the acid table (6<sup>th</sup> from the bottom).  $\text{HCO}_3^-$  is *amphiprotic* (it is found on BOTH sides of the acid table.) To find out whether it acts as an acid or base in a water solution, we have to compare its  $K_a$  and its  $K_b$ . To find its  $K_a$ , we look for  $\text{HCO}_3^-$  on the *left* side of the table (6<sup>th</sup> from the bottom). Its  $K_a = 5.6 \times 10^{-11}$ . To find its  $K_b$ , we look for  $\text{HCO}_3^-$  on the *right* side of the table, then use the equation:

$$K_b = K_w / K_a \text{ (conjugate acid)} \quad \text{so} \quad K_b (\text{HCO}_3^-) = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8}$$

So we see that  $K_b > K_a$  for  $\text{HCO}_3^-$ , therefore it acts as a *base* in water. (Making water act as the acid.) So II and III only are bases (Water acts as an acid in II and III only.)

24. Which of the following 1.0M solutions will have the greatest electrical conductivity?

- A. HI
- B.  $\text{H}_2\text{S}$
- C. HCN
- D.  $\text{H}_3\text{PO}_4$

Explanation: The Answer is “**A**”. Look on the Acid Table. HI is a strong acid while the others are all weak acids. The strong acid (HI) is 100% ionized so it will be the best electrical conductor. The other acids are weak and are only partially ionized. They exist mostly as neutral molecules in solution, so conductivity is low.

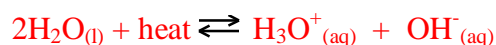
25. An acid is added to water and a new equilibrium is established. The new equilibrium can be described by
- A.  $\text{pH} < \text{pOH}$  and  $K_w = 1 \times 10^{-14}$
  - B.  $\text{pH} < \text{pOH}$  and  $K_w < 1 \times 10^{-14}$
  - C.  $\text{pH} > \text{pOH}$  and  $K_w = 1 \times 10^{-14}$
  - D.  $\text{pH} > \text{pOH}$  and  $K_w > 1 \times 10^{-14}$

The answer is **A**



**Explanation:**

$K_w$  is the equilibrium constant for the ionization of water:



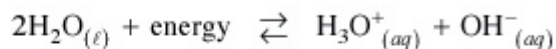
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

The value of  $K_w$  can ONLY be changed by a change in temperature and not by addition of an acid or base to the solution. Therefore answers “B” and “D” cannot be true since no mention is made of a change in temperature.

When an acid is added, this increases the  $[\text{H}_3\text{O}^+]$ . This causes the equilibrium above to shift to the left. The net change will be that  $[\text{H}_3\text{O}^+]$  will be slightly higher and  $[\text{OH}^-]$  will be lower than it was in neutral water. Since  $[\text{H}_3\text{O}^+]$  goes up, pH goes down. And because  $[\text{OH}^-]$  decreases, the pOH goes up. This will make  $\text{pH} < \text{pOH}$ .

Another way to look at this problem is to look at the  $K_w$  expression:  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$  As acid is added,  $[\text{H}_3\text{O}^+]$  increases, and since  $K_w$  is constant, the  $[\text{OH}^-]$  must decrease. Since  $[\text{H}_3\text{O}^+]$  goes up, pH goes down. And because  $[\text{OH}^-]$  decreases, the pOH goes up. This will make  $\text{pH} < \text{pOH}$ . Also, you could use the fact that  $\text{pH} + \text{pOH} = \text{p}K_w$ .

26. Consider the following equilibrium:



The  $[\text{H}_3\text{O}^+]$  will decrease and the  $K_w$  will remain constant when

- A. a strong acid is added.
- B. a strong base is added.
- C. the temperature is increased.
- D. the temperature is decreased.

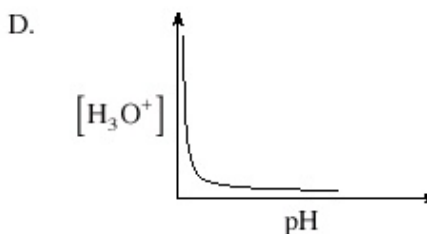
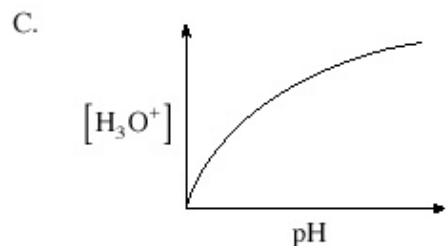
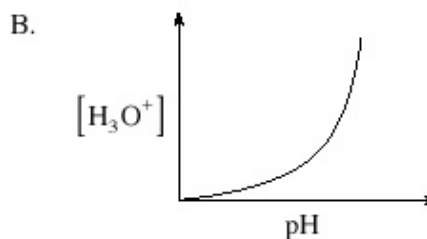
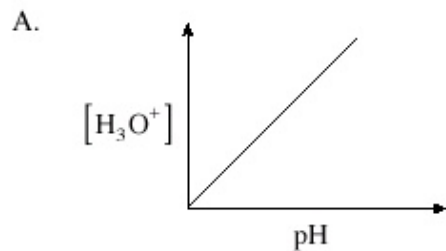
The answer is **B**



Explanation:

When a strong base is added, the  $[\text{OH}^-]$  increases. This causes the equilibrium to shift to the left and  $[\text{H}_3\text{O}^+]$  will decrease. The value of  $K_w$  can ONLY be changed by a change in temperature and not by addition of an acid or base to the solution. With answers “C” or “D”, the value of  $K_w$  would change.

27. Which of the following graphs describes the relationship between  $[\text{H}_3\text{O}^+]$  and pH ?



The answer is **D**. Don't let graphs scare you! Since  $\text{pH} = -\log [\text{H}_3\text{O}^+]$ , as pH goes up,  $[\text{H}_3\text{O}^+]$  will go down. The only graph which shows this is answer “D”. If you are a math student, you might recognize this curve as logarithmic.

28. When the  $[\text{H}_3\text{O}^+]$  in a solution is increased to twice the original concentration, the change in pH could be from
- A. 1.7 to 1.4
  - B. 2.0 to 4.0
  - C. 5.0 to 2.5
  - D. 8.5 to 6.5

The answer is **A**. EXPLANATION:

First of all, if  $[\text{H}_3\text{O}^+]$  is increased, the pH will go *down*, since  $\text{pH} = -\log [\text{H}_3\text{O}^+]$ . So answer “B” is eliminated. A foolproof method of doing this question is to convert pH’s into  $[\text{H}_3\text{O}^+]$  for answers A, C and D, until you hit one in which the  $[\text{H}_3\text{O}^+]$  doubles. Use the equation:  $[\text{H}_3\text{O}^+] = \text{antilog} (-\text{pH})$

For “A”  $\rightarrow [\text{H}_3\text{O}^+]_1 = \text{antilog} (-1.7) = 0.02 \text{ M}$       $[\text{H}_3\text{O}^+]_2 = \text{antilog} (-1.4) = 0.04 \text{ M}$

So, we have our answer already! When pH goes from 1.7 to 1.4, the  $[\text{H}_3\text{O}^+]$  changes from 0.02M to 0.04M, which is *doubled*.

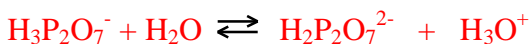
29. The relationship  $\frac{[\text{H}_2\text{P}_2\text{O}_7^{2-}][\text{H}_3\text{O}^+]}{[\text{H}_3\text{P}_2\text{O}_7^-]}$  is the
- A.  $K_a$  for  $\text{H}_3\text{P}_2\text{O}_7^-$
  - B.  $K_b$  for  $\text{H}_3\text{P}_2\text{O}_7^-$
  - C.  $K_a$  for  $\text{H}_2\text{P}_2\text{O}_7^{2-}$
  - D.  $K_b$  for  $\text{H}_2\text{P}_2\text{O}_7^{2-}$

The answer is **A**. EXPLANATION:

In any  $K_{\text{eq}}$  expression, the Products are On Top (POT) and the reactants are on the bottom. So if we write an equation to go along with this:



I think you can see that to make this balance we need to put an  $\text{H}_2\text{O}$  on the left side:



The  $\text{H}_3\text{P}_2\text{O}_7^-$  is losing a proton as it forms  $\text{H}_2\text{P}_2\text{O}_7^{2-}$  so it must be acting as an acid. The  $K_{\text{eq}}$  for the reaction of any acid with water to form  $\text{H}_3\text{O}^+$  and its conjugate base is called the  $K_a$  for that acid. So the expression above is the  $K_a$  for  $\text{H}_3\text{P}_2\text{O}_7^-$ . For this type of question, take the time to reason it out like I have done here.

30. Which of the following describes the relationship between acid strength and  $K_a$  value for weak acids?

|    | Acid Strength | $K_a$            |
|----|---------------|------------------|
| A. | increases     | increases        |
| B. | increases     | decreases        |
| C. | decreases     | increases        |
| D. | decreases     | remains constant |

The answer is **A**. EXPLANATION:

If you look on the Acid Table, you will see that the strength of the acid increases as you move toward the *top* (The long arrow on the left). Also, you will see that the values of  $K_a$  increase as you move toward the top of the table. Also, you should know that for a weak acid,  $K_a$  is the equilibrium constant for its ionization. The greater the  $K_a$ , the greater the degree of ionization, hence the greater the *strength* of the acid. (PS. Don't get "strength" confused with "concentration" of acids.)

31. The value of  $K_b$  for  $\text{HPO}_4^{2-}$  is

- A.  $2.2 \times 10^{-13}$
- B.  $6.2 \times 10^{-8}$
- C.  $1.6 \times 10^{-7}$
- D.  $4.5 \times 10^{-2}$

The answer is **C**. EXPLANATION:

Be careful because  $\text{HPO}_4^{2-}$  is amphoteric. To find  $K_b$  of a species on the acid table, remember to use the following sequence.

Locate the species ( $\text{HPO}_4^{2-}$  in this case) on the **RIGHT** side of the acid table.

If you look to the right of it, the  $K_a$  for its conjugate acid ( $\text{H}_2\text{PO}_4^-$ ) is  $6.2 \times 10^{-8}$

Use the equation:  $K_b = K_w / K_{a(\text{conjugate acid})}$

$$K_b (\text{HPO}_4^{2-}) = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7}$$

32. Which of the following 1.0 M solutions would have a pH greater than 7.00?

- A. HCN
- B.  $\text{KNO}_3$
- C.  $\text{NH}_4\text{Cl}$
- D.  $\text{NaCH}_3\text{COO}$

The answer is **D**. EXPLANATION:

Something with a  $\text{pH} > 7.00$  would have to be *basic*. So we need to find which one of the following is basic. Looking at “A”, HCN is a weak acid so a solution of it would be acidic ( $\text{pH} < 7$ ). The rest of the choices are *salts*. In order to find out whether a salt is acidic, basic or neutral we must use the concept of *hydrolysis*. Recall that for a salt we:

**Dissociate** :  $\text{KNO}_3 \rightarrow \text{K}^+_{(\text{aq})} + \text{NO}_3^-_{(\text{aq})}$

**Eliminate** (the spectators) : Both  $\text{K}^+$  and  $\text{NO}_3^-$  are spectators:

**Evaluate** (the active ions and what they are by their location on the acid table): Both ions are spectators so  $\text{KNO}_3$  is neutral, so its  $\text{pH}=7$

**Dissociate** :  $\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$

**Eliminate** (the spectators) :  $\text{Cl}^-$  is a spectator

**Evaluate** (the active ions and what they are by their location on the acid table): The ion  $\text{NH}_4^+$  is found on the *left* side of the acid table, so it is *acidic*. So this salt,  $\text{NH}_4\text{Cl}$ , is *acidic* and its  $\text{pH}$  would be  $< 7$ .

**Dissociate** :  $\text{NaCH}_3\text{COO} \rightarrow \text{Na}^+_{(\text{aq})} + \text{CH}_3\text{COO}^-_{(\text{aq})}$

**Eliminate** (the spectators) :  $\text{Na}^+$  is a spectator

**Evaluate** (the active ions and what they are by their location on the acid table): The ion  $\text{CH}_3\text{COO}^-$  is found on *right* side of the acid table, near the middle. So it is a weak *base*. Therefore the salt  $\text{NaCH}_3\text{COO}$  is *basic* and its  $\text{pH} > 7$ .

33. What is the pH at the transition point for an indicator with a  $K_a$  of  $2.5 \times 10^{-4}$  ?

- A.  $2.5 \times 10^{-4}$
- B. 3.60
- C. 7.00
- D. 10.40

The answer is **B**. EXPLANATION:

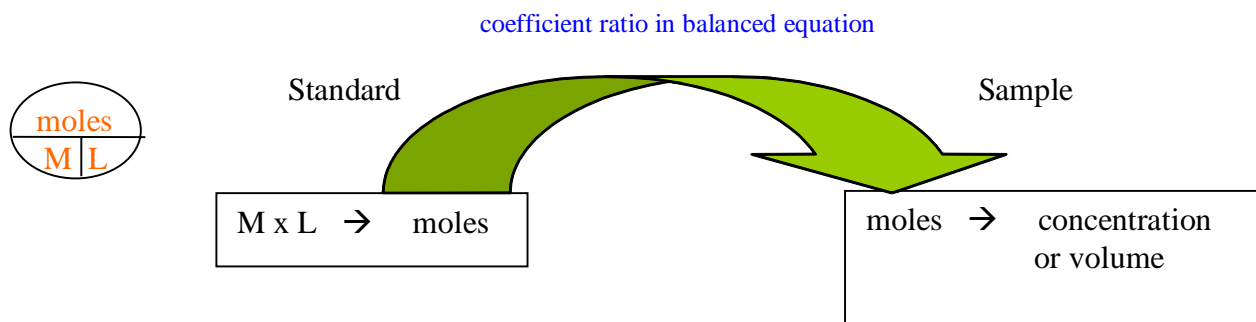
Recall that at the *transition point* of an indicator  $\text{Ka}(\text{of indicator}) = [\text{H}_3\text{O}^+]$  or  $\text{pKa} = \text{pH}$

$\text{pKa} = -\log \text{Ka}$  so at the transition point  $\rightarrow \text{pH} = \text{pKa} = -\log (2.5 \times 10^{-4}) = 3.60$

34. What volume of 0.100 M NaOH is required to completely neutralize 15.00 mL of 0.100 M  $\text{H}_3\text{PO}_4$  ?
- A. 5.00 mL  
 B. 15.0 mL  
 C. 30.0 mL  
 D. 45.0 mL

The answer is **D**. EXPLANATION:

This is a *titration* question involving neutralization. You might recall the saying: “*Standard, Sample, Moles in the Middle*”. The following diagram can be used for this type of titration question:



The  $\text{H}_3\text{PO}_4$  is the standard because you are given the volume and the concentration of it.

$$\text{Moles of H}_3\text{PO}_4 = 0.100 \text{ M} \times 0.01500 \text{ L} = 0.001500 \text{ moles of H}_3\text{PO}_4$$

The balanced equation is:  $\text{H}_3\text{PO}_4 + 3\text{NaOH} \rightarrow 3 \text{H}_2\text{O} + \text{Na}_3\text{PO}_4$

$$\text{Moles of NaOH} = 0.001500 \text{ moles of H}_3\text{PO}_4 \times \frac{3 \text{ moles NaOH}}{1 \text{ mole H}_3\text{PO}_4} = 0.00450 \text{ moles NaOH}$$

$$\text{Volume of NaOH} = \frac{0.00450 \text{ moles}}{0.100 \text{ M}} = 0.0450 \text{ L} = 45.0 \text{ mL}$$

35. What is the pH of the solution formed when 0.060 moles NaOH is added to 1.00 L of 0.050 M HCl?
- A. 2.00  
 B. 7.00  
 C. 12.00  
 D. 12.78

The answer is **C**. **EXPLANATION:**

This is a “Strong Acid-Strong Base Mixture Problem”:

1. Find the initial moles of each
2. Find the moles (of acid or base) in excess
3. Divide the excess moles by the total volume to find the  $[\text{H}_3\text{O}^+]$  or  $[\text{OH}^-]$
4. Use this to calculate the pH of the solution formed (mixture)

$$\begin{array}{l} \text{Initial moles of NaOH (given)} = 0.060 \text{ moles NaOH} \\ \text{Initial moles of HCl} = 0.050 \text{ M} \times 1.00 \text{ L} = 0.050 \text{ moles of HCl} \\ \hline \text{Excess (moles of NaOH)} = 0.010 \text{ moles NaOH} \end{array}$$

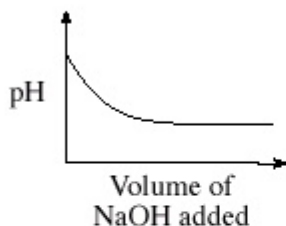
$$[\text{OH}^-] = [\text{NaOH}] = \frac{0.010 \text{ moles}}{1.00 \text{ L}} = 0.010 \text{ M}$$

$$\text{pOH} = -\log(0.010) = 2.00$$

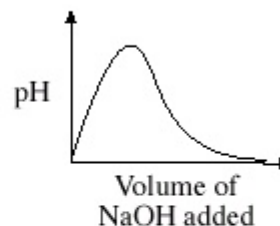
$$\text{pH} = 14.00 - \text{pOH} = 12.00$$

36. Which of the following graphs describes the relationship between the pH of a buffer and the volume of NaOH added to the buffer?

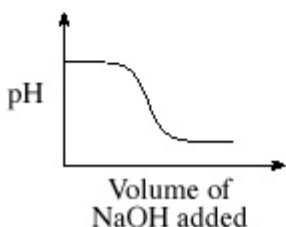
A.



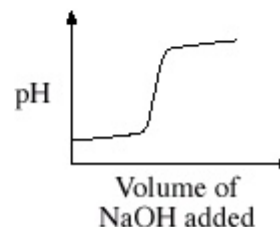
B.



C.



D.



The answer is **D**. EXPLANATION:

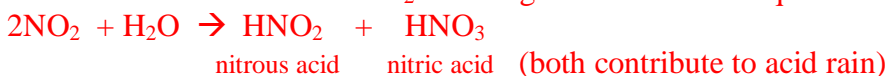
The function of a buffer is to keep the pH relatively constant when an acid or base is added to the solution. A buffer will do this fairly effectively until its capacity is exceeded (the weak acid or weak base is all used up). NaOH is a strong base. As NaOH is added to a buffer, the pH will go up quite slowly as long as the buffer is effective. As soon as the capacity is exceeded, the pH will start going up a lot more quickly. The pH will NOT go down as long as NaOH is being added. The only graph which displays this behavior is graph "D". (Notice that the curve resembles a titration curve – which is really what is happening here.)

37. A gas which is produced by internal combustion engines and contributes to the formation of acid rain is

- A. H<sub>2</sub>
- B. O<sub>3</sub>
- C. CH<sub>4</sub>
- D. NO<sub>2</sub>

The Answer is **D**. EXPLANATION:

Remember that in an internal combustion engine the extreme heat causes N<sub>2</sub> and O<sub>2</sub> to combine and form NO and NO<sub>2</sub>. You might also recall the equation:



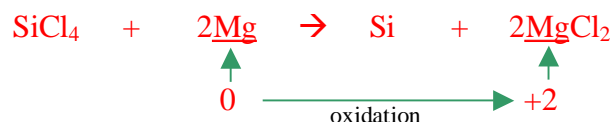
38. Which of the following represents a redox reaction?

- A.  $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$   
 B.  $\text{SiCl}_4 + 2\text{Mg} \rightarrow \text{Si} + 2\text{MgCl}_2$   
 C.  $2\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow 2\text{H}_2\text{O} + \text{Na}_2\text{SO}_4$   
 D.  $\text{AgBr} + 2\text{S}_2\text{O}_3^{2-} \rightarrow \text{Ag}(\text{S}_2\text{O}_3)_2^{3-} + \text{Br}^-$

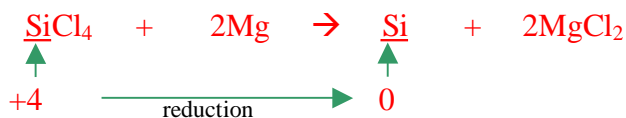
The Answer is **B**. EXPLANATION:

A “redox” reaction is one in which there is *both* an *oxidation* and a *reduction*. A hint in this type of question is to look for a reaction in which there are *elements and compounds*, because the oxidation number of an element in elemental form is “0” and in compounds or ions, the oxidation numbers of elements are usually something other than “0”.

Notice in answer “B” that Mg’s oxidation # goes from “0” to “+2”:



Remember the rule that *halogens* when at the *end* of a formula, have an oxidation # of “-1”. That means that Si’s oxidation number goes from “+4” to “0”:



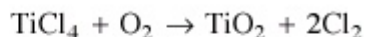
As a good review of oxidation numbers, you could check the other answers and see that there are no changes in oxidation #'s. Remember that unless there is  $\text{O}_2$ ,  $\text{H}_2$ ,  $\text{H}_2\text{O}_2$ , or a metallic hydride (like  $\text{NaH}$ ), that oxygen and hydrogen don’t change in oxidation #'s. Go over the following and make sure you understand how I got them!

In “A”, Ca remains as “+2” and C remains as “+4”.

In “C”, Na remains as “+1” and S remains as “+6”

In “D”, Ag remains as “+1”, S remains as “+2” and Br remains as “-1”

39. Consider the following reaction:



Each oxygen atom is

- A. reduced and loses  $2e^-$
- B. reduced and gains  $2e^-$
- C. oxidized and loses  $2e^-$
- D. oxidized and gains  $2e^-$

The Answer is **B**. EXPLANATION:

The best way to tackle this question is to use *oxidation numbers*. Oxygen in its *elemental* form ( $\text{O}_2$ ) has an oxidation number of “0”, while in compounds (eg.  $\text{TiO}_2$ ), oxygen has an oxidation # of “-2”:



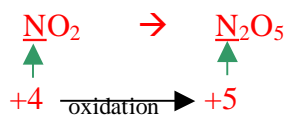
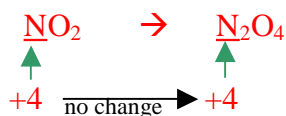
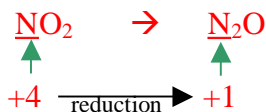
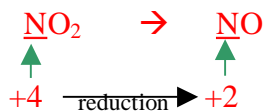
Each oxygen atom is changing oxidation # from “0” to “-2”, which is a *reduction* of two, or *gaining*  $2e^-$ 's. (Remember GER!)

40. When  $\text{NO}_2$  acts as a reducing agent, a possible product is

- A. NO
- B.  $\text{N}_2\text{O}$
- C.  $\text{N}_2\text{O}_4$
- D.  $\text{N}_2\text{O}_5$

The Answer is **D**. EXPLANATION:

This one can be a little tricky if you don't read it correctly! Remember that a **reducing agent is oxidized** (undergoes oxidation itself). So if  $\text{NO}_2$  is *oxidized*, the oxidation number of N must *increase*. (We can see that oxygen is always in compounds and there is no  $\text{H}_2\text{O}_2$  so we can assume that the oxidation # of *oxygen* does not change.) So we just find the changes in oxidation #'s of N ( $\Delta \text{ON}$ ) as  $\text{NO}_2$  forms each of the following compounds in the answers. In “D”, N is *oxidized* so  $\text{NO}_2$  is the *reducing agent*.



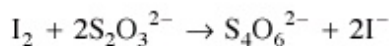
41. Which of the following 1.0M solutions will react spontaneously with lead?

- A. KCl
- B.  $\text{CuCl}_2$
- C.  $\text{ZnCl}_2$
- D.  $\text{MgCl}_2$

The Answer is **B**. EXPLANATION:

We need to use our Reduction Table for this question. By “lead”, they mean the *neutral atom* “Pb” which is found on the *right* side at “-0.13 volts”. This Pb will react spontaneously with any species on the *left* side *above* it. All of the compounds shown are *ionic* and contain the chloride ion “Cl<sup>-</sup>”. Cl<sup>-</sup> is found (by itself) only on the *right* side at “+1.36 volts”. Therefore the Cl<sup>-</sup> will not react with the Pb. Dissociating the four ionic compounds given in the answers yield the ions: K<sup>+</sup>, Cu<sup>2+</sup>, Zn<sup>2+</sup>, and Mg<sup>2+</sup>. If you look for these ions on the *left* side, you will see that they are all *below* Pb except the Cu<sup>2+</sup> which is found on the *left* side at “+0.34 volts”. This means the reaction:  $\text{Pb} + \text{Cu}^{2+} \rightarrow$  is *spontaneous*.

42. Consider the following redox reaction:

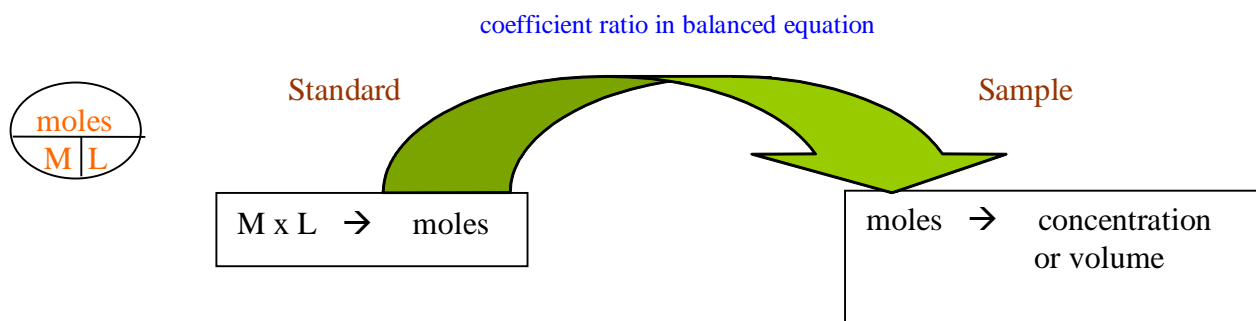


In a titration, 40.00 mL of  $\text{Na}_2\text{S}_2\text{O}_3$  is needed to react completely with  $4.0 \times 10^{-3}$  mol  $\text{I}_2$ . What is the concentration of  $\text{Na}_2\text{S}_2\text{O}_3$  ?

- A. 0.10 M
- B. 0.16 M
- C. 0.20 M
- D. 0.32 M

The Answer is **C**. EXPLANATION:

This is another *titration* question. You have to realize that each  $\text{Na}_2\text{S}_2\text{O}_3$  dissociates into *one*  $\text{S}_2\text{O}_3^{2-}$  ion, so the  $[\text{Na}_2\text{S}_2\text{O}_3] = [\text{S}_2\text{O}_3^{2-}]$ . Use the *titration* diagram:



The  $\text{I}_2$  is the *standard* because you are given the *moles* of it.

We use the moles of  $\text{I}_2$  and the coefficient ratio in the balanced redox equation to calculate the moles of  $\text{S}_2\text{O}_3^{2-}$  (the *sample*).

$$4.0 \times 10^{-3} \text{ moles I}_2 \times \frac{2 \text{ moles S}_2\text{O}_3^{2-}}{1 \text{ mole I}_2} = 8.0 \times 10^{-3} \text{ moles S}_2\text{O}_3^{2-}$$

Now, we calculate the  $[\text{S}_2\text{O}_3^{2-}]$  using  $\text{M} = \text{moles/L}$ :

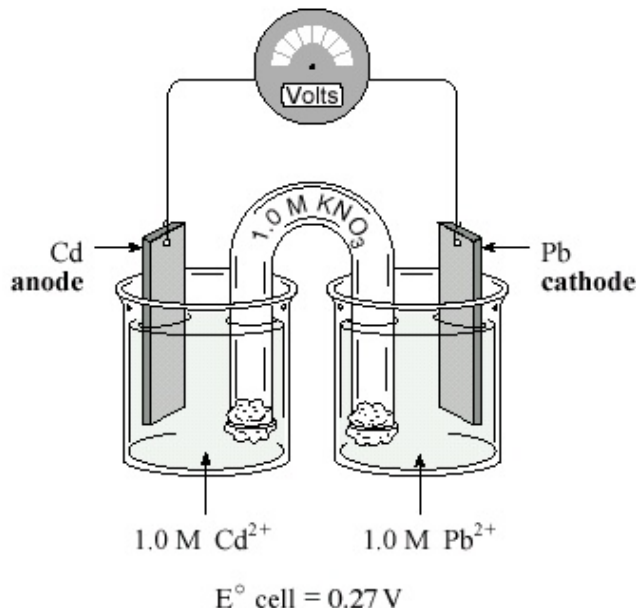
$$[\text{Na}_2\text{S}_2\text{O}_3] = [\text{S}_2\text{O}_3^{2-}] = \frac{8.0 \times 10^{-3} \text{ moles S}_2\text{O}_3^{2-}}{0.04000 \text{ L}} = 0.20 \text{ M}$$

43. In an operating electrochemical cell the function of a salt bridge is to
- allow hydrolysis to occur.
  - allow a non-spontaneous reaction to occur.
  - permit the migration of ions within the cell.
  - transfer electrons from the cathode to the anode.

The Answer is **C**. EXPLANATION:

Just knowing the function of each part of the electrochemical cell will tell you that “C” is the answer. Just going through the other answers: Looking at “A”, “*hydrolysis*” is the reaction of a salt, or ions of a salt with water to form an acid or base solution – nothing to do with electrochemical cells. Answer “B” is wrong because in an ECC (electrochemical cell) the reaction is *spontaneous*. Looking at “D”, remember that  $e^-$ 's move from the *anode* (where *oxidation* takes place) to the *cathode* (where *reduction* takes place) in the *wire*.

Use the following diagram to answer questions 44 and 45.



44. As the cell operates, electrons flow toward
- the Pb electrode, where Pb is oxidized.
  - the Cd electrode, where Cd is oxidized.
  - the Pb electrode, where  $\text{Pb}^{2+}$  is reduced.
  - the Cd electrode, where  $\text{Cd}^{2+}$  is reduced.

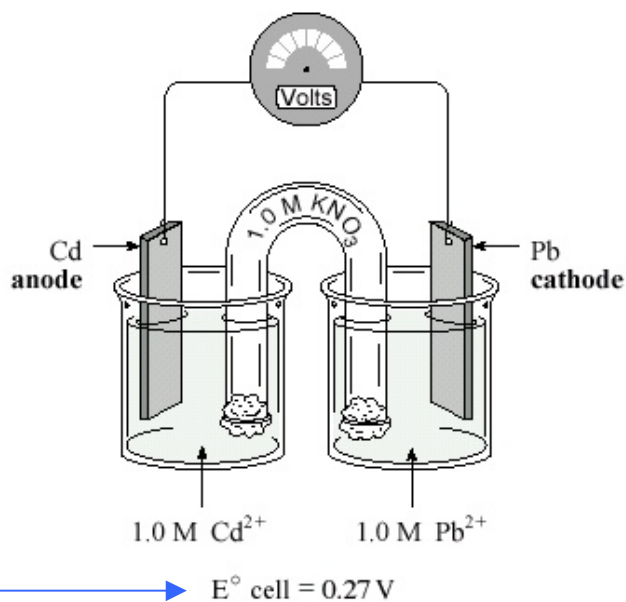
The Answer is **C**. EXPLANATION:

Remember that  $e^-$ 's move from the *anode* (Cd) (where *oxidation* takes place) toward the *cathode* (Pb) (where *reduction* takes place) in the *wire*. The half reaction at the cathode (GERC) is:  $\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}_{(aq)}$  Here, as you can see the  $\text{Pb}^{2+}$  is being *reduced*.

45. The  $E^{\circ}$  value for the reduction of  $\text{Cd}^{2+}$  is

- A.  $-0.40 \text{ V}$
- B.  $-0.27 \text{ V}$
- C.  $+0.14 \text{ V}$
- D.  $+0.40 \text{ V}$

Use the following diagram to answer questions 44 and 45.



The Answer is **A**. EXPLANATION:

Cd is the anode so it is being oxidized:



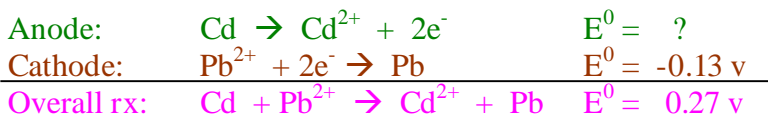
The half-reaction at the cathode is:



We know that the reduction potential for  $\text{Pb}^{2+}$  is  $-0.13 \text{ v}$  (see reduction table)

We also know that the  $E^{\circ}$  for the overall redox reaction is  $0.27 \text{ v}$  (see the bottom of the cell diagram.  $\rightarrow E^{\circ} \text{ cell} = 0.27 \text{ V}$ )

We can set up the half-reactions with the  $E^{\circ}$ 's adding up to the overall reaction:



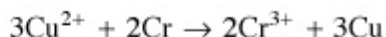
From this we can calculate “?” (the oxidation potential for Cd)

Oxidation potential for Cd (?) =  $0.27\text{v} - (-0.13 \text{ v}) = 0.27 \text{ v} + 0.13\text{v} = +0.40\text{v}$ .



You have to be really careful here, because the question asks for the **reduction potential** of  $\text{Cd}^{2+}$ ! That would be the  $E^{\circ}$  for the half-reaction:  $\text{Cd}^{2+} + 2e^{-} \rightarrow \text{Cd}$ . Since the half-reaction  $\text{Cd}^{2+} + 2e^{-} \rightarrow \text{Cd}$  is the *reverse* of the half-reaction  $\text{Cd} \rightarrow \text{Cd}^{2+} + 2e^{-}$  The  $E^{\circ}$  for the reduction of  $\text{Cd}^{2+}$  (the reduction potential of  $\text{Cd}^{2+}$ ) would be  $-0.40\text{v}$ . Remember, when a half-reaction is reversed, the sign on the  $E^{\circ}$  is also reversed.

46. The following reaction occurs in an electrochemical cell:



The  $E^\circ$  for the cell is

- A. 0.40 V
- B. 0.75 V
- C. 1.08 V
- D. 2.50 V

The Answer is **C**. EXPLANATION:

What you must do is look on the reduction table and find two half-reactions which will add up to give this as the overall redox reaction. Remember that for the one that is reversed, the sign on the  $E^0$  will also be reversed.

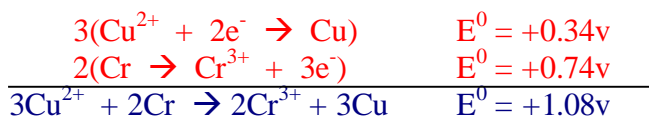
The  $\text{Cu}^{2+}$  is on the left side of the overall reaction so it must also be on the left side of the half-reaction.  $\text{Cu}^{2+}$  changing to  $\text{Cu}$  is shown by the half-reaction at +0.34v on the reduction table:



The  $\text{Cr}$  in the overall reaction is on the left side and the  $\text{Cr}^{3+}$  is on the right side, so the half-reaction at -0.74v on the table must be reversed. The sign on the  $E^0$  is also reversed:



In order for these half-reactions to add up to the overall reaction, the first reaction must be multiplied by 3 and the second one by 2 so that the  $\text{e}^-$ 's cancel out. REMEMBER: When a half-reaction is multiplied, the value of the  $E^0$  does NOT change!



47. During the corrosion of magnesium, the anode reaction is

- A.  $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$   
 B.  $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$   
 C.  $4\text{OH}^- \rightarrow \text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^-$   
 D.  $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$

The Answer is **A**. EXPLANATION:

Corrosion of a metal always involves the *oxidation* of that metal. The *oxidation of Mg* is shown in answer “A”. We can eliminate answers “B” and “E” because they are *reductions* (the  $\text{e}^-$ 's are on the left). “C” is the oxidation of  $\text{OH}^-$ , which is NOT the corrosion of magnesium.

48. A molten binary salt,  $\text{ZnCl}_2$ , undergoes electrolysis. The cathode reaction is

- A.  $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$   
 B.  $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$   
 C.  $\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$   
 D.  $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$

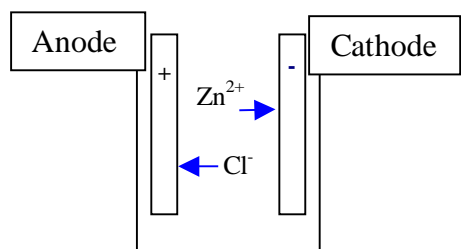
The Answer is **D**. EXPLANATION:

When a binary salt is melted, it dissociates into its ions:  $\text{ZnCl}_2 \rightarrow \text{Zn}^{2+} + 2\text{Cl}^-$   
 These ions are in *liquid* form so they are mobile.

Remember in an electrolytic cell, the **Anode is +** and the **Cathode is -**.

Also: The **Anode** is where **Oxidation** takes place and the **Cathode** is where **Reduction** takes place. (LEOA and GERC)

Since the salt is molten, there is *no water* present to be oxidized or reduced, so all that is oxidized and reduced are the ions in the salt.



The  $\text{Zn}^{2+}$  ions, being positive are attracted to the negative cathode. At the cathode, they are reduced:  
 $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$  (Cathode reaction)

The  $\text{Cl}^-$  ions, being negative are attracted to the positive anode. At the anode, they are oxidized to  $\text{Cl}_2$ :  
 $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$  (Anode reaction)

**This is the end of the multiple choice section of this examination. Make sure you obtain the key from the teacher or the URL:**

**<http://www.bced.gov.bc.ca/exams/search/exsection.htm> . Go over the key answers carefully and ASK the teacher if you have any problem understanding their solution!**

**Good luck on your final exam! Work hard and keep at it!**