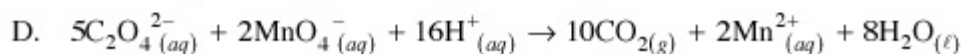
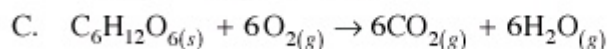
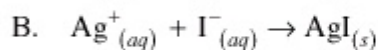
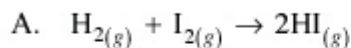


APRIL 2000
 PROVINCIAL EXAMINATION
Multiple Choice Explanations

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. At 25°C, which of the following reactions is fastest?

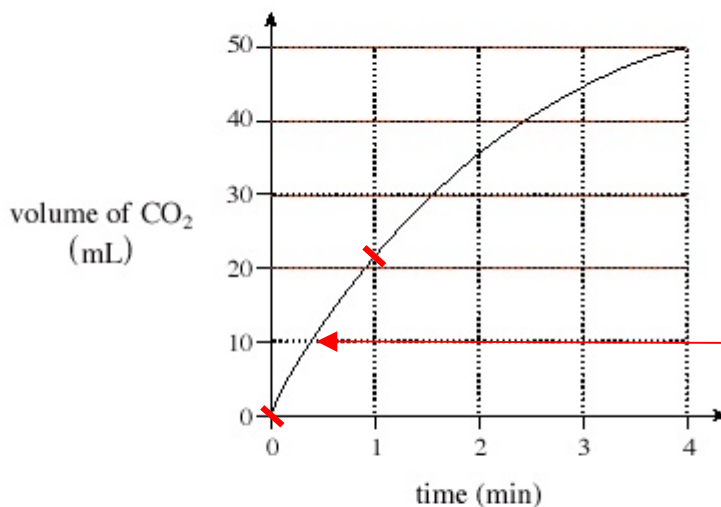
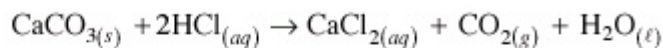


The Answer is **B**

EXPLANATION:

“B” is the fastest reaction because aqueous ions are highly mobile and more concentrated than molecules in a gas. The aqueous ions have a high probability of colliding. Also, when Ag^+ and I^- react, there are **no** bonds to break. They simply collide and bond to form the solid. (The activation energy is very low.) “A” is slow at room temperature because the *covalent* bonds in the diatomic molecules of H_2 and I_2 must be broken before they can react to form HI. This takes energy (fairly high activation energy), and is therefore slow at room temperature. In answer “C”, many C—C covalent bonds (in the sugar) must be broken (and also the diatomic bond in O_2). Covalent bonds require a high activation energy and are therefore slow to break. Answer “D” is slow at room temperature because there are 23 reacting particles, therefore the reaction would require numerous steps and hence is quite slow. (“B” is a 2 particle collision which takes place in a single step).

2. Consider the graph for the following reaction:



The average rate of reaction is greatest in the time interval

- A. 0 – 1 minute.
- B. 0 – 2 minutes.
- C. 0 – 3 minutes.
- D. 0 – 4 minutes.

The Answer is **A**

EXPLANATION:

Rate is defined as $\Delta\text{amount}/\Delta\text{time}$, which in this case would be $\Delta\text{Volume CO}_2/\Delta\text{time}$. Looking at the graph above, so can see that $\Delta\text{Volume CO}_2/\Delta\text{time}$ is just the slope of the graph. The slope of the graph is *steepest* in the time interval between 0 and 1 minute. Therefore the rate of the reaction is fastest during this interval. This makes sense because generally as a reaction proceeds (starting with reactants), the rate *starts out fast* and then *slows down* as the reactants are used up and their concentration(s) decrease.

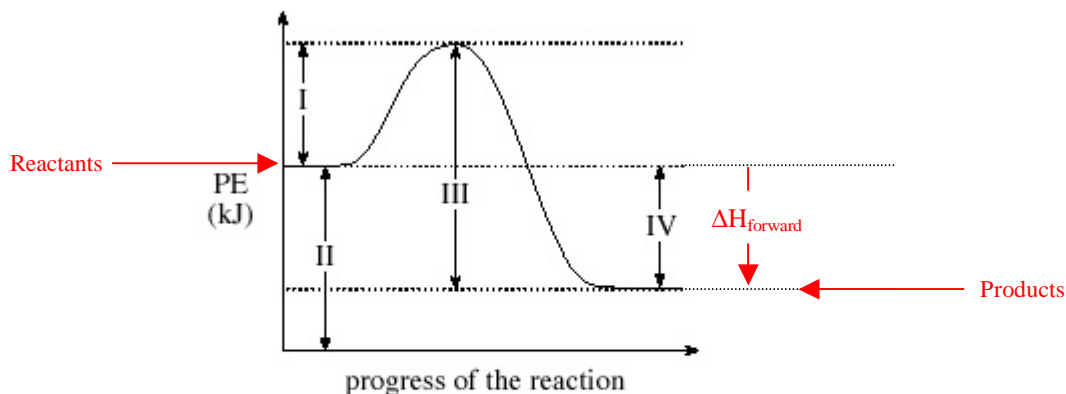
3. Which of the following changes will increase the average kinetic energy of reactant molecules?
- adding a catalyst
 - increasing the temperature
 - increasing the surface area
 - increasing the concentration

The Answer is **B**

EXPLANATION:

The *only* thing that can change the average kinetic energy of molecules is a change in the *temperature*. Adding a *catalyst* simply decreases the *activation energy* (minimum energy required for a successful collision) by providing an *alternate mechanism* (pathway) for the reaction. Increasing the *surface area* allows for *more possible collisions* between reactant molecules in *two different phases* (heterogeneous). Increasing the *concentration*, again *increases the probability of collisions* but does not affect the kinetic energy.

4. Consider the following potential energy diagram:



Which of the following represents the heat of reaction, ΔH , for the forward reaction?

- I
- II
- III
- IV

The Answer is **D**

EXPLANATION:

Looking at the red markers on the graph above, the ΔH for the forward reaction is the net energy difference between the products and the reactants (arrow IV). Because PE or enthalpy is *decreasing* as the reaction proceeds from left to right, the value of $\Delta H_{\text{forward}}$ would be a *negative* number (exothermic). The $\Delta H_{\text{reverse}}$ would be the same amount but a *positive* value (endothermic). Just to review, arrow I is $E_{a(\text{forward})}$, arrow II is really nothing because the starting point of the graph is arbitrary. (Only the energy differences among reactants, products and activated complex is relevant in these PE graphs.) Arrow III is $E_{a(\text{reverse})}$.

5. When a catalyst is added to a reaction

I.	the heat of reaction increases
II.	a new mechanism is provided
III.	the equilibrium constant increases

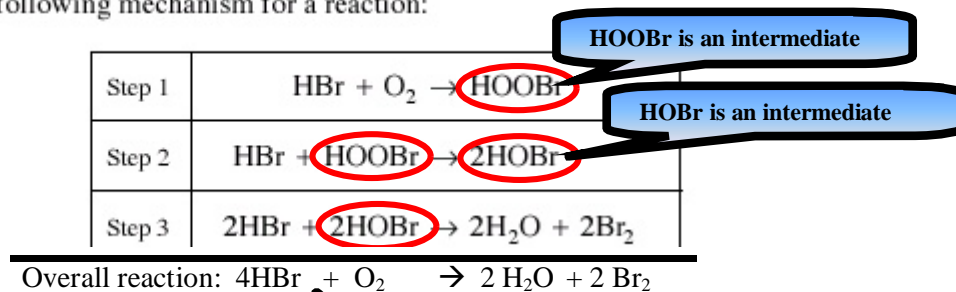
- A. II only
- B. I and II only
- C. II and III only
- D. I, II and III

The Answer is **A**

EXPLANATION:

A catalyst only provides an *alternate reaction mechanism* which has a *lower activation energy*. The heat of reaction (ΔH) is the net energy difference between the reactants and the products. The energies of the reactants and the products *cannot* be changed with a catalyst, so ΔH cannot be changed with a catalyst. Since a catalyst lowers both the $E_{a(\text{forward})}$ and the $E_{a(\text{reverse})}$ by the same amount, both the forward and the reverse reactions will be speeded up by the catalyst. Therefore the equilibrium will *not* shift and the equilibrium constant (K_{eq}) will remain the same. Remember, the **only** thing that can change the value of K_{eq} is the **temperature** (or when the Ministry of Education comes out with a new Data Booklet for Chemistry 12!).

6. Consider the following mechanism for a reaction:



Which of the following statements is correct?

- A. Br_2 is a reactant.
- B. HBr is a product.
- C. HOBr is a catalyst.
- D. HOBr is a reaction intermediate.

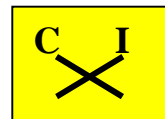
HBr and O_2 are reactants

H_2O and Br_2 are products

The Answer is **D**

EXPLANATION:

Remember that a reaction intermediate is produced in one step and used up in a later step. (So it is on the right in a step and on the left in a step somewhere below). If you look at the question above you will see that HOBr and HOBr are both reaction intermediates. When the reaction intermediates are cancelled out and the rest of the equations are added, you obtain the overall reaction which shows the reactants and the products. Remember that a catalyst is used up (on the left) of one step and regenerated (on the right) in a later step. One way to remember this is the CI(A) and the "X Files":



7. Which of the following applies to a chemical equilibrium?

I.	Forward and reverse reaction rates are equal
II.	Equilibrium can be achieved from either direction
III.	Macroscopic properties are constant

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

The Answer is **D**

EXPLANATION:

According to the Chemistry 12 PLO's from the Ministry:

It is expected that students will:

D4. describe chemical equilibrium as a closed system at constant temperature:

1. whose macroscopic properties are constant
2. where the forward and reverse reaction rates are equal
3. that can be achieved from either direction
4. where the concentrations of reactants and products are constant

You can see that I, II and III in the box above are contained in the characteristics of equilibrium as described by the Ministry. Recall that concentrations of reactants and products are *constant* but NOT usually equal to each other. *Macroscopic* means *observable*.

8. In which of the following will the driving forces of minimum enthalpy and maximum entropy oppose one another?

- | | |
|--|-------------------------------|
| A. $2\text{C}_{(s)} + \text{O}_{2(g)} \rightarrow 2\text{CO}_{(g)}$ | $\Delta H = -221 \text{ kJ}$ |
| B. $2\text{N}_{2(g)} + \text{O}_{2(g)} \rightarrow 2\text{N}_2\text{O}_{(g)}$ | $\Delta H = +164 \text{ kJ}$ |
| C. $2\text{CO}_{(g)} + \text{O}_{2(g)} \rightarrow 2\text{CO}_{2(g)}$ | $\Delta H = -566 \text{ kJ}$ |
| D. $4\text{CO}_{2(g)} + 6\text{H}_2\text{O}_{(g)} \rightarrow 2\text{C}_2\text{H}_{6(g)} + 7\text{O}_{2(g)}$ | $\Delta H = +3122 \text{ kJ}$ |

The Answer is **C**

EXPLANATION:

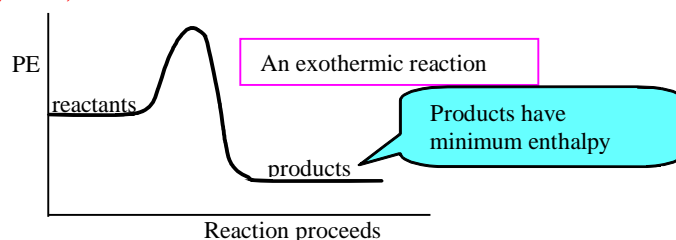
For the answers "A" and "C" in which ΔH is negative, the reactions are *exothermic* and the tendency toward *minimum enthalpy* favors the *products*:

Reaction "A" has 1 mole of gas (mog) in the reactants and 2 mog's in the products, so *products* have *maximum entropy*.

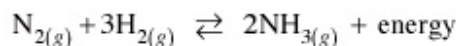
Reaction "C" has 3 mog's in the reactants and 2 mog's in the products, so *maximum entropy* favors the *reactants*. So in reaction "C" *minimum enthalpy favors the products* and *maximum entropy favors the reactants* so these two driving forces are **opposing** each other.

If you check reaction "B" so will see that because it is *endothermic*, *minimum enthalpy* favors the *reactants*, and because it goes from 3 mog's to 2 mog's, *maximum entropy* also favors the *reactants*.

In reaction "D", *minimum enthalpy* favors the *reactants* (*endothermic*) and *maximum entropy* also favors the *reactants*. (10 mog's in reactants and 9 mog's in products.)



10. Ammonia, NH_3 , is produced by the following reaction:



Which of the following would result in the highest concentration of ammonia at equilibrium?

- A. increasing the temperature and increasing the pressure
- B. decreasing the temperature and increasing the pressure
- C. increasing the temperature and decreasing the pressure
- D. decreasing the temperature and decreasing the pressure

The Answer is **B**

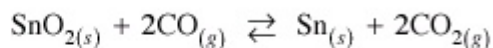
EXPLANATION:

To get the highest concentration of NH_3 at equilibrium, we want the reaction to be shifted as far as possible to the **right**.

Since the energy (heat) term is on the right side, *decreasing the temperature* (taking away heat) would cause a shift to the **right** according to LCP (LeChatelier's Principle), so we can eliminate "A" and "C" as answers.

There are 4 moles of gas (mog) on the left side and 3 mog on the right side. *Increasing the pressure* would cause the reaction to shift to the **right** (least mog) according to LCP. So answer "B" would be the correct one.

11. What is the K_{eq} expression for the following reaction?



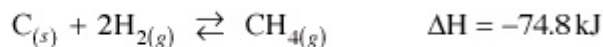
- A. $K_{eq} = \frac{[\text{CO}_2]}{[\text{CO}]}$
- B. $K_{eq} = \frac{[\text{CO}_2]^2}{[\text{CO}]^2}$
- C. $K_{eq} = \frac{[\text{Sn}][\text{CO}_2]^2}{[\text{CO}]^2}$
- D. $K_{eq} = \frac{[\text{Sn}][\text{CO}_2]^2}{[\text{SnO}_2][\text{CO}]^2}$

The Answer is **B**

EXPLANATION:

The $\text{SnO}_{2(s)}$ and the $\text{Sn}_{(s)}$ are left out of the K_{eq} expression because they are solids and have a constant concentration. Remember the products are on top and the reactants are on the bottom of the expression. Also, don't forget that a coefficient in the balanced equation becomes an exponent in the K_{eq} expression!

12. Consider the following reaction:



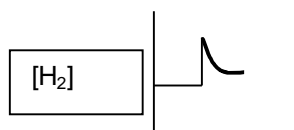
Which of the following will cause an increase in the value of K_{eq} ?

- A. increasing $[\text{H}_2]$
- B. decreasing the volume
- C. finely powdering the $\text{C}_{(s)}$
- D. decreasing the temperature

The Answer is **D**

EXPLANATION:

Be careful with this one! The changes in "A" and "B" will cause the equilibrium to shift to the *right* but this does NOT alter the value of K_{eq} ! When $[\text{H}_2]$ is increased, the equilibrium shifts to the right just enough to compensate and keep the value of K_{eq}



constant. If we decrease the volume (increase the pressure) as in answer "B", again the equilibrium shifts to counteract our change and keep the K_{eq} value constant. Powdering the solid C as in answer "C" will increase the *surface area* and hence the *rate* of the reaction. But it increases *both* the forward and the reverse rate, giving no difference in the equilibrium mixture and hence does not affect K_{eq} .

The reaction is *exothermic* and can be written with the heat term on the right:



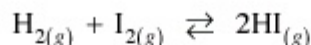
Taking away heat (decreasing the temperature) will cause this reaction to shift to the *right*. This will cause $[\text{CH}_4]$ to increase and $[\text{H}_2]$ to decrease. The K_{eq} expression is:

$$K_{eq} = \frac{[\text{CH}_4]}{[\text{H}_2]^2}$$

So this will cause the ratio (and the value of K_{eq}) to increase.

Remember the temperature is the only thing that can change K_{eq} !

13. Consider the following equilibrium:



At equilibrium $[\text{H}_2] = 0.00220 \text{ mol/L}$, $[\text{I}_2] = 0.00220 \text{ mol/L}$ and $[\text{HI}] = 0.0156 \text{ mol/L}$.
The value of K_{eq} is

- A. 3.10×10^{-4}
- B. 1.99×10^{-2}
- C. 5.03×10^1
- D. 3.22×10^3

The Answer is **C**

EXPLANATION:

Notice that all concentrations given are at *equilibrium*. Therefore we just write the K_{eq} expression and insert the *equilibrium concentration* values. (This is not an “ICE” problem!)

$$K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{[0.0156]^2}{[0.00220][0.00220]} = 5.03 \times 10^1$$

14. Which of the following will dissolve in water to produce a molecular solution?

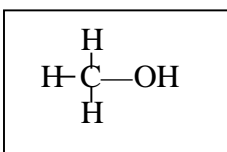
- A. CaCl₂
- B. NaOH
- C. CH₃OH
- D. Sr(OH)₂

Ca, Na and Sr are all metals,
so these compounds are ionic.

The Answer is **C**

EXPLANATION:

“A”, “B” and “D” are all *ionic* because they have a *metal* in the formula. “B” and “D” are *bases* because they are *ionic* and form the OH⁻ ion in aqueous solution. When -OH is bonded directly to a *carbon* (C) atom, the bond is a *strong covalent* bond and does not break to release OH⁻ ions in water.



Remember that compounds with “C’s and “H’s are *molecular* except for ones that end with the *carboxyl group* (COOH) or (COO⁻). CH₃OH is an example of an *alcohol*, which is a *molecular* compound. Go back and read Tutorial 7.

April 2000 Chemistry 12 Provincial

Multiple Choice Explanations

15. In a solubility equilibrium, the
- rate of dissolving equals the rate of crystallization.
 - neither dissolving nor crystallization are occurring.
 - concentration of solute and solvent are always equal.
 - mass of dissolved solute is greater than the mass of the solution.

The Answer is **A**

EXPLANATION:

Answer "A" is the primary condition for solubility equilibrium. Remember the solution is also called "saturated". "B", of course is wrong because these opposing processes are continuing to happen at equilibrium (remember "dynamic"), but their rates are equal so the concentration of the dissolved ions remains constant. Looking at answer "C", a solute is a solid (*remember by definition there has to be some solid present to have a true saturated solution*). The concentration of a solid is huge (*ions are packed tightly together*). There is also *no relation* between the concentration of a solute and of a solvent. Hopefully you can see that "D" is impossible because the dissolved solute is only part of the solution and the mass of a part cannot be greater than the mass of the whole.

16. Which of the following solutions would have $[\text{Fe}^{3+}] = 0.020 \text{ M}$?

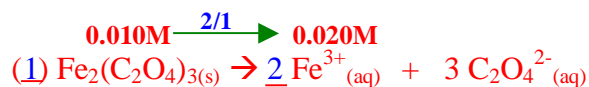
- 0.40 L of 0.050 M $\text{Fe}(\text{NO}_3)_3$
- 0.80 L of 0.020 M $\text{Fe}_2(\text{SO}_4)_3$
- 0.50 L of 0.040 M $\text{FeC}_6\text{H}_5\text{O}_7$
- 0.50 L of 0.010 M $\text{Fe}_2(\text{C}_2\text{O}_4)_3$

These are irrelevant!

The Answer is **D**

EXPLANATION:

I hope you can see the devious trick in this one! Since they're asking for the concentration of Fe^{3+} and the concentrations of the solutions are given (eg. 0.050 M), the volumes given for each solution (eg. 0.40 L) are completely irrelevant! If we do a balanced dissociation equation for "D":



With the same method, if you dissociate the equation:

- in "A" $\rightarrow [\text{Fe}^{3+}] = 0.050 \text{ M}$
- in "B" $\rightarrow [\text{Fe}^{3+}] = 0.040 \text{ M}$
- in "C" $\rightarrow [\text{Fe}^{3+}] = 0.040 \text{ M}$

Make sure you can figure out how I got these.

17. Which of the following substances has the lowest solubility?

- A. BaS
- B. CuS
- C. FeS
- D. ZnS

The Answer is **B**

EXPLANATION:

Whenever you're asked to make any comparisons of solubility you should immediately turn to your solubility table and your Ksp table in the Data booklet. You may need to use one or both of these tables.

If you check your solubility table, you will find that BaS is soluble and the others all have low solubilities. So "A" cannot be our answer.

Since CuS, FeS and ZnS all have low solubilities, we use the Ksp table to find out which one is lowest.

Remember that if all compounds are the same type (AB or AB₂) (These compounds are all type "AB" → ie. they have one of each ion) :

- The higher the Ksp, the higher the solubility.

So looking up the Ksp's on the table:

Compound	Ksp
CuS	6.0×10^{-37}
FeS	6.0×10^{-19}
ZnS	2.0×10^{-25}

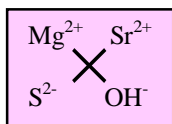
We can see that CuS (Answer "B") has the lowest Ksp and therefore the lowest solubility.

18. The complete ionic equation for the reaction between MgS and Sr(OH)₂ is

- A. $\text{MgS}_{(aq)} + \text{Sr}(\text{OH})_{2(aq)} \rightarrow \text{Mg}(\text{OH})_{2(s)} + \text{SrS}_{(s)}$
- B. $\text{MgS}_{(aq)} + \text{Sr}(\text{OH})_{2(aq)} \rightarrow \text{Mg}(\text{OH})_{2(s)} + \text{SrS}_{(aq)}$
- C. $\text{Mg}^{2+}_{(aq)} + \text{S}^{2-}_{(aq)} + \text{Sr}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Mg}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} + \text{SrS}_{(s)}$
- D. $\text{Mg}^{2+}_{(aq)} + \text{S}^{2-}_{(aq)} + \text{Sr}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Mg}(\text{OH})_{2(s)} + \text{Sr}^{2+}_{(aq)} + \text{S}^{2-}_{(aq)}$

The Answer is **D**

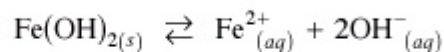
EXPLANATION:



If you use an "ion-box", you will see that the two possible products are SrS and Mg(OH)₂. Looking these up on the solubility table we see that SrS is *soluble* and Mg(OH)₂ has *low solubility*. Therefore Mg(OH)₂ will be the *solid* (precipitate) and SrS will be *aqueous* and dissociated. You have to be careful how you read this question. They are asking for the *complete ionic equation*.

Remember that "A" is not correct because they have the SrS as a solid. "B" is the correct *formula equation* but that's not what they're asking for. "C" has the wrong precipitate and "D" is the *correct complete ionic equation*. Recall that if they asked for the *net ionic equation*, it would be: $\text{Mg}^{2+}_{(aq)} + 2\text{OH}^{-}_{(aq)} \rightarrow \text{Mg}(\text{OH})_{2(s)}$

19. Consider the following equilibrium:



Which of the following will cause the equilibrium to shift to the right?

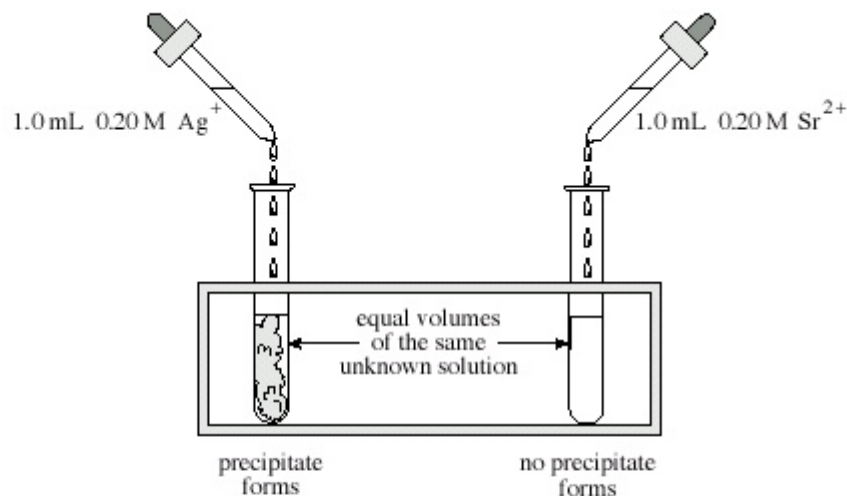
- A. adding KOH
- B. adding Na_2S
- C. adding Fe(OH)_2
- D. adding $\text{Fe(NO}_3)_2$

The Answer is **B**

EXPLANATION:

“A” and “C” both dissociate to form OH^- ions. Increasing the $[\text{OH}^-]$ will shift the equilibrium to the *left* due to LeChatelier’s Principle (LCP). Remember, this is called the *Common Ion Effect*. “D” dissociates into Fe^{2+} and NO_3^- . Increasing the $[\text{Fe}^{2+}]$ will again cause a shift to the *left* due to LCP. “B” dissociates into Na^+ and S^{2-} . Looking at the solubility table beside “Sulphide”, we see that S^{2-} forms a *precipitate* with Fe^{2+} . ($\text{Fe}^{2+}_{(aq)} + \text{S}^{2-}_{(aq)} \rightarrow \text{FeS}_{(s)}$). The addition of Na_2S , therefore will *decrease* the $[\text{Fe}^{2+}]$ and cause the equilibrium above to shift to the *right*. This, of course will *increase* the *solubility* of the $\text{Fe(OH)}_{2(s)}$.

20. Consider the following experiment:



The unknown solution could contain

- A. 0.20 M OH^-
- B. 0.20 M NO_3^-
- C. 0.20 M PO_4^{3-}
- D. 0.20 M SO_4^{2-}

The Answer is **A**

EXPLANATION:

The diagram in this question simply gives us the information that our unknown forms a *precipitate* with Ag^+ but *not* with Sr^{2+} . So again we consult our trustworthy old *solubility table* and go through the suggested answers. Looking at “Hydroxide”, we see that OH^- forms a precipitate (compound of low solubility) with Ag^+ but not with Sr^{2+} . Therefore, *this* is our answer. In “B” NO_3^- does not precipitate with *any* cation. In “C”, PO_4^{3-} precipitates *both* Ag^+ and Sr^{2+} . In answer “D”, SO_4^{2-} also precipitates *both* Ag^+ and Sr^{2+} .

21. A compound has a solubility of $7.1 \times 10^{-5} \text{ M}$ at 25°C . The compound is

- A. CuS
- B. AgBr
- C. CaCO_3
- D. CaSO_4

The Answer is **C**

EXPLANATION:

Please, NEVER confuse the solubility (which is the moles of a compound which will dissolve in a Litre of solution) with the Ksp (the solubility product). We first have to calculate the Ksp of the compound and then see which one matches on the Ksp table.

The answers are all type AB compounds (one of each ion), so we can temporarily call our mystery compound $\text{AB}_{(s)}$

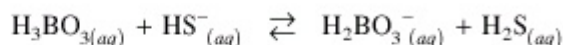
Equation for the Solubility equilibrium: The " $7.1 \times 10^{-5} \text{ M}$ " written on top is the molar solubility: (" $7.1 \times 10^{-5} \text{ M}$ " is written above the solid AB because the solubility is how many moles of AB will dissolve per Litre. Remember the actual concentration of a solid doesn't change.)



The Ksp expression is: $\text{Ksp} = [\text{A}^{2+}] [\text{B}^{2-}]$
 So $\text{Ksp} = (7.1 \times 10^{-5})^2$
 $\text{Ksp} = 5.0 \times 10^{-9}$

Looking on the Ksp table, 5.0×10^{-9} is the Ksp for CaCO_3 . So "C" is the correct answer.

22. Consider the following reaction:



The order of Brønsted-Lowry acids and bases in this equation is

- A. acid, base, base, acid.
- B. acid, base, acid, base.
- C. base, acid, acid, base.
- D. base, acid, base, acid.

The Answer is **A**

EXPLANATION:

Remember a conjugate acid has one more H and one more "+" charge than it's conjugate base. So in one conjugate acid/base pair, H_3BO_3 is the conjugate acid and H_2BO_3^- is it's conjugate base. In the other conjugate acid/base pair, H_2S is the conjugate acid and HS^- is it's conjugate base. So the order in this case is A B B A.

23. The conjugate base of an acid is produced by
- adding a proton to the acid.
 - adding an electron to the acid.
 - removing a proton from the acid.
 - removing an electron from the acid.

The Answer is **C** EXPLANATION:

To find the conjugate base of any species, remove an "H" and remove one "+" charge. (in other words remove an "H⁺" or a proton.)

24. Which of the following represents the predominant reaction between HCO_3^- and water?
- $2\text{HCO}_3^- \rightarrow \text{H}_2\text{O} + 2\text{CO}_2$
 - $\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3 + \text{OH}^-$
 - $\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{CO}_3^{2-}$
 - $2\text{HCO}_3^- + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{CO}_3^{2-} + \text{OH}^- + \text{CO}_2$

The Answer is **B** EXPLANATION:

The ion HCO_3^- starts with an "H" and ends with a negative charge. Remember this means it is *amphiprotic* (can act as an acid and donate a proton or act as a base and accept a proton). If you look up HCO_3^- on the acid table, you will find it on **both** sides. To find the *predominant hydrolysis reaction* of an *amphiprotic ion*, find its K_a and its K_b and compare them.

To find the K_a , look for HCO_3^- on the *left* side and read the K_a on the right:

$$K_a(\text{HCO}_3^-) = 5.6 \times 10^{-11}$$

To find the K_b , look up HCO_3^- on the *right* side, find the K_a for its conjugate acid beside it: (4.3×10^{-7}). Then use the formula:

$$K_b = K_w / K_a(\text{conjugate acid})$$

$$\text{So } K_b(\text{HCO}_3^-) = \frac{1.0 \times 10^{-14}}{4.3 \times 10^{-7}} = 2.3 \times 10^{-8}$$

So we see that the K_b of HCO_3^- is bigger than its K_a . Therefore the predominant hydrolysis is **base** hydrolysis. The HCO_3^- *accepts a proton* from water, forming H_2CO_3 and OH^- . The equilibrium equation is: $\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{CO}_3 + \text{OH}^-$. Remember the term "hydrolysis" means reacting with water to form H_3O^+ (*acid hydrolysis*) or OH^- (*base hydrolysis*).

25. Water acts as an acid when it reacts with which of the following?

When water act AS and acid, the other species is acting AS A BASE in water!

I.	CN ⁻	weak base
II.	NH ₃	weak base
III.	HClO ₄	strong acid
IV.	CH ₃ COO ⁻	weak base

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

The Answer is **C**

EXPLANATION:

Using your acid table, you can see that CN⁻ and CH₃COO⁻ are found only on the *right* side of the table, so they are *weak bases*. The NH₃ is found on both sides, but remember the NH₃ on the bottom of the left does not include NH₃ as an acid. It never acts as an acid in aqueous solution (it is only there because it is the conjugate acid of NH₂⁻.) So NH₃ acts as a *base* in water. The HClO₄ is on the top of the left side, so it is a *strong acid*. So the three bases are I, III and IV. It is with these *bases* that *water acts as an acid*.

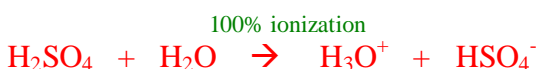
26. In a solution of 0.10 M H_2SO_4 , the ions present in order of decreasing concentration are

- A. $[\text{H}_3\text{O}^+] > [\text{HSO}_4^-] > [\text{SO}_4^{2-}] > [\text{OH}^-]$
 B. $[\text{H}_3\text{O}^+] > [\text{SO}_4^{2-}] > [\text{HSO}_4^-] > [\text{OH}^-]$
 C. $[\text{OH}^-] > [\text{SO}_4^{2-}] > [\text{HSO}_4^-] > [\text{H}_3\text{O}^+]$
 D. $[\text{SO}_4^{2-}] > [\text{HSO}_4^-] > [\text{OH}^-] > [\text{H}_3\text{O}^+]$

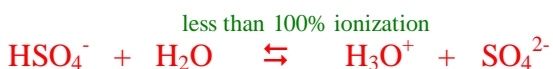
The Answer is **A**

EXPLANATION:

H_2SO_4 is a strong acid. That means it is 100% ionized in aqueous solution. So a solution which is 0.10M H_2SO_4 ionizes completely into H_3O^+ and HSO_4^- :



When this happens, there are *no* H_2SO_4 molecules left and the $[\text{H}_3\text{O}^+]$ and the $[\text{HSO}_4^-]$ are both 0.10 M. However, this is not all that happens. If you look on the left side of the acid table, you will see that HSO_4^- is a *weak acid*. Therefore, it is *partially ionized* to form more H_3O^+ . Remember when a weak acid ionizes, most of it remains as the acid and only a *small fraction* of it ionizes. Notice the *double arrow*:

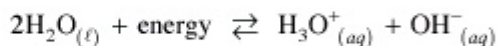


As this occurs, the HSO_4^- which was 0.10M will go *down* a small amount as some of it is ionized. The $[\text{H}_3\text{O}^+]$ will go up slightly as a small amount of it is *formed* from the ionization of HSO_4^- . So in the final result, the $[\text{H}_3\text{O}^+]$ will be slightly higher than the $[\text{HSO}_4^-]$. Also there will be a small amount of SO_4^{2-} formed as a result of the ionization of HSO_4^- . So in this solution the $[\text{H}_3\text{O}^+]$ is a little higher than 0.10M. If it was 0.10M, we could find the $[\text{OH}^-]$ using the K_w expression ($[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$):

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{0.10} = 1.0 \times 10^{-13} \text{ M} \text{ which is } \textit{extremely low}. \text{ (Much lower than all the other ions in this solution!)}$$

So the ions in order of most to least concentrated are: $[\text{H}_3\text{O}^+] > [\text{HSO}_4^-] > [\text{SO}_4^{2-}] > [\text{OH}^-]$

27. Consider the following equilibrium:



What will cause the pH to increase and the K_w to decrease?

- A. adding a strong acid
- B. adding a strong base
- C. increasing the temperature
- D. decreasing the temperature

The Answer is **D**

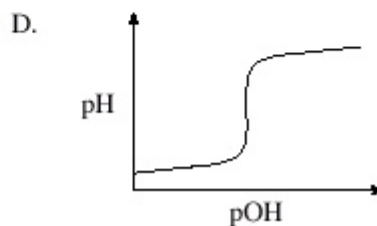
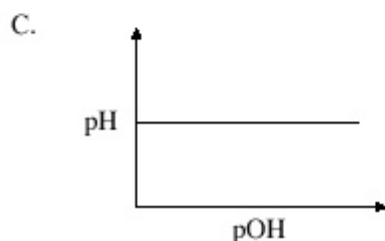
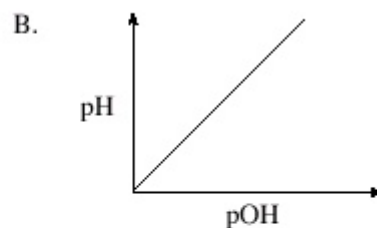
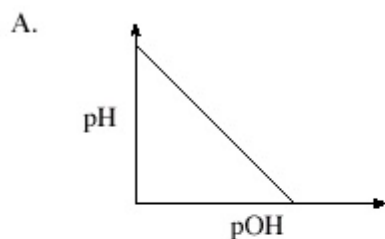
EXPLANATION:

Don't forget, when the pH *increases*, it means the $[\text{H}_3\text{O}^+]$ is *decreasing*. ($\text{pH} = -\log [\text{H}_3\text{O}^+]$). So in order to increase the pH, we must *decrease* the $[\text{H}_3\text{O}^+]$. Adding a strong acid as in "A" will only *increase* $[\text{H}_3\text{O}^+]$, so this is not correct. Adding a strong base as in "B" will cause the $[\text{H}_3\text{O}^+]$ to decrease but it will **NOT** change the value of K_w ! (Make sure you carefully read the whole question!) **REMEMBER, THE ONLY THING THAT CAN CHANGE THE VALUE OF A K_{eq} IS THE TEMPERATURE!!!**

Remember the K_w expression is: $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$

If we *decrease* the temperature as in "D", the equilibrium above will shift to the *left* (*the heat term is on the left*). This will *decrease* $[\text{H}_3\text{O}^+]$ *and* $[\text{OH}^-]$. Therefore it will increase pH and decrease the value of K_w . So this is our answer. (Note: Because $[\text{OH}^-]$ decreases, the pOH will also increase.)

28. Which of the following graphs describes the relationship between pH and pOH?



The Answer is **A** EXPLANATION:

Don't be frightened by these graphs! Just know your equations well. Remember that: $\text{pH} + \text{pOH} = \text{pK}_w$. K_w is a *constant* at constant temperature, so pK_w must also be *constant*. There is no mention of temperature so we can assume that it is held constant. Since $\text{pH} + \text{pOH} = \text{a constant}$, if pH *increases*, that means that pOH *decreases*. The only graph which shows this relationship is graph "A". (Notice that it is a "+", not a "x", so it is not an *inverse* relation (which would be a curve))

29. The relationship $\frac{[\text{H}_3\text{BO}_3][\text{OH}^-]}{[\text{H}_2\text{BO}_3^-]}$ is the expression for

- A. K_a for H_3BO_3
- B. K_b for H_3BO_3
- C. K_a for H_2BO_3^-
- D. K_b for H_2BO_3^-

The Answer is **D** EXPLANATION:

Remember in any K_{eq} expression, the products are on top and reactants are on the bottom. K_a and K_b are K_{eq} 's for the reactions of species with H_2O to produce H_3O^+ or OH^- . So the reaction would be: $\text{H}_2\text{BO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{BO}_3 + \text{OH}^-$. You can see in this reaction that the H_2BO_3^- is accepting a proton from the water, so it is undergoing base hydrolysis. The expression, therefore is the K_b for H_2BO_3^- , which is answer "D". If it was a K_a expression, there would be H_3O^+ on top.

April 2000 Chemistry 12 Provincial

Multiple Choice Explanations

30. The value of K_b for H_2PO_4^- is

- A. 1.3×10^{-12}
- B. 6.2×10^{-8}
- C. 1.6×10^{-7}
- D. 7.5×10^{-3}

The Answer is **A**

EXPLANATION:

To find the K_b , look up H_2PO_4^- on the *right* side, find the K_a for its conjugate acid beside it: (7.5×10^{-3}). Then use the formula:

$K_b = K_w / K_a(\text{conjugate acid})$

$$\text{So } K_b(\text{H}_2\text{PO}_4^-) = \frac{1.0 \times 10^{-14}}{7.5 \times 10^{-3}} = 1.3 \times 10^{-12} \quad (\text{which is answer "A"})$$

31. Which of the following solutions has a pH less than 7.00?

- A. NaCl
- B. LiOH
- C. NH_4NO_3
- D. KCH_3COO

The Answer is **C**

EXPLANATION:

This question is about *hydrolysis*. These are all *ionic* compounds.

"B" is a *base* (ends in "OH") and "A", "C" and "D" are all *salts*. They are asking to identify the one with a $\text{pH} < 7.00$. In other words, the one which is *acidic*. "B" is definitely not the answer because LiOH (an alkali metal hydroxide) is a *strong base*. For "A", "C" and "D" we use the process:

Dissociate → Eliminate(spectators) → Evaluate (the remaining species to see if they are acidic or basic)

$\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$ (both ions are neutral spectators so NaCl is neutral or $\text{pH} = 7.00$)

$\text{NH}_4\text{NO}_3 \rightarrow \text{NH}_4^+ + \text{NO}_3^-$ (NH_4^+ is a weak acid (found on the left side of the acid table) and NO_3^- is a neutral spectator, so NH_4NO_3 is acidic or $\text{pH} < 7.00$)

$\text{KCH}_3\text{COO} \rightarrow \text{K}^+ + \text{CH}_3\text{COO}^-$ (K^+ is a neutral spectator and CH_3COO^- is a weak base (found on the right side of the acid table), so KCH_3COO is basic or $\text{pH} > 7.00$)

So you can see that NH_4NO_3 (Answer "C") is the correct answer.

32. Which of the following will form a basic aqueous solution?

- A. HSO_3^-
- B. HSO_4^-
- C. HPO_4^{2-}
- D. HC_2O_4^-

The Answer is **C**

EXPLANATION:

These are all *amphiprotic anions*. Since this is a multiple choice question (only one correct answer) and it asks for the “basic” one, look all of these up on the *right* side of the acid table and find the one that is the “*strongest*” base, i.e. the one closest to the bottom of the right side of the table. The lowest one on the right side is HPO_4^{2-} . If any of them are going to act as a base in water it would be this one. To make sure, find it’s K_a and it’s K_b and compare them as shown below:

To find out whether an amphiprotic anion is acidic or basic in aqueous solution, we have to find the K_a and the K_b of each one of these and compare them. If $K_a > K_b$, the ion will be acidic and if $K_b > K_a$, it will be basic.

To find the K_a , look for the ion on the *left* side and read the K_a on the right:

eg.) $K_a(\text{HSO}_3^-) = 1.0 \times 10^{-7}$

To find the K_b , look up the ion on the *right* side, find the K_a for it’s conjugate acid beside it: (eg. K_a for the conjugate acid of $\text{HSO}_3^- = 1.5 \times 10^{-2}$). Then use the formula:

$K_b = K_w / K_a(\text{conjugate acid})$:

So $K_b(\text{HSO}_3^-) = \frac{1.0 \times 10^{-14}}{1.5 \times 10^{-2}} = 6.7 \times 10^{-13}$. So for HSO_3^- , $K_a > K_b$ so this ion is acidic.

Doing similar calculations for the other ions, we get the following results:

Ion	K_a	K_a (conj. acid)	K_b	Acidic or Basic
HSO_4^-	1.2×10^{-2}	Very large	So $K_b(\text{HSO}_4^-) = \frac{1.0 \times 10^{-14}}{\text{very large}} = \text{very small}$	$K_a > K_b$ (acidic)
HPO_4^{2-}	2.2×10^{-13}	6.2×10^{-8}	So $K_b(\text{HPO}_4^{2-}) = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7}$	$K_b > K_a$ (basic)
HC_2O_4^-	6.4×10^{-5}	5.9×10^{-2}	So $K_b(\text{HC}_2\text{O}_4^-) = \frac{1.0 \times 10^{-14}}{5.9 \times 10^{-2}} = 1.7 \times 10^{-13}$	$K_a > K_b$ (acidic)

So the only basic ion is the HPO_4^{2-} , which is answer “C”.

33. What is the approximate K_a value for the indicator chlorophenol red?

- A. 1×10^{-14}
- B. 1×10^{-8}
- C. 1×10^{-6}
- D. 1×10^{-3}

The Answer is **C**

EXPLANATION:

The equilibrium equation for any indicator can be written as:



At the *transition point* of the indicator $[\text{HInd}] = [\text{Ind}^-]$, so these cancel out of the K_a expression:

$$K_a = \frac{[\text{H}_3\text{O}^+]{\cancel{[\text{Ind}^-}]}}{\cancel{[\text{HInd}]}} = [\text{H}_3\text{O}^+]$$

Looking at the Indicator Table, chlorophenol red changes colour in the pH range (5.2-6.8)

The pH at the transition point (also called the pK_a) is calculated by: $\frac{5.2 + 6.8}{2} = 6.0$

Since the $\text{pH} = 6.0$, the $[\text{H}_3\text{O}^+] = \text{antilog}(-6.0) = 1 \times 10^{-6}$ and as it says three lines up $K_a = [\text{H}_3\text{O}^+]$ at the transition point.

34. What is the pH of the solution formed when 0.040 mol $\text{NaOH}_{(s)}$ is added to 2.00 L of 0.020 M HCl?

- A. 0.00
- B. 1.40
- C. 1.70
- D. 7.00

The Answer is **D**

EXPLANATION:

This is a problem where you have to find the pH of a mixture of a strong acid and strong base. Follow these steps:

1. Find the initial moles of both acid and base (if the mole ratio is not 1/1, convert these to moles of H^+ and moles of OH^- -remember that 1 mol of $\text{H}_2\text{SO}_4 \rightarrow 2$ moles of H^+)
2. Calculate the moles of the reactant in excess
3. Divide these moles by the total volume (M=moles/L) to calculate the $[\text{H}_3\text{O}^+]$ or $[\text{OH}^-]$
4. Use the result to find the pH

The mole ratio of NaOH to HCl is 1/1 so we don't have to go to H^+ and OH^- .

The initial moles of NaOH is already given 0.040 mol NaOH

Initial moles HCl = 0.020 M x 2.00L 0.040 mol HCl

Excess none

You can see that neither HCl nor NaOH is in excess (We have reached the *equivalence point!*). Since there is no NaOH or HCl left over, all we have is H_2O (neutral) and NaCl (a neutral salt). So the pH of the final mixture would be 7.00!

35. Which of the following titrations will always have an equivalence point at a pH > 7.00?
- weak acid with a weak base
 - strong acid with a weak base
 - weak acid with a strong base
 - strong acid with a strong base

The Answer is **C** EXPLANATION:
It really helps to know from your notes the following:

Type of Titration	pH at Equivalence Point
Strong Acid/Strong Base	= 7.0
Weak Acid/Strong Base	> 7.0
Strong Acid/Weak Base	< 7.0

One way to remember this is the stronger thing “dominates” over the other one. For example in a WA/SB titration, the base is stronger than the acid, so the pH is basic (>7).

This is *not really why* the results are obtained. Read the following:

When a SA and a SB are reacted, they form water and a neutral salt:

e.g.) $\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}$ so at EP, you only have H_2O and the neutral salt NaCl so the pH is 7.0

When a WA and a SB are reacted, they form water and a basic salt:

e.g.) $\text{HF} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaF}$ so at EP, you only have H_2O and the basic salt NaF so the pH is > 7.0

When a SA and a WB are reacted, they form an acidic salt:

e.g.) $\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4\text{Cl}$ so at EP, you only have NH_4^+ (a weak acid) and Cl^- (a neutral spectator), so the pH is < 7.0

36. A buffer solution may contain equal moles of
- weak acid and strong base.
 - strong acid and strong base.
 - weak acid and its conjugate base.
 - strong acid and its conjugate base.

The Answer is **C** EXPLANATION:
Remember, there are two types of *buffer* solutions:

- WASCB** (weak acid and (the salt of) it's conjugate base (eg. CH_3COOH and NaCH_3COO))
- WBSCA** (weak base and (the salt of) it's conjugate acid (eg. NH_3 and NH_4Cl))

Also remember: A buffer *cannot* be created using a strong acid or strong base!

Answers “A”, “B” and “D” all have SA's or SB's in them.

Answer “C” is a WASCB buffer. (the “salt” is optional-it's the *source* of the conjugate base of the weak acid)

37. A gas which is produced by burning coal and also contributes to the formation of acid rain is

- A. H_2
- B. O_3
- C. SO_2
- D. C_3H_8

The Answer is **C**

EXPLANATION:

This is a “knowledge” question. Read over your notes on Acid Rain and Acidic Anhydrides. When coal containing sulphur is burned, the following chain of reactions happens:

- $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$
- $2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$
- $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$ (sulphurous acid)
- $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$ (sulphuric acid)

38. Manganese has an oxidation number of +4 in

- A. MnO
- B. MnO_2
- C. Mn_2O_3
- D. Mn_2O_7

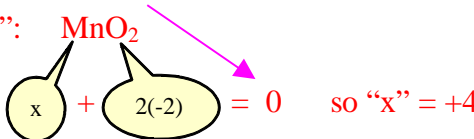
The Answer is **B**

EXPLANATION:

Remember that “oxygen” (ie. one “O” atom) in all compounds except H_2O_2 and Na_2O_2 has an oxidation number of “-2”. (O_2 in its elemental form has an oxidation number of “0”.)

Also the oxidation numbers all must add up to the total ionic charge (which in a neutral molecule is “0”).

So looking at answer “B”: MnO_2



$x + 2(-2) = 0$ so “x” = +4

For practice, work out the oxidation numbers of Mn in “A”, “C” and “D”

- In MnO the oxidation number of Mn is +2
- In Mn_2O_3 the oxidation number of Mn is +3
- In Mn_2O_7 the oxidation number of Mn is +7

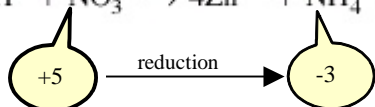
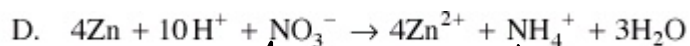
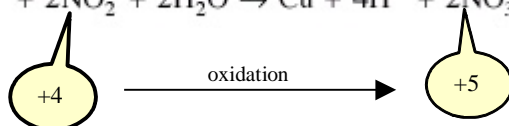
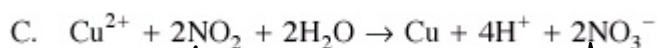
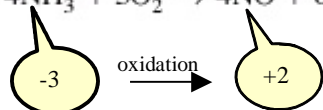
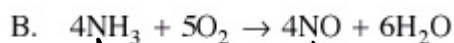
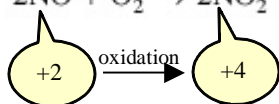
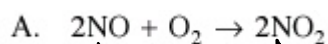
39. In which reaction is nitrogen reduced?

- A. $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$
 B. $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$
 C. $\text{Cu}^{2+} + 2\text{NO}_2 + 2\text{H}_2\text{O} \rightarrow \text{Cu} + 4\text{H}^+ + 2\text{NO}_3^-$
 D. $4\text{Zn} + 10\text{H}^+ + \text{NO}_3^- \rightarrow 4\text{Zn}^{2+} + \text{NH}_4^+ + 3\text{H}_2\text{O}$

The Answer is **D**

EXPLANATION:

Remember *reduction* means the *oxidation number is decreasing* (e^- s are being added)
 So just find the ΔON (change in oxidation number) on "N" in each of the reactions. Go back and review your rules for oxidation numbers. NOW!



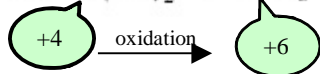
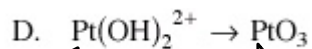
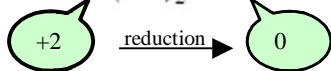
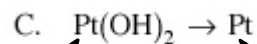
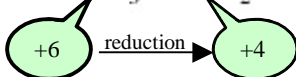
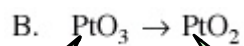
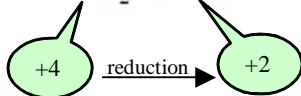
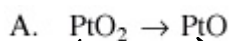
40. An oxidizing agent will cause which of the following changes?

- A. $\text{PtO}_2 \rightarrow \text{PtO}$
- B. $\text{PtO}_3 \rightarrow \text{PtO}_2$
- C. $\text{Pt(OH)}_2 \rightarrow \text{Pt}$
- D. $\text{Pt(OH)}_2^{2+} \rightarrow \text{PtO}_3$

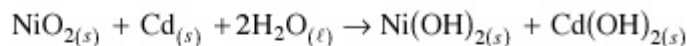
The Answer is **D**

EXPLANATION:

An *oxidizing agent* will oxidize something (or cause something to be oxidized). What we need to do is figure out which one of these changes is an *oxidation*. There is no O_2 or H_2O_2 , so we see that Pt is the element that we determine ΔON for each case. Remember in *oxidation* the oxidation number *increases*.



41. Consider the overall reaction of the rechargeable nickel-cadmium battery:



Which of the following occurs at the **anode** as the reaction proceeds?

- A. Cd loses $2e^-$ and forms $\text{Cd}(\text{OH})_{2(s)}$
- B. Cd gains $2e^-$ and forms $\text{Cd}(\text{OH})_{2(s)}$
- C. NiO_2 loses $2e^-$ and forms $\text{Ni}(\text{OH})_{2(s)}$
- D. NiO_2 gains $2e^-$ and forms $\text{Ni}(\text{OH})_{2(s)}$

The Answer is **A** EXPLANATION:

Remember “LEOA” (losing e^- s is oxidation at the anode). So we can immediately cross out “B” and “D” because these are *reductions*. We can use *oxidation numbers* in “A” and “C” to find out which one really is the *oxidation*.

In “A” the oxidation number of Cd is “0” since it is in elemental form. In $\text{Cd}(\text{OH})_2$ the oxidation number of Cd is “+2” (remember each “OH” has an oxidation number of “-1”). So “A” is the *oxidation*.

Checking “C”: The oxidation number of Ni in NiO_2 is “+4” and in $\text{Ni}(\text{OH})_2$ is “+2”, so that is actually a *reduction*, so the NiO_2 would be gaining electrons, not losing them.

Therefore this one is wrong and can't be the answer.

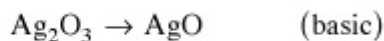
42. Which of the following will oxidize Fe^{2+} ?

- A. $\text{I}_{2(s)}$
- B. $\text{Ni}_{(s)}$
- C. $\text{Zn}_{(s)}$
- D. $\text{Br}_{2(\ell)}$

The Answer is **D** EXPLANATION:

If something is going to *oxidize* Fe^{2+} , we must look for the Fe^{2+} on the *right* side and look for the *oxidizing agent above it on the left side*. Look at your reduction table. The Fe^{2+} on the *right* side is at +0.77V. So any OA on the *left above that* will *oxidize* it. Looking at the possible answers, you can see that Br_2 is *above it on the left*, so that is our answer. I_2 is *below* it on the left, so it will *not* oxidize Fe^{2+} . Ni is on the *same side* (it is a *reducing agent*) so it will *not* oxidize Fe^{2+} . The Zn is a little tricky. You can see that Zn and Fe^{2+} actually *do* react. But Zn reacts with the Fe^{2+} on the *left* side (at -0.45v). In this reaction, the Zn is *reducing* (not oxidizing) the Fe^{2+} . Got it?

43. Consider the following half-reaction in a basic solution:



The balanced half-reaction is

- A. $\text{Ag}_2\text{O}_3 + 4\text{H}^+ + 4\text{e}^- \rightarrow \text{AgO} + 2\text{H}_2\text{O}$
- B. $\text{Ag}_2\text{O}_3 + 2\text{H}^+ + 2\text{e}^- \rightarrow 2\text{AgO} + \text{H}_2\text{O}$
- C. $\text{Ag}_2\text{O}_3 + \text{H}_2\text{O} + 2\text{e}^- \rightarrow 2\text{AgO} + 2\text{OH}^-$
- D. $\text{Ag}_2\text{O}_3 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow \text{AgO} + 4\text{OH}^-$

The Answer is **C**

EXPLANATION:

The best way to approach a question like this is to start from scratch balancing the half-reaction in *basic* solution. Then see which answer your's matches with. In balancing half-reactions remember the sequence for *basic* ones:

Major \rightarrow "O" \rightarrow "H" \rightarrow charge \rightarrow "basify"

Balancing *major* atoms (Ag in this case) first: $\text{Ag}_2\text{O}_3 \rightarrow 2\text{AgO}$

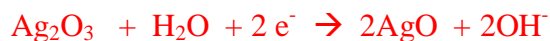
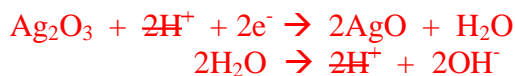
Then balance "O"s by adding H_2O 's: $\text{Ag}_2\text{O}_3 \rightarrow 2\text{AgO} + \text{H}_2\text{O}$

Then balance "H"s by adding H^+ 's: $\text{Ag}_2\text{O}_3 + 2\text{H}^+ \rightarrow 2\text{AgO} + \text{H}_2\text{O}$

Next balance charges by adding e^- s to the more "+" side:



Now, make it basic by adding the reaction: $2\text{H}_2\text{O} \rightarrow 2\text{H}^+ + 2\text{OH}^-$ to cancel the H^+ s and simplify the H_2O 's:



The answer is the same as their answer "C".

44. The concentration of $\text{Fe}^{2+}_{(aq)}$ can be determined by a redox titration using
- KBr
 - SnCl_2
 - KMnO_4 (basic)
 - KBrO_3 (acidic)

The Answer is **D**

EXPLANATION:

We have to find something which *reacts spontaneously* with Fe^{2+} using the reduction table. Remember, there are 2 Fe^{2+} 's on the table: one on the *left* at -0.45v and one on the *right* at $+0.77\text{v}$. What you need to do is *dissociate* all of the ionic compounds given in the answers and find an ion which *reacts spontaneously* with Fe^{2+} .

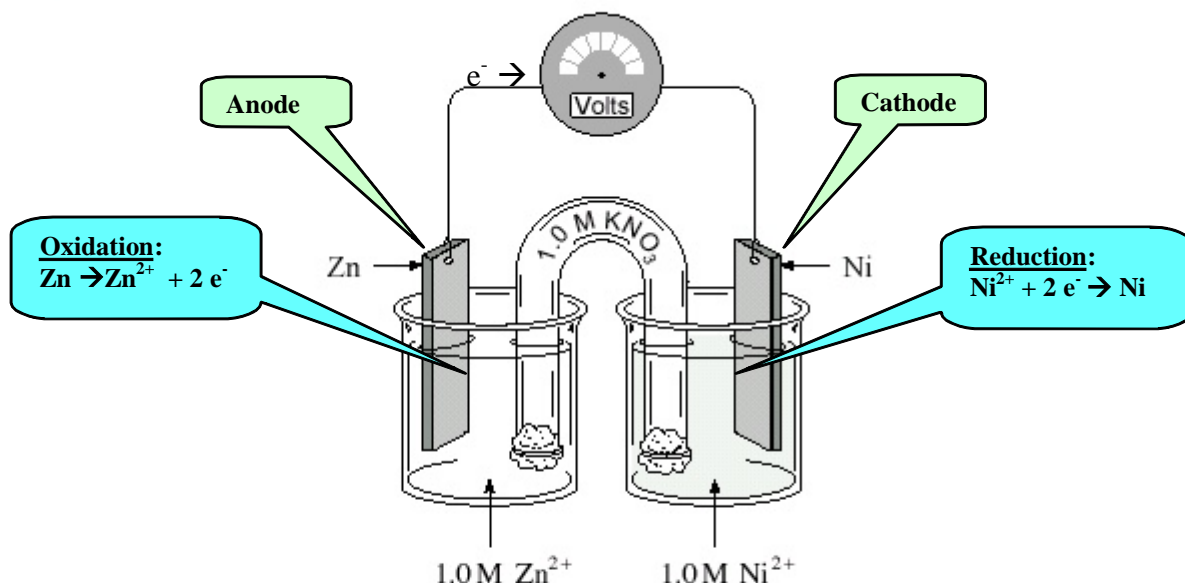
Dissociating:

- $\text{KBr} \rightarrow \text{K}^+ + \text{Br}^-$
- $\text{SnCl}_2 \rightarrow \text{Sn}^{2+} + 2\text{Cl}^-$
- $\text{KMnO}_4 \rightarrow \text{K}^+ + \text{MnO}_4^-$ (basic means OH^- is present)
- $\text{KBrO}_3 \rightarrow \text{K}^+ + \text{BrO}_3^-$ (acidic mean H^+ is present)

Looking at the Fe^{2+} on the *left* at -0.45v , we see that *none* of these ions are present on the right side below it, so we look at the other one.

Looking at the Fe^{2+} on the *right* side $+0.77\text{v}$, *anything above it on the left side* will react with it. Acidified bromate (BrO_3^- with H^+) is found on the left at 1.48v . This is the only thing here that will react with Fe^{2+} . There is MnO_4^- on the left at 1.51v , but that one is in *acidic* solution—answer “C” says the KMnO_4 is in *basic* solution (MnO_4^- in *basic* solution is down at $+0.60\text{v}$ —too low to react with Fe^{2+}) So the correct answer is acidified BrO_3^- (on the left at $+1.48\text{v}$). This is answer “D”. In this case the “ K^+ ” in KBrO_3 is a *spectator*.

45. Consider the following electrochemical cell:



Which of the following occurs as the cell operates?

- A. Zinc electrode is reduced and increases in mass.
- B. Zinc electrode is reduced and decreases in mass.
- C. Zinc electrode is oxidized and increases in mass.
- D. Zinc electrode is oxidized and decreases in mass.

The Answer is **D**

EXPLANATION:

We first identify the cathode (the one higher on the table). The cathode is nickel and the anode is zinc. The half-reaction at the Zn electrode is the oxidation of Zn: ($\text{Zn} \rightarrow \text{Zn}^{2+} + 2 \text{e}^-$) As solid Zn is reacting to form Zn^{2+} ions (which are dissolved), the mass of the Zn electrode is decreasing. So "D" is the correct answer.

More review about this Electrochemical Cell (ECC):

- The e^- s are moving *from the anode (Zn) toward the cathode (Ni) through the wire.*
- *Cations (K^+ , Zn^{2+}) are migrating toward the cathode (Ni) in the salt bridge.*
- *Anions (NO_3^-) are migrating toward the anode (Zn) in the salt bridge.*
- *The cathode (Ni electrode) gains mass as Ni^{2+} ions are reduced to Ni solid:*
 $\text{Ni}^{2+} + 2 \text{e}^- \rightarrow \text{Ni}$
- *The initial cell voltage is $(-0.26 + 0.76 = 0.50)$ volts)*
- *The equilibrium cell voltage is 0.00 volts.*

April 2000 Chemistry 12 Provincial

Multiple Choice Explanations

46. Which of the following reactants would produce an E° of $+0.63\text{ V}$?

- A. $\text{Ag}^+ + \text{I}_2$
- B. $\text{Pb}^{2+} + \text{Zn}$
- C. $\text{Mg}^{2+} + \text{Ca}$
- D. $\text{Zn}^{2+} + \text{Mn}$

The Answer is **B** EXPLANATION:

We can calculate the E° of each reaction as follows:

$\text{Ag}^+ + \text{I}_2$: This reaction is *impossible* since both reactants are on the same side only.

$\text{Pb}^{2+} + \text{Zn}$: $\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$ $E^\circ = -0.13\text{ v}$
 $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ $E^\circ = +0.76\text{ v}$ (oxidation so sign on E° on table is reversed)
 Overall redox: $E^\circ = \underline{+0.63\text{ v}}$

$\text{Mg}^{2+} + \text{Ca}$: $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$ $E^\circ = -2.37\text{ v}$
 $\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$ $E^\circ = +2.87\text{ v}$ (oxidation so sign on E° on table is reversed)
 Overall redox: $E^\circ = \underline{+0.50\text{ v}}$

$\text{Zn}^{2+} + \text{Mn}$: $\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$ $E^\circ = -0.76\text{ v}$
 $\text{Mn} \rightarrow \text{Mn}^{2+} + 2\text{e}^-$ $E^\circ = +1.19\text{ v}$ (oxidation so sign on E° on table is reversed)
 Overall redox: $E^\circ = \underline{+0.43\text{ v}}$

So you can see why the reaction in "B" is the correct answer.

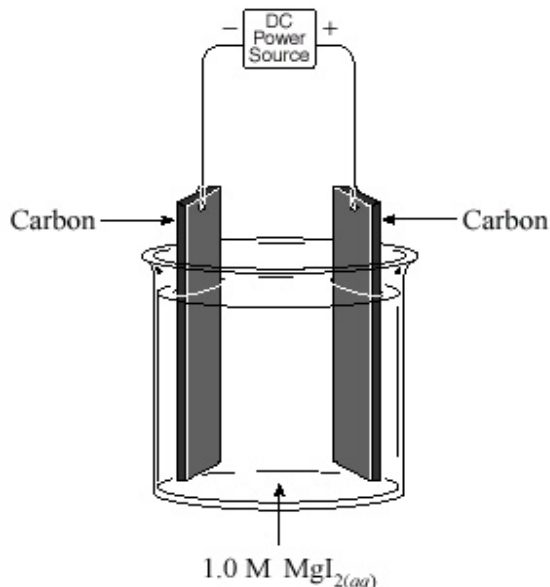
47. The process of applying an electric current through a cell to produce a chemical change is called

- A. corrosion.
- B. ionization.
- C. hydrolysis.
- D. electrolysis.

The Answer is **D** EXPLANATION:

This is just the *definition* of *electrolysis*. It's as simple as that!

48. Consider the following electrolytic cell:



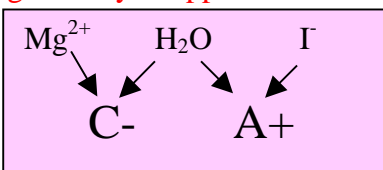
The cathode reaction is

- A. $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$
- B. $\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$
- C. $\text{H}_2\text{O} \rightarrow \frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2\text{e}^-$
- D. $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$

The Answer is **D**

EXPLANATION:

Notice that we have an *aqueous* solution and the electrodes (Carbon) are *unreactive*. A good way to approach electrolytic cells like this one is to make a diagram like the box shown here.



To find out whether the Mg^{2+} or the H_2O is reduced at the cathode, we look at the “overpotential arrow” on the left side of the table. Remember any ion *above* this arrow *will be reduced* in aqueous solution while any ion *below* this arrow *will not be reduced* in

aqueous solution—the H_2O *will be reduced* instead. We see that Mg^{2+} is *below* the arrow, so Mg^{2+} is *not* reduced. H_2O is reduced. The half-reaction for the reduction of H_2O is on the bottom shaded line: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$. This is answer “D”.

Just a little more review:

The *products at the cathode* would be H_2 gas(bubbles) and OH^- so the solution around the cathode would get more *basic* (pH would *increase*)

At the *anode* we consider the “overpotential arrow” on the *right* side. Anything (ion or solid metal) *below* this arrow *will be oxidized* in aqueous solution while anything *above* the arrow *won't*. The H_2O would be oxidized instead. The half-reaction for the oxidation of water is the *reverse* of the equation in the top shaded line: $\text{H}_2\text{O} \rightarrow \frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2\text{e}^-$

In this cell, the I^- is *below* the arrow so the half-reaction at the anode would be:

