CHAPTER 5 QUESTIONS

Multiple-Choice Questions

Use the following solubility rules to answer questions 1-4.

Salts containing halide anions are soluble except for those containing Ag^+, Pb^{2+}, and Hg_2^{2+}.

Salts containing carbonate anions are insoluble except for those containing alkali metals or ammonium.

1. If solutions of iron (III) nitrate and sodium carbonate are mixed, what would be the formula of the precipitate?

    (A) Fe_3CO_3
    (B) $Fe_2(CO_3)_3$
    (C) $NaNO_3$
    (D) No precipitate would form.

2. If solutions containing equal amounts of $AgNO_3$ and KCl are mixed, what is the identity of the spectator ions?

    (A) Ag^+, NO_3^-, K^+, and Cl^-
    (B) Ag^+ and Cl^-
    (C) K^+ and Ag^+
    (D) K^+ and NO_3^-

3. If equimolar solutions of $Pb(NO_3)_2$ and NaCl are mixed, which ion will NOT be present in significant amounts in the resulting solution after equilibrium is established?

    (A) Pb^{2+}
    (B) NO_3^-
    (C) Na^+
    (D) Cl^-

4. Choose the correct net ionic equation representing the reaction that occurs when solutions of potassium carbonate and copper (I) chloride are mixed.

    (A) $K_2CO_3(aq) + 2CuCl(aq) \rightarrow 2KCl(aq) + Cu_2CO_3(s)$
    (B) $K_2CO_3(aq) + 2CuCl(aq) \rightarrow 2KCl(s) + Cu_2CO_3(aq)$
    (C) $CO_3^{2-} + 2Cu^+ \rightarrow Cu_2CO_3(s)$
    (D) $CO_3^{2-} + Cu^{2+} \rightarrow CuCO_3(s)$

5. A strip of metal X is placed into a solution containing Y^{2+} ions and no reaction occurs. When metal X is placed in a separate solution containing Z^{2+} ions, metal Z starts to form on the strip. Which of the following choices organizes the reduction potentials for metals X, Y, and Z from greatest to least?

   (A) X > Y > Z
   (B) Y > Z > X
   (C) Z > X > Y
   (D) Y > X > Z

6. In which of the following compounds is the oxidation number of chromium the greatest?

   (A) CrO_4^{2-}
   (B) CrO
   (C) Cr^{3+}
   (D) Cr(*s*)

7. What is the mass of oxygen in 148 grams of calcium hydroxide ($Ca(OH)_2$)?

   (A) 24 grams
   (B) 32 grams
   (C) 48 grams
   (D) 64 grams

8. A sample of a compound known to consist of only carbon, hydrogen, and oxygen is found to have a total mass of 29.05 g. If the mass of the carbon is 18.02 g and the mass of the hydrogen is 3.03 g, what is the empirical formula of the compound?

   (A) C_2H_4O
   (B) C_3H_6O
   (C) $C_2H_6O_3$
   (D) $C_3H_8O_2$

Use the following information to answer questions 9-11.

When heated in a closed container in the presence of a catalyst, potassium chlorate decomposes into potassium chloride and oxygen gas via the following reaction:

$$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$$

9. If 12.25 g of potassium chlorate decomposes, how many grams of oxygen gas will be generated?

(A) 1.60 g
(B) 3.20 g
(C) 4.80 g
(D) 18.37 g

10. Approximately how many liters of oxygen gas will be evolved at STP?

(A) 2.24 L
(B) 3.36 L
(C) 4.48 L
(D) 22.4 L

11. If the temperature of the gas is doubled while the volume is held constant, what will happen to the pressure exerted by the gas and why?

(A) It will also double, because the gas molecules will be moving faster.
(B) It will also double, because the gas molecules are exerting a greater force on each other.
(C) It will be cut in half, because the molecules will lose more energy when colliding.
(D) It will increase by a factor of 4, because the kinetic energy will be four times greater.

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12. A sample of a hydrate of $CuSO_4$ with a mass of 250 grams was heated until all the water was removed. The sample was then weighed and found to have a mass of 160 grams. What is the formula for the hydrate?

(A) $CuSO_4 \bullet 10H_2O$
(B) $CuSO_4 \bullet 7H_2O$
(C) $CuSO_4 \bullet 5H_2O$
(D) $CuSO_4 \bullet 2H_2O$

13. A gaseous mixture at 25°C contained 1 mole of CH_4 and 2 moles of O_2 and the pressure was measured at 2 atm. The gases then underwent the reaction shown below.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

What was the pressure in the container after the reaction had gone to completion and the temperature was allowed to return to 25°C?

(A) 1 atm
(B) 2 atm
(C) 3 atm
(D) 4 atm

14. During a chemical reaction, NO(g) gets reduced and no nitrogen-containing compound is oxidized. Which of the following is a possible product of this reaction?

(A) $NO_2(g)$
(B) $N_2(g)$
(C) $NO_3^-(aq)$
(D) $NO_2^-(aq)$

15. Which expression below should be used to calculate the mass of copper that can be plated out of a 1.0 M $Cu(NO_3)_2$ solution using a current of 0.75 A for 5.0 minutes?

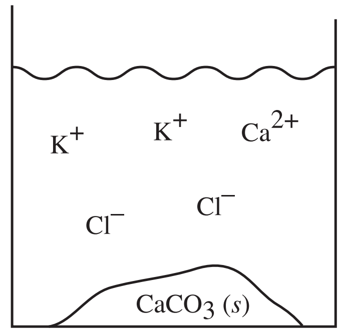
(A) $\dfrac{(5.0)(60)(0.75)(63.55)}{(96500)(2)}$

(B) $\dfrac{(5.0)(60)(63.55)(2)}{(0.75)(96500)}$

(C) $\dfrac{(5.0)(60)(96500)(0.75)}{(63.55)(2)}$

(D) $\dfrac{(5.0)(60)(96500)(63.55)}{(0.75)(2)}$

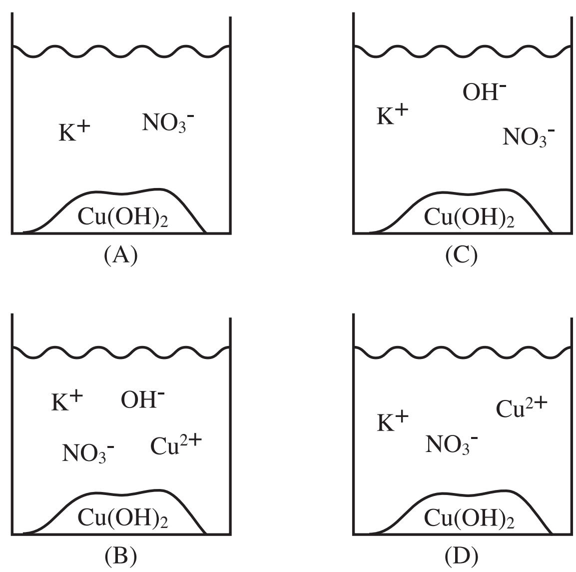
16. Solutions of potassium carbonate and calcium chloride are mixed, and the particulate representation below shows which are present in significant amounts after the reaction has gone to completion.



Which of the two original solutions is the limiting reagent and why?

(A) The potassium carbonate, because of the polyatomic anion
(B) The potassium carbonate, because there is no carbonate left after the reaction
(C) The calcium chloride, because there is an excess of calcium ions post-reaction
(D) The calcium chloride, because the component ions are smaller than those in potassium carbonate

17. A student mixes equimolar amounts of KOH and $Cu(NO_3)_2$ in a beaker. Which of the following particulate diagrams correctly shows all species present after the reaction occurs?



Use the following information to answer questions 18-20.

$14H^{+}(aq) + Cr_2O_7^{2-}(aq) + 3Ni(s) \rightarrow$
$2Cr^{3+}(aq) + 3Ni^{2+}(aq) + 7H_2O(l)$

In the above reaction, a piece of solid nickel is added to a solution of potassium dichromate.

18. Which species is being oxidized and which is being reduced?

	Oxidized	Reduced
(A)	$Cr_2O_7^{2-}(aq)$	$Ni(s)$
(B)	$Cr^{3+}(aq)$	$Ni^{2+}(aq)$
(C)	$Ni(s)$	$Cr_2O_7^{2-}(aq)$
(D)	$Ni^{2+}(aq)$	$Cr^{3+}(aq)$

19. How many moles of electrons are transferred when 1 mole of potassium dichromate is mixed with 3 moles of nickel?

(A) 2 moles of electrons
(B) 3 moles of electrons
(C) 5 moles of electrons
(D) 6 moles of electrons

20. How does the pH of the solution change as the reaction progresses?

(A) It increases until the solution becomes basic.
(B) It increases, but the solution remains acidic.
(C) It decreases until the solution becomes basic.
(D) It decreases, but the solution remains acidic.

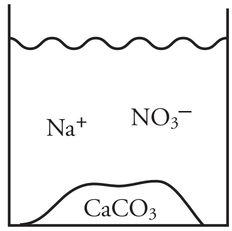
Use the following information to answer questions 21-24.

20.0 mL of 1.0 *M* Na_2CO_3 is placed in a beaker and titrated with a solution of 1.0 *M* $Ca(NO_3)_2$, resulting in the creation of a precipitate.

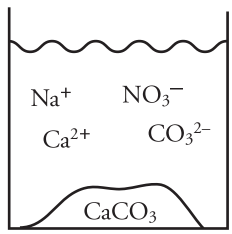
21. How much $Ca(NO_3)_2$ must be added to reach the equivalence point?

(A) 10.0 mL
(B) 20.0 mL
(C) 30.0 mL
(D) 40.0 mL

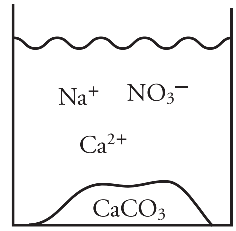
22. Which of the following diagrams correctly shows the species present in the solution in significant amounts at the equivalence point?



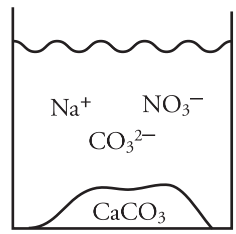
(A)



(B)



(C)



(D)

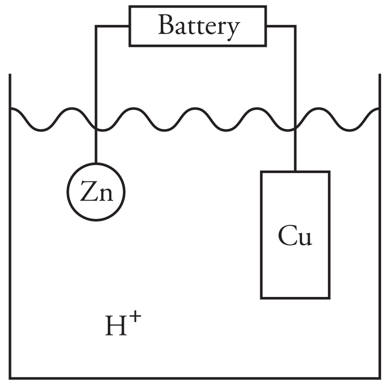
23. What will happen to the conductivity of the solution after additional $Ca(NO_3)_2$ is added past the equivalence point?

(A) The conductivity will increase as additional ions are being added to the solution.
(B) The conductivity will stay constant as the precipitation reaction has gone to completion.
(C) The conductivity will decrease as the solution will be diluted with the addition of additional $Ca(NO_3)_2$.
(D) The conductivity will stay constant as equilibrium has been established.

24. If the experiment were repeated and the Na_2CO_3 was diluted to 40.0 mL with distilled water prior to the titration, how would that affect the volume of $Ca(NO_3)_2$ needed to reach the equivalence point?

(A) It would be cut in half.
(B) It would decrease by a factor of 1.5.
(C) It would double.
(D) It would not change.

Use the following information to answer questions 25-27.



Pennies are made primarily of zinc, which is coated with a thin layer of copper through electroplating, using a setup like the one above. The solution in the beaker is a strong acid (which produces H^+ ions), and the cell is wired so that the copper electrode is the anode and the zinc penny is the cathode. Use the following reduction potentials to answer questions 25-27.

Half-Reaction	Standard Reduction Potential
$Cu^{2+} + 2e \rightarrow Cu(s)$	+0.34 V
$2H^+ + 2e \rightarrow H_2(g)$	0.00 V
$Ni^{2+} + 2e \rightarrow Ni(s)$	–0.25 V
$Zn^{2+} + 2e \rightarrow Zn(s)$	–0.76 V

25. When the cell is connected, which of the following reactions takes place at the anode?
 (A) $Cu^{2+} + 2e^- \rightarrow Cu(s)$
 (B) $Cu(s) \rightarrow Cu^{2+} + 2e$
 (C) $2H^+ + 2e^- \rightarrow H_2(g)$
 (D) $H_2(g) \rightarrow 2H^+ + 2e$

26. What is the required voltage to make this cell function?
 (A) 0.34 V
 (B) 0.42 V
 (C) 0.76 V
 (D) 1.10 V

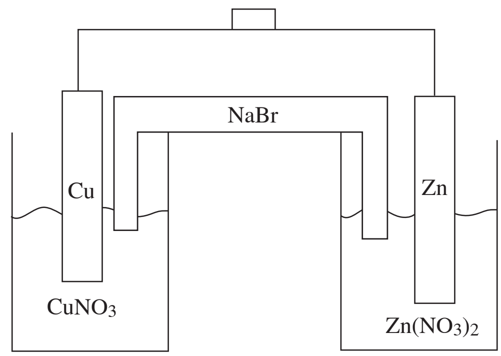
27. If, instead of copper, a nickel bar were to be used, could nickel be plated onto the zinc penny effectively? Why or why not?
 (A) Yes, nickel's SRP is greater than that of zinc, which is all that is required for nickel to be reduced at the cathode.
 (B) Yes, nickel is able to take electrons from the H^+ ions in solution, allowing it to be reduced.
 (C) No, nickel's SRP is lower than that of H^+ ions, which means the only product being produced at the cathode would be hydrogen gas.
 (D) No, nickel's SRP is negative, meaning it cannot be reduced in an electrolytic cell.

Use the following information to answer questions 28-32.

Two half-cells are set up as follows:

Half-Cell A: Strip of Cu(*s*) in $CuNO_3$(*aq*)
Half-Cell B: Strip of Zn(*s*) in $Zn(NO_3)_2$(*aq*)

When the cells are connected according to the diagram below, the following reaction occurs:



$2Cu^+(aq) + Zn(s) \rightarrow 2Cu(s) + Zn^{2+}(aq)$ $E° = +1.28$ V

28. Correctly identify the anode and cathode in this reaction as well as where oxidation and reduction are taking place.
    (A) Cu is the anode where oxidation occurs, and Zn is the cathode where reduction occurs.
    (B) Cu is the anode where reduction occurs, and Zn is the cathode where oxidation occurs.
    (C) Zn is the anode where oxidation occurs, and Cu is the cathode where reduction occurs.
    (D) Zn is the anode where reduction occurs, and Cu is the cathode where oxidation occurs.

29. How many moles of electrons must be transferred to create 127 g of copper?
    (A) 1 mole of electrons
    (B) 2 moles of electrons
    (C) 3 moles of electrons
    (D) 4 moles of electrons

30. If the $Cu^+ + e^- \rightarrow Cu(s)$ half-reaction has a standard reduction potential of +0.52 V, what is the standard reduction potential for the $Zn^{2+} + 2e^- \rightarrow Zn(s)$ half-reaction?
    (A) +0.76 V
    (B) –0.76 V
    (C) +0.24 V
    (D) –0.24 V

31. As the reaction progresses, what will happen to the overall voltage of the cell?
    (A) It will increase as $[Zn^{2+}]$ increases.
    (B) It will increase as $[Cu^+]$ increases.
    (C) It will decrease as $[Zn^{2+}]$ increases.
    (D) The voltage will remain constant.

32. What will happen in the salt bridge as the reaction progresses?
    (A) The Na^+ ions will flow to the Cu/Cu^+ half-cell.
    (B) The Br^- ions will flow to the Cu/Cu^+ half-cell.
    (C) Electrons will transfer from the Cu/Cu^+ half-cell to the Zn/Zn^{2+} half-cell.
    (D) Electrons will transfer from the Zn/Zn^{2+} half-cell to the Cu/Cu^+ half-cell.

Free-Response Questions

1. 2.54 g of beryllium chloride are completely dissolved into 50.00 mL of water inside a beaker.
   (a) Draw a particulate representation of all species in the beaker after the solute has dissolved. Your diagram should include at least one beryllium ion, one chloride ion, and four water molecules. Make sure the atoms and ions are correctly sized and oriented relative to each other.
   (b) What is the concentration of beryllium and chloride ions in the beaker?

   A solution of 0.850 *M* lead nitrate is then titrated into the beaker, causing a precipitate of lead (II) chloride to form.
   (c) Identify the net ionic reaction occurring in the beaker.
   (d) What volume of lead nitrate must be added to the beaker to cause the maximum precipitate formation?
   (e) What is the theoretical yield of precipitate?
   (f) Students performing this experiment suggested the following techniques to separate the precipitate from the water. Their teacher rejected each idea. Explain why the teacher may have done so, and name the inherent errors of
      (i) boiling off the water
      (ii) decanting (pouring off) the water

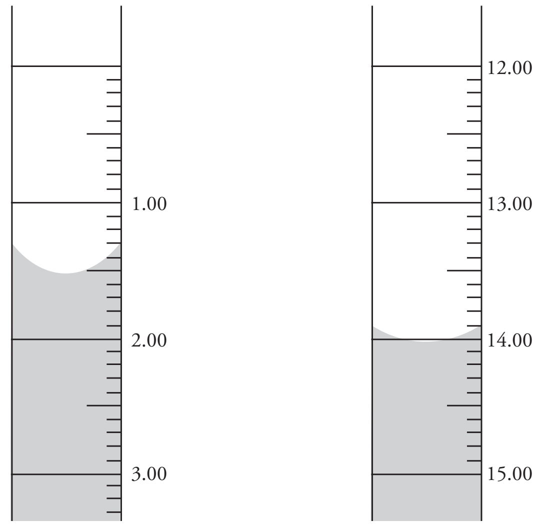
2. Hydrogen peroxide, H_2O_2, is a common disinfectant. Pure hydrogen peroxide is a very strong oxidizer, and as such, it is diluted with water to low percentages before being bottled and sold. One method to determine the exact concentration of H_2O_2 in a bottle of hydrogen peroxide is to titrate a sample with a solution of acidified potassium permanganate. This causes the following redox reactions to occur:

   Reduction: $8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O(l)$

   Oxidation: $H_2O_2(aq) \rightarrow 2H^+ + O_2(g) + 2e$

   During a titration, a student measures out 5.0 mL of hydrogen peroxide solution into a graduated cylinder, and he pours it into a flask, diluting it to 50.0 mL with water. The student then titrates 0.150 *M* potassium permanganate solution into the flask with constant stirring.
   (a) Write out the full, balanced redox reaction that is taking place during the titration.
   (b) List two observations that the student will see as the titration progresses that are indicative of chemical reactions.

Diagrams of the permanganate in the buret at the start and end of the titration are as follows:



(c) (i) What volume of $KMnO_4$ was titrated?
(ii) What is the concentration of hydrogen peroxide in the original sample?

(d) How would the precision of the student's results have changed if the hydrogen peroxide sample were measured out in a 50 mL beaker instead of a graduated cylinder?

(e) How would each of the following errors affect the student's final calculated hydrogen peroxide concentration?
(i) Not filling the buret tip with solution prior to the titration
(ii) Not rinsing down the sides of the flask during titration

3. The **unbalanced** reaction between potassium permanganate and acidified iron (II) sulfate is a redox reaction that proceeds as follows:

$$H^+(aq) + Fe^{2+}(aq) + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + Fe^{3+}(aq) + H_2O(l)$$

(a) Provide the equations for both half-reactions that occur below:
(i) Oxidation half-reaction
(ii) Reduction half-reaction

(b) What is the balanced net ionic equation?

A solution of 0.150 *M* potassium permanganate is placed in a buret before being titrated into a flask containing 50.00 mL of iron (II) sulfate solution of unknown concentration. The following data describes the colors of the various ions in solution:

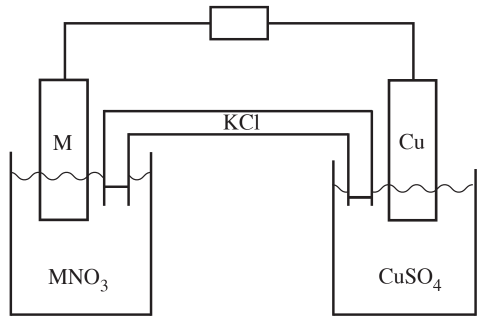
Ion	Color in solution
H^+	Colorless
Fe^{2+}	Pale Green
MnO_4^-	Dark Purple
Mn^{2+}	Colorless
Fe^{3+}	Yellow
K^+	Colorless
SO_4^{2-}	Colorless

(c) Describe the color of the solution in the flask at the following points:
(i) Before titration begins
(ii) During titration prior to the endpoint
(iii) At the endpoint of the titration
(d) (i) If 15.55 mL of permanganate are added to reach the endpoint, what is the initial concentration of the iron (II) sulfate?
(ii) The actual concentration of the $FeSO_4$ is 0.250 *M*. Calculate the percent error.
(e) Could the following errors have led to the experimental result deviating in the direction that it did? You must justify your answers quantitatively.
(i) 55.0 mL of $FeSO_4$ was added to the flask prior to titration instead of 50.0 mL.
(ii) The concentration of the potassium permanganate was actually 0.160 *M* instead of 0.150 *M*.

4. $$2Mg(s) + 2CuSO_4(aq) + H_2O(l) \rightarrow 2MgSO_4(aq) + Cu_2O(s) + H_2(g)$$

(a) If 1.46 grams of Mg(*s*) are added to 500 milliliters of a 0.200-molar solution of $CuSO_4$, what is the maximum molar yield of $H_2(g)$?
(b) When all of the limiting reagent has been consumed in (a), how many moles of the other reactant (not water) remain?
(c) What is the mass of the Cu_2O produced in (a)?
(d) What is the value of $[Mg^{2+}]$ in the solution at the end of the experiment? (Assume that the volume of the solution remains unchanged.)

5. A student performs an experiment in which a bar of unknown metal M is placed in a solution with the formula MNO_3. The metal is then hooked up to a copper bar in a solution of $CuSO_4$ as shown below. A salt bridge that contains aqueous KCl links the cell together.



The cell potential is found to be +0.74 V. Separately, when a bar of metal M is placed in the copper sulfate solution, solid copper starts to form on the bar. When a bar of copper is placed in the MNO_3 solution, no visible reaction occurs.

The following gives some reduction potentials for copper:

Half-reaction	E
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34 V
$Cu^{2+} + e^- \rightarrow Cu^+$	0.15 V
$Cu^+ + e^- \rightarrow Cu(s)$	0.52 V

(a) Write the net ionic equation that takes place in the Cu/M cell.
(b) What is the standard reduction potential for metal M?
(c) Which metal acted as the anode and which as the cathode? Justify your answer.
(d) On the diagram of the cell, indicate which way the electrons are flowing in the wire. Additionally, indicate any ionic movement occurring in the salt bridge.
(e) What would happen to the voltage of the reaction in the Cu/M cell if the concentration of the $CuSO_4$ increased while the concentration of the MNO_3 remained constant? Justify your answer.

6. Two electrodes are inserted into a solution of nickel (II) fluoride and a current of 2.20 A is run through them. A list of standard reduction potentials is as follows:

Half-reaction	$E°$
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow H_2O(l)$	1.23 V
$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	2.87 V
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	–0.83 V
$Ni^{2+} + 2e^- \rightarrow Ni(s)$	–0.25 V

(a) Write the net ionic equation that takes place during this reaction.
(b) Qualitatively describe what an observer would see taking place at each electrode.
(c) Will the solution become acidic, basic, or remain neutral as the reaction progresses?
(d) How long would it take to create 1.2 g of Ni(*s*) at the cathode?