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CH1020

EXAM 2 (70 points)

November 14th, 2018

B

There is a total of 10 pages in the exam (including this page).
There is a total of 13 questions.
Show your work to get full credit.
The activity series table, the precipitation rules, and the periodic table are on the last 3 pages

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1. (6 Points) Complete and balance each equation. If no precipitation reaction occurs, write "NR".

a. $\_1\_ FeSO_4(aq) + \_1\_ Sr(OH)_2(aq) \rightarrow$ $Fe(OH)_2(s)$ (1)

$+$ $SrSO_4(s)$ (1)

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b. $\_3\_ CaCl_2(aq) + \_2\_ Cs_3PO_4(aq) \rightarrow$ $Ca_3(PO_4)_2(s)$ (1/2)

$+ 6\,CsCl(aq)$ (1/2)

c. $\_\_ Na_2SO_4(aq) + \_\_ NiCl_2(aq) \rightarrow$ NR (2)

$+$ _____

2. (4 Points) Write balanced net ionic equations for each reaction:

a. $MgS(aq) + CuCl_2(aq) \rightarrow$

$Cu^{2+}(aq) + S^{2-}(aq) \rightarrow CuS(s)$ (2)

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b. $K_2SO_4(aq) + CaI_2(aq) \rightarrow$

$Ca^{2+}(aq) + SO_4^{2-}(aq) \rightarrow CaSO_4(s)$ (2)

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3. (6 Points) A solution contains Ag^+ and Hg^{2+} ions. The addition of 0.100 L of 1.22 M NaI solution is just enough to precipitate all the ions as AgI and HgI_2. The total mass of the precipitate is 28.1 g. Find the moles of AgI that precipitate.
$M_w(AgI) = 234.8$ g/mol ; $M_w(HgI_2) = 454.4$ g/mol

$\text{moles}(I^-) = 0.100\,L \cdot 1.22\,M = 0.122\,\text{mol}$ (1)

$AgI = x \qquad HgI_2 = y$

$x + 2y = 0.122\,\text{mol} \Rightarrow y = 0.061\,\text{mol} - \frac{1}{2}x$ (1)

$x \cdot 234.8\,g/\text{mol} + y \cdot 454.4\,g/\text{mol} = 28.1\,g$ (1)

$234.8\,g/\text{mol} \cdot x + (0.061\,\text{mol} - \frac{1}{2}x)\,454\,g/\text{mol} = 28.1\,g$

$234.8\,g/\text{mol} \cdot x + 27.72\,g - 227.2\,g/\text{mol} \cdot x = 28.1\,g$ (1)

$7.6\,g/\text{mol} \cdot x = 28.1\,g - 27.72\,g = 0.4\,g$

$x = 5.26 \cdot 10^{-2}\,\text{mol}$ (2)

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4. (4 Points) A solution contains one or more of the following ions: Hg_2^{2+}, Ba^{2+}, and Fe^{2+} When you add potassium chloride to the solution, a precipitate forms. The precipitate is filtered off, and you add potassium sulfate to the remaining solution, forming no precipitate. When you add potassium carbonate to the remaining solution, a precipitate forms. Which ions were present in the original solution? Write net ionic equations for the formation of each of the precipitates observed.

1. $Hg_2^{2+}(aq) + 2Cl^- \rightarrow Hg_2Cl_2(s)$ (Hg_2^{2+} present) (1) (1)
2. no precipitate with SO_4^{2-} => Ba^{2+} not present
3. $Fe^{2+}(aq) + CO_3^{2-}(aq) \rightarrow FeCO_3(s)$ (Fe^{2+} present) (1) (1)

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5. (9 Points) A sample of $4.05*10^{-4}$ mol K_3PO_4 is added to 22.0 mL of 0.050 M $FeCl_2$ solution, resulting in the formation of a precipitate. (a) Write the net ionic equation for the reaction. (b) What is the mass of the precipitate that forms? (c) What is the remaining concentration of the ion in solution that is in excess? $M_w(K_3PO_4) = 212.27$ g/mol ; $M_w(FeCl_2) = 126.75$ g/mol; (you need to determine the M_w of the precipitate yourself once you have identified its formula)

a). $3Fe^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Fe_3(PO_4)_2(s)$ (1)

b). $\text{moles}(PO_4^{3-}) = 4.05 \cdot 10^{-4}\,\text{mol}$

$\text{moles}(Fe_3(PO_4)_2)_{\text{produced}} = 4.05 \cdot 10^{-4}\,\text{mol} \cdot \frac{1\,\text{mol}}{2\,\text{mol}} = 2.03 \cdot 10^{-4}\,\text{mol}$ (1)

$\text{moles}(FeCl_2) = 0.022\,L \cdot 0.050\,M = 1.10 \cdot 10^{-3}\,\text{mol}$

$\text{moles}(Fe_3(PO_4)_2) = 1.10 \cdot 10^{-3}\,\text{mol} \cdot \frac{1\,\text{mol}}{3\,\text{mol}} = 3.67 \cdot 10^{-4}\,\text{mol}$ (1)

$\Rightarrow PO_4^{3-}$ limiting (1)

$\text{mass}(Fe_3(PO_4)_2) = 2.03 \cdot 10^{-4}\,\text{mol} \cdot 357\,g/\text{mol} = 0.075\,g$ (2)

c). $Fe^{2+}(\text{used}) = 2.03 \cdot 10^{-4}\,\text{mol} \cdot \frac{3\,\text{mol}}{1\,\text{mol}} = 6.09 \cdot 10^{-4}\,\text{mol}$

$Fe^{2+}(\text{remaining}) = 1.10 \cdot 10^{-3}\,\text{mol} - 6.09 \cdot 10^{-4}\,\text{mol} = 4.91 \cdot 10^{-4}\,\text{mol}$ (1)

$[Fe^{2+}] = \frac{4.91 \cdot 10^{-4}\,\text{mol}}{0.022\,L} = 0.0223\,M$ (1)

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6. (5 Points) What is the definition of an Arrhenius acid and base? What is the definition of a Brønstedt-Lowry acid and base? Give an example of a Brønstedt-Lowry acid that is not an Arrhenius acid.

Arrhenius acid: produces H_3O^+ (1)

base: produces OH^- (1)

Brønsted acid: donates H^+ (1)

base accepts H^+ (1)

e.g. $HCl(g)$ (1)

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7. (5 points) A 75.00 mL sample of an unknown phosphoric acid solution is titrated with a 0.100 M sodium hydroxide solution. The equivalence point is reached when 41.50 mL sodium hydroxide solution is added. Provide the balanced equation and calculate the concentration of the unknown phosphoric acid solution.

$H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(\ell)$ (2)

$\text{moles}(OH^-) = 0.04150\ L \cdot 0.100\ M = 4.15 \cdot 10^{-4}\ mol$ (1)

$\text{moles}(H_3PO_4) = 4.15 \cdot 10^{-4}\ mol \cdot \frac{1\ mol\ (H_3PO_4)}{3\ mol\ (OH^-)} = 1.38 \cdot 10^{-4}\ mol$ (1)

$M(H_3PO_4) = \frac{1.38 \cdot 10^{-4}\ mol}{0.075\ L} = 1.84 \cdot 10^{-3}\ M$ (1)

$(0.0184\ M)$

8. (3 points) Provide the chemical formulas for each of the following acids:

a. carbonic acid H_2CO_3 (1)

b. phosphoric acid H_3PO_4 (1)

c. hydrocyanic acid HCN (1)

9. (5 points) Determine the oxidation numbers for the indicated elements in the following substances:

a. $NaClO_4$ Na: +1 (1) Cl: +7 (1) O: −2 (1)

b. NH_4NO_3 First "N": −3 (1) Second "N": +5 (1)

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10. (6 points) Write balanced chemical equations for the following redox reactions. If no reaction occurs, write NR.

    a. Fe is added to a solution of Nickel(II)sulfate

    $Fe(s) + NiSO_4(aq) \rightarrow FeSO_4(aq) + Ni(s)$

    b. Manganese is added to a solution of iron(II)chloride

    $Mn(s) + FeCl_2(aq) \rightarrow MnCl_2(aq) + Fe(s)$

    c. Silver is added to a solution of Mercury(II)nitrate

    NR

11. (5 points) A mixture contains only $SrCl_2$ and NH_4NO_3. A 0.638 g of this mixture is dissolved in water, and an excess of $AgNO_3$ is added, producing a precipitate of AgCl. The precipitate is filtered, dried, and weighed. Its mass is 0.216 g. Calculate the mass percent of $SrCl_2$ in the mixture.

$\#\text{ moles}(AgCl) = 0.216\,g \cdot \frac{1\,mol}{143.3\,g} = 1.51 \cdot 10^{-3}\,mol$ (1)

$\#\text{ moles}(SrCl_2) = 1.51 \cdot 10^{-3}\,mol \cdot \frac{1\,mol}{2\,mol} = 7.54 \cdot 10^{-4}\,mol$ (1)

$\text{mass}(SrCl_2) = 7.54 \cdot 10^{-4}\,mol \cdot 158.5\,g/mol = 0.119\,g$ (1)

$\% = \frac{0.119\,g}{0.638\,g} \cdot 100\% = 18.7\%$ (2)

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12. (2 points) Complete and balance the following gas evolution reaction:

$HCl(aq) + KHCO_3(aq) \rightarrow$ $H_2CO_3(aq) + KCl(aq)$

$\rightarrow H_2O(l) + CO_2(g) + KCl(aq)$

13. (10 points) Complete and balance the following redox reaction in <u>acidic</u> solution

$S_2O_3^{2-}(aq) + Cl_2(g) \rightarrow SO_4^{2-}(aq) + Cl^-(aq)$

$S_2O_3^{2-} \rightarrow SO_4^{2-}$ $\quad Cl_2 \rightarrow Cl^-$

$S_2O_3^{2-} \rightarrow 2SO_4^{2-}$ $\quad Cl_2 \rightarrow 2Cl^-$

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-}$ $\quad Cl_2 \rightarrow 2Cl^-$

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+$ $\quad Cl_2 \rightarrow 2Cl^-$

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 8e^-$

$Cl_2 + 2e^- \rightarrow 2Cl^-$ $\quad (\times 4)$

$S_2O_3^{2-} + 5H_2O \rightarrow 2SO_4^{2-} + 10H^+ + 8e^-$

$4Cl_2 + 8e^- \rightarrow 8Cl^-$

$S_2O_3^{2-} + 5H_2O + 4Cl_2 + 8e^- \rightarrow 2SO_4^{2-} + 10H^+ + 8e^- + 8Cl^-$

$S_2O_3^{2-}(aq) + 5H_2O(l) + 4Cl_2(g) \rightarrow$

$2SO_4^{2-}(aq) + 10H^+(aq) + 8Cl^-(aq)$