Chemistry 192
Recitation Session Questions
Monday January 22, 2018
Solutions

1. Calculate the mass of solid magnesium oxide (MgO) that forms when 10.0 grams of magnesium are burned in 14.0 grams of pure oxygen.
Answer:

$$n_{Mg} = \frac{10.0 \text{ g}}{24.30 \text{ g mol}^{-1}} = 0.411 \text{ mol} \qquad n_{O_2} = \frac{14.0 \text{ g}}{32.00 \text{ g mol}^{-1}} = 0.438 \text{ mol}$$

$$2\text{Mg}_{(s)} + \text{O}_{2(g)} \longrightarrow 2\text{MgO}_{(s)}$$

Magnesium is the limiting reagent.

$$n_{MgO} = n_{Mg} = 0.411 \text{ mol}$$

$$(0.411 \text{ mol})(40.30 \text{ g mol}^{-1}) = 16.6 \text{ g}$$

2. Calculate the final concentrations of potassium ions and dichromate ions when 0.0151 L of 0.100 M aqueous potassium dichromate ($K_2Cr_2O_7$) are diluted with water to produce a solution of final volume 2.00 L.
Answer:

$$n_{Cr_2O_7^{2-}} = (0.100 \text{ mol L}^{-1})(0.0151 \text{ L}) = 0.00151 \text{ mol}$$

$$n_{K^+} = 2n_{Cr_2O_7^{2-}} = 0.00302 \text{ mol}$$

$$[\text{K}^+] = \frac{0.00302 \text{ mol}}{2.00 \text{ L}} = 0.00151 \text{ M}$$

$$[\text{Cr}_2\text{O}_7^{2-}] = \frac{0.00151 \text{ mol}}{2.00 \text{ L}} = 7.55 \times 10^{-4} \text{ M}$$

3. Consider the reaction between 0.513 L of an aqueous 0.100 M cadmium chloride ($CdCl_2$) solution and 0.362 L of an aqueous 0.275 M sodium sulfate (Na_2SO_4) solution. Given that $CdSO_4$ is insoluble in water, give the chemical reaction when the two solutions are mixed. Make sure to include only the ions taking part in the chemistry. The spectator ions should

not be included in your balanced chemical reaction. Calculate the amount of cadmium sulfate that forms and the final concentrations of cadmium ions and sulfate ions.
Answer:

$$Cd^{2+}_{(aq)} + SO^{2-}_{4(aq)} \longrightarrow CdSO_{4(s)}$$

$n_{Cd^{2+}} = (0.100\,\text{mol L}^{-1})(0.513\,\text{L}) = 0.0513\,\text{mol} \quad n_{SO_4^{2-}} = (0.275\,\text{mol L}^{-1})(0.362\,\text{L}) = 0.0996\,\text{mol}$

Cadmium is the limiting reagent

$$n_{Cd^{2+}} = n_{CdSO_4} = 0.0513 \text{ mol}$$

$$[\text{Cd}^{2+}] \approx 0$$

$$n_{SO_4^{2-}} = 0.0996 \text{ mol} - 0.0513 \text{ mol} = 0.0483 \text{ mol}$$

$$[\text{SO}_4^{2-}] = \frac{0.0483 \text{ mol}}{0.513 \text{ L} + 0.362 \text{ L}} = 5.52 \times 10^{-2} \text{ M}$$

4. When 1.45 grams of holmium chloride ($HoCl_3$) react with excess aqueous silver nitrate, 2.30 grams of silver chloride form. Given that silver nitrate and holmium chloride both dissolve completely in water and are both fully ionized, and given the molar masses of chlorine and silver are respectively $M_{Cl} = 35.34$ g mol^{-1} and $M_{Ag} = 107.9$ g mol^{-1}, calculate the molar mass of holmium.
Answer:

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \longrightarrow AgCl_{(s)}$$

$$n_{Ag^+} = n_{Cl^-} = \frac{2.30 \text{ g}}{143.24 \text{ g mol}^{-1}} = 1.60 \times 10^{-2} \text{ mol} = 3n_{HoCl_3}$$

$$n_{HoCl_3} = 5.35 \times 10^{-3} \text{ mol}$$

$$M_{HoCl_3} = \frac{1.45 \text{ g}}{5.35 \times 10^{-3} \text{ mol}} = 271. \text{ g mol}^{-1}$$

$$M_{Ho} = 271. \text{ g mol}^{-1} - 3(35.34 \text{ g mol}^{-1}) = 165. \text{ g mol}^{-1}$$

5. Calculate the mass of zinc produced and the final concentration of zinc ions when 0.050 grams of aluminum dissolve in 1.00 L of a solution that is originally 1.00 M in Zn^{2+} ions, assuming the reaction results in no change in the volume of the solution.
Answer:

$$3Zn^{2+}_{(aq)} + 2Al_{(s)} \longrightarrow 2Al^{3+}_{(aq)} + 3Zn_{(s)}$$

$$n_{Al} = \frac{0.050 \text{ g}}{26.98 \text{ g mol}^{-1}} = 1.85 \times 10^{-3} \text{ mol} \qquad n_{Zn} = \frac{3}{2} n_{Al} = 2.78 \times 10^{-3} \text{ mol}$$

$$m_{Zn} = (2.78 \times 10^{-3} \text{ mol})(65.39 \text{ g mol}^{-1}) = 0.182 \text{ g}$$

$$\text{Initial} : n_{Zn^{2+}} = (1.00 \text{ mol L}^{-1})(1.00 \text{ L}) = 1.00 \text{ mol}$$

$$\text{Final} : n_{Zn^{2+}} = 1.00 \text{ mol} - 2.78 \times 10^{-3} \text{ mol} = 0.997 \text{ mol}$$

$$[\mathrm{Zn}^{2+}] = \frac{0.997 \text{ mol}}{1.00 \text{ L}} = 0.997 \text{ M}$$

6. It is found that 0.125 L of 0.246 M hydrochloric acid exactly neutralizes 0.250 L of an aqueous sodium hydroxide solution. Calculate the original concentration of the sodium hydroxide solution.

Answer:

$$n_{H^+} = (0.246 \text{ mol L}^{-1})(0.125 \text{ L}) = 3.08 \times 10^{-2} \text{ mol} = n_{OH^-}$$

$$[\mathrm{OH}^-] = \frac{3.08 \times 10^{-2} \text{ mol}}{0.250 \text{ L}} = 0.123 \text{ M}$$

7. Calculate the volume of hydrogen gas produced when 2.00 grams of solid magnesium dissolve in excess hydrochloric acid at a pressure of 0.95 bar and a temperature of 25.0° C.

Answer:

$$\mathrm{Mg}_{(s)} + 2\mathrm{H}^+_{(aq)} \longrightarrow \mathrm{H}_{2(g)} + \mathrm{Mg}^{2+}_{(aq)}$$

$$n_{Mg} = \frac{2.00 \text{ g}}{24.30 \text{ g mol}^{-1}} = 8.23 \times 10^{-2} \text{ mol} = n_{H_2}$$

$$V = \frac{nRT}{P} = \frac{(8.23 \times 10^{-2} \text{ mol})(0.08314 \text{ L bar mol}^{-1}\text{K}^{-1})(298 \text{ K})}{0.95 \text{ bar}} = 2.2 \text{ L}$$

8. When 10.0 grams of argon gas and 15.0 grams of nitrogen gas are combined in a closed container, the total pressure is found to be 3.27 bar. Calculate the mole fraction of each gas and the partial pressure of each gas in the system.

Answer:

$$n_{Ar} = \frac{10.0 \text{ g}}{39.95 \text{ g mol}^{-1}} = 0.250 \text{ mol} \qquad n_{N_2} = \frac{15.0 \text{ g}}{28.00 \text{ g mol}^{-1}} = 0.536 \text{ mol}$$

$$\chi_{Ar} = \frac{0.250}{0.250 + 0.536} = 0.318 \qquad \chi_{N_2} = 1 - \chi_{Ar} = 0.682$$

$$P_{Ar} = \chi_{Ar} P = (0.318)(3.27 \text{ bar}) = 1.04 \text{ bar} \qquad P_{N_2} = \chi_{N_2} P = (0.682)(3.27 \text{ bar}) = 2.23 \text{ bar}$$