

Problem Set 2 Solutions (For January 20)

1. Since the net reaction is $\text{SrBr}_2 + \text{Cl}_2 \rightarrow \text{SrCl}_2 + \text{Br}_2$, every mole of SrBr_2 destroyed in the reaction will produce exactly one mole of SrCl_2 . 14.3 g of SrBr_2 is

$$\frac{(14.3 \text{ g SrBr}_2)}{(247.44 \text{ g mol}^{-1} \text{ SrBr}_2)} = 57.8 \text{ mmol SrBr}_2$$

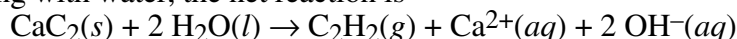
and thus 57.8 mmol of SrCl_2 will be formed. This amount would weigh

$$(57.8 \text{ mmol SrCl}_2)(158.53 \text{ g mol}^{-1} \text{ SrCl}_2) = 9.16 \text{ g SrCl}_2$$

We would need an equal number of moles of Cl_2 at a minimum. They would weigh

$$(57.8 \text{ mmol Cl}_2)(70.906 \text{ g mol}^{-1} \text{ Cl}_2) = 4.10 \text{ g Cl}_2$$

2. For CaC_2 reacting with water, the net reaction is



We see that every mole of acetylene produced requires one mole of CaC_2 ; so, we begin by calculating the number of moles of C_2H_2 in 5.00 g of C_2H_2 . The molar mass of C_2H_2 is 26.04 g mol^{-1} (to a good approximation worth memorizing, the atomic masses of H, C, N, and O are 1, 12, 14, and 16 g mol^{-1} , respectively), and we find

$$\frac{(5.00 \text{ g C}_2\text{H}_2)}{(26.04 \text{ g mol}^{-1} \text{ C}_2\text{H}_2)} = 0.192 \text{ mol C}_2\text{H}_2$$

This amount of CaC_2 (with molar mass 64.10 g mol^{-1}) corresponds to a mass of

$$(0.192 \text{ mol CaC}_2)(64.10 \text{ g mol}^{-1} \text{ CaC}_2) = 12.3 \text{ g CaC}_2$$

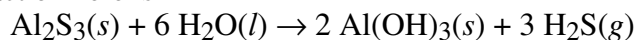
For heavy water, just write "D" instead of "H" in the net reaction:



If we generate C_2D_2 , we follow the same logic using the molar mass of C_2D_2 , which is (using 2.00 g mol^{-1} for the atomic mass of deuterium) 28.04 g mol^{-1} . Using the same mass of CaC_2 , 12.3 g, we would generate *the same number of moles* of C_2D_2 as we did C_2H_2 , but *the mass would be greater* than 5.00 g:

$$(0.192 \text{ mol C}_2\text{D}_2)(28.04 \text{ g mol}^{-1} \text{ C}_2\text{D}_2) = 5.39 \text{ g C}_2\text{D}_2$$

3. The balanced net reaction here is



We first calculate and collect the molar masses we will need: for Al_2S_3 , $150.16 \text{ g mol}^{-1}$; for H_2O , $18.015 \text{ g mol}^{-1}$; and for H_2S , $34.082 \text{ g mol}^{-1}$. Next, we turn the initial masses of the reactants into corresponding initial numbers of moles:

$$n_{\text{Al}_2\text{S}_3}^{\circ} = \frac{15.28 \text{ g Al}_2\text{S}_3}{150.16 \text{ g mol}^{-1} \text{ Al}_2\text{S}_3} = 0.1018 \text{ mol Al}_2\text{S}_3$$

$$n_{\text{H}_2\text{O}}^{\circ} = \frac{8.51 \text{ g H}_2\text{O}}{18.015 \text{ g mol}^{-1} \text{ H}_2\text{O}} = 0.4727 \text{ mol H}_2\text{O}$$

Then, we divide each initial amount by its corresponding stoichiometric coefficient, 1 for Al_2S_3 and 6 for H_2O :

$$\frac{n_{\text{Al}_2\text{S}_3}^{\circ}}{1} = \frac{0.1018 \text{ mol Al}_2\text{S}_3}{1} = 0.1018 \text{ mol} \quad \frac{n_{\text{H}_2\text{O}}^{\circ}}{6} = \frac{0.4727 \text{ mol H}_2\text{O}}{6} = 0.07873 \text{ mol}$$

The *smaller* result is for H_2O ; so, water must be the limiting reagent and Al_2S_3 is in excess. We call this smaller amount x_{max} , the maximum extent of reaction (note that the *minimum* amount of

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reactant controls the *maximum* extent of reaction: we can only go as far as the limiting reagent allows). We can use x_{\max} to calculate first the number of moles of H_2S product produced and then the mass of H_2S that amount corresponds to, taking care to incorporate the stoichiometric factor of 3 for H_2S in the net reaction:

$$x_{\max} = 0.07873 \text{ mol} = \frac{n_{\text{H}_2\text{S}}}{3} \quad \text{or} \quad n_{\text{H}_2\text{S}} = 3 \times 0.07873 \text{ mol} = 0.2362 \text{ mol H}_2\text{S}$$

$$m_{\text{H}_2\text{S}} = (0.2362 \text{ mol H}_2\text{S})(34.082 \text{ g mol}^{-1} \text{ H}_2\text{S}) = 8.05 \text{ g}$$

Finally, we calculate the number of moles of Al_2S_3 left unreacted and its corresponding mass:

$$x_{\max} = 0.07873 \text{ mol} = \frac{n_{\text{Al}_2\text{S}_3}^{\circ} - n_{\text{Al}_2\text{S}_3}}{1} = \frac{0.1018 \text{ mol Al}_2\text{S}_3 - n_{\text{Al}_2\text{S}_3}}{1}$$

$$n_{\text{Al}_2\text{S}_3} = 0.1018 \text{ mol} - 0.07873 \text{ mol} = 0.02307 \text{ mol Al}_2\text{S}_3 \text{ remaining}$$

and

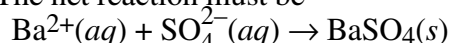
$$m_{\text{Al}_2\text{S}_3} = (0.02307 \text{ mol Al}_2\text{S}_3)(150.16 \text{ g mol}^{-1}) = 3.46 \text{ g Al}_2\text{S}_3 \text{ remaining}$$

4. Step by step: a solution formed from 0.700 mol Na_2SO_4 dissolved in 350 mL = 0.350 L of water has a total “ Na_2SO_4 ” concentration of $(0.700 \text{ mol})/(0.350 \text{ L}) = 2.00 \text{ mol L}^{-1} = 2.00 \text{ M}$. But Na_2SO_4 is a strong electrolyte, which means it dissociates completely into $\text{Na}^+(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$ (sulfate) ions. Thus, we can say the solution composition is better expressed in terms of concentrations of these two ions: $[\text{Na}^+] = 4.00 \text{ M}$ (because each mole of Na_2SO_4 produces *two* moles of Na^+) and $[\text{SO}_4^{2-}] = 2.00 \text{ M}$. Note as well that our 350 mL of solution contains 1.40 mol Na^+ and 0.700 mol SO_4^{2-} . Next, consider the 550 mL = 0.550 L of the 1.6 M solution of BaCl_2 . This is also a solution of a strong electrolyte, and the BaCl_2 will dissociate completely into Ba^{2+} and Cl^- ions. We can write $[\text{Ba}^{2+}] = 1.6 \text{ M}$ and $[\text{Cl}^-] = 3.2 \text{ M}$, but of greater interest to us is the amounts of each ion:

$$n_{\text{Cl}^-} = [\text{Cl}^-]V_{\text{solution}} = (3.2 \text{ M})(0.550 \text{ L}) = 1.76 \text{ mol}$$

$$n_{\text{Ba}^{2+}} = [\text{Ba}^{2+}]V_{\text{solution}} = (1.6 \text{ M})(0.550 \text{ L}) = 0.880 \text{ mol}$$

Now we mix. Note first that our final volume will be 350 mL + 550 mL = 900 mL = 0.900 L. Then, figure out the chemistry: the precipitate must be BaSO_4 , because NaCl , the only other option, is a strong electrolyte. The net reaction must be



Which ion, Ba^{2+} or SO_4^{2-} , is limiting? The simple 1:1 stoichiometry makes this an easy question to answer: we have more Ba^{2+} (0.880 mol) than we have SO_4^{2-} (0.700 mol). Thus, we precipitate 0.700 mol $\text{BaSO}_4(\text{s})$, leaving *no* sulfate ions in solution. This solid weighs (the molar mass of BaSO_4 is $233.39 \text{ g mol}^{-1}$)

$$m_{\text{BaSO}_4} = (0.700 \text{ mol})(233.39 \text{ g mol}^{-1}) = 163. \text{ g}$$

Some Ba^{2+} is left behind in solution: $(0.880 \text{ mol}) - (0.700 \text{ mol}) = 0.180 \text{ mol}$. This is in a total volume of 900 mL, and the final $[\text{Ba}^{2+}]$ is $(0.180 \text{ mol})/(0.900 \text{ L}) = 0.200 \text{ M}$. *All* the Na^+ and Cl^- is still in solution, but now in a 900 mL volume. These ions have final concentrations given by $[\text{Na}^+] = (1.40 \text{ mol})/(0.900 \text{ L}) = 1.56 \text{ M}$ and $[\text{Cl}^-] = (1.76 \text{ mol})/(0.900 \text{ L}) = 1.96 \text{ M}$. As a final check, note that the total amount of negative charges in solution, 1.76 mol from the Cl^- ions only, equals the total amount of positive charges from the Na^+ ions and the left-over Ba^{2+} ions (which have, of course, a charge twice that of Na^+): $1.40 \text{ mol} + 2 \times 0.180 \text{ mol} = 1.76 \text{ mol}$, as expected!

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5. We start with our friend BaSO_4 , featured in the previous problem. We are asked to show that the mass composition data, 58.84% Ba, 13.74% S, and 27.42% O, is consistent with this formula. 1 mol of BaSO_4 weighs (1 mol) $(233.39 \text{ g mol}^{-1}) = 233.4 \text{ g}$ (to four significant figures). This amount has 1 mol of Ba (that would weigh 137.3 g), 1 mol of S (that would weigh 32.06 g) and 4 mol O (that would weigh $4 \times 16.00 \text{ g} = 64.00 \text{ g}$). Thus the mass fractions are

$$\frac{137.33 \text{ g}}{233.4 \text{ g}} = 0.5884 = 58.84 \% \text{ Ba}$$

$$\frac{32.06 \text{ g}}{233.4 \text{ g}} = 0.1374 = 13.74 \% \text{ S}$$

$$\frac{64.00 \text{ g}}{233.4 \text{ g}} = 0.2742 = 27.42 \% \text{ O}$$

as expected. Now for what's left behind in solution. The H_2S replaces Ag^+ with H^+ ; the $\text{Ba}(\text{NO}_3)_2$ replaces sulfate ion, SO_4^{2-} with nitrate ion, NO_3^- ; and, since the barium nitrate was added in excess, there will also be excess Ba^{2+} in solution. The molar mass of Ag_2SO_4 is $311.794 \text{ g mol}^{-1}$. Each mole of Ag_2SO_4 produces one mole of Ag_2S (molar mass $247.7964 \text{ g mol}^{-1}$). Thus we have a mass of Ag_2S given by

$$\left(\frac{4.52 \text{ g Ag}_2\text{SO}_4}{311.794 \text{ g mol}^{-1} \text{ Ag}_2\text{SO}_4} \right) \left(\frac{1 \text{ mol Ag}_2\text{S}}{\text{mol Ag}_2\text{SO}_4} \right) (247.7964 \text{ g mol}^{-1} \text{ Ag}_2\text{S}) = 3.49 \text{ g Ag}_2\text{S}$$