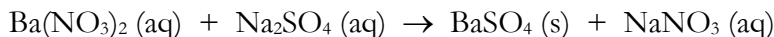
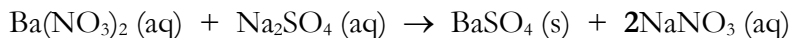


Solutions to the Extra Problems for Chapter 9

1. The reaction is $\text{Ba}(\text{NO}_3)_2(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + 2\text{NaNO}_3(\text{aq})$. First, we need to figure out the reactants. Barium is Ba^{2+} in ionic compounds, and nitrate is NO_3^- . That makes one reactant $\text{Ba}(\text{NO}_3)_2$. Sodium is Na^+ in ionic compounds, and sulfate is SO_4^{2-} , so the other reactant is Na_2SO_4 . They are both aqueous, since they are both in solution with water. We are told that barium sulfate, which would be BaSO_4 , is insoluble, so it is the precipitate. The other compound is made with the other two ions (Na^+ and NO_3^-), so it is NaNO_3 , and it is aqueous, since we aren't told it is insoluble. Thus, the reaction is:



The Ba's are balanced, and to balance the N's, we need to realize that the "2" in $\text{Ba}(\text{NO}_3)_2$ goes with the N and the O's, so there are two N's. That means we must multiply NaNO_3 by two on the products side:



Now everything is balanced.

2. The concentration is 0.706 M. Moles per liter is molarity, so we need to use Equation (9.1). For that, we need moles and liters. We have to convert to get both of those.

$$\frac{15.0 \text{ g } \cancel{\text{AgNO}_3}}{1} \times \frac{1 \text{ mole AgNO}_3}{169.88 \text{ g } \cancel{\text{AgNO}_3}} = 0.0883 \text{ moles AgNO}_3$$

$$\frac{125.0 \cancel{\text{ mL}}}{1} \times \frac{0.001 \text{ L}}{1 \cancel{\text{ mL}}} = 0.1250 \text{ L}$$

Now we can use Equation (9.1):

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{0.0883 \text{ moles}}{0.1250 \text{ L}} = 0.706 \frac{\text{moles}}{\text{L}}$$

3. You will need 110 g. In this case, we know the molarity and want to figure out grams. We can't relate molarity to grams, but we can relate it to moles, as soon as we get liters of solution:

$$\frac{500.0 \cancel{\text{ mL}}}{1} \times \frac{0.001 \text{ L}}{1 \cancel{\text{ mL}}} = 0.5000 \text{ L}$$

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$3.0 \frac{\text{moles}}{\text{L}} = \frac{\text{moles of solute}}{0.5000 \text{ L}}$$

Multiplying both sides of the equation by 0.5000 L gives us:

$$\text{moles of solute} = 3.0 \frac{\text{moles}}{\cancel{\text{L}}} \times 0.5000 \cancel{\text{L}} = 1.5 \text{ moles}$$

Now that we have moles, we can get grams:

$$\frac{1.5 \cancel{\text{moles KCl}}}{1} \times \frac{74.55 \text{ g KCl}}{1 \cancel{\text{mole KCl}}} = 110 \text{ g KCl}$$

4. There will be 32 g formed. We know the limiting reactant is K_2S , because $\text{Bi}(\text{NO}_3)_3$ is in excess. We need to know the number of moles of K_2S , and we can get that by multiplying concentration by the number of liters. Of course, we are given mL, so we have to first convert:

$$\frac{67.1 \cancel{\text{mL}}}{1} \times \frac{0.001 \text{ L}}{1 \cancel{\text{mL}}} = 0.0671 \text{ L}$$

$$2.9 \frac{\text{moles}}{\cancel{\text{L}}} \times 0.0671 \cancel{\text{L}} = 0.19 \text{ moles}$$

Now that we have the moles of the limiting reactant, we can use stoichiometry to calculate the moles of product:

$$\frac{0.19 \cancel{\text{moles K}_2\text{S}}}{1} \times \frac{1 \text{ mole Bi}_2\text{S}_3}{3 \cancel{\text{moles K}_2\text{S}}} = 0.063 \text{ moles Bi}_2\text{S}_3$$

Now that we have moles, we can convert to grams:

$$\frac{0.063 \cancel{\text{moles Bi}_2\text{S}_3}}{1} \times \frac{514.14 \text{ g Bi}_2\text{S}_3}{1 \cancel{\text{mole Bi}_2\text{S}_3}} = 32 \text{ g Bi}_2\text{S}_3$$

5. She needs 12.0 liters of solution. We know how much H_2 the chemist wants to make, so we need to convert that to moles:

$$\frac{50.0 \cancel{\text{g H}_2}}{1} \times \frac{1 \text{ mole H}_2}{2.02 \cancel{\text{g H}_2}} = 24.8 \text{ moles H}_2$$

We can convert that to moles of the reactant in question:

$$\frac{24.8 \cancel{\text{moles H}_2}}{1} \times \frac{2 \text{ moles HCl}}{1 \cancel{\text{mole H}_2}} = 49.6 \text{ moles HCl}$$

Now that we know the number of moles, we can calculate the number of liters using Equation (9.1):

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$4.12 \frac{\text{moles}}{\text{L}} = \frac{49.6 \text{ moles}}{\text{liters of solution}}$$

Multiplying both sides of the equation by “liters of solution” and dividing both sides by 4.12 moles/liter gives us:

$$\text{liters of solution} = \frac{49.6 \cancel{\text{moles}}}{4.12 \frac{\cancel{\text{moles}}}{\text{L}}} = 12.0 \text{ L}$$

6. The concentration is 0.564 m. To get molality, we need moles of solute and kilograms of solvent. Thus, we need to do two conversions:

$$\frac{40.3 \cancel{\text{g MgCl}_2}}{1} \times \frac{1 \text{ mole MgCl}_2}{95.21 \cancel{\text{g MgCl}_2}} = 0.423 \text{ moles MgCl}_2$$

$$\frac{750.0 \cancel{\text{g}}}{1} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 0.7500 \text{ kg}$$

Now we can use Equation (9.2).

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{0.423 \text{ moles}}{0.7500 \text{ kg}} = 0.564 \frac{\text{moles}}{\text{kg}}$$

7. There were 20.6 g of Na₂S used. We can relate molality to moles of the solute if we have the kilograms of solvent:

$$\frac{150.0 \cancel{\text{g}}}{1} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 0.1500 \text{ kg}$$

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

$$1.76 \frac{\text{moles}}{\text{kg}} = \frac{\text{moles}}{0.1500 \text{ kg}}$$

Multiplying both sides of the equation by 0.1500 kg:

$$\frac{1.76 \text{ moles}}{\cancel{\text{kg}}} \times 0.1500 \cancel{\text{kg}} = 0.264 \text{ moles}$$

Now that we know the moles of the solute (Na_2S), we can determine the grams.

$$\frac{0.264 \text{ moles } \cancel{\text{Na}_2\text{S}}}{1} \times \frac{78.04 \text{ g } \text{Na}_2\text{S}}{1 \cancel{\text{ mole } \text{Na}_2\text{S}}} = 20.6 \text{ g } \text{Na}_2\text{S}$$

8. The $(\text{NH}_4)_3\text{P}$ solution will have the lowest freezing point, because there are four ions in $(\text{NH}_4)_3\text{P}$. If the molalities are all the same, the only other property of the solute that matters is the number of particles it splits into in solution (i). $\text{Ba}(\text{NO}_3)_2$ splits into three ions, as do Na_2SO_4 and SrCl_2 . NaCl splits into two ions, while $\text{C}_2\text{H}_6\text{O}$ is covalent and therefore splits into individual molecules, making $i = 1$.

9. The freezing temperature of the solution is -13.6°C . To get the freezing point depression of a solution, we need i , K_f , and m . The problem gives us K_f , and since we know KCl has one K^+ and one Cl^- ion, we know $i = 2$. Now we just have to calculate m :

$$\frac{150.0 \cancel{\text{ g } \text{KCl}}}{1} \times \frac{1 \text{ mole } \text{KCl}}{74.55 \cancel{\text{ g } \text{KCl}}} = 2.012 \text{ moles } \text{KCl}$$

$$\frac{550.0 \cancel{\text{ g}}}{1} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{ g}}} = 0.5500 \text{ kg}$$

Now we can use Equation (9.2).

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{2.012 \text{ moles}}{0.5500 \text{ kg}} = 3.658 \text{ m}$$

Now that we have i , K_f , and m , we can use Equation (9.3):

$$\Delta T = -(2) \cdot \left(1.86 \frac{^\circ\text{C}}{\text{m}} \right) \cdot (3.658 \text{ m}) = -13.6^\circ\text{C}$$

Since the freezing point of water is 0°C (which is exact), that means the freezing point of the solution is -13.6°C . Remember, when we add, we look at precision. Since 0°C is exact, the only thing that's important is that -13.6°C is reported to the tenths place, so the answer must be as well.

10. The boiling point is 103.75°C . We already know the molality for that solution. It's 3.658 m . We also know that $i = 2$, and we are given K_b . Now we can just use Equation (9.4):

$$\Delta T = (2) \cdot \left(0.512 \frac{^\circ\text{C}}{\text{m}} \right) \cdot (3.658 \text{ m}) = 3.75^\circ\text{C}$$

That's not the answer, though. That just tells us the boiling point was raised by 3.75°C . Since the boiling point of water is 100°C (exactly), that means the new boiling point is 103.75°C . Remember, when we add, we look at precision. Since 100°C is exact, the only thing that's important is that 3.75°C is reported to the hundredths place, so the answer must be as well.

11. The concentration is 2.05 m. The boiling point was raised from 78.37 °C to 80.56 °C. That allows us to determine ΔT :

$$\Delta T = 80.56 \text{ }^{\circ}\text{C} - 78.37 \text{ }^{\circ}\text{C} = 2.19 \text{ }^{\circ}\text{C}$$

We are given K_b , and we know that water is a covalent compound, so $i = 1$. That gives us everything we need to calculate molality:

$$\Delta T = (i) \cdot K_b \cdot \text{molality}$$

$$2.19 \text{ }^{\circ}\text{C} = (1) \cdot \left(1.07 \frac{\text{ }^{\circ}\text{C}}{\text{m}}\right) \cdot (\text{molality})$$

Dividing both sides by 1 and 1.07 °C/m gives us:

$$\text{molality} = \frac{2.19 \text{ }^{\circ}\text{C}}{(1) \cdot \left(1.07 \frac{\text{ }^{\circ}\text{C}}{\text{m}}\right)} = 2.05 \text{ m}$$