

1. Suppose a piece of iron (specific heat capacity = $0.449 \text{ J/g}\cdot\text{K}$) with a mass of 21.5 g at a temperature of 100.0°C is dropped into an insulated container of water. The mass of the water is 132.0 g and its temperature before adding the iron is 20.0°C . What will be the final temperature of the system once thermal equilibrium is reached? (specific heat capacity of water = $4.184 \text{ J/g}\cdot\text{K}$)

$$q_{Fe} + q_{H_2O} = 0$$

$$q_{Fe} = -q_{H_2O}$$

$$(21.5 \text{ g}) \left(0.449 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (\Delta T) = - \left[(132.0 \text{ g}) \left(4.184 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (\Delta T) \right]$$

$$(21.5 \text{ g}) \left(0.449 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (T_f - 100.0^\circ\text{C}) = - \left[(132.0 \text{ g}) \left(4.184 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \right) (T_f - 20.0^\circ\text{C}) \right]$$

$$9.6535T_f - 965.35 = -552.288T_f + 11045.76$$

$$561.9415T_f = 12011.11$$

$$T_f = 21.4^\circ\text{C}$$

2. What mass of ice can be melted with the same quantity of heat as required to raise the temperature of 3.50 mol H₂O (l) by 50.0°C? (Δ_{Hfusion} for H₂O (s) = 6.01 kJ mol⁻¹)

First, find the amount of energy needed to raise the temperature of liquid water.

$$q = mc\Delta T$$

$$3.50 \text{ mol H}_2\text{O} * \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 63.07 \text{ g H}_2\text{O}$$

$$q = (63.07 \text{ g}) \left(4.184 \frac{\text{J}}{\text{g}^\circ\text{C}} \right) (50.0^\circ\text{C}) = 13194.244 \text{ J}$$

$$13194.244 \text{ J} * \frac{1 \text{ mol H}_2\text{O}}{6010 \text{ J}} * \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 39.6 \text{ g H}_2\text{O}$$

3. Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005-g sample of menthol is combusted, producing 0.2829 g of CO₂ and 0.1159 g of H₂O. What is the empirical formula for menthol? If menthol has a molar mass of 156 g/mol, what is its molecular formula?

$$0.1159 \text{ g H}_2\text{O} * \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} * \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} = 0.01286 \text{ mol H} * \frac{1.008 \text{ g H}}{1 \text{ mol H}} \\ = 0.01297 \text{ g H}$$

$$0.2829 \text{ g CO}_2 * \frac{1 \text{ mol CO}_2}{44.01 \text{ g CO}_2} * \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} = 0.006428 \text{ mol C} * \frac{12.011 \text{ g C}}{1 \text{ mol C}} \\ = 0.07721 \text{ g C}$$

$$\text{Mass O} = \text{Mass sample} - \text{Mass C} - \text{Mass H} = 0.1005 \text{ g} - 0.07721 \text{ g} - 0.01297 \text{ g} \\ = 0.01032 \text{ g O}$$

$$0.01032 \text{ g O} * \frac{1 \text{ mol O}}{15.9994 \text{ g O}} = 0.000645 \text{ mol O}$$

Take mole ratios to find empirical formula!

$$\text{C: } \frac{0.006428}{0.000645} = 10$$

$$H: \frac{0.01286}{0.000645} = 20$$

$$O: \frac{0.000645}{0.000645} = 1$$

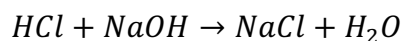
Empirical formula: $C_{10}H_{20}O$

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{156 \frac{g}{mol}}{156 \frac{g}{mol}} = 1$$

Molecular formula: $C_{10}H_{20}O$

4. It takes 83 mL of a 0.45 M NaOH solution to neutralize 235 mL of an HCl solution.
What is the concentration of the HCl solution?

Always write the balanced equation first!



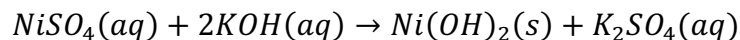
$$\begin{aligned} \frac{0.45 \text{ mol NaOH}}{1 \text{ L}} * 0.083 \text{ L} &= 0.03735 \text{ mol NaOH} * \frac{1 \text{ mol OH}^-}{1 \text{ mol NaOH}} * \frac{1 \text{ mol HCl}}{1 \text{ mol OH}^-} \\ &= 0.03735 \text{ mol HCl} \end{aligned}$$

0.03735 mol of HCl was neutralized.

$$[HCl] = \frac{0.037535 \text{ mol HCl}}{0.235 \text{ L}} = 0.16 \text{ M}$$

5. A solution of 100.0 mL of 0.200 M KOH is mixed with a solution of 200.0 mL of 0.150 M NiSO₄

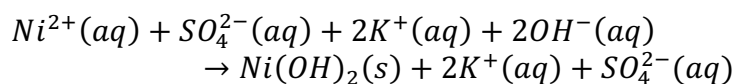
a. Write the balanced chemical equation for the reaction that occurs.



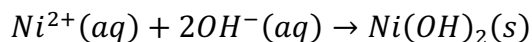
b. What precipitate forms? Write the net ionic equation

Nickel hydroxide is the precipitate!

Total ionic:



Net ionic:



c. How many grams of this precipitate form?

Remember to calculate the limiting reactant!

$$\begin{aligned} 0.200 \text{ M KOH} &= \frac{0.200 \text{ mol KOH}}{1 \text{ L}} * 0.100 \text{ L} * \frac{1 \text{ mol Ni}(\text{OH})_2}{2 \text{ mol KOH}} \\ &= 0.01 \text{ mol Ni}(\text{OH})_2 \end{aligned}$$

$$\begin{aligned} 0.150 \text{ M NiSO}_4 &= \frac{0.150 \text{ mol NiSO}_4}{1 \text{ L}} * 0.200 \text{ L} * \frac{1 \text{ mol Ni}(\text{OH})_2}{1 \text{ mol NiSO}_4} \\ &= 0.03 \text{ mol Ni}(\text{OH})_2 \end{aligned}$$

KOH is the limiting reactant!

$$0.01 \text{ mol Ni}(\text{OH})_2 * \frac{92.71 \text{ g Ni}(\text{OH})_2}{1 \text{ mol Ni}(\text{OH})_2} = 0.927 \text{ g Ni}(\text{OH})_2$$

d. What is the concentration of each ion that remains in solution?

OH⁻: Is part of the limiting reactant, and all goes into forming the precipitate

$$[OH^-] = 0$$

K⁺: Is part of the limiting reactant, but is simply a spectator ion.

$$\frac{0.200 \text{ mol KOH}}{1 \text{ L}} * 0.100 \text{ L} * \frac{1 \text{ mol K}^+}{1 \text{ mol KOH}} = 0.02 \text{ mol K}^+$$

$$[K^+] = \frac{0.02 \text{ mol K}^+}{0.300 \text{ L}} = 0.0667 \text{ M}$$

NiSO₄ is not the limiting reactant. First, with respect to SO₄²⁻, it is a spectator ion, so its initial mole amount will equal its final mole amount.

$$0.150 \text{ M NiSO}_4 = \frac{0.150 \text{ mol NiSO}_4}{1 \text{ L}} * 0.200 \text{ L} * \frac{1 \text{ mol SO}_4^{2-}}{1 \text{ mol NiSO}_4} = 0.03 \text{ mol SO}_4^{2-}$$

$$[SO_4^{2-}] = \frac{0.03 \text{ mol SO}_4^{2-}}{0.300 \text{ L}} = 0.100 \text{ M}$$

For Ni, some of it is going into forming the precipitate, but there will be some left over. First, find out how much is being used up:

$$0.02 \text{ mol KOH} * \frac{1 \text{ mol NiSO}_4}{2 \text{ mol KOH}} * \frac{1 \text{ mol Ni}^{2+}}{1 \text{ mol NiSO}_4} = 0.01 \text{ mol Ni}^{2+} \text{ used up}$$

$$\begin{aligned} 0.150 \text{ M NiSO}_4 &= \frac{0.150 \text{ mol NiSO}_4}{1 \text{ L}} * 0.200 \text{ L} * \frac{1 \text{ mol Ni}^{2+}}{1 \text{ mol NiSO}_4} \\ &= 0.03 \text{ mol Ni}^{2+} \text{ initially} \end{aligned}$$

Thus, 0.02 mol remain.

$$[Ni^{2+}] = \frac{0.02 \text{ mol Ni}^{2+}}{0.300 \text{ L}} = 0.0667 \text{ M}$$

6.

- a. You have a stock solution of 14.8 M NH_3 . How many milliliters of this solution should you dilute to make 1000.0 mL of 0.250 M NH_3 ?

$$0.250 \text{ M } \text{NH}_3 = \frac{0.250 \text{ mol } \text{NH}_3}{1 \text{ L}} * 1 \text{ L} = 0.250 \text{ mol } \text{NH}_3 \text{ required}$$

$$0.250 \text{ mol } \text{NH}_3 * \frac{1 \text{ L}}{14.8 \text{ mol } \text{NH}_3} * \frac{1000 \text{ mL}}{1 \text{ L}} = 16.89 \text{ mL for dilution}$$

- b. If you take a 10.0-mL portion of the stock solution and dilute it to a total volume of 0.500 L, what will be the concentration of the final solution?

First, find out how many moles you have taken out in that 10.0 mL portion.

$$\frac{14.8 \text{ mol } \text{NH}_3}{1 \text{ L}} * 0.01 \text{ L} = 0.148 \text{ mol } \text{NH}_3$$

$$\text{Final solution concentration} = \frac{0.148 \text{ mol } \text{NH}_3}{0.500 \text{ L}} = 0.296 \text{ M } \text{NH}_3$$