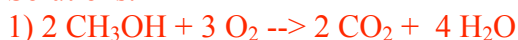


Chem 101 practice problems:

Dear students,

Please do practice problems before looking at the key below. As a policy, keys to practice problems will no longer be provided in the future if students are found not to do the problems *before* looking at the key. This key relates to the first set of practice problems.

Solutions:



2) FW of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ is $63.5 + 32.1 + 4(16.0) + 5(2.0 + 16.0) = 249.6 \text{g/mol}$

$$\% \text{H}_2\text{O} = \frac{\text{g H}_2\text{O}}{\text{g CuSO}_4 \cdot 5\text{H}_2\text{O}} \times 100\% \Rightarrow \text{it is convenient to assume 1 mole:}$$

$$\% \text{H}_2\text{O} = \frac{5(2.0 + 16.0) \text{g H}_2\text{O}}{249.6 \text{g CuSO}_4 \cdot 5\text{H}_2\text{O}} \times 100\% = 36.1\%$$

3) assume 100g: $\# \text{mol N} = 35.0 \text{g N} \frac{1 \text{mol N}}{14.0 \text{g N}} = 2.50 \text{ mol N}$

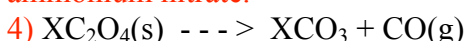
$$\# \text{mol H} = 5.0 \text{ g H} \frac{1 \text{mol H}}{1.0 \text{ g H}} = 5.0 \text{ mol H}$$

$$\# \text{mol O} = 60.0 \text{g O} \frac{1 \text{mol O}}{16.0 \text{g O}} = 3.75 \text{ mol O}$$

so tentative empirical formula is: $\text{N}_{2.50}\text{H}_{5.0}\text{O}_{3.75} \Rightarrow$ divide by least subscript: $\text{NH}_2\text{O}_{1.5}$

Multiply subscripts by 2 to make them all integers: $\text{N}_2\text{H}_4\text{O}_3$

Since this is an ionic compound, there must be a cation and an anion. The most likely cation is NH_4^+ and that leaves us with NO_3^- as the anion. The compound: NH_4NO_3 ammonium nitrate.



Strategy: we can determine g CO produced, then mol CO produced then mol XC_2O_4 present. That would give us the moles of X present. Then deduct the mass of that many moles of C_2O_4 from XC_2O_4 to get the grams of X. And from the grams of X and the moles of X we can get the g/mol of X. In this problem it is "convenient" to work in milligrams (mg) and millimoles (mmol). To illustrate this:

$$\# \text{ mmol CO evolved} = (64.0 \text{mg} - 50.0 \text{mg}) \left(\frac{1 \text{ mmol CO}}{28.0 \text{ mg}} \right) = 0.500 \text{ mmol CO}$$

mmol X = 0.500 mmol. The mg of $\text{C}_2\text{O}_4^{2-}$ in XC_2O_4 is:

$$0.500 \text{ mmol C}_2\text{O}_4^{2-} \left(\frac{88.0 \text{ mg}}{1 \text{ mmol}} \right) = 44.0 \text{ mg}; \text{ so mg X} = 64.0 \text{ mg} - 44.0 \text{ mg} = 20.0 \text{ mg. The}$$

$$\text{atomic weight (AW) of the metal X is: } \frac{20.0 \text{ mg}}{0.500 \text{ mmol}} = 40.0 \text{ mg/mmol} = 40.0 \text{ g/mol}$$

so X = Ca (it's the closest AW in the periodic table. And we know that Ca^{2+} is the stable cation of Ca.

5) We can solve for g H_2SO_4 taking each of the reactants to be limiting. The reactant yielding the least # mol H_2SO_4 will be the one limiting.

Assuming SO₂ limiting: $\# \text{mol H}_2\text{SO}_4 = 100.0 \text{ g SO}_2 \frac{1 \text{ mol SO}_2}{64.1 \text{ g SO}_2} \frac{2 \text{ mol H}_2\text{SO}_4}{2 \text{ mol SO}_2} = 1.56 \text{ mol}$

Assuming O₂ limiting: $\# \text{mol H}_2\text{SO}_4 = 20.0 \text{ g O}_2 \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \frac{2 \text{ mol H}_2\text{SO}_4}{1 \text{ mol O}_2} = 1.25 \text{ mol}$

Assuming H₂O limiting: $\# \text{mol H}_2\text{SO}_4 = 25.0 \text{ g H}_2\text{O} \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \frac{2 \text{ mol H}_2\text{SO}_4}{2 \text{ mol H}_2\text{O}} = 1.39 \text{ mol}$

Since O₂ produces the least amount of sulfuric acid, it must be the limiting reagent.

6) This problem gives the % yield. It also gives the actual yield. We know the formula

for the %yield = $\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$. From the formula for %yield, we can write:

$$\text{theoretical yield} = \frac{\text{actual yield}}{\% \text{ yield}} \times 100\% = \frac{920 \times 10^3 \text{ g}}{75\%} \times 100\% = 1.2 \times 10^6 \text{ g}$$

We use this to calculate the g N₂ needed. The balanced equation: N₂ + 3 H₂ → 2 NH₃

$$\# \text{ g N}_2 \text{ needed} = 1.2 \times 10^6 \text{ g NH}_3 \frac{1 \text{ mol NH}_3}{17.0 \text{ g NH}_3} \frac{1 \text{ mol N}_2}{2 \text{ mol NH}_3} \frac{28.0 \text{ g N}_2}{\text{mol N}_2} = 9.9 \times 10^5 \text{ g}$$

7) Balanced equation: 4 Al + 3 O₂ → 2 Al₂O₃

First calculate g Al₂O₃ formed if Al is limiting:

$$\# \text{ g Al}_2\text{O}_3 = 5.40 \text{ g Al} \left(\frac{1 \text{ mol}}{27.0} \right) \left(\frac{2 \text{ mol Al}_2\text{O}_3}{4 \text{ mol Al}} \right) (102.0 \text{ g/mol Al}_2\text{O}_3) = 10.2 \text{ g}$$

Or, if O₂ is limiting:

$$\# \text{ g Al}_2\text{O}_3 = 5.00 \text{ g O}_2 \left(\frac{1 \text{ mol}}{32.0} \right) \left(\frac{2 \text{ mol Al}_2\text{O}_3}{3 \text{ mol O}_2} \right) (102.0 \text{ g/mol Al}_2\text{O}_3) = 10.6 \text{ g}$$

That means that Al is limiting. So 0 g of Al will remain. And 10.2 g Al₂O₃ is formed.

$$\# \text{ g O}_{2,\text{remaining}} = 5.00 \text{ g} - (5.40 \text{ g Al}) \left(\frac{1 \text{ mol}}{27.0 \text{ g}} \right) \left(\frac{3 \text{ mol O}_2}{4 \text{ mol Al}} \right) (32.0 \text{ g O}_2/\text{mol O}_2) = 5.00 - 4.80 = 0.20 \text{ g O}_2$$

8) a) H₂CO₃ + 2 NaOH → Na₂CO₃ + 2 H₂O

b) H₂CO₃ + 2 OH⁻ → CO₃²⁻ + 2 H₂O

9) Write balanced equation: H₃PO₄ + 3 NaOH → 3 H₂O + Na₃PO₄

Calculate # mol Na₃PO₄ formed assuming either the acid or the base is limiting:

$$\# \text{ mol Na}_3\text{PO}_4 = 2 \text{ mol H}_3\text{PO}_4 \left(\frac{1 \text{ mol Na}_3\text{PO}_4}{1 \text{ mol H}_3\text{PO}_4} \right) = 2 \text{ mol Na}_3\text{PO}_4$$

$$\text{or, } \# \text{ mol Na}_3\text{PO}_4 = 3 \text{ mol NaOH} \left(\frac{1 \text{ mol Na}_3\text{PO}_4}{3 \text{ mol NaOH}} \right) = 1 \text{ mol Na}_3\text{PO}_4$$

so, the phosphoric acid is limiting.

10) Titration reaction: HNO₃ + NaOH → H₂O + NaNO₃ :

$$[\text{HNO}_3] = \frac{20.0 \text{ mL NaOH} \frac{1 \text{ L}}{1000 \text{ mL}} \frac{0.400 \text{ mol}}{\text{L}} \frac{1 \text{ mol HNO}_3}{\text{mol NaOH}}}{25.0 \text{ mL} \frac{1 \text{ L}}{1000 \text{ mL}}} = 0.320 \text{ M}$$

Or, we can say: at equiv. pt: # mol H⁺ donated = # mol H⁺ accepted

$$\# \text{ mol HNO}_3 = \# \text{ mol NaOH} \Rightarrow M_{\text{HNO}_3} V_{\text{HNO}_3} = M_{\text{NaOH}} V_{\text{NaOH}}$$

$$\Rightarrow M_{\text{HNO}_3} = (.400 \text{ M})(20.0 \text{ mL}) / (25.0 \text{ mL}) = 0.320 \text{ M}$$

11) titration equation: $\text{H}_2\text{SO}_4 + 2 \text{NaOH} \rightarrow 2 \text{H}_2\text{O} + \text{Na}_2\text{SO}_4$

$$[\text{H}_2\text{SO}_4] = \frac{20.0\text{mL NaOH} \frac{1\text{L}}{1000\text{mL}} \frac{0.400\text{mol}}{\text{L}} \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}}{25.0\text{mL} \frac{1\text{L}}{1000\text{mL}}} = 0.160\text{M}$$

Or, we can say: at equiv pt: : # mol H^+ donated = # mol H^+ accepted
mol $\text{H}_2\text{SO}_4 \times 2 = \# \text{ mol NaOH} \Rightarrow 2M_{\text{H}_2\text{SO}_4}V_{\text{H}_2\text{SO}_4} = M_{\text{NaOH}}V_{\text{NaOH}}$
 $\Rightarrow M_{\text{H}_2\text{SO}_4} = \frac{.400\text{M}(20.0\text{mL})}{2(25.0\text{mL})} = 0.320\text{M}$