Chemical equations describe mole ratios/proportions of reactants and products

$$Fe_2O_{3(s)} + \tfrac{3}{2}C_{(s)} \rightarrow 2Fe_{(s)} + \tfrac{3}{2}CO_{2(g)}$$

or

$$2Fe_2O_{3(s)} + 3C_{(s)} \rightarrow 4Fe_{(s)} + 3CO_{2(g)}$$

If we want 10 moles Fe, how much Fe_2O_3 and C are needed

$$10 \text{ } \cancel{\text{moles Fe}} \times \frac{2 \text{ moles } Fe_2O_3}{4 \text{ } \cancel{\text{moles Fe}}} = 5 \text{ moles } Fe_2O_3 \text{ needed}$$

$$10 \text{ } \cancel{\text{moles Fe}} \times \frac{3 \text{ moles C}}{4 \text{ } \cancel{\text{moles Fe}}} = 7.5 \text{ moles C needed}$$

If we make 10 moles Fe, how much CO_2 is made

$$10 \text{ } \cancel{\text{moles Fe}} \times \frac{3 \text{ moles } CO_2}{4 \text{ } \cancel{\text{moles Fe}}} = 7.5 \text{ moles } CO_2 \text{ made}$$

If we want 100 kg of Fe, how many kg of Fe_2O_3 and C do we need

* Need to be working in moles

$$100 \text{ } \cancel{\text{kg}} \text{ Fe} \times \frac{1000 \cancel{g}}{1 \cancel{\text{kg}}} \times \frac{1 \text{ mole Fe}}{55.85 \cancel{g \text{ Fe}}} = 1792 \text{ moles Fe}$$

$$1792 \text{ } \cancel{\text{moles Fe}} \times \frac{2 \text{ } \cancel{\text{mole } Fe_2O_3}}{4 \text{ } \cancel{\text{moles Fe}}} \times \frac{159.6 \text{ } \cancel{g} \text{ } Fe_2O_3}{1 \text{ } \cancel{\text{moles } Fe_2O_3}} \times \frac{1 \text{ kg}}{1000 \cancel{g}} = 143 \text{ kg } Fe_2O_3$$

$$1792 \text{ mole Fe} \times \frac{3 \text{ moles C}}{4 \text{ moles Fe}} \times \frac{12 \text{ g C}}{1 \text{ moles C}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 16.1 \text{ kg C}$$

Limiting reactant situation – run out of one reactant before we run out of other(s)

– limiting reactant controls how much product is made

Sandwich – 2 slices of bread
1 piece of cheese
3 slices of turkey

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47 kg Fe_2O_3 and 4.3 kg C – how much Fe can we make

1. Find limiting reactant
2. How much product made
3. How much leftover reactant

$$47 \text{ kg } Fe_2O_3 \times \frac{1000 g}{1 kg} \times \frac{1 \text{ mol } Fe_2O_3}{159.6 g} = 294.5 \text{ moles } Fe_2O_3$$

$$4.3 \text{ kg } C \times \frac{1000 g}{1 kg} \times \frac{1 \text{ mol } C}{12 gC} = 358 \text{ moles } C$$

$$2 Fe_2O_{3(s)} + 3 C_{(s)} \rightarrow 4 Fe_{(s)} + 3 CO_{2(g)}$$

How much C needed to react w/ 294.5 moles Fe_2O_3?

$$294.5 \text{ moles } Fe_2O_3 \times \frac{3 \text{ moles } C}{2 \text{ moles } Fe_2O_3} = 441.8 \text{ moles } C \text{ needed}$$
to react w/ 294.5 moles Fe_2O_3

* Not enough C to react with all Fe_2O_3
C runs out with some Fe_2O_3 unreacted
C is limiting reactant

How much Fe? – based on amount of C

$$358 \text{ moles } C \times \frac{4 \text{ moles } Fe}{3 \text{ moles } C} = 477 \text{ moles } Fe \times \frac{55.85 g}{1 \text{ mole } Fe} \times \frac{1 kg}{1000 g} = 27 \text{ kg } Fe$$

How much Fe_2O_3 left over?
Initially 294.5 moles Fe_2O_3

$$477 \text{ moles Fe made} \times \frac{2 \text{ moles } Fe_2O_3}{4 \text{ moles } Fe} = 238.5 \text{ moles } Fe_2O_3$$

$294.5 - 238.5 = 56$ moles Fe_2O_3 left over

$56 \text{ moles } Fe_2O_3 \times \frac{159.6 \text{ g}}{1 \text{ mole}} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 8.9 \text{ kg } Fe_2O_3$ left over

$$\text{Percent yield} = \frac{\text{what you actually got}}{\text{what you thought you'd get}} \times 100$$

$$\frac{20 \text{ cupcakes}}{24 \text{ cupcakes}} \times 100 = 83\% \text{ yield}$$

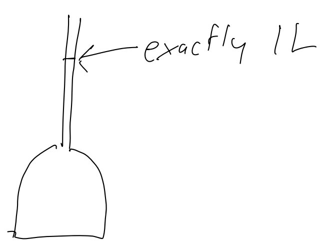
$$\frac{\overbrace{24 \text{ kg Fe}}^{\text{actual}}}{\underbrace{27 \text{ kg Fe}}_{\text{expected}}} \times 100 = 89\% \text{ yield}$$

To keep track of moles when we have solutions we use "molarity"

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

not the same as liters of solvent

2 moles $C_6H_{12}O_6$ dissolved in enough water to make 1 L of solution



$$\frac{2 \text{ moles sugar}}{1 \text{ L sol'n}} = 2\,\overset{\text{"molar"}}{M}$$

$$[C_6H_{12}O_6] = 2\,M$$

In 0.005 L of the 2M sugar solution

$$0.005 \text{ L} \times \frac{2 \text{ moles sugar}}{1 \text{ L sol'n}} = 0.01 \text{ moles sugar}$$

Ex: 25.3 g NaCl dissolved in enough water to make 0.5 L of solution

$$25.3\,g\ NaCl \times \frac{1\ mol}{58.4\,g} = 0.433\ \text{moles}\ NaCl$$

$$[NaCl] = \frac{0.433\ \text{moles}}{0.5\ L} = 0.866\ M$$

Ex: 11.1 g $Mg(NO_3)_2$ dissolved in enough water to make 0.25 L of solution

$$11.1\,g\ Mg(NO_3)_2 \times \frac{1\ mol}{148.3\,g} = 0.075\ mol\ Mg(NO_3)_2$$

$$[Mg(NO_3)_2] = \frac{0.075\ mol}{0.25\ L} = 0.3\ M$$

$$Mg(NO_3)_{2\,(s)} \rightarrow Mg^{2+}_{(aq)} + 2\,NO^-_{3\,(aq)}$$

0.075 mol $Mg(NO_3)_2$ dissolves to give 0.075 mol Mg^{2+} and 0.150 mol NO_3^-

In our 0.3 M $Mg(NO_3)_2$ sol'n

$$\frac{0.075\ mol\ Mg^{2+}}{0.25\ L} \rightarrow [Mg^{2+}] = 0.3\ M$$

$$\frac{0.150\ mol\ NO_3^-}{0.25\ L} \rightarrow [NO_3^-] = 0.6\ M$$

$$\underline{3}\,Co^{2+}_{(aq)} + 6\,Cl^-_{(aq)} + 6\,Na^+_{(aq)} + \underline{2}\,PO_4^{3-}{}_{(aq)} \rightarrow Co_3(PO_4)_{2\,(s)} + 6\,Na^+_{(aq)} + 6\,Cl^-_{(aq)}$$

20 mL of 0.2 M $CoCl_2$ sol'n
15 mL of 0.25 M Na_3PO_4 sol'n } mix → precipitate → how many grams of precipitate?

$$0.02\ L\ CoCl_2\ \text{sol'n} \times \frac{0.2\ \text{moles}\ CoCl_2}{1\ L} \times \frac{1\ \text{mole}\ Co^{2+}}{1\ \text{mole}\ CoCl_2} = 0.004\ mol\ Co^{2+}$$

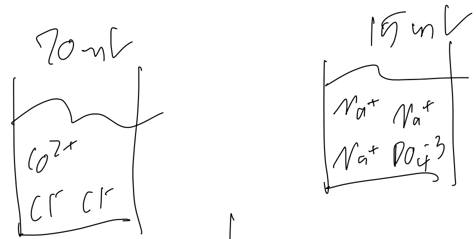
$$0.015\ L\ Na_3PO_4\ \text{sol'n} \times \frac{0.25\ mol\ Na_3PO_4}{1\ L} \times \frac{1\ \text{mole}\ PO_4^{3-}}{1\ \text{mole}\ Na_3PO_4} = 0.00375\ \text{mole}\ PO_4^{3-}$$

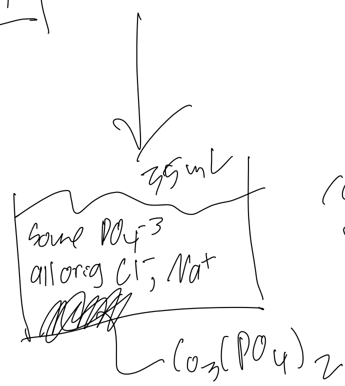
How much PO_4^{3-} needed to react w/ all 0.004 moles Co^{2+}

$$0.004 \text{ mol } Co^{2+} \times \frac{2 \text{ mol } PO_4^{-3}}{3 \text{ mol } Co^{2+}} = 0.0027 \text{ mol } PO_4^{-3}$$

we have more PO_4^{-3} than we need
so Co^{2+} is limiting

$$0.004 \text{ mol } Co^{2+} \times \frac{1 \text{ mol } Co_3(PO_4)_2}{3 \text{ mol } Co^{2+}} \times \frac{367 \text{ g } Co_3(PO_4)_2}{1 \text{ mol } Co_3(PO_4)_2} = 0.49 \text{ g } Co_3(PO_4)_{2(s)} \text{ made}$$





could figure out molarity of PO_4^{-3}, Cl^-, Na^+ in new combined solution

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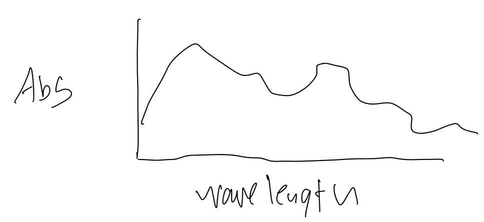
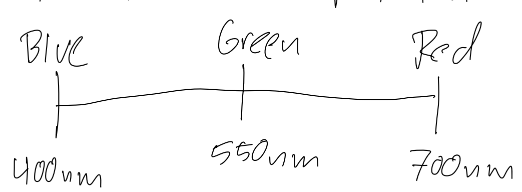
Spectrophotometry – some colors (wavelengths) of light pass thru a material and some are absorbed

$$\text{Absorbance} = -\log_{10}(\text{Transmission})$$

$$\text{Transmission} = \frac{\text{amount of light that passed thru}}{\text{amount of light that was put in}}$$





$$Abs = \varepsilon \times l \times c$$

c – molarity of solution

l – distance thru sample "pathlength" in cm

ε – extinction coefficient $M^{-1}cm^{-1}$