Chemical equations - tells you proportions or ratios
in terms of moles
- not absolute amounts

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

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

or

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

If we want 10 moles Fe, how much Fe_2O_3, C do we need?

* Mole ratios are the key

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

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

If we want 100 kg Fe, how many kg C, how many kg Fe_2O_3 are needed?

$$100 \text{ kg Fe} \times \frac{1000g}{1kg} \times \frac{1 \text{ mole}}{55.8g} = 1792 \text{ moles Fe}$$

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

$$1792 \text{ mole Fe} \times \frac{3 \text{ mole C}}{4 \text{ moles Fe}} \times \frac{12.01g\ C}{1 \text{ mole C}} \times \frac{1kg}{1000g} = 16.1 \text{ kg C}$$

Limiting Reactant - usually we run out of one reactant before others
- once you run out of one of the components you can't make any more products

Ex: 47 kg Fe_2O_3
~~5.3~~ 4.3 kg C

1. what is limiting reactant
2. How much Fe
3. How much reactant left over

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

$\cancel{5300} 4300 g \times \frac{1 mol}{12.01 g} = \cancel{441.3}\ 358.0$ moles C

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

How much C needed to react all 294.5 moles Fe_2O_3?

$294.5\ mol\ Fe_2O_3 \times \frac{3\ moles\ C}{2\ mole\ Fe_2O_3} = 441.8$ moles C needed

but we only have 358 moles C
so we will run out of C before running out of Fe_2O_3
C is limiting reactant

Use mole ratio from LR to get product

$358\ moles\ C \times \frac{4\ moles\ Fe}{3\ moles\ C} = 477.3$ moles Fe $\times \frac{55.8g}{mol} \rightarrow$ 26.6 kg Fe

Started w/ 294.5 moles Fe_2O_3

$477.3\ moles\ Fe \times \frac{2\ moles\ Fe_2O_3}{4\ moles\ Fe} = 238.7$ moles Fe_2O_3 needed to make 477.3 moles Fe

Left over $294.5 - 238.7 = 55.8$ moles $Fe_2O_3 \rightarrow 8.9$ kg Fe_2O_3
unreacted

In reality things don't always work as well as the recipe (balanced eq) says

Percent yield $\dfrac{\text{amount of product actual}}{\text{amount of product expected}} \times 100$

If we only made 410 moles of Fe instead of expected 477

$$\frac{410}{477} \times 100 = 86\% \text{ yield}$$

For things dissolved in solution we use "molarity"

Molarity of a solute $= \dfrac{\text{moles of solute}}{\text{1 liter of solution}}$

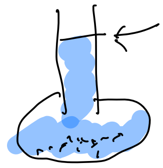
1 liter of solution $\neq$ 1 liter of solvent

[X] = "the molarity of X"

sugar solution with 2 moles of sugar per 1 liter of solution molarity $= \dfrac{2 \text{ mol}}{1 L}$ = 2 molar solution
2M

[sugar] = 2M

We dissolve 25.3g NaCl in enough water to make 0.5L of solution



$$25.3g\ NaCl \times \frac{1\ mol\ NaCl}{58.44g\ NaCl} = 0.433\ mol\ NaCl$$

$$\frac{0.433\ mol}{0.5L} = 0.866M\ NaCl$$

11.1 g $Mg(NO_3)_2$ dissolved in enough water to make 250 mL of solution

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

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

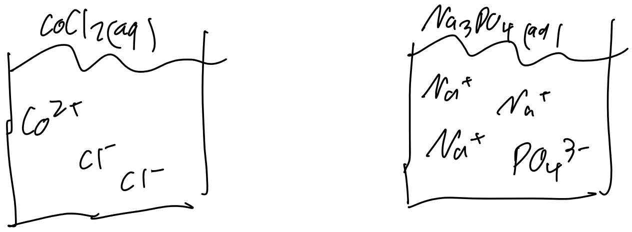
moles of $Mg^{2+} \rightarrow 0.075$ $\quad [Mg^{2+}] = \frac{0.075\ mol}{.25L} = 0.3M$

mole of $NO_3^- \rightarrow 0.15$ $\quad [NO_3^-] = \frac{0.15\ mol}{.25L} = 0.6M$

If we had 0.3M $Mg(NO_3)_2$ solution
and we wanted 0.05 moles of $Mg(NO_3)_2$
what volume of sol'n needed

$$0.05\ moles\ Mg(NO_3)_2 \times \frac{1\ \ L\ of\ solution}{0.3\ moles\ Mg(NO_3)_2} = 0.167\ L\ of\ solution$$

$$molarity = \frac{moles}{liters}$$



$$3Co^{2+}_{(aq)} + \cancel{2}6Cl^-_{(aq)} + \cancel{3}6Na^+_{(aq)} + 2PO_4^{3-}_{(aq)} \rightarrow Co_3(PO_4)_2 + \cancel{3}6Na^+_{(aq)} + \cancel{2}6Cl^-_{(aq)}$$

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

20 mL of 0.2 M $CoCl_{2(aq)}$ + 15 mL of 0.25 M $Na_3PO_{4(aq)}$

How many grams $Co_3(PO_4)_2$ form

$$0.02\,L \times \frac{0.2\text{ mol }CoCl_2}{L} \times \frac{1\text{ mol }Co^{2+}}{1\text{ mol }CoCl_2} = 0.004\text{ moles }Co^{2+}_{(aq)}$$

$$0.015\,L \times \frac{0.25\text{ mol }Na_3PO_4}{L} \times \frac{1\text{ mol }PO_4^{3-}}{1\text{ mol }Na_3PO_4} = 0.00375\text{ moles }PO_4^{3-}{}_{(aq)}$$

If we have 0.004 moles Co^{2+}, how much PO_4^{3-} needed

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

we have > 0.0027 mol PO_4^{3-}

so Co^{2+} is LR

If we have 0.00375 mol PO_4^{3-} how much Co^{2+} needed

$$0.00375\text{ mol }PO_4^{3-} \times \frac{3\text{ mol }Co^{2+}}{2\text{ mol }PO_4^{3-}} = 0.0056\text{ moles }Co^{2+}$$

we have < 0.0056 mol Co^{2+}

so Co^{2+} LR

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

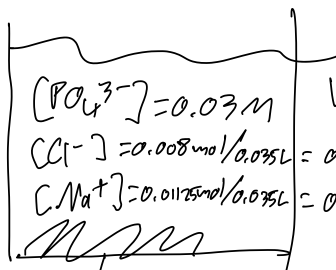
or 0.0013 mol $Co_3(PO_4)_2$

Left over PO_4^{3-} Initially 0.00375 moles

$$\text{In } 0.0013\text{ mol }Co_3(PO_4)_2 \times \frac{2\text{ mol }PO_4^{3-}}{1\text{ mol }Co_3(PO_4)_2} = 0.0027\text{ mol }PO_4^{3-}\text{ in precipitate}$$

$0.00375 - 0.0027 = 0.00105$ mol PO_4^{3-} still dissolved

$[PO_4^{3-}]$ after precipitation $= \dfrac{0.00105 \text{ mol}}{(0.02 + 0.015) L} = 0.03 M \; PO_4^{3-}$



$V_{tot} = 35$ mL

0.0013 mol solid $Co_3(PO_4)_2$