

Example: You need 1.75 moles of iron. What mass of iron do you need to weigh out on the balance?

$$55.85 \text{ g Fe} = 1 \text{ mol Fe}$$

$$1.75 \text{ mol Fe} \times \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 97.7 \text{ g Fe}$$

WHAT ABOUT COMPOUNDS? FORMULA WEIGHT

Example: 25.0 g of WATER contain how many MOLES of water molecules?

$$\text{H}_2\text{O} : \quad \text{H} : 2 \times 1.008 = 2.016$$

$$\quad \quad \quad \text{O} : 1 \times 16.00 = 16.00$$

18.016 ← FORMULA WEIGHT of water

FORMULA WEIGHT is the mass of one mole of either an element OR a compound.

$$18.016 \text{ g H}_2\text{O} = 1 \text{ mol H}_2\text{O}$$

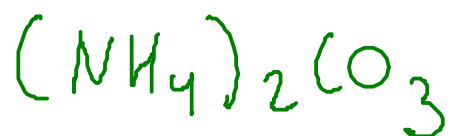
$$25.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.016 \text{ g H}_2\text{O}} = 1.39 \text{ mol H}_2\text{O}$$

Formula weight goes by several names:

- For atoms, it's the same thing as ATOMIC WEIGHT
- For molecules, it's called MOLECULAR WEIGHT
- Also called "MOLAR MASS"

Example: How many grams of ammonium carbonate do we need to weigh out to get 3.65 moles of ammonium carbonate?

Find the formula of ammonium carbonate!



Find formula weight!

$$\text{N} : 2 \times 14.01$$

$$\text{H} : 8 \times 1.008$$

$$\text{C} : 1 \times 12.01$$

$$\text{O} : 3 \times 16.00$$

$$96.094$$

$$96.094 \text{ g } (\text{NH}_4)_2\text{CO}_3 = \text{mol } (\text{NH}_4)_2\text{CO}_3$$

$$3.65 \text{ mol } (\text{NH}_4)_2\text{CO}_3 \times \frac{96.094 \text{ g } (\text{NH}_4)_2\text{CO}_3}{\text{mol } (\text{NH}_4)_2\text{CO}_3} =$$

$$= 351 \text{ g } (\text{NH}_4)_2\text{CO}_3$$

PERCENTAGE COMPOSITION

- sometimes called "percent composition" or "percent composition by mass"
- the percentage of each element in a compound, expressed in terms of mass

Example: Find the percentage composition of ammonium nitrate.

$$\text{NH}_4\text{NO}_3 : \text{N} : 2 \times 14.01 = 28.02$$

$$\text{H} : 4 \times 1.008 = 4.032$$

$$\text{O} : 3 \times 16.00 = 48.00$$

$$\underline{80.052 \text{ g NH}_4\text{NO}_3 = 1 \text{ mol NH}_4\text{NO}_3}$$

These numbers are the masses of each element in a mole of the compound!

$$\% \text{ N} : \frac{28.02 \text{ g N}}{80.052 \text{ g total}} \times 100\% =$$

$$35.0\% \text{ N}$$

$$\% \text{ H} : \frac{4.032 \text{ g H}}{80.052 \text{ g total}} \times 100\% =$$

$$5.0\% \text{ H}$$

$$\% \text{ O} : \frac{48.00 \text{ g O}}{80.052 \text{ g total}} \times 100\% =$$

$$60.0\% \text{ O}$$

Check: All the percentages should sum to 100% - taking roundoff error into account!

So far, we have

- looked at how to determine the composition by mass of a compound from a formula
- converted from MASS to MOLES (related to the number of atoms/molecules)
- converted from MOLES to MASS

Are we missing anything?

- What about SOLUTIONS, where the desired chemical is not PURE, but found DISSOLVED IN WATER?
- How do we deal with finding the moles of a desired chemical when it's in solution?

MOLAR CONCENTRATION *

- unit: MOLARITY (M): moles of dissolved substance per LITER of solution

$$M = \text{molarity} = \frac{\text{moles of SOLUTE}}{\text{L SOLUTION}}$$

↖ dissolved substance

$$6.0 \text{ M HCl solution} = \frac{6.0 \text{ mol HCl}}{\text{L}}$$

If you have 0.250 L (250 mL) of 6.0 M HCl, how many moles of HCl do you have?

$$6.0 \text{ mol HCl} = \text{L}$$

$$0.250 \text{ L} \times \frac{6.0 \text{ mol HCl}}{\text{L}} = \boxed{1.5 \text{ mol HCl}}$$

*See SECTIONS 4.7 - 4.10 for more information about MOLARITY and solution calculations (p 154 - 162 - 9th edition) (p 156-164 - 10th edition)

If you need 0.657 moles of hydrochloric acid, how many liters of 0.0555 M HCl do you need to measure out?

$$0.0555 \text{ mol HCl} = 1 \text{ L}$$

$$0.657 \text{ mol HCl} \times \frac{1 \text{ L}}{0.0555 \text{ mol HCl}} = \boxed{11.8 \text{ L}}$$

11800 mL

This is too large a volume for lab-scale work. To get a more reasonable volume, we should use a more concentrated solution ... like the 6.00 M HCl described below!

What if we used 6.00 M HCl?

$$6.00 \text{ mol HCl} = 1 \text{ L}$$

$$0.657 \text{ mol HCl} \times \frac{1 \text{ L}}{6.00 \text{ mol HCl}} = \boxed{0.110 \text{ L}}$$

110. mL

110 mL is a reasonable volume for lab-scale work. We can measure this out easily in a 250 mL cylinder.