

Stoichiometry & Density

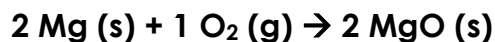
Density can be included in stoichiometry problems because it can be used to calculate **mass** (grams) or **volume** (liters). Density units often are shown as **g/ml**, **g/cm³**, **g/L**, and **g/dm³**. It is important to remember that **cm³** and **ml** are interchangeable, just like **liters** and **dm³** are interchangeable.

For any gas at STP, the molar volume is 22.4 Liters.

Example #1: Find the density of oxygen gas using molar mass and molar volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\left(\frac{32 \text{ grams}}{1 \text{ mole}}\right)}{\left(\frac{22.4 \text{ Liters}}{1 \text{ mole}}\right)} = 1.429 \text{ g/L}$$

Example #2: Using the density of oxygen gas from above, calculate the mass of magnesium required to react with 100. ml of oxygen.



$$\left(\frac{100 \text{ ml O}_2}{1}\right) \left(\frac{1 \text{ Liter O}_2}{1000 \text{ ml O}_2}\right) \left(\frac{1.429 \text{ g O}_2}{1 \text{ Liter O}_2}\right) \left(\frac{1 \text{ mole O}_2}{32 \text{ g O}_2}\right) \left(\frac{2 \text{ moles Mg}}{1 \text{ moles O}_2}\right) \left(\frac{24.3 \text{ g Mg}}{1 \text{ mole Mg}}\right) = 0.106 \text{ g Mg}$$

Notice in the purple box above that the density multiplied by the inverse of the molar mass equals the inverse of the molar volume.

$$\left(\frac{1.429 \text{ g O}_2}{1 \text{ Liter O}_2}\right) \left(\frac{1 \text{ mole O}_2}{32 \text{ g O}_2}\right) = \left(\frac{1 \text{ mole}}{22.4 \text{ Liters}}\right)$$

So technically, you can do this same problem without using density...

$$\left(\frac{100 \text{ ml O}_2}{1}\right) \left(\frac{1 \text{ Liter O}_2}{1000 \text{ ml O}_2}\right) \left(\frac{1 \text{ mole O}_2}{22.4 \text{ Liters O}_2}\right) \left(\frac{2 \text{ moles Mg}}{1 \text{ moles O}_2}\right) \left(\frac{24.3 \text{ g Mg}}{1 \text{ mole Mg}}\right) = 0.106 \text{ g Mg}$$

WARNING! Don't assume the density given is at STP, unless specifically told so in the problem.

Example #3: What mass of sodium azide must be included in an air bag to generate 68.0 L of N₂? Use 0.916 g/L as the density of nitrogen gas.



$$\left(\frac{68 \text{ L N}_2}{1 \text{ air bag}} \right) \left(\frac{0.916 \text{ g N}_2}{1 \text{ Liter N}_2} \right) \left(\frac{1 \text{ mole N}_2}{28 \text{ g N}_2} \right) \left(\frac{2 \text{ moles NaN}_3}{3 \text{ moles N}_2} \right) \left(\frac{65 \text{ g NaN}_3}{1 \text{ mole NaN}_3} \right) = 96.4 \text{ g NaN}_3$$

The actual density of nitrogen gas is 1.25 g/L at STP. This word problem as it is given, is not at STP. You can verify this because molar mass divided by density does not equal molar volume.

$$\frac{\left(\frac{28 \text{ g N}_2}{1 \text{ mole N}_2} \right)}{\left(\frac{0.916 \text{ g N}_2}{1 \text{ Liter N}_2} \right)} = \left(\frac{30.6 \text{ L N}_2}{1 \text{ mole N}_2} \right) \quad \text{NOT} \quad \left(\frac{22.4 \text{ Liters}}{1 \text{ mole}} \right)$$