

Chem 11 finals review #3 balancing and moles

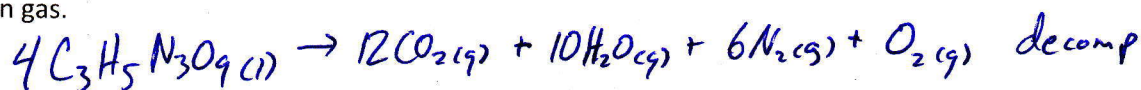
Part 1: Balance and classify the following equations (you may have to predict the products.)

*** No net ionic equations at this level. If you don't know what that is, don't worry about it. ***

- 1) Solid magnesium reacts with gaseous oxygen to form solid magnesium oxide.



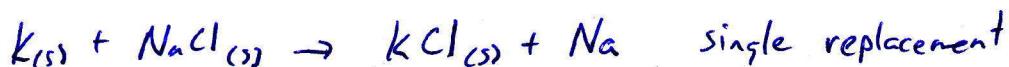
- 2) Liquid nitroglycerine ($\text{C}_3\text{H}_5\text{N}_3\text{O}_9$) breaks down into carbon dioxide gas, water vapor, nitrogen gas and oxygen gas.



- 3) Solid lead is mixed with potassium chloride powder.

no reaction (activity series)

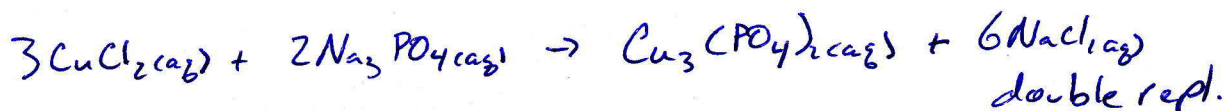
- 4) Solid potassium is mixed with table salt.



- 5) ^{solid} potassium chloride is mixed with fluorine gas



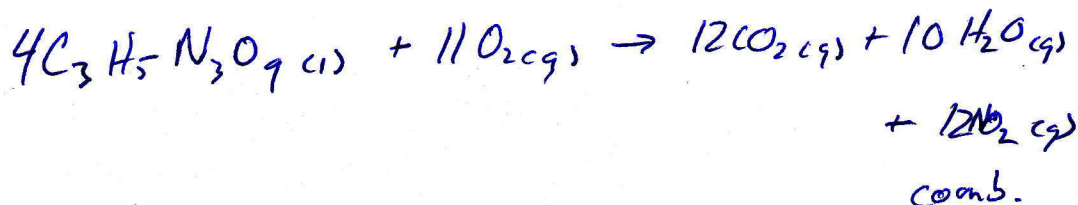
- 6) A solution of copper (II) chloride is mixed with a solution of sodium phosphate



- 7) Phosphoric acid is mixed with aqueous sodium hydroxide



- 8) Liquid nitroglycerin is ignited.



Part 2: Percent composition, empirical and molecular formula

- 1) What is the percent composition of each element in Nitroglycerine?



$C \rightarrow 3 \times 12.0 g = 36.0 g \quad \frac{36.0}{227.0} \times 100\% = 15.8\%$

$H \rightarrow 5 \times 1.0 g = 5.0 g \quad \frac{5.0}{227.0} \times 100\% = 2.2\%$

$N \rightarrow 3 \times 14.0 g = 42.0 g \quad \frac{42.0}{227.0} \times 100\% = 18.5\%$

$O \rightarrow 9 \times 16.0 g = 144.0 g \quad \frac{144.0}{227.0} \times 100\% = 63.4\%$

- 2) An imaginary compound consists of 15.96% C, 2.33% H, 21.28% O and 26.367% Br. What is its empirical formula?

Assume 100g so % = g

$C \quad 15.96g \times \frac{1 mol}{12.0g} = 1.33$

$H \quad 2.33 \times \frac{1 mol}{1.0g} = 2.33$

$O \quad 21.28g \times \frac{1 mol}{16.0g} = 1.33$

$Br \quad 26.367g \times \frac{1 mol}{79.9g} = 0.33$

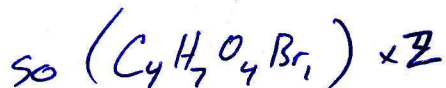
$C_{1.33} H_{2.33} O_{1.33} Br_{0.33} \quad 0.33 \approx \frac{1}{3}$
so $\times 3$



empirical mass = $198.9 \frac{g}{mol}$

- 3) If the compound from #2 has a molecular mass of $397.8 \frac{g}{mol}$, what is its molecular formula?

$\frac{397.8 \frac{g}{mol}}{198.9 \frac{g}{mol}} = 2$



Part 2: Basic mole calculations

- 1) What is the mass of 12.2 moles of naturally occurring Bromine?

$12.2 mol Br_2 \times \frac{159.8 g}{1 mol} = 1.95 \times 10^3 g$

- 2) How many moles of iron are there in a 1.24 g pure iron nail?

$$1.24 \text{ g} \times \frac{1 \text{ mol}}{55.8 \text{ g}} = 0.0222 \text{ mol}$$

- 3) How many molecules of Iodine are there in an 8.1 mole sample?

$$8.1 \text{ mol} \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol}} = 4.9 \times 10^{24} \text{ molecules}$$

- 4) How many moles of copper are there in a 3.10 g pure copper penny?

$$3.10 \text{ g} \times \frac{1 \text{ mol}}{63.5 \text{ g}} = 0.0488 \text{ mol}$$

- 5) How many moles of oxygen gas are there in a 61.2 L sample at STP?

$$61.2 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 2.73 \text{ mol}$$

- 6) How many liters of volume does a 12.1 mole sample of gas take up at STP?

$$12.1 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 271 \text{ L}$$

- 7) What is the molarity of the solution if 2.7 moles of solute is dissolved in 1.3 L of solvent?

$$M = \frac{n}{V} = \frac{2.7 \text{ mol}}{1.3 \text{ L}} = 2.1 \text{ M} \quad \text{or } 2.1 \frac{\text{mol}}{\text{L}}$$

- 8) How many moles of solute must you dissolve in 0.76 L of solvent to get a 1.7 M solution?

$$M = \frac{n}{V} \rightarrow 1.7 \text{ M} = \frac{n}{0.76} \rightarrow n = 1.3 \text{ mol}$$

- 9) How many liters of solvent must you dissolve 1.88 moles of solute in to get a 1.41 $\frac{\text{mol}}{\text{L}}$ solution?

$$M = \frac{n}{V} \rightarrow 1.41 \frac{\text{mol}}{\text{L}} = \frac{1.88 \text{ mol}}{V} \rightarrow V = \frac{1.88 \text{ mol}}{1.41 \frac{\text{mol}}{\text{L}}} \rightarrow V = 1.33 \text{ L}$$

Part 3: Multistep mole calculations:

- 1) How many liters of volume does 8.21 g of nitrogen gas take up at STP?

$$8.21 \text{ g} \times \frac{1 \text{ mol } \text{N}_2}{28.0 \text{ g}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 6.56 \text{ L}$$

- 2) What is the mass of 1.74×10^{12} molecules of Rhenium (IV) oxide?



$$1.74 \times 10^{12} \text{ molecules} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times \frac{218.29}{1 \text{ mol}} = 6.30 \times 10^{-10} \text{ g}$$

- 3) 11.0 L of Hydrogen Chloride gas is at STP is bubbled through 500.0 mL of water. Assuming all of the Hydrogen Chloride is absorbed, what is the molarity of Hydrochloric acid that is made?

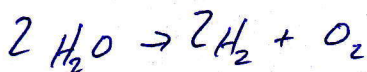
$$11.0 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.491071428 \text{ mol}$$

$$M = \frac{n}{V} \Rightarrow \frac{0.491071428 \text{ mol}}{0.5000 \text{ L}} = 0.982 \text{ M}$$

- 4) 350.0 mL of water (density = $1.00 \frac{\text{g}}{\text{mL}}$) is decomposed into its elements:

- a. How many grams of oxygen gas are created?

$$D = \frac{m}{V}$$



$$1.00 \frac{\text{g}}{\text{mL}} = \frac{m}{350.0 \text{ mL}}$$

$$350 \text{ g H}_2\text{O} \times \frac{1}{18} = 19.44444 \text{ mol H}_2\text{O}$$

$$350 \text{ g} = m$$

$$19.44444 \text{ mol H}_2\text{O} \times \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}} \times \frac{32.0 \text{ g}}{1 \text{ mol O}_2} = 311 \text{ g O}_2$$

- b. How many L of volume would the Hydrogen gas take up at STP?

$$19.44444 \text{ mol H}_2\text{O} \times \frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 436 \text{ L}$$

Remember the fives rounding rule

$$435.55 \rightarrow 435.6$$

round to even #