Stoichiometry

Answers and Explanations

1. B

The atomic weight of carbon is 12.0u, and the atomic weight hydrogen is 1.0u. Rounding to the whole number value, the molecular weight of C_2H_6 is 30g/mol.

$$MW = \frac{30\text{ g}}{\text{mol}}$$

A real key to stoichiometric problem solving is to see a molecular weight as a conversion factor. A conversion factor is a variation on the number 1. 30g C_2H_6 is the same thing as 1 mol. You can always turn a conversion factor upside down.

$$\frac{1\text{ mol}}{30\text{ g}}$$

Now it's just the simple matter of converting the mass of C_2H_6 given in the problem into moles.

$$\left(\frac{1\text{ mol}}{30\,\cancel{g}}\right) 0.6\,\cancel{g} = \left(\frac{6\times10^{-1}}{3\times10^{1}}\right)\text{mol}$$

$$= 2\times10^{-2}\text{ mol}$$

2. B

With the atomic weights of bromine (79.9u), carbon (12.0u), and hydrogen (1.0u), we have the molecular weights of Br_2 (160g/mol) and acetylene, which is C_2H_2 (26g/mol). We know the stoichiometric ratios of the reagents as they comprise the final compound. A big part of stoichiometrical problem solving is to think on the page through the conversion factors. Lay it out and convert grams Br_2 into moles Br_2 into moles $C_2H_2Br_4$ into moles C_2H_2 into grams C_2H_2. Use mental math wherever you can, ie.. 500 ÷ 160 ~ 3.

$$500\cancel{g\ Br_2}\left(\frac{1\,\cancel{mol\ Br_2}}{160\,\cancel{g\ Br_2}}\right)\left(\frac{1\,\cancel{mol\ C_2H_2Br_4}}{2\,\cancel{mol\ Br_2}}\right)\left(\frac{1\,\cancel{mol\ C_2H_2}}{1\,\cancel{mol\ C_2H_2Br_4}}\right)\left(\frac{26g\ C_2H_2}{1\,\cancel{mol\ C_2H_2}}\right)$$

$$= 41g\ C_2H_2$$

3. C

Think on the page with conversion factors, and because the answer choices are spaced well, feel free to use mental math (35.5 ÷ 72.6 ~ ½). Grams to moles and then with Avogadro's number from moles to atoms.

$$35.5\cancel{g\ Ge}\left(\frac{1\,\cancel{mol}}{72.6\,\cancel{g\ Ge}}\right)\left(\frac{6.02\times10^{23}\text{ atoms}}{\cancel{mol}}\right) = 2.94\times10^{23}\text{ atoms}$$

4. C

A chemical equation needs to be balanced to follow the law of conservation of mass. The equation is balanced when the moles of elements in the reactants side is equal to that of the products side. There is no real secret to balancing chemical equations. It's a process of trial and error.

$$6HCl + 2Cr \longrightarrow 2CrCl_3 + 3H_2$$

5. C

With the atomic weights of carbon (12.0u), chlorine (35.5u), and hydrogen (1.0u), for the *empirical formula* C_2ClH, we have a *formula weight* of 60.5u.

On the mass spec we see a cluster of peaks near 180u. The reason we don't see a single obvious molecular peak is that there are two stable isotopes of chlorine, ^{35}Cl (75.77%) and ^{37}Cl (24.23%). The relative abundance of these two is why chlorine has a standard atomic weight of 35.45.

Given the molecular peaks clustered just above 180u, the molecular weight of our compound must be 181.5 and the *molecular formula* of our compound must be $C_6Cl_3H_3$.

6. A

Empirical formula from percent composition is a classic Chem 101 problem. The first step is to assume you have 100g of the substance. Then you determine how many moles of each component that represents. Then for the subscripts you figure out the simplest whole number ratio among the moles of components.

$$39\text{g C}\left(\frac{1\text{ mol}}{12.0\text{ g C}}\right) = 3.3\text{ mol}$$

$$16\text{g H}\left(\frac{1\text{ mol}}{1.0\text{ g H}}\right) = 16.0\text{ mol}$$

$$45\text{g N}\left(\frac{1\text{ mol}}{14.0\text{ g N}}\right) = 3.2\text{ mol}$$

The simple whole number ratio is 1:5:1, so the correct empirical formula is CH_5N.

7. D

Heating the hydrate drives off the water. The molecular weight of magnesium sulfate ($MgSO_4$) is 120.4 g/mol. The molecular weight of water is 18.0 g/mol. To determine n in the expression $MgSO_4\text{–}nH_2O$ from the given information, we can construct and solve the following equation.

$$\frac{120.4}{120.4 + n\,18.0} = 0.49$$

$$n = 7$$

8. B

The masses of water and carbon dioxide produced in the combustion reaction correspond to two moles of each. From this we know that four moles of hydrogen and two moles of carbon were consumed in the reaction. Figuring out how many moles of oxygen were in the original compound is more difficult because we don't know how much reacted from the air. However, we do know the mass of the original sample. If we subtract out the weight of four moles of hydrogen and two of carbon, we are left with 16 grams of oxygen, or one mole.

9. A

From the reaction it can be seen that one mole of acetylene is produced for every two moles of water consumed. Remember that one ml of water equals one gram, so 18 ml would equal one mole. Therefore, 0.5 moles (13g) of acetylene is produced.

10. B

If n equals the number of carbons in an unbranched alkane, $2n + 2$ equals the number of hydrogens. Therefore, we can express the stoichiometry of the combustion reaction as follows:

$$C_nH_{2n+2} + (3n + 1)/2\ 0_2\,(g) \longrightarrow nCO_2\,(g) + (n + 1)H_2O\,(g)$$

One mole of ideal gas at STP occupies 22.4 liters, so 157 L gas equals seven moles of gas. For n moles of CO_2 plus $n + 1$ moles of H_2O to equal 7 moles, n must equal 3.

11. B

Think on the page with conversion factors and the problem will unfold for you and give you the atomic weight of the metal, Ru (101.07u). After setting out the conversion factors, you should be able to use mental math to see that the atomic weight will be very close to 100 g/mol.

$$\left(\frac{6.75\text{ g X}}{1.6\ \cancel{\text{g O}}}\right)\left(\frac{16\ \cancel{\text{g O}}}{\cancel{\text{mol O}}}\right)\left(\frac{3\ \cancel{\text{mol O}}}{2\text{ mol X}}\right) = \left(\frac{101\text{ g X}}{\text{mol X}}\right)$$

12. C

The first step is to write out the balanced reaction.

$$CO + 2H_2 \longrightarrow CH_3OH$$

The stoichiometry dictates that 1 mol CO reacts with 2 mol H_2 to form 1 mol CH_3OH, but the amounts of our reagents in the vessel aren't at these molar ratios.

$$14\text{ g CO}\left(\frac{1\text{ mol}}{28\text{ g}}\right) = 0.5\text{ mol CO}$$

$$10\text{g }H_2\left(\frac{1\text{ mol}}{2\text{ g}}\right) = 5\text{ mol }H_2$$

Instead of a 1:2 molar ratio between CO and H_2, what's actually present is a 1:10 ratio. We have H_2 in excess. This means that CO is the *limiting reagent*. It will dictate the amount of CH_3OH per the stoichiometry. If 0.5 mol CO are consumed, 0.5 mol CH_3OH will form, or 16 g (MW CH_3OH = 32 g/mol).