

## 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.

$$\left( \frac{1 \text{ mol}}{30 \text{ g}} \right) 0.6 \text{ g} = \left( \frac{6 \times 10^{-1}}{3 \times 10^1} \right) \text{ mol} \\ = 2 \times 10^{-2} \text{ mol}$$

#### 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..  $500 \div 160 \sim 3$ .

$$500 \text{ g } Br_2 \left( \frac{1 \text{ mol } Br_2}{160 \text{ g } Br_2} \right) \left( \frac{1 \text{ mol } C_2H_2Br_4}{2 \text{ mol } Br_2} \right) \left( \frac{1 \text{ mol } C_2H_2}{1 \text{ mol } C_2H_2Br_4} \right) \left( \frac{26 \text{ g } C_2H_2}{1 \text{ mol } C_2H_2} \right) \\ = 41 \text{ g } C_2H_2$$

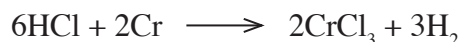
#### 3. C

Think on the page with conversion factors, and because the answer choices are spaced well, feel free to use mental math ( $35.5 \div 72.6 \sim 1/2$ ). Grams to moles and then with Avogadro's number from moles to atoms.

$$35.5 \text{ g } Ge \left( \frac{1 \text{ mol}}{72.6 \text{ g } Ge} \right) \left( \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \right) = 2.94 \times 10^{23} \text{ atoms}$$

#### 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.



#### 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,  $^{35}\text{Cl}$  (75.77%) and  $^{37}\text{Cl}$  (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  $\text{CH}_5\text{N}$ .

### 7. D

Heating the hydrate drives off the water. The molecular weight of magnesium sulfate ( $\text{MgSO}_4$ ) is 120.4 g/mol. The molecular weight of water is 18.0 g/mol. To determine  $n$  in the expression  $\text{MgSO}_4 \cdot n\text{H}_2\text{O}$  from the given information, we can construct and solve the following equation.

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

$$n = 7$$

### 8. B

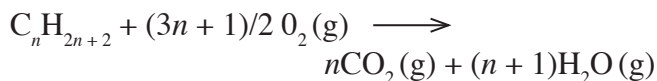
The masses of water and carbon dioxide produced in the combustion reaction correspond to two moles of each. From this we know that four moles of hydrogen and two moles of carbon were consumed in the reaction. Figuring out how many moles of oxygen were in the original compound is more difficult because 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:



One mole of ideal gas at STP occupies 22.4 liters, so 157 L gas equals seven moles of gas. For  $n$  moles of  $\text{CO}_2$  plus  $n + 1$  moles of  $\text{H}_2\text{O}$  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.

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

### 12. C

The first step is to write out the balanced reaction.



The stoichiometry dictates that 1 mol CO reacts with 2 mol  $\text{H}_2$  to form 1 mol  $\text{CH}_3\text{OH}$ , 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  $\text{H}_2$ , what's actually present is a 1:10 ratio. We have  $\text{H}_2$  in excess. This means that CO is the *limiting reagent*. It will dictate the amount of  $\text{CH}_3\text{OH}$  per the stoichiometry. If 0.5 mol CO are consumed, 0.5 mol  $\text{CH}_3\text{OH}$  will form, or 16 g (MW  $\text{CH}_3\text{OH} = 32 \text{ g/mol}$ ).