

AP Chem Unit 1 - Test Key

1. ANS: B

Expect easy math! No need to calculate since the abundance of the other two isotopes is just over 1%, so while the atomic mass isn't exactly 12 it is only slightly above 12. Ignoring the ridiculously low abundance of C-14 and doing the math yields an average atomic mass of 12.0107 amu.

OBJ: 1.2, SP 6.1 TOP: Stoichiometry

2. ANS: C

Expect easy math and estimate! Assume it's a 100 g sample so the percentages become grams. Remember, EF's are simply mole to mole ratios. Hf has molar mass of 178.5 call it 180. Cl has molar mass of 35.5 call it 36.

$$\frac{62.2 \text{ g}}{178.5 \frac{\text{g}}{\text{mol}}} \approx \frac{60 \text{ g}}{180 \frac{\text{g}}{\text{mol}}} \approx 3 \text{ moles Hf} \quad \text{and} \quad \frac{37.4 \text{ g}}{35.5 \frac{\text{g}}{\text{mol}}} \approx 1 \text{ mole Cl, so the empirical formula is HfCl}_3$$

DIF: Medium OBJ: 1.2; SP 2.2 TOP: Stoichiometry
MSC: 1999 #47 NOT: 50% answered correctly

3. ANS: B

Think empirical formula and empirical mass. The smallest value given is 38, yet the other values are not whole number multiples of 38. What to do? Divide 38 by 2, then 3, etc. until you find a value that can be multiplied by simple whole numbers to obtain the remaining values.

$$19 \div 2 = 38$$

$$19 \div 3 = 57$$

$$19 \div 4 = 76$$

$$19 \div 6 = 114$$

DIF: Medium OBJ: 1.1, SP 6.1 TOP: Stoichiometry
MSC: 1989 #24 NOT: 64% answered correctly

Given:

$$Q_1 \gamma = 38 \text{ g} \approx 40$$

$$Q_2 \gamma = 57 \text{ g} \approx 60$$

$$Q_3 \gamma = 76 \text{ g} \approx 80$$

$$Q_4 \gamma = 114 \text{ g} \approx 120$$

Soln: 38 doesn't divide into any of the masses.
Subscript (?) or 1st can't be 1

$$\therefore 38/2 = 19 \approx 20$$

$$40/20 = 2$$

$$60/20 = 3 \quad \therefore 19 \text{ g}$$

$$80/20 = 4$$

$$120/20 = 6$$

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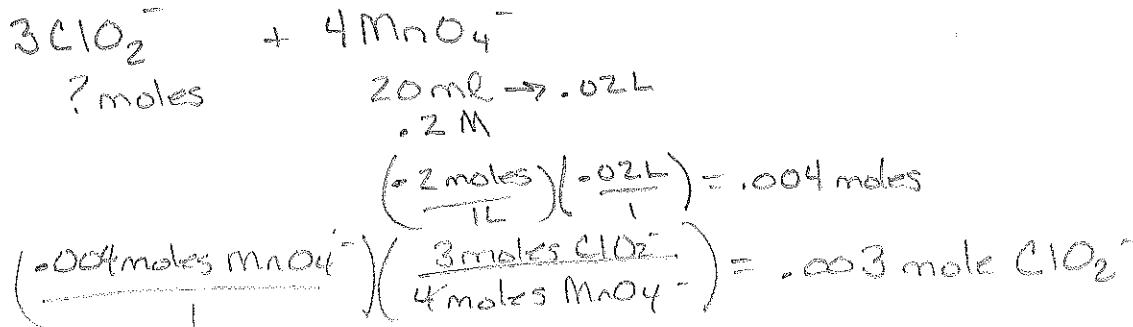
4.

ANS: A

Expect easy math! Calculate moles of potassium permanganate (purple). # moles = $(0.20 \text{ mol/L})(0.020 \text{ L}) = 0.004 \text{ mol KMnO}_4$. Look at the coefficients in the balanced equation. Divide 0.004 mol by 4, to get 0.001 mol, then multiply by 3 to get 0.003 mol ClO_2^- .

DIF: Medium
MSC: 2002 #60

OBJ: 1.4, SP 7.1 TOP: Stoichiometry
NOT: 53% answered correctly

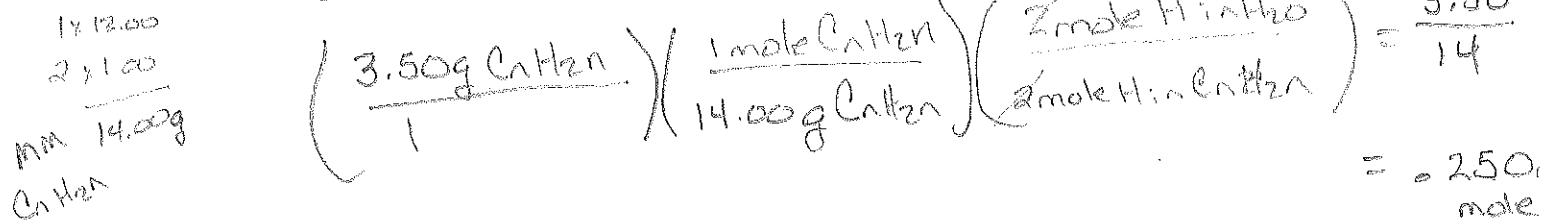
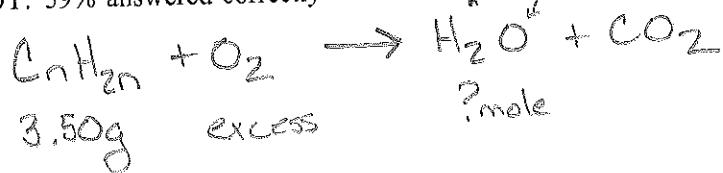


5.

ANS: A

Expect easy math!
If you think of C_nH_{2n} as a "unit", then the molar mass of the "unit" is $(12 + 2) = 14 \text{ g/mol}$. Look for a simple mathematical relationship: 3.50 g is ? of a mole and each unit has 2 H per mole, AND each water molecules requires 2 H as well, so you have ? of a mole of water formed.

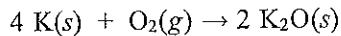
OBJ: 1.1, SP 6.1 TOP: Stoichiometry MSC: 1984 #45
NOT: 59% answered correctly



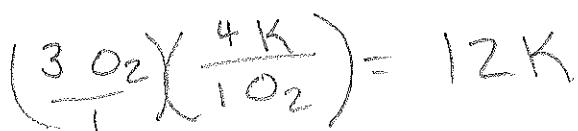
6.

ANS: C

The flask contains 3 molecules of oxygen as opposed to the one molecule given in the balanced equation. So, you need to triple the amount of K needed. There are 4 atoms of K in the balanced equation so if you triple that you'll need 3×4 or 12 K atoms as shown in answer c.



OBJ: 1.4, SP 7.1 TOP: Stoichiometry
ANS: B



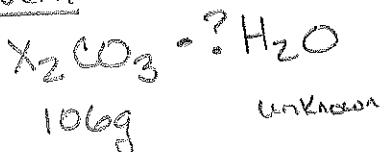
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ANS: A

7. Expect easy math! Estimate! 54% is close to 50% so, if the anhydrous carbonate's mass is 106, the water's mass is in that same neighborhood plus a bit. 180 g of water would be 10 waters of hydration, therefore 90 g would be 5 waters of hydration, so 106+ g of water falls around 7 waters of hydration.

DIF: Medium OBJ: 1.2, SP 6.1 TOP: Stoichiometry
MSC: 1989 #39 NOT: 51% answered correctly

Given:



MM of 1 H₂O is 18.00g

Soln:

lose 54% mass = 50% when Δ = original MASS is 212+g



so H₂O is 106+g in Hydrate

10 H₂O is 180g

$$\frac{-18}{162g} 9\text{H}_2\text{O}$$

$$\frac{-18}{144g} 8\text{H}_2\text{O}$$

Best

$$\frac{-18}{126g} 7\text{H}_2\text{O}$$

108g 6H₂O not an option

$$\frac{-18}{90g} 5\text{H}_2\text{O}$$

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8.

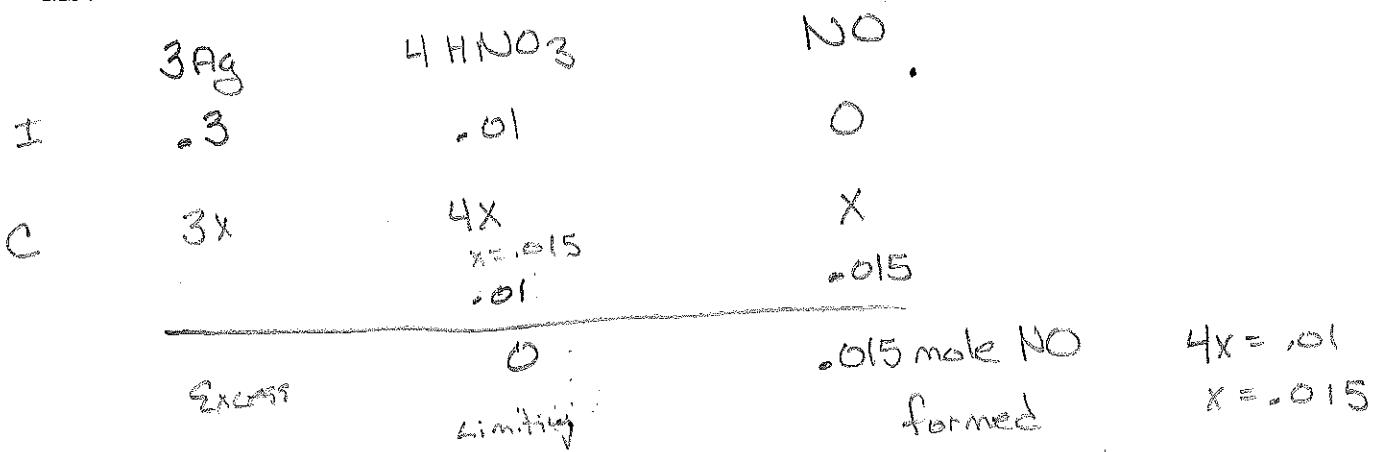
ANS: A

The "trick" to getting this one correct is to recognize that you have entered the "land of limiting reagent"! You were given the number of moles of silver, but must calculate the moles of nitric acid, it's two starting amounts either way! Remember that molarity \times liters = moles. Determine the limiting reagent and calculate subsequent moles from that limiting amount of moles using the mole:mole.

3 Ag	+ 4 HNO ₃	$\xrightarrow{3}$ AgNO ₃	+ NO	+ 2 H ₂ O
mole:mole 3	4	3	1	2
# moles 0.10 divide by 3 = 0.033	= (0.010 liter)(6.0 mol/L) = 0.060 mol divide by 4 = 0.015, compare to 0.033 LIMITING! work from this now...		If 4 = 0.060, what's "1" equal? 0.015 moles NO formed	

DIF: Medium
MSC: 1984 #52

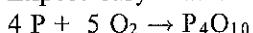
OBJ: 1.4, SP 7.1 TOP: Stoichiometry
NOT: 63% answered correctly



9.

ANS: D

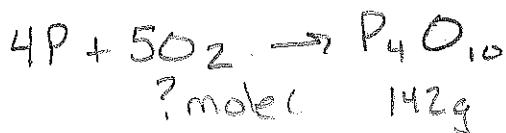
Expect easy math!



142 grams of P₄O₁₀ is ? a mole. So, you need ? of 5 moles of O₂ or 2.5 moles.

DIF: Easy
MSC: 1984 #44

OBJ: 1.4, SP 7.1 TOP: Stoichiometry
NOT: 67% answered correctly



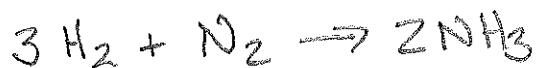
$$\frac{(142g P_4O_{10})}{(284g P_4O_{10})} \times \frac{1 \text{ mole } P_4O_{10}}{1 \text{ mole } P_4O_{10}} \times \frac{5 \text{ mole } O_2}{1 \text{ mole } P_4O_{10}} = \frac{5}{2} = 2.50$$

mole O₂

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10. **ANS: B**
 The balanced equation is $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$. We started with 2 molecules of nitrogen and excess hydrogen, so the equation yields $2\text{N}_2 + 6\text{H}_2 \rightarrow 4\text{NH}_3$. Therefore, we are looking for a product flask containing 4 ammonia molecules, so Flasks A,B & D qualify, but B is the only one accounting for the correct number of excess (unreacted) hydrogen molecules.

OBJ: 1.4, SP 7.1 TOP: Stoichiometry



8 moles 2 moles

I	8	2	0
R	-3x	-x	2x
	-6	-2	4
E	2	0	4 NH_3
	excess	unknown	

$$x = ?$$

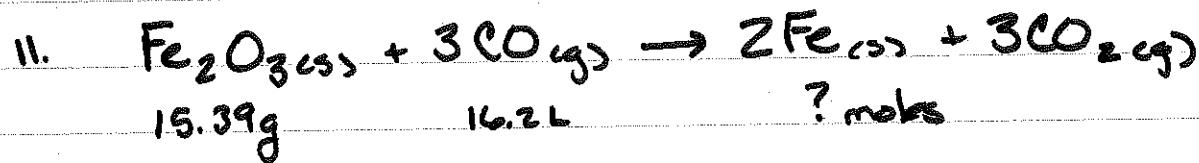
$$3x = 8$$

$$x = 2.67$$

so 2.00 H_2 left over

4 NH_3 produced

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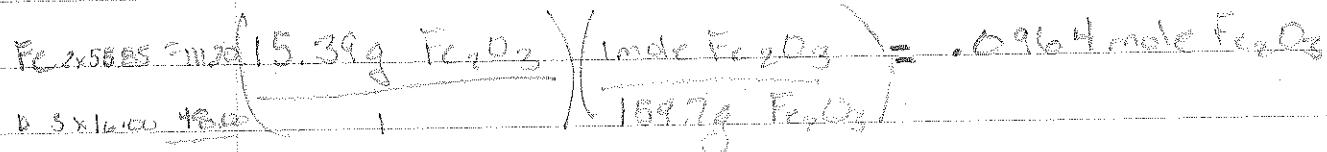
1.50 Atm
200°C = 473 K

? moles

$$PV = nRT$$

$$n = \frac{PV}{RT} = \frac{(1.50 \text{ atm})(16.2 \text{ L})}{(473 \text{ K})(0.08206 \frac{\text{L}}{\text{atm}\cdot\text{K}})} = 0.26 \text{ mole CO}$$

Available



(1) 183.2



A 0.964 mol 0.26 mol

C -x -3x 2x 0.264 mol

-0.0464 -0.1392 0.1428 0.264 mol

O	0.337 mol	192.8
Limiting	Excess	Produced

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12/ Tagatose $M_m = 180.15 \text{ g/mole}$
40.00% C 6.73% H 53.27% O

$$\text{C} \left(\frac{40.00\% \text{ C}}{12.00 \text{ g/C}} \right) \left(\frac{1 \text{ mole C}}{1 \text{ mole C}} \right) = \frac{3.333 \text{ mole C}}{3.329} \approx 1$$

$$\text{H} \left(\frac{6.73\% \text{ H}}{1.008 \text{ g/H}} \right) \left(\frac{1 \text{ mole H}}{1 \text{ mole H}} \right) = \frac{6.68 \text{ g H}}{3.329} \approx 2$$

$$\text{O} \left(\frac{53.27\% \text{ O}}{16.00 \text{ g/O}} \right) = \frac{3.329 \text{ g O}}{3.329} \approx 1$$

CH_2O empirical Formula

$$\text{CH}_2\text{O} \quad \text{C} = 1 \times 12.00 = 12.00$$

$$\text{H} = 2 \times 1.008 = 2.016$$

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

$$30.02 \text{ g/mole } \text{CH}_2\text{O}$$

$$\frac{180.15 \text{ g/mole Tagatose}}{30.02 \text{ g/mole CH}_2\text{O}} = 6$$

$\text{C}_6\text{H}_{12}\text{O}_6$ molecular formula