Practice MC Test unit D (Ch 10) Gas Laws

- ***This is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.***
- ***DO NOT USE A CALCULATOR. You may use ONLY the green periodic table.***
- ***Try to work at a pace of 1.5 min per question. Time yourself. It is important that you practice working for speed.***
- ***Then when time is up, continue working and finish as necessary.***

1. Two flexible containers for gases are at the same temperature and pressure. One holds 14 g of nitrogen and the other holds 22 g of carbon dioxide. Which of the following statements about these gas samples is true?
   a. The volume of the carbon dioxide container is the same as the volume of the nitrogen container.
   b. The number of molecules in the carbon dioxide container is greater than the number of molecules in the nitrogen container.
   c. The density of the carbon dioxide sample is the same as that of the nitrogen sample.
   d. The average kinetic energy of the carbon dioxide molecules is greater than the average kinetic energy of the nitrogen molecules.
   e. The average speed of the carbon dioxide molecules is greater than the average speed of the nitrogen molecules.

2. A sample of 0.010 mole of nitrogen dioxide gas is confined at 127ºC and 2.5 atmospheres. What would be the pressure of this sample at 27ºC and the same volume
   a. 0.033 atm
   b. 0.33 atm
   c. 0.53 atm
   d. 1.25 atm
   e. 1.88 atm

3. A hydrocarbon gas with the empirical formula CH_2 has a density of 1.3 g/L at 0ºC and 1.00 atm. A possible formula for the hydrocarbon is:
   a. CH_2
   b. C_2H_4
   c. C_3H_6
   d. C_4H_8
   e. C_5H_{10}

4. A real gas would act most ideal at
   a. 1 atm and 273 K
   b. 10 atm and 547 K
   c. 10 atm and 273 K
   d. 0.5 atm and 546 K
   e. 0.5 atm and 273 K

5. The reaction below takes place in a closed, rigid vessel. The initial pressure of $N_{2(g)}$ is 1.0 atm, and that of $O_{2(g)}$ is 1.5 atm. No $N_2O_{4(g)}$ is initially present. The experiment is carried out at a constant temperature. What is the total pressure in the container when the partial pressure of N_2O_4 reaches 0.75 atm?

   $$N_2 + 2O_2 \rightarrow N_2O_4$$

   a. 0.25 atm
   b. 0.75 atm
   c. 1.0 atm
   d. 2.0 atm
   e. 2.5 atm

6. A 0.33 mole sample of $CaCO_{3(s)}$ is placed in a 1 L evacuated flask, which is then sealed and heated. The $CaCO_{3(s)}$ decomposes completely according to the balanced equation below. The total pressure in the flask, measured at 300 K is closest to which of the following? (The gas constant, R = 0.082 L atm/mol K)

   $$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

   a. 2.0 atm
   b. 4.1 atm
   c. 8.1 atm
   d. 16 atm
   e. 18 atm

7. Equal numbers of moles of $CO_{2(g)}$, $N_{2(g)}$, and $NH_{3(g)}$ are placed in a sealed vessel at room temperature. If the vessel has a pinhole-size leak, which of the following will be true after some of the gas mixture has effused?
   a. All gases will effuse at the same rate since the temperature is held constant.
   b. The N_2 will effuse the fastest since it is the lightest.
   c. At any time during the effusion, there will always be an equal number of moles remaining in the vessel.
   d. The mole fraction of CO_2 in the sample will increase.

8. When a sample of carbon dioxide gas in a closed container of constant volume at 0.5 atm and 200 K is heated until its temperature reaches 400 K, its new pressure is closest to
   a. 0.25 atm
   b. 0.50 atm
   c. 1.0 atm
   d. 1.5 atm
   e. 2.0 atm

9. Liquid nitrogen has a boiling point of -196ºC. This corresponds to
   a. −469 K
   b. 77 K
   c. 153 K
   d. 469 K
   e. −196 K

10. 1.5 atm is the same pressure as
   a. 1140 mmHg
   b. 847 mmHg
   c. 507 mmHg
   d. 410 mmHg
   e. 182 mmHg

11. For an ideal gas, which pair of variables are inversely proportional to each other (if all other factors remain constant)?
   a. P, V
   b. P, T
   c. V, T
   d. n, P
   e. density, molar mass

12. The coldest possible temperature of any substance is
   a. 0ºC
   b. 273 K
   c. -273 K
   d. -273ºC
   e. In theory there is no limit as to how cold any theoretical gas can be.

13. A 7 Liter container will hold about 12 g of which of the following gases at 0ºC and 1 atm?
   a. F_2
   b. Cl_2
   c. Br_2
   d. N_2
   e. NO

14. The pressure of 4.0 L of an ideal gas in a flexible container is decreased to one-third of its original pressure and its absolute temperature is decreased by one-half. The volume then is
   a. 1.0 L
   b. 4.0 L
   c. 6.0 L
   d. 8.0 L
   e. 24 L

15. A given mass of gas in a rigid container is heated from 100 °C to 300 °C. Which of the following best describes what will happen to the pressure of the gas? The pressure will
   a. decrease by a factor of three
   b. increase by a factor of three.
   c. increase by a factor less than three.
   d. decrease by a factor greater than three.
   e. unable to be determined without information about volume.

16. A given mass of a gas occupies 5.00 L at 65 °C and 480 mmHg. What is the volume of the gas at 630 mmHg and 85 °C?
   a. $5.00 \times \frac{65}{85} \times \frac{480}{630}$
   b. $5.00 \times \frac{338}{358} \times \frac{480}{630}$
   c. $5.00 \times \frac{358}{338} \times \frac{480}{630}$
   d. $5.00 \times \frac{358}{338} \times \frac{630}{480}$
   e. $5.00 \times \frac{338}{358} \times \frac{630}{480}$

17. Consider the reaction below, in which 6 atm of ammonium nitrite is added to an evacuated flask with a catalyst and then heated.

$$NH_4NO_{2(g)} \rightarrow N_{2(g)} + 2\ H_2O_{(g)}$$

At equilibrium the total pressure is 14 atm. Calculate the partial pressure of the water vapor at equilibrium
   a. 2.0 atm
   b. 4.0 atm
   c. 6.0 atm
   d. 8.0 atm
   e. 10.0 atm

18. A 5.0 L sample of gas is collected at 400. mmHg at 727ºC. What is the volume if the temperature were cooled to 77ºC and the pressure increased to 700. mmHg?
    a. 2.50 L
    b. 250 ml
    c. 2.0 L
    d. 50. L
    e. 1.0 L

19. A sealed flask at 20ºC contains 1 molecule of carbon dioxide, CO_2 for every 3 atoms of helium, He. If the total pressure is 800 mmHg, the partial pressure of helium is
    a. 200 mmHg
    b. 300 mmHg
    c. 400 mmHg
    d. 600 mmHg
    e. 800 mmHg

20. Real gases vary from the ideal gas laws gases at conditions of
    a. high temperature and low pressure
    b. both high temperature and high pressure
    c. both low temperature and low pressure
    d. low temperature and high pressure
    e. both high density and low pressure

21. If 2.0 moles of gas in a sealed glass flask is heated from 25ºC to 50ºC. Select the conditions that are true.

	kinetic energy	pressure	number of moles
a.	increases	increases	stays the same
b.	stays the same	increases	stays the same
c.	decreases	increases	stays the same
d.	increases	increases	increases
e.	stays the same	increases	increases

22. A flask with a pressure of 2000 mmHg contains 6 mol of helium with a partial pressure of 1500 mmHg. The remaining gas is hydrogen, what is the mass of hydrogen in the flask?
    a. 0.50 g
    b. 2.0 g
    c. 4.0 g
    d. 8.0 g
    e. 16 g

23. The temperature of a sample of an ideal gas confined in a 2.0 L container was raised from 27ºC to 77ºC. If the initial pressure of the gas was 1,200 mmHg, what was the final pressure of the gas?
    a. 600 mmHg
    b. 1,400 mmHg
    c. 2,400 mmHg
    d. 2,100 mmHg
    e. 3,600 mmHg

24. A sealed container containing 8.0 grams of oxygen gas and 7.0 g of nitrogen gas is kept at a constant temperature and pressure. Which of the following is true?
    a. The volume occupied by oxygen is greater than the volume occupied by nitrogen.
    b. The volume occupied by oxygen is equal to the volume occupied by nitrogen.
    c. The volume occupied by nitrogen is greater than the volume occupied by oxygen.
    d. The density of nitrogen gas is greater than the density of oxygen.
    e. The average molecular speeds if the two gases are the same.

25. A gas sample contains 0.1 mole of oxygen and 0.4 mole of nitrogen. If the sample is at standard temperature and pressure, what is the partial pressure due to nitrogen?
    a. 0.1 atm
    b. 0.2 atm
    c. 0.5 atm
    d. 0.8 atm
    e. 1.0 atm

26. A mixture of gases contains 1.5 moles of oxygen, 3.0 moles of nitrogen, and 0.5 mole of water vapor. If the total pressure is 700 mmHg, what is the partial pressure of the nitrogen gas?
    a. 70 mmHg
    b. 210 mmHg
    c. 280 mmHg
    d. 350 mmHg
    e. 420 mmHg

27. A mixture of helium and neon gases has a total pressure of 1.2 atm. if the mixture contains twice as many moles of helium as neon, what is the partial pressure due to neon?
    a. 0.2 atm
    b. 0.3 atm
    c. 0.4 atm
    d. 0.8 atm
    e. 0.9 atm

28. Nitrogen gas was collected over water at 25ºC. If the vapor pressure of water at 25ºC is 23 mmHg, and the total pressure in the container is measured at 781 mmHg, what is the partial pressure of the nitrogen gas?
    a. 23 mmHg
    b. 46 mmHg
    c. 551 mmHg
    d. 735 mmHg
    e. 758 mmHg

29. When 4.0 moles of oxygen are confined in a 24-liter vessel at 176ºC, the pressure is 6.0 atm. If the oxygen is allowed to expand isothermally until it occupies 36 liters, what will be the new pressure?
    a. 2 atm
    b. 3 atm
    c. 4 atm
    d. 8 atm
    e. 9 atm

30. A gas sample is confined in a 5-liter container. Which of the following will occur if the temperature of the container is increased?
    I. The kinetic energy of the gas will increase.
    II. The pressure of the gas will increase
    III. The density of the gas will increase.
    a. I only
    b. II only
    c. I and II only
    d. I and III only
    e. I, II, and III

31. A 14.0 gram sample of an unknown gas occupies 11.2 liters at standard temperature and pressure. Which of the following could be the identity of the gas?
    I. N_2
    II. CO_2
    III. CO
    a. I only
    b. II only
    c. I and III only
    d. II and III only
    e. I, II, and III

32. Consider the combustion of 6.0 g of ethane. What volume of carbon dioxide will be formed at STP?
    a. 0.20 L
    b. 0.40 L
    c. 2.2 L
    d. 9.0 L
    e. 22.4 L

33. A gaseous mixture at a constant temperature contains O_2, CO_2, and He. Which of the following lists the three gases in order of increasing average molecular speeds?
    a. O_2, CO_2, He
    b. O_2, He , CO_2
    c. He, CO_2, O_2
    d. He, O_2, CO_2
    e. CO_2, O_2, He

34. Which of the following conditions would be most likely to cause the ideal gas laws to fail?
    I. High pressure
    II. High temperature
    III. Large volume
    a. I only
    b. II only
    c. I and II only
    d. I and III only
    e. II, and III only

35. An ideal gas contained in 5.0 liter chamber at a temperature of 37ºC. If the gas exerts a pressure of 2.0 atm on the walls of the chamber, which of the following expressions is equal to the number of moles of the gas?
    a. $\frac{(2.0)(5.0)}{(0.082)(37)}\,mol$
    b. $\frac{(2.0)(0.082)}{(5.0)(37)}\,mol$
    c. $\frac{(2.0)(0.082)}{(5.0)(310)}\,mol$
    d. $\frac{(2.0)(310)}{(0.082)(5.0)}\,mol$
    e. $\frac{(2.0)(5.0)}{(0.082)(310)}\,mol$

36. A sample of 18.0 g of aluminum metal is added to excess hydrochloric acid. The volume of hydrogen gas produced at 0.0ºC and 1 atm pressure is approximately
    a. 67 L
    b. 45 L
    c. 22 L
    d. 11 L
    e. 7 L

37. A gaseous mixture of oxygen and nitrogen is maintained at a constant temperature. Which of the following MUST be true regarding the two gases?
    a. Their average kinetic energies will be the same.
    b. Their average molecular speeds will be the same.
    c. Their partial pressures will be the same.
    d. Their total masses will be the same
    e. Their densities will be the same.

38. An ideal gas fills a ballon at a temperature of 27ºC and 1 atm pressure. By what factor will the volume of the balloon change if the gas in the balloon is heated to 127ºC at constant pressure?
    a. $\frac{27}{127}$
    b. $\frac{3}{4}$
    c. $\frac{4}{3}$
    d. $\frac{2}{1}$
    e. $\frac{127}{27}$

39. Which of the following assumption(s) is (are) valid based on kinetic molecular theory.
    I. Gas molecules have negligible volume.
    II. Gas molecules have no attractive forces on each other.
    III. The temperature of a gas is directly proportional to its kinetic energy.
    a. I only
    b. III only
    c. I and III only
    d. II and III only
    e. I, II, and III

40. A gas sample with a mass of 10 grams occupies 6.0 liters and exerts a pressure of 2.0 atm at a temperature of 26ºC. Which of the following expressions is equal to the molecular mass of the gas?
    a. $\frac{(10)(0.0821)(299)}{(2.0)(6.0)} g \bullet mol^{-1}$
    b. $\frac{(0.0821)(299)}{(10)(2.0)(6.0)} g \bullet mol^{-1}$
    c. $\frac{(299)(2.0)(6.0)}{(10)(0.0821)} g \bullet mol^{-1}$
    d. $\frac{(10)(2.0)(6.0)}{(299)(0.0821)} g \bullet mol^{-1}$
    e. $\frac{(2.0)(6.0)}{(10)(299)(0.0821)} g \bullet mol^{-1}$

41. Chlorine gas and fluorine gas will combine to form one gaseous product. One L of Cl_2 reacts with 3 L of F_2 to produce 2 L of product. assuming constant temperature and pressure condition, what is the formula of the product?
    a. ClF
    b. Cl_2F_2
    c. ClF_2
    d. Cl_2F
    e. ClF_3

42. A 1.1 L container will hold about 1.6 g of which of the following gases at 0ºC and 1 atm?
    a. SO_2
    b. O_2
    c. NH_3
    d. CH_4
    e. F_2

43. What is the total volume of products formed at STP when 1.2 g of carbon is burned?
    a. 0.1 L
    b. 2 L
    c. 12 L
    d. 22 L
    e. 44 L

1. a It is important to convert to moles to consider the situation presented. 14 g of N_2 is 0.5 mol and 22 g of CO_2 is 0.5 mol. You are meant to realize that two containers with the same number of moles at the same temperature and same pressure will have to be the same volume.

2. e This problem is a "before and after" problem, which means you need to break out the combined gas law. $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$.

   Reread the problem and realize that n and V remain constant, thus the problem reduces to $\frac{2.5}{400} = \frac{P_2}{300}$. With a bit of algebra, you should realize you are looking for ¾ of 2.5, half of 2.5 is 1.25, and it must be larger than this, thus the only option is (e).

3. b In this problem, we need a molar mass to determine the molecular formula. Notice that the conditions in the problem, 0ºC and 1 atm are STP conditions, and this allows us to use 22.4 L/mol. $\frac{1.3g}{1L} \times \frac{22.4L}{1mol}$, which means you have to do a roush estimate of this calculation. To make the math easy, consider 22.4 to be ~21, and ⅓ of 21 is 7, thus a MM of ~29. Of the options given, the only formula with a molar mass of ~29 is (b), C_2H_4 with a molar mass of 28.

4. d Gases follow the gas laws best (most mathematically correct) when they are not be pushed to turn in to a liquid. Gases will liquify at high pressure and low temperature, thus gases behave most ideally at lower pressure and higher temperatures.

5. c Set up a RICE box. When the temp and volume are constant, the pressure of a gas is proportional to its number of moles, so you can treat the pressure values, the same way you would treat mole values. Set up a rice box as shown, which allows you to calculate the pressure at the end using Dalton's Law of Partial Pressures; $P_{total} = P_1 + P_2 + P_3 + ...$ In the RCE box shown below, you can see the pressure of the nitrogen and oxygen given in the I row, and the problem tells us there is no product. The problem also tells us there is 0.75 of the product at the end. This lets us calculate how much oxygen was used up to produce the 0.75 atm of product: $0.75atm \times \frac{2O_2}{1N_2O_4} = 1.5atm$ for O_2 and how much nitrogen was used to produce the 0.75 atm of product: $0.75atm \times \frac{1N_2}{1N_2O_4} = 0.75atm$ for N_2. Put these in the change row and then do the subtraction to find the End row.

R	N_2 +	2 O_2 →	N_2O_4
I	1	1.5	0
C	-0.75	-1.5	+0.75
E	0.25	0	0.75

6. c The word "evacuated in the problem means emptied out. Thus there is no carbon dioxide at the start. The simple stoichiometry tells us 0.33 mol of calcium carbonate will form 0.33 mol of carbon dioxide. Then use the ideal gas law to calculate pressure. $P = \frac{nRT}{V}$ Substitute $P = \frac{(0.33mol)\left(\frac{0.082atm \bullet L}{mol \bullet K}\right)(300K)}{(1L)}$ This looks really bad to try and solve without a calculator. Sometimes you will have an easier time using fractions to find the common factors. Look what happens when you change 0.33 to ⅓, and I will leave off all the units. $P = \frac{1}{3}\frac{(0.082)(300)}{1}$

7. d As the gases effuse out of the vessel, the gases that have a slowest velocity will leak out more slowly, thus carbon dioxide with the largest molar mass will go out the slowest, as the other gases move out faster, and the carbon dioxide's moles proportional to the other gases will become higher than at the start. The other statements are flat out false; N_2 is not the lightest gas, NH_3 is the lightest. The gases will effuse at different rates because they all have different molar masses.

8. c The should appear to be a "before and after" problem, thus use the combined gas law without *n*, since the container is sealed, and without volume since you are told that there is a constant volume. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ Substitute $\frac{0.5}{200} = \frac{P_2}{400}$ and solve.

9. b K = ºC + 273.
10. a atm is changed to mmHg by multiplying by 760.
11. a This is Boyle's Law, part of the combined gas law.
12. d When you cool a substance down, the molecules slow down, and molecular motion would stop at absolute zero, 0 K or −273ºC. When this occurs, there is no way to slow the molecules further, thus there is no way to cool the temperature further, making absolute zero the coldest possible temperature.
13. a Don't miss the fact that this problem is also at STP, thus 7 L is ~⅓ of a liter. Calculate molar mass

$$MM = \frac{12g}{⅓mol} = 36g/mol$$

14. c Use the combined gas law with n removed because without any other information, you can assume the container is sealed.

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ Substitute with pressure and temp with values of 1 $\frac{(1)(4L)}{(1)} = \frac{\left(\frac{1}{3}\right)V_2}{\left(\frac{1}{2}\right)}$, though you may find it far less complicated by picking the original pressure as 3 and temperature as 2 $\frac{(3)(4L)}{(2)} = \frac{(1)V_2}{(1)}$ and then solve.

15. c Because of the combined gas law, $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, which you can rearrange, $\frac{T_2P_1}{T_1} = P_2$, thus you may be tempted to say the pressure would be triple, but you must change to Kelvin temperature $\frac{(573)P_1}{(373)}$ which does not triple the pressure.

16. c Rearrange the combined gas law and change to Kelvin, again assuming the container is sealed and moles remains constant. $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, thus $V_1\frac{T_2}{T_1}\frac{P_1}{P_2} = V_2$

17. d It's a good idea to use a RICE box to solve this problem. The volume is constant, and we can assume that at equilibrium (the end) that the temperature will return to the starting temperature, thus T is constant, which allows us to substitute with pressure values.

R	NH_4NO_2	$\rightarrow N_2$ +	$2\ H_2O$
I	6	0	0
C	− x	+ x	+ 2x
E	6 − x	+ x	+ 2x

We know that the pressure at the end must add up to 14. Thus (6 − x) + x + x = 14 and solve for x = 4 but the question asks for the pressure of water vapor, which is 2x.

18. e Use the combined gas law $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, remember to convert to Kelvin, and assume moles are constant.

$\frac{(400)(5)}{(1000)} = \frac{(V_2)(700)}{(350)}$ Rearrange and look for common factors $\frac{(350)(400)(5)}{(700)(1000)} = V_2$

19. d The helium must be ¾ of the total pressure.
20. d Remember that the gas laws "work best" and a gas behaves most ideally when their molecules are far apart and moving very fast. Gases vary from the ideal when their molecules begin to slow down enough, lower temps, and get pushed close together enough, high pressure, that they start to feel their intermolecular attractions for one another. This makes them "sticky" to each other and eventually they will turn to a liquid. Even before condensation can occur, the gas law proportions will not mathematically work out so well.
21. a When temperature increases, kinetic energy increases, which in turn causes pressure to increase since the flask has constant volume. The number of moles is constant because the flask is sealed.
22. c We know that each gas causes a portion of the total pressure. Thus we can make a proportion using moles and pressures. $\frac{n_1}{P_1} = \frac{n_{total}}{P_{total}}$ to solve for the total number of moles. $\frac{P_{total}n_1}{P_1} = n_{total}$ $\frac{(2000)(6)}{(1500)} = N_{total}$ Thus the total number of moles is 8, and ¼ of that is 2 moles of hydrogen with a molar mass of 2 g/mol.
23. b Using the combined gas law with moles and volume constant. $\frac{P_1}{T_1} = \frac{P_2}{T_2}$ Substitute $\frac{(350)(1200)}{(300)} = P_2$ and look for common factors.
24. b When give masses it is important to convert to moles so you can compare the quantities of gas molecules. 8 g of O_2 is 0.25 mol and 7 g of N_2 is also 0.25 mol, Thus at constant temperature and pressure, for equal moles of gas, the volume would remain constant.
25. d The mole fraction (symbolized N_1) of a gas causes that same fraction of the total pressure. $N_1P_{total} = P_1$ The mole fraction of the nitrogen is $\frac{(0.4)}{(0.5)}$ Thus $\frac{(4)}{(5)}1 = 8$
26. e The is the same problem as the previous problem, solve for the mole fraction of nitrogen, $\frac{(3)}{(5)}700$. You will likely find it easiest to divide the 5 into the 700 to get 140, then triple that to get 420.

27. c The information tells us that the neon is ⅓ of the total amount of gas, thus $\frac{(1)}{(3)}1.2 = 0.4$

28. e Just like our lab work, the pressure for any gas collected over water is caused by the gas and the water vapor pressure. A simple subtraction allows the calculation of the gas itself. 781 − 23 = 758.

29. c The term, *isothermally*, means occurs at the *same temperature*. Thus this is a combined gas law problem with equal moles and equal temperature. $P_1V_1 = P_2V_2$ Solve for P_2 and substitute and look for common factors. $\frac{P_1V_1}{V_2} = P_2$ $\frac{(6)(24)}{(36)} = P_2$

30. c When you increase the temperature of a sealed rigid container, the mass nor the volume will change, thus the density will not change. Kinetic energy will increase, which in turn increased the pressure.

31. c Take the easy route by realizing this problem is at STP, and you know that the volume of a mole of any gas at STP is 22.4 L/mol. Thus 11.2 L is a ½ mol of gas. $\frac{(14g)}{(0.5mol)} = 28g/mol$

32. d This is a stoichiometry problem that requires a balanced equation. Ethane is C_2H_6 thus $C_2H_6 + 7/2O_2 \rightarrow 2CO_2 + 3H_2O$ Convert the ethane to moles, but keep it as a fraction, ⅕ Thus $\frac{1}{5} \times \frac{2CO_2}{1C_2H_6} \times 22.4 =$ While this math may not look so smooth, consider the 22.4 to be ~20, to get an answer of ~8 and survey the answers to see that a value of 9 L compared to the rest of the choices, is close enough to select and move on.

33. e As you must know by now, molecular speed of gases is inversely proportional to their molar mass, thus list the gases from largest MM to smallest MM.

34. a Gas laws don't work out as well when the molecules are pushed too close together, causing them "notice" each other and "feel the stickiness."

35. e Use the ideal gas law and solve for n. $n = \frac{PV}{RT}$ Substitute in the values given in the problem.

36. c Again, you must write a balanced equation: $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$ Convert the 18 g of Al to moles $\frac{18g}{27g/mol} = \frac{2}{3}$

$\frac{2}{3}mol \times \frac{3H_2}{2Al} \times \frac{22.4L}{1mol} =\sim 22L$

37. a If substances are at the same temperature, they must have the same amount of kinetic energy.

38. c (The original answer sheet said the answer is b, it is actually c) This is a before and after problem. Use the combined gas law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ and be sure and convert to Kelvin. Solve $\frac{T_2V_1}{T_1} = V_2$ and substitute $\frac{400V_1}{300} = V_2$

39. e These are the three most important assumptions of the KMT.

40. a You can revise the ideal gas law $PV = nRT$ substituting $n = \frac{m}{MM}$, OR use the Molar Mass Kitty Kat! $MM = \frac{DRT}{P}$, but remember that $D = \frac{m}{V}$, so you can substitute it back in $MM = \frac{mRT}{VP}$. Then substitute in the data.

41. e Since the reaction is at constant temp and pressure, you can write an equation using the volumes as moles. Knowing the stoichiometry of the reactants and products will tell you what the formula must be. $Cl_2 + 3F_2 \rightarrow 2\ Cl_?F_?$ thus ClF_3

42. b Don't miss the fact that this problem is at STP. To figure out which gas, you need to calculate a molar mass. $\frac{1.6g}{1.1L} \times \frac{22.4L}{1mol}$. This may look "messy" but the 1.1 divide into the 22.4 ~20 times, and thus 1.6 × 20 is just over 30, thus oxygen.

43. b Write a balanced equation: $C + O_2 \rightarrow CO_2$ $1.2g \times \frac{1mol}{12g} \times \frac{1C}{1CO_2} \times \frac{22.4L}{1mol} =\sim 2L$