STUDY GUIDE
AP Chemistry

Chapter Ten- Gases

Homework

#1 Pg 432 #3, 15, 17, 21, 23, 24, 26, 29, 33, 37

#2 Pg 432 #5, 43, 45, 47, 51, 52, 58, 59, 63,

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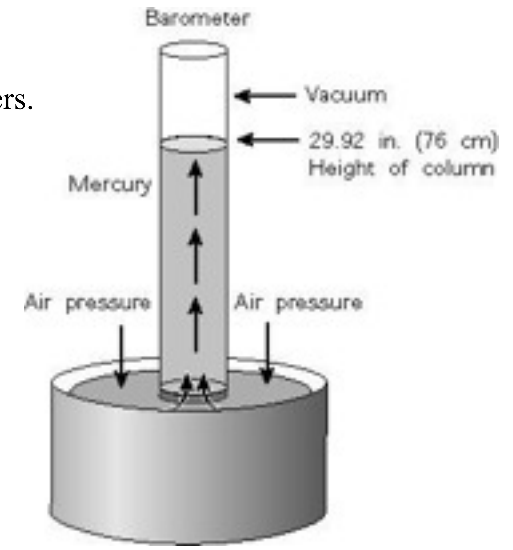
Lecture Notes

10.1 Characteristics of Gases

•All substances have three phases: solid, liquid and gas.
•Substances that are liquids or solids under ordinary conditions may also exist as gases.
  •These are often referred to as **vapors**.
•Many of the properties of gases differ from those of solids and liquids:
  •Gases are highly compressible and occupy the full volume of their containers.
  •When a gas is subjected to pressure, its volume decreases.
  •Gases always form homogeneous mixtures with other gases.
•Gases only occupy a small fraction of the total volume of their containers.
  •As a result, each molecule of gas behaves largely as though other molecules were absent.



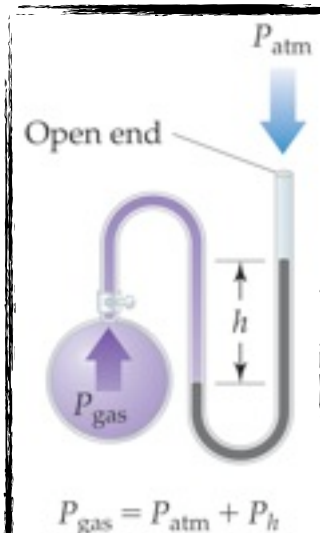
10.2 Pressure

•**Pressure** is the force acting on an object per unit area: $P = \frac{F}{A}$

Atmospheric Pressure and the Barometer

•The SI unit of force is the *newton* (N).
  •$1 \text{ N} = 1 \text{ kg·m/s}^2$
•The SI unit of pressure is the **pascal** (Pa).
  •$1 \text{ Pa} = 1 \text{ N/m}^2$
  •A related unit is the **bar,** which is equal to 10^5 Pa.
•Gravity exerts a force on the Earth's atmosphere.
  •A column of air 1 m^2 in cross section extending to the upper atmosphere exerts a force of 10^5 N.
  •Thus, the pressure of a 1 m^2 column of air extending to the upper atmosphere is 100 kPa.
    •Atmospheric pressure at sea level is about 100 kPa or 1 bar.
      •The actual atmospheric pressure at a specific location depends on the altitude and the weather conditions.
•Atmospheric pressure is measured with a *barometer*.
  •If a tube is completely filled with mercury and then inverted into a container of mercury open to the atmosphere, the mercury will rise 760 mm up the tube.
  •**Standard atmospheric pressure** is the pressure required to support 760 mm of Hg in a column.
  •Important non SI units used to express gas pressure include:
    •**atmospheres** (atm)
    • ***millimeters of mercury*** (mm Hg) or **torr**

1 atm = 760 mm Hg = 760 torr = 1.01325 x 10^5 Pa = 101.325 kPa



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10.3 The Gas Laws

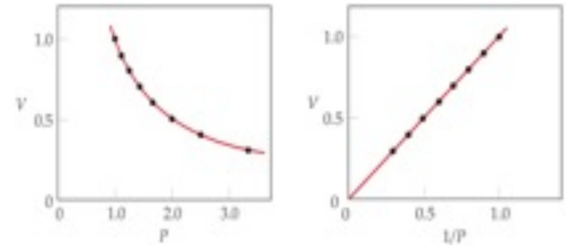
•The equations that express the relationships among *T* (temperature), *P* (pressure), *V* (volume), and *n* (number of moles of gas) are known as the *gas laws*.

The Pressure-Volume Relationship: Boyle's Law

•Weather balloons are used as a practical application of the relationship between pressure and volume of a gas.

•As the weather balloon ascends, the volume increases.

•As the weather balloon gets further from Earth's surface, the atmospheric pressure decreases.



•**Boyle's law**: The volume of a fixed quantity of gas, at constant temperature, is inversely proportional to its pressure.

•Mathematically:

•A plot of V versus P is a hyperbola.

•A plot of V versus $1/P$ must be a straight line passing through the origin.

•The working of the lungs illustrates that: $V = \text{constant} \times \frac{1}{P}$ or $PV = \text{constant}$

•as we breathe in, the diaphragm moves down & the ribs expand; therefore, the volume of the lungs increases.

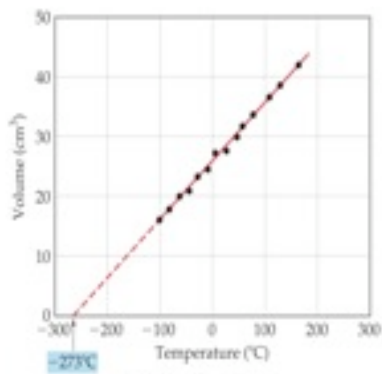
•according to Boyle's law, when the volume of the lungs increases, the pressure decreases; therefore, the pressure inside the lungs is less than the atmospheric pressure.

•atmospheric pressure forces air into the lungs until the pressure once again equals atmospheric pressure.

•as we breathe out, the diaphragm moves up and the ribs contract; therefore, the volume of the lungs decreases.

•By Boyle's law, the pressure increases and air is forced out.

The Temperature-Volume Relationship: Charles's Law



•We know that hot-air balloons expand when they are heated.

•**Charles's law:** The volume of a fixed quantity of gas at constant pressure is directly proportional to its absolute temperature.

•Mathematically: $V = \text{constant} \times T$ or $\frac{V}{T} = \text{constant}$

•Note that the value of the constant depends on the pressure and the number of moles of gas.

•A plot of V versus T is a straight line.

•When T is measured in °C, the intercept on the temperature axis is –273.15°C.

•We define *absolute zero*, 0 K = –273.15°C.

The Quantity-Volume Relationship: Avogadro's Law

•Gay-Lussac's *law of combining volumes*: At a given temperature and pressure the volumes of gases that react with one another are ratios of small whole numbers.

•**Avogadro's hypothesis:** Equal volumes of gases at the same temperature and pressure contain the same number of molecules.

•**Avogadro's law**: The volume of gas at a given temperature and pressure is directly proportional to the number of moles of gas.

•Mathematically: $V = \text{constant} \times n$

•We can show that 22.4 L of any gas at 0°C and 1 atmosphere contains 6.02×10^{23} gas molecules.

10.4 The Ideal-Gas Equation

•Summarizing the gas laws:

• Boyle: $V \propto 1/P$ (constant n, T) • Charles: $V \propto T$ (constant n, P)

• Avogadro: $V \propto n$ (constant P, T) • Combined: $V \propto nT/P$

•**Ideal gas equation**: $PV = nRT$

•An **ideal gas** is a hypothetical gas whose P, V, and T behavior is completely described by the ideal-gas equation.

•R = **gas constant** = 0.08206 L·atm/mol·K

•Other numerical values of R in various units are given in Table 10.2.

•Define **STP** (**standard temperature and pressure**) = 0°C, 273.15 K, 1 atm.
  •The molar volume of 1 mol of an ideal gas at STP is 22.41 L.

Relating the Ideal-Gas Equation and the Gas Laws

•If $PV = nRT$ and n and T are constant, then PV is constant and we have Boyle's law.
  •Other laws can be generated similarly.
•In general, if we have a gas under two sets of conditions, then $\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2}$
•We often have a situation in which P, V, and T all change for a fixed number of moles of gas.
  •For this set of circumstances, $\frac{PV}{T} = nR = \text{constant}$
    •Which gives $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

10.5 Further Applications of the Ideal-Gas Equation

Gas Densities and Molar Mass

•Density has units of mass over volume.

$$\frac{n}{V} = \frac{P}{RT} \qquad \frac{nM}{V} = \frac{PM}{RT} \qquad \therefore d = \frac{PM}{RT}$$

•Rearranging the ideal-gas equation with M as molar mass we get $M = \frac{dRT}{P}$
•The molar mass of a gas can be determined as follows:

Volumes of Gases in Chemical Reactions

•The ideal-gas equation relates P, V, and T to number of moles of gas.
•The n can then be used in stoichiometric calculations.

10.6 Gas Mixtures and Partial Pressures

•Since gas molecules are so far apart, we can assume that they behave independently.
•Dalton observed:
  •The total pressure of a mixture of gases equals the sum of the pressures that each would exert if present alone.
  •**Partial pressure** is the pressure exerted by a particular component of a gas mixture.
•**Dalton's law of partial pressures**: In a gas mixture the total pressure is given by the sum of partial pressures of each component:

$$P_t = P_1 + P_2 + P_3 + \ldots$$

•Each gas obeys the ideal gas equation.
  •Thus,

$$P_t = (n_1 + n_2 + n_3 + \cdots)\frac{RT}{V} = n_t\frac{RT}{V}$$

Partial Pressures and Mole Fractions

•Let n_1 be the number of moles of gas 1 exerting a partial pressure P_1, then $P_1 = X_1P_t$
  •where X_1 is the **mole fraction** (n_1/n_t).
  •Note that a mole fraction is a dimensionless number.

Collecting Gases over Water

•It is common to synthesize gases and collect them by displacing a volume of water.
•To calculate the amount of gas produced, we need to correct for the partial pressure of the water:

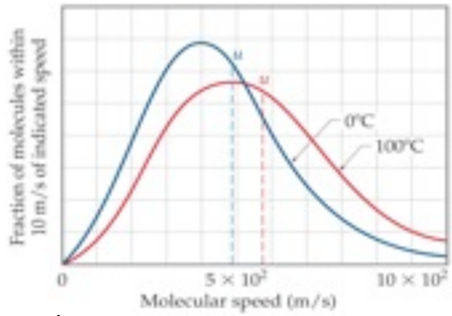
$$P_{total} = P_{gas} + P_{water}$$

•The vapor pressure of water varies with temperature.
  •Values can be found in Appendix B.

10.7 Kinetic-Molecular Theory

•The **kinetic-molecular theory** was developed to *explain* gas behavior.
  •It is a theory of moving molecules.

•Summary:
  •Gases consist of a large number of molecules in constant random motion.
  •The combined volume of all the molecules is negligible compared with the volume of the container.
  •Intermolecular forces (forces between gas molecules) are negligible.
    •Energy can be transferred between molecules during collisions, but the average kinetic energy is constant at constant temperature.
  •The collisions are perfectly elastic.
  •The average kinetic energy of the gas molecules is proportional to the absolute temperature.



•Kinetic molecular theory gives us an *understanding* of pressure and temperature on the molecular level.
  •The pressure of a gas results from the collisions with the walls of the container.
  •The magnitude of the pressure is determined by how often and how hard the molecules strike.

•The absolute temperature of a gas is a measure of the average kinetic energy.
  •Some molecules will have less kinetic energy or more kinetic energy than the average (distribution).
    •There is a spread of individual energies of gas molecules in any sample of gas.
    •As the temperature increases, the average kinetic energy of the gas molecules increases.

•As kinetic energy increases, the velocity of the gas molecules increases.
  •**Root-mean-square (rms) speed**, u, is the speed of a gas molecule having average kinetic energy.

•Average kinetic energy, ε, is related to rms speed:

$$\varepsilon = \tfrac{1}{2}mu^2$$

  •where m = mass of the molecule.

Application to the Gas-Laws

•We can understand empirical observations of gas properties within the framework of the kinetic-molecular theory.

•The effect of an increase in volume (at constant temperature) is as follows:
  •As volume increases at constant temperature, the average kinetic energy of the gas remains constant.
  •Therefore, u is constant.
  •However, volume increases, so the gas molecules have to travel further to hit the walls of the container.
  •Therefore, pressure decreases.

•The effect of an increase in temperature (at constant volume) is as follows:
  •If temperature increases at constant volume, the average kinetic energy of the gas molecules increases.
  • Therefore, u increases.
  •There are more collisions with the container walls.
  •The change in momentum in each collision increases (molecules strike harder).
  •Therefore, pressure increases. TWO REASONS-More collisions and harder collisions

10.8 Molecular Effusion and Diffusion

•The average kinetic energy of a gas is related to its mass:

$$KE = \tfrac{1}{2}mv^2$$

•Consider two gases at the same temperature: the lighter gas has a higher rms speed than the heavier gas.
  •Mathematically:

$$u = \sqrt{\frac{3RT}{M}}$$

  •The lower the molar mass, M, the higher the rms speed for that gas at a constant temperature.

•Two consequences of the dependence of molecular speeds on mass are:
  • **Effusion** is the escape of gas molecules through a tiny hole into an evacuated space (aka. A VACUUM).
  • **Diffusion** is the spread of one substance throughout a space or throughout a second substance.

Graham's Law of Effusion

•The rate of effusion can be quantified.

•Consider two gases with molar masses, M_1 and M_2, and with effusion rates, r_1 *and* r_2, respectively.

•The relative rate of effusion is given by **Graham's law**: $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

•Only those molecules which hit the small hole will escape through it.

•Therefore, the higher the rms speed the more likely it is that a gas molecule will hit the hole.

•We can show $\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{M_2}{M_1}}$

Diffusion and Mean Free Path

•Diffusion is faster for light gas molecules.

•Diffusion is significantly slower than the rms speed.

•Diffusion is slowed by collisions of gas molecules with one another.

•Consider someone opening a perfume bottle: It takes awhile to detect the odor, but the average speed of the molecules at 25°C is about 515 m/s (1150 mi/hr).

•The average distance traveled by a gas molecule between collisions is called the **mean free path**.

•At sea level, the mean free path for air molecules is about 6×10^{-6} cm.

10.9 Real Gases: Deviations from Ideal Behavior

Gases behave most ideally at LOW PRESSURE & HIGH TEMPERATURE

•From the ideal gas equation: $\frac{PV}{RT} = n$

•For 1 mol of an ideal gas, $PV/RT = 1$ for all pressures.

•In a real gas, PV/RT varies from 1 significantly.

•The higher the pressure the more the deviation from ideal behavior.

•For 1 mol of an ideal gas, $PV/RT = 1$ for all temperatures.

•As temperature increases, the gases behave more ideally.

•The assumptions in the kinetic-molecular theory show where ideal gas behavior breaks down:

•The molecules of a gas *have* finite volume. •Molecules of a gas *do* attract each other.

•As the pressure on a gas increases, the molecules are forced closer together.

•As the molecules get closer together, the free space in which the molecules can move gets smaller.

•The smaller the container, the more of the total space the gas molecules occupy.

•Therefore, the higher the pressure, the less the gas resembles an ideal gas.

•As the gas molecules get closer together, the intermolecular distances decrease.

•The smaller the distance between molecules, the more likely that attractive forces will develop between them.

•Therefore, the less the gas resembles an ideal gas.

•As temperature increases, the gas molecules move faster and further apart.

•Also, higher temperatures mean more energy is available to break intermolecular forces.

•As temperature increases, the negative departure from ideal-gas behavior disappears.

The Van der Waals Equation

•We add two terms to the ideal gas equation to correct for

•the volume of molecules: $(V - nb)$ •for molecular attractions: $\left(\frac{n^2 a}{V^2}\right)$

•The correction terms generate the **van der Waals equation**:

$$\left(P + \frac{n^2 a}{V^2}\right)(V - nb) = nRT$$

•where *a* and *b* are empirical constants that differ for each gas.

•Table 10.3 shows Van der Waals constants for common gases.

•To understand the effect of intermolecular forces on pressure, consider a molecule that is about to strike the wall of the container.

•The striking molecule is attracted by neighboring molecules.

•Therefore, the impact on the wall is lessened.

IDEAL GASES have
NO Intermolecular forces
NO size (So they don't exist!)

REAL GASES deviate from ideal behavior most when they have
STRONG intermolecular forces
LARGE size

Practice Problems

Review the properties of gases! How do they differ from liquids and solids?
Know the "ideal" conditions (under what conditions do gases behave ideally?

1) A sample of gas (24.2 g) initially at 4.00 atm was compressed from 8.00 L to 2.00 L at constant temperature. After the compression, the gas pressure was __________ atm.

A) 4.00 B) 2.00 C) 1.00
D) 8.00 E) 16.0

2) A balloon originally had a volume of 4.39 L at 44 °C and a pressure of 729 torr. The balloon must be cooled to _______°C to reduce its volume to 3.78 L (at constant pressure).

A) 38 B) 0 C) 72.9
D) 273 E) 546

3) A sample of H_2 gas (12.28 g) occupies 100.0 L at 400.0 K and 2.00 atm. A sample weighing 9.49 g occupies __________ L at 353 K and 2.00 atm.

A) 109 B) 68.2 C) 54.7
D) 147 E) 77.3

4) A sample of a gas originally at 25 °C and 1.00 atm pressure in a 2.5 L container expands until the pressure is 0.85 atm and the temperature is 15 °C. The final volume of the gas is ____ L.

A) 3.0 B) 2.8 C) 2.6
D) 2.1 E) 0.38

5) The reaction of 50 mL of Cl_2 gas with 50 mL of CH_4 gas via the equation: $Cl_2(g) + C_2H_4(g) \rightarrow C_2H_4Cl_2(g)$ will produce a total of __________ mL of products if pressure and temperature are kept constant.

A) 100 B) 50 C) 25
D) 125 E) 150

6) At a temperature of __________ °C, 0.444 mol of CO gas occupies 11.8 L at 889 torr.

A) 379 B) 73 C) 14
D) 32 E) 106

7) The volume of 0.25 mol of a gas at 72.7 kPa and 15 °C is __________ m^3.

A) 8.1×10^{-5} B) 1.2×10^{-4}
C) 4.3×10^{-4} D) 8.2×10^{-3}
E) 2.2×10^{-1}

8) 1.3 moles of gas will occupy ______ L at 22 °C and 2.5 atm.

A) 0.079 B) 0.94 C) 13
D) 31 E) 3.2×10^{-2}

9) The density of ammonia gas in a 4.32 L container at 837 torr and 45.0 °C is __________ g/L.

A) 3.86 B) 0.717 C) 0.432
D) 0.194 E) 4.22×10^{-2}

10) The molecular weight of a gas that has a density of 6.70 g/L at STP is __________ g/mol.

A) 496 B) 150 C) 73.0
D) 3.35 E) 0.298

11) Automobile air bags use the decomposition of sodium azide as their source of gas for rapid inflation:
$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$
What mass (g) of NaN_3 is required to provide 40.0 L of N_2 at 25.0 °C and 763 torr?

A) 1.64 B) 1.09 C) 160
D) 71.1 E) 107

12) A sample of He gas (3.0 L) at 5.6 atm and 25 °C was combined with 4.5 L of Ne gas at 3.6 atm and 25 °C at constant temperature in a 9.0 L flask. The total pressure in the flask was ____ atm. Assume the initial pressure in the flask was 0.00 atm.

A) 2.6 B) 9.2 C) 1.0
D) 3.7 E) 24

13) A flask contains a mixture of He and Ne at a total pressure of 2.6 atm. There are 2.0 mol of He and 5.0 mol of Ne in the flask. The partial pressure of He is __________ atm.

A) 9.1 B) 6.5 C) 1.04
D) 0.74 E) 1.86

14) SO_2 (5.00 g) and CO_2 (5.00 g) were placed in a 750.0 mL container at 50.0 °C. The total pressure was _______ atm.

A) 0.192 B) 4.02 C) 2.76
D) 6.78 E) 1.60

15) SO_2 (5.00 g) and CO_2 (5.00 g) were placed in a 750.0 mL container at 50.0 °C. The partial pressure of CO_2 in the container was __________ atm.

A) 6.78 B) 2.76 C) 1.60
D) 0.192 E) 4.02

16) A sample of N_2 gas (2.0 mmol) effused through a pinhole in 5.5 s. It will take __________ s for the same amount of CH_4 to effuse under the same conditions.

A) 7.3 B) 5.5 C) 3.1
D) 4.2 E) 9.6

17) A sample of He gas (2.0 mmol) effused through a pinhole in 53 s. The same amount of an unknown gas, under the same conditions, effused through the pinhole in 248 s. The molecular mass of the unknown gas is __________ g/mol.
A) 0.19 B) 5.5 C) 88
D) 19 E) 350

18) Gaseous mixtures __________.
A) can only contain molecules
B) are all heterogeneous
C) can only contain isolated atoms
D) are all homogeneous
E) must contain both isolated atoms and molecules

19) Which of the following equations shows an incorrect relationship between pressures given in terms of different units?
A) 1.20 atm = 122 kPa
B) 152 mmHg = 2.03×10^4 Pa
C) 0.760 atm = 578 mmHg
D) 1.0 torr = 2.00 mmHg
E) 1.00 atm = 760 torr

20) The pressure exerted by a column of liquid is equal to the product of the height of the column times the gravitational constant times the density of the liquid, $P = ghd$. How high a column of water (d = 1.0g/mL) would be supported by a pressure that supports a 713 mm column of mercury (d = 13.6g/mL)?
A) 14 mm B) 52 mm C) 713 mm
D) 1.2×10^4 mm E) 9.7×10^3 mm

21) Which statement about atmospheric pressure is false?
A) As air becomes thinner, its density decreases.
B) Air actually has weight.
C) With an increase in altitude, atmospheric pressure increases as well.
D) The warmer the air, the lower the atmospheric pressure.
E) Atmospheric pressure prevents water in lakes, rivers, and oceans from boiling away.

22) In ideal gas equation calculations, expressing pressure in Pascals (Pa), necessitates the use of the gas constant, R, equal to __________.
A) 0.08206 atm L $mol^{-1}K^{-1}$ B) 8.314 J $mol^{-1}K^{-1}$
C) 62.36 L torr $mol^{-1}K^{-1}$ D) 1.987 cal $mol^{-1}K^{-1}$
E) none of the above

23) Of the following, __________ is a correct statement of Boyle's law.
A) PV = constant B) $\frac{P}{V}$ = constant
C) $\frac{V}{P}$ = constant D) $\frac{V}{T}$ = constant
E) $\frac{n}{P}$ = constant

24) "Isothermal" means __________.
A) at constant pressure
B) at constant temperature
C) at variable temperature and pressure conditions
D) at ideal temperature and pressure conditions
E) that $\Delta H_{rxn} = 0$

25) The volume of an ideal gas is zero at __________.
A) 0 °C B) -45 °F C) -273 K
D) -363 K E) -273 °C

26) Standard temperature and pressure (STP), in the context of gases, refers to __________.
A) 298 K and 1 atm B) 273 K and 1 atm
C) 298 K and 1 torr D) 273 K and 1 pascal
E) 273 K and 1 torr

27) The volume of a sample of gas (2.49 g) was 752 mL at 1.98 atm and 62 °C. The gas is __________.
A) SO_2 B) SO_3 C) NH_3
D) NO_2 E) Ne

28) The average kinetic energy of the particles of a gas is directly proportional to __________.
A) the rms speed
B) the square of the rms speed
C) the square root of the rms speed
D) the square of the particle mass
E) the particle mass

29) The kinetic-molecular theory predicts that pressure rises as the temperature of a gas increases because _____.
A) the average kinetic energy of the gas molecules decreases
B) the gas molecules collide more frequently with the wall
C) the gas molecules collide less frequently with the wall
D) the gas molecules collide more energetically with the wall
E) both the gas molecules collide more frequently with the wall and the gas molecules collide more energetically with the wall

30) Which of the following is not part of the kinetic-molecular theory?
A) Atoms are neither created nor destroyed by ordinary chemical reactions.
B) Attractive and repulsive forces between gas molecules are negligible.
C) Gases consist of molecules in continuous, random motion.
D) Collisions between gas molecules do not result in the loss of energy.
E) The volume occupied by all of the gas molecules in a container is negligible compared to the volume of the container.

31) Of the following gases, __________ will have the greatest rate of effusion at a given temperature.
A) NH_3 B) CH_4 C) Ar
D) HBr E) HCl

32) Arrange the following gases in order of increasing average molecular speed at 25 °C.
He, O_2, CO_2, N_2

A) $He < N_2 < O_2 < CO_2$
B) $He < O_2 < N_2 < CO_2$
C) $CO_2 < O_2 < N_2 < He$
D) $CO_2 < N_2 < O_2 < He$
E) $CO_2 < He < N_2 < O_2$

33) A sample of oxygen gas (O_2) was found to effuse at a rate equal to three times that of an unknown gas. The molecular weight of the unknown gas is __________ g/mol.
A) 288 B) 96 C) 55
D) 4 E) 10.7

34) A real gas will behave most like an ideal gas under conditions of __________.
A) high temperature and high pressure
B) high temperature and low pressure
C) low temperature and high pressure
D) low temperature and low pressure
E) STP

35) Which one of the following gases would deviate the least from ideal gas behavior?
A) Ne B) CH_3Cl C) Kr
D) CO_2 E) F_2

36) When gases are treated as real, via use of the van der Waals equation, the actual volume occupied by gas molecules __________ the pressure exerted and the attractive forces between gas molecules __________ the pressure exerted, as compared to an ideal gas.
A) decreases, increases
B) increases, increases
C) increases, decreases
D) does not affect, decreases
E) does not affect, increases

37) A fixed amount of gas at 25.0 °C occupies a volume of 10.0 L when the pressure is 667 torr. Use Boyle's law to calculate the pressure (torr) when the volume is reduced to 7.88 L at a constant temperature of 25.0 °C.
A) 846 B) 0.118 C) 5.26×10^4
D) 526 E) 1.11

38) A fixed amount of gas at 25.0 °C occupies a volume of 10.0 L when the pressure is 629 torr. Use Charles's law to calculate the volume (L) the gas will occupy when the temperature is increased to 121 °C while maintaining the pressure at 629 torr.
A) 10.9 B) 13.2 C) 2.07
D) 7.56 E) 48.4

39) The density of nitric oxide (NO) gas at 1.21 atm and 54.1 °C is __________ g/L.
A) 0.0451 B) 0.740 C) 1.35
D) 0.273 E) 8.2

40) A 1.44-g sample of an unknown pure gas occupies a volume of 0.335 L at a pressure of 1.00 atm and a temperature of 100.0 °C. The unknown gas is _______.
A) argon B) helium C) krypton
D) neon E) xenon

41) The total pressure exerted by a mixture of 1.50 g of H_2 and 5.00 g of N_2 in a 5.00-L vessel at 298 K is _ atm.
A) 1.06 B) 9.08 C) 4.54
D) 32.4 E) 5.27

42) Zinc reacts with aqueous sulfuric acid to form hydrogen gas: $Zn\ (s) + H_2SO_4\ (aq) \rightarrow ZnSO_4\ (aq) + H_2\ (g)$
In an experiment, 225 mL of wet H_2 is collected over water at 27 °C and a barometric pressure of 748 torr. How many grams of Zn have been consumed? The vapor pressure of water at 27 °C is 26.74 torr.
A) 4.79×10^6 B) 0.567 C) 567
D) 431 E) 4.31×10^5

43) Given the equation
$$C_2H_6\ (g) + O_2\ (g) \rightarrow CO_2\ (g) + H_2O\ (g)$$
Determine the number of liters of CO_2 formed at STP when 240.0 grams of C_2H_6 is burned in excess oxygen gas.

1. Answer: E
2. Answer: B
3. Answer: B
4. Answer: B
5. Answer: B
6. Answer: E
7. Answer: D
8. Answer: C
9. Answer: B
10. Answer: B
11. Answer: D
12. Answer: D
13. Answer: D
14. Answer: D
15. Answer: E
16. Answer: D
17. Answer: C
18. Answer: D
19. Answer: D
20. Answer: E
21. Answer: C
22. Answer: B
23. Answer: A
24. Answer: B
25. Answer: E
26. Answer: B
27. Answer: D
28. Answer: B
29. Answer: E
30. Answer: A
31. Answer: B
32. Answer: C
33. Answer: A
34. Answer: B
35. Answer: A
36. Answer: C
37. Answer: A
38. Answer: B
39. Answer: C
40. Answer: E
41. Answer: C
42. Answer: B
43. Answer: 358