**Chemistry 20: Gases**

A. **Properties of Gases**

***Chemical Properties*** (reactivity)

1. Halogen gases : **Very Reactive**
2. Nobel gases :\_**Non reactive - inert**
3. Other gases : \_**reactivity varies – ie) oxygen reacts with metals (rust)**

***Physical Properties***

1. Gases do not have a fixed volume or shape. Gases fill their containers
2. Gases are highly compressible
3. Gases diffuse
4. Three variables affect gases :
5. Pressure: Force over a fixed area measure by a barometer

The SI unit for pressure is kPa (kilo Pascal)

**IMPORTANT**: At sea level, P=101.325kPa = 1atm = 760mmHg

1. Volume: occupied space measured by displacement in a graduated cylinder

SI unit for volume is L (litres)

1. Temperature: kinetic energy measured by a thermometer.

SI unit for temperature is oC (degrees Celsius) or K (Kelvin)

B. **Behavior of Gases**

1. ***Effect of Adding or Removing Gas*** (Changing **PRESSURE**)

ADDING GAS:

* increases the pressure
* increases the number of collisions & temperature
* if you double the amount of gas, double the pressure (direct)
* example: pumping a tire

REMOVING GAS:

* decreases the pressure, collisions & temperature
* remove half the gas - results in half the pressure

2. ***Effect of Changing the size of the container*** (Changing **VOLUME**)

SMALLER CONTAINER:

* decreases the volume
* increases the pressure & temperature
* a container that is compressed to half its size will double the pressure (indirect)

LARGER CONTAINER:

* increases the volume & decreases the pressure & temperature
* a container that is double in size will half the pressure

1. ***Effect of Heating or*** ***Cooling the Gas*** (Changing **TEMPERATURE**)

HEATING: \* increases the temperature

* increases the number of collisions & pressure
* Double the amount of heat will double the pressure

COOLING: \*decreases the temperature, collisions & pressure

* Decrease the amount of heat by half will half the pressure

C. **Dalton’s Law of Partial Pressure**

***Introduction***:

\* most gases are **mixture** (**homogeneous**)

**Air at Sea Level percent mmHg atmospheres**

Nitrogen 78% 593.4 0.7 81

Oxygen 21% 159.2 0.209

Carbon dioxide 0.04% 0.3 0.001

Others 0.96% 7.1 0.009

TOTAL 100% 760.0 1.000

\* molecules of gases at the same temperature will have **same** kinetic energy

\* pressure depends on the **number** of particles, not the **type** of particles

\* if you increase your altitude then the pressure of the air **decreases**

\* each gas exerts its own **partial pressure**

***Dalton’s Law***: At constant **volume**  & **pressure**, the total **pressure** exerted by a mixture of gases is equal to the sum of the partial **pressures**.

* ***Formula***: Ptotal = P1 + P2 + P3 NOTE: Significant Digits: least number of ***decimals*** from question

***Example***:

1. Determine the pressure of a gas mixture if the pressure of the oxygen is 150 mmHg, the nitrogen is 350 mmHg and the helium is 200 mmHg.

**P1 + P2 + P3= Ptotal**

**150 + 350 + 200 = 700 mmHg**

2. What is the pressure of oxygen in a gas mixture if the total pressure is 2.5 atm and the pressure of hydrogen is 1.7 atm?

**P1= Ptotal - P2**

**2.5atm - 1.7atm = 0.8 atm**

**(1 dec. for sig digs)**

3. Why do planes have to pressurize the cabin?

**The air at high altitudes has less pressure, therefore it has lower O2 amount and it is harder to breath**

**D. Boyle’s Law of Pressure-Volume Changes**

***Introduction***:

\* STP: **Standard temperature (OC), pressure (101.325 kPa), molar volume =22.4L/mol**

\* SATP: **Standard ambient temp (25C), pressure (100kPa), molar volume =24.8L/mol**

\* 1kPa = \_**1000N/m2**

\* Gas Laws are important for a Scuba Diver.

- When a person dives into water, the deeper they go the **greater** the pressure. The scuba equipment provides air to the lungs at a pressure that matches the pressure of the water.

- If a diver ascends to the surface without exhaling steadily, the air in the lungs will **expand**  as the pressure drops and the lungs will **rupture**.

- gases under pressure are **poisonous** so the length & depth of dive are important

- If the diver ascends to quickly, he can get the **BENDS or decompression sickness**. What happens is the gases are less soluble when the pressure drops and the gases form **bubbles**.

\* Robert Boyle’s (1627-1691) 1662 experiment:

**Problem**: What is the mathematical relationship between the pressure & volume of a gas?

**Design**: Pressure is exerted on a gas and the change in volume is measured.

**Manipulated variable**: **Pressure** **Responding variable**: **Volume**

**Data**: **Pressure (kPa)** **Volume (L)** **Graph**:

100 5.00

110 4.55 Vol (L)

120 4.16

130 3.85

**Analysis**:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Pressure (kPa)

***Boyle’s Law*for a given mass at a constant temp, the volume of a gas varies inversely with pressure (indirect relationship)**

***Formula***:\_**P1V1 = k; P2V2 = k**

Comparing two sets of measurements on the same gas: **P1V1 = P2V2**

***Example***:

1. The pressure on 2.50 L **(V1)** of anesthetic gas is changed from 760 mmHg to 304 mmHg. What will the new volume be, if the temperature remains constant?

P1V1=P2V2

V2=760mmHg x 2.50L/304mmHg; V2= **6.25 L**

2. A balloon is filled with 30.0 L of helium at 1 atm. What is the volume when the balloon rises to an altitude where the pressure is 0.250 atm.

P1V1 = P2V2

V2=30.0L x 1atm/0.250atm

V2= 120 L = **1 x 102 L** (1 sig dig)

E. **Charles’ Law for Temperature-Volume Changes**

***Introduction***:

\* Absolute zero: **Lowest temperature theoretically possible (-273.15 C or O K) (particles stop moving at this temperature.)**

\* Kelvin temperature: **a temperature scale that starts zero being absolute zero. K = C + 273.15**

100 C = **373.15 K (100 + 273.15)**

25 C = **298.15 K (25 + 273.15)**

0 C = **273.15 K (0 + 273.15)**

-273.15 C = **0 K (absolute zero) (-273.15 – 273.15)**

\*Jacques Charles’(1746-1823) 1787 experiment:

**Problem**: What is the mathematical relationship between the temperature & volume of a gas?

**Design**: Temperature is varied and the change in volume of the gas is measured.

**Manipulated variable**: **Temperature** **Responding variable**: **Volume**

**Data**: **Temperature (C)** **Volume (L)** **Graph**:

25 ( **298.15** K) 5.00

50 (**323.15** K) 5.42 **Volume (L)**

75 ( **348.15** K) 5.84 deviates

100 ( **373.15**K) 5.26

**Analysis**:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

**Temperature (K)**

***Charles’ Law***: **for a given mass of gas at a constant pressure, the volume of the gas varies directly with temperature (MUST be in KELVIN)**

***Formula***: **\_ V1 /T1 = k V2/T2 = k**

Comparing two sets of measurements on the same gas: **V1/T1 = V2/T2 or V1T2 = V2T1**

***Example***:

1) A balloon is inflated in an air-conditioned room at 27 C, and has a volume of 4.0 L. If it is heated to 57 C, what is the new volume of the balloon if the pressure is constant. (Must use Kelvin.)**T1=273.15 + 27 = 300.15K; T2=273.15 + 57= 330.15K**

**V1/T1=V2/T2; V2= V1T2/T1**

**V2= 4.0L x 330.15K / 300.15K**

**V = 4.399…L**

**V = 4.4 L**

2) If a sample of gas occupies 6.9 L at 327 C, what temperature is required to reduce the gas to 3.4 L? (Convert 327 to K and then convert answer back to C.)

**T1=273.15 + 327 = 600.15K**

**V1/T1 = V2/T2 ; T2= V2T1/V1**

**T2=3.4L x 600.15K / 6.9L**

**T2=295.72 K (-273.15)**

**T2=23 C**

F. **Gay-Lussac’s Law for Temperature-Pressure Changes**

***Introduction***:

\* On a hot summer day the pressure in a car tire increases. Why?

\* In 1802 Joseph Gay-Lussac (1778-1850) explained this relationship

***Gay Lussac’s Law***: **for a given mass of gas at a constant volume, the pressure of the gas varies directly with temperature (in Kelvin).**

***Formula***: **P1 /T1 = k P2/T2 = k (direct relationship)**

Comparing two sets of measurements on the same gas: **P1/T1 = P2/T2 or P1T2 = P2T1**

***Graph***:  **Temperature vs Pressure**

**710**

**deviates at high temp or pressure**

**Pressure**

**(kPa)**

**0 298 373**

**Temperature (K)**

***Example***:

1) A gas in an aerosol can at a pressure of 1 atm at 27 C. The can is thrown in the fire. What is the pressure if the temperature reaches 927 C?

**T1=273.15 + 27 = 300.15 K**

**T2=273.15 + 927 = 1200.15 K**

**P1/T1 = P2/T2 ; P2= T2P1/T1**

**P2=1200.15K x 1atm**

**300.15K**

**P2=3.998… atm**

**P2=4 atm (1 sig dig)**

2) A gas has a pressure of 50.0 mmHg at 540 K. What is the temperature when the pressure is 18.5 mmHg?

**P1/T1 = P2/T2 ; T2= P2T1/P1**

**T2=18.5mmHg x 540K**

**50.0mmHg**

**T2=199.8… K**

**T2=200 K (3 sig digs**)

G. **Combined Gas Law**

***Introduction***:

\* Boyle’s, Charles’ and Gay-Lussac’s laws can be combined

***Combined Gas Law***: **for a given gas, the product of the pressure & the volume is directly proportional to the temperature (in Kelvin)**

* ***Formula***: **P1 V1)/T1=k (P2 V2)/T2= k**

Comparing two sets of measurements on the same gas: **(P1 V1)/T1 = (P2 V2)/T2**

**or P1 V1T2 = P2 V2T1**

***Example***:

1) A balloon containing hydrogen gas at 20 C and a pressure of 100 kPa has a volume of 7.50 L. Calculate the volume of the balloon after it rises 10 km where the temperature is -36 C and the pressure is 28 kPa. (Assume that no hydrogen gas escapes.

**Step 1) Organize data: T1 = 273.15 + 20 = 293.15 K**

**P1 = 100 kPa; V1 = 7.50 L**

**T2 = 273.15 + -36 = 237.15 K**

**P2 = 28 kPa; V2 = ?**

**Step 2) Chose formula & plug in data: V1P1/T1=V2P2/T2 or V1P1T2=V2P2T1**

**V2=7.50L x 100kPa x 237.15K**

**(28kPa x 293.15K)**

**Step 3) Solve: V2 = 21.669L; V2 = 22 L (2 sig digs)**

2) A cylinder of compressed oxygen has a volume of 30 L and a pressure of 100 atm at 27 C. The cylinder is cooled until the pressure is 5.0 atm. What is the new temperature in C?

**STEP 1) T1 = 273.15 + 27 = 300.15 K P1 = 100 atm; V1 = 30 L**

**T2 = ?; P2 = 5.0 atm; V2 = 30L**

**STEP 2) V1P1/T1 = V2P2/T2  or P1/T1 = P2/T2**

**100 atm = 5.0 atm (cross multiply)**

**300.15 K T2**

**STEP 3) T2 = 15.008 K (can leave out volume because they remain the same)**

**T2 = 15 K (-258.15C or -2.6x102 C)**

**(Temperature dropped – This is actually Gay’s Lusac’s Law)**

G. **Ideal Gas Law**

***Introduction***:

\* Ideal gas: a hypothetical gas that \_**obeys**\_ all the gas laws **Perfectly** under **all conditions**

- does not condense into a **liquid** when cooled

- graphs have **perfectly straight** lines (**don’t deviate or curve)**

- composed of particles of **no** size that do not **attract each other.**

- real gases deviate from ideal gases at **low** temperatures & **high** pressure

\* One can calculate the moles of a gas at **STP**. (n=v/V where V is 22.4 L/mol)

- The molar volume of an ideal gas is 22.414 L/mol. The molar volume of helium is 22.426 L/mol. The molar volume of chlorine gas is 22.063 L/mol

- Therefore the **smaller** the particle the closer the gas resembles an ideal gas

\* Avogadro stated that equal volumes of gases at the same **pressure**  and **temperature**\_ contain equal numbers of molecules (moles)

\* Summary

- Boyle: volume of a gas is inversely proportional to pressure

-Charles: volume of a gas in directly proportional to Kelvin temperature

- Avogadro: volume of a gas is directly proportional to the number of moles

- COMBINATION: v n T 1/P

- Ideal gas law constant(R): At STP R = PV/nT = (101.3kPa x 22.4L)/(1mol x 273K)

R = ***8.314*** L x kPa or **\_62.4** L x mmHg or **0.0821**  L x atm

K x mol K x mol K x mol

***Ideal Gas Law***: the product of the pressure and volume is directly proportional to the amount and absolute temperature of the gas.

***Formula***: **\_: PV = nRT**

***Example***:1) What mass of neon should be introduced into an evacuated 0.88L tube to produce a pressure of 90 kPa at 30 C?

**Step 1) V = 1.00L; P = 200 kPa**

**n = 1.00 mol R = 8.314; T = PV/nR**

**Step 2) T = (200kPa x 1.00L)**

**(1.00mol x 8.314kPaL/Kmol)**

**T = 24.05… K T = 24.05… K – 273.15 = -248 C**

2) A cylinder containing 20.0 L of nitrogen gas is pressurized to 200 atm at 27 C. How many moles of nitrogen gas are present?

**1) n=PV/RT \*T=315.15K**

**2) n= (15.0atm x2.24E6L)**

**(0.0821atmL/Kmol x 315.15K)**

**n=1.2986 x 106 mol**

**3) m=nM; m=1.2…mol x 16.05g/mo m =2.1 x 107 g of methane**

3) A cylinder contains 2.24 x 106 L of methane gas at a pressure of 15.0 atm and a temperature of 42 C. How many grams of methane are in the container?

**1) V=nRT/P**

**V=(2.24molx8.314kPaL/Kmolx273.15K)**

**101.325kPa**

**2) V=50.2L (at STP) CHECK:v=nV v=2.24x22.4=50.2L**

H. Gas Stoichiometry

Steps of Stoichiometry

1. **Balanced Chemical equation; ID the Required and Given(s)**

2a. **Convert the first given to moles (n=PV/RT is the new formula)**

2b. **Convert the second given to moles**

3. **Mulitply by mole ratio (R/G) (Do it twice to find the limiting)**

4. **Convert the required from moles (P =nRT/V; V=nRT/P; T=PV/nR)**

5. **Find %yield (AY/TY x 100)**

6. **Find % error (TY-AY/TY x 100)**

Example:

1) 2.00 mol of methane burns at a 373K, and 100 kPa. What volume of water vapour is produced? CH4(g) + 2O2(g) =>CO2(g) + 2H2O(g)

**2) 2.00 mol ?**

**3) 2.00mol/1mol = x/2mol**

**x = 4.00 mol**

**4) V = nRT/P**

**V = (4.00 x 8.314 x 373)/100**

**V = 124 L of water vapour**

2) In an industrial application known as the Harber process ammonia to be used as fertilizer results from the reaction of nitrogen and hydrogen. What is the percent yield of ammonia, if 12kL is produced at 450 kPa pressure and 80°C from the reaction of 7.5 kg of hydrogen with 7.0x1026 particles of nitrogen?

**G1 G2 TY = R**

1. **3 H2(g) + 1 N2(g) 🡪 2 NH3(g)**
2. **7.5E3g/2.02 7.0E26/6.02E23 AY = 12kL**

**= 3712…mol = 1162...mol**

**3) X 2mol/3mol x 2mol/1mol**

**= 2475…mol = 2325…mol (limiting nr)**

**4) V = nRT/P**

**= 2325… x 8.314x 353.15 K /450 kPa**

**=15173...L**

1. **% yield = AY/TY x 100**

**= 12000L / 15173…Lx 100 = 79 %**