Order: 1 and 2, 5 and 6 (and 7/8?), 13

Gas Law / Solutions review!

Assume that any acids in this worksheet are strong acids.
You'll need the vapor pressure charts for some of these problems.

The VP in the chart is the max VP / Equilibrium VP / saturation VP at that temp.

1. (You'll need a vapor pressure chart for these)

a. If the temperature is 24.0 °C when the relative humidity is 54%, find the vapor pressure of water in the atmosphere.

$RH = \frac{\text{actual VP}}{\text{max VP (chart)}} \times 100 \qquad 54 = \frac{VP}{22.38\ \text{mmHg}} \times 100$

VP = 12 mmHg

b. What is the relative humidity on a 29°C day, if the vapor pressure of water is 25.4 mmHg?

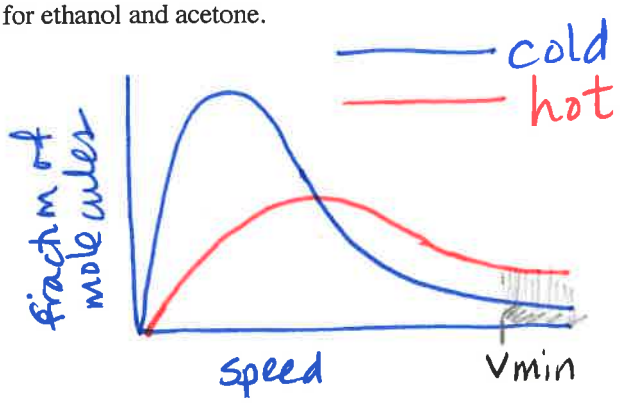
$\frac{25.4\ \text{mmHg}}{30.04\ \text{mmHg}} \times 100 = 84.55 \rightarrow$ **84.6 %**

c. What is the boiling point of water at a very high elevation, where air pressure is 400. mmHg?

≈84 °C since when T = 84°C, the VP is 400. mmHg

2. Consider this chart of Vapor Pressure vs. Temperature, for ethanol and acetone.

Temperature °C	Ethanol VP mmHg	Acetone VP mmHg
-2.3	10	71
19.0	40	188
34.9	100	354
48.0	200	571
56.5	300	761
63.5	400	958
78.4	760	1510
97.5	1520	2560
126.0	3800	5100



a. Explain why the equilibrium vapor pressure of a liquid increases as temperature increases.

As temp increases, a larger fraction of the liquid molecules will have enough speed / enough kinetic energy to overcome the attractive forces between molecules, so more molecules will enter the gas phase, so vapor pressure will increase.

b. Estimate the normal boiling point of acetone. 56 °C / 56.5 °C

c. Estimate the boiling point of ethanol if it is in a pressure chamber, where air pressure is 5 atm. 126 °C

5 × 760 = 3800 mmHg

d. Which liquid has stronger intermolecular forces? Explain how you can tell from the data.

Ethanol has stronger forces.

- Ethanol has a normal boiling pt of 78.4°C, while acetone's normal boiling point of 56 °C. Since ethanol requires a higher temp to reach 1 atm of VP, ethanol's molecules must be more strongly attracted to each other.

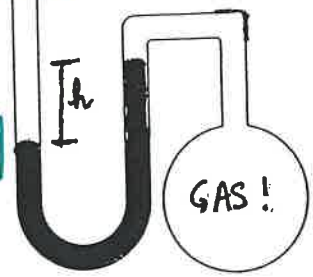
OR

- At a given temp, acetone has higher VP than ethanol, which means acetone's molecules are less strongly attracted to each other.

I subtracted, since the gas is at lower P than the atmosphere.

3. Consider the following mercury manometers. Each are open to the atmosphere on the left side. Assume that the atmospheric pressure is 745 mmHg. Determine the pressure of the gas sample in each case.

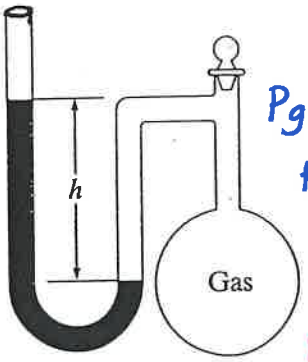
a. h = 3.5 cm = 35 mm



$P_{gas} = 745 - 35$

$P_{gas} =$ **710. mmHg**

b. h = 95 mm



$P_{gas} = 745 + 95$

$P_{gas} =$ **840. mmHg**

(I added since the gas is at a higher P than the atmosphere.

c. In part (a), what would the height be if the fluid in the tubing is water, instead of mercury?

Density of $H_2O_{(l)}$ = 1.00 g/mL

Density of Hg (l) = 13.6 g/mL

$$(35\ mm\ Hg)\left(\frac{13.6\ mm\ H_2O}{1\ mm\ Hg}\right) = 476$$

480 mm tall if H_2O

4. $Ba(OH)_2$ and $C_2H_4(OH)_2$ can both dissolve into water.

a. Compare these two compounds in terms of:

Total concentration of solute particles in a 0.02 M (aq) solution.

$Ba(OH)_2$: $Ba^{+2} + 2OH^{-1}$ ions. $(0.02\ M)\left(\frac{3\ ions}{1\ Ba(OH)_2}\right) =$ **.06 M particles** $Ba(OH)_2$

$C_2H_4(OH)_2$ won't split up since covalent, so just **0.02M $C_2H_4O_2$ particles**

Electrical conductivity of their 0.02 M (aq) solutions.

The $Ba(OH)_2$ (aq) will conduct. $Ba(OH)_2$ is a strong electrolyte.

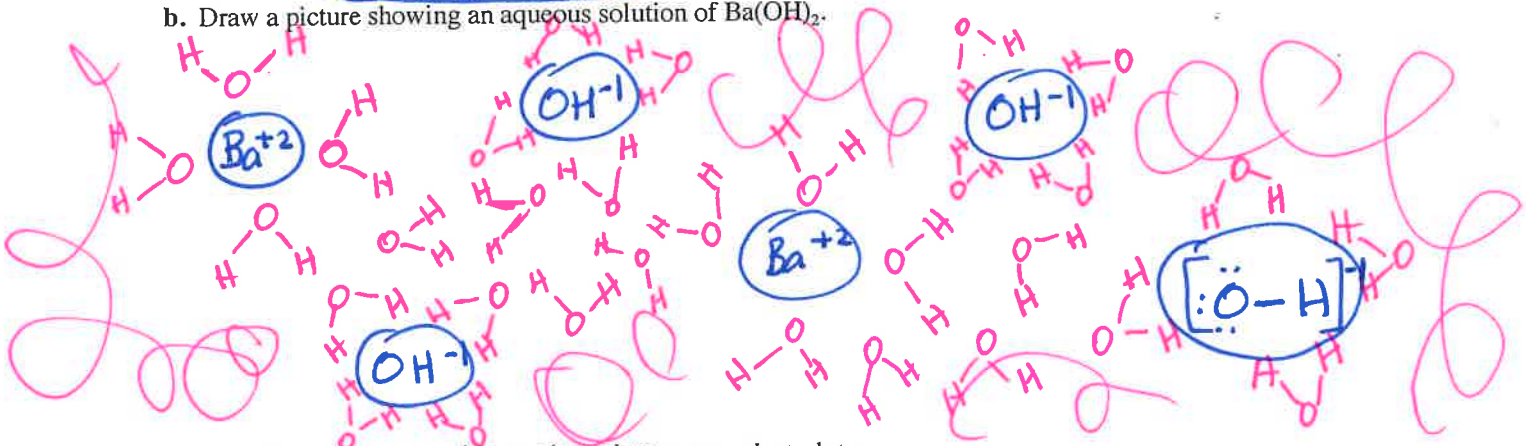
The $C_2H_4(OH)_2$ will not conduct. It is a non electrolyte.

Type(s) of bonding involved in the compound.

$Ba(OH)_2$ is ionic, so it has ionic bonds. (there is also a covalent bond between O and H in the hydroxide)

$C_2H_4(OH)_2$ is covalent, so contains covalent bonds.

b. Draw a picture showing an aqueous solution of $Ba(OH)_2$.



c. Classify each compound as an electrolyte or non-electrolyte:

$C_4H_{10}O$	$CaBr_2$	NaBrO	HNO_3	LiOH	CH_3OH	HBr
NON	Electrolyte	Electrolyte	Electrolyte	Electrolyte	NON	electrolyte

5. Consider two samples of gases:
10.0 liters of fluorine at 1.0 atm and 5.0 °C
10.0 liters of butane (C_4H_{10}) at 1.0 atm and 5.0 °C

F_2: 37.9968 g/mole, 2 atoms per molecule
C_4H_{10}: 58.123 g/mole, 10 atoms per molecule

For each letter, determine which sample has the higher value of the quantity listed (or indicate whether they are equal)
For the ones when one sample has a higher value of the quantity, find the ratio of the higher value to the lower value of that quantity. Assume that the gases are behaving ideally.

a. Density	fluorine	butane	they are equal	$\frac{58.123}{37.9968} = 1.5297$ Butane is 1.5297 times as dense
b. rate of effusion	fluorine	butane	they are equal	$r_{F_2}/r_{butane} = \left(\frac{58.123}{37.9968}\right)^{1/2} = 1.2368$ F_2 is effusion rate is 1.2368 x higher
c. number of atoms	fluorine	butane	they are equal	14/2 = 7 butane has 7x more atoms
d. number of molecules	fluorine	butane	they are equal	$n = PV/RT$ same P, V, T so same n and molecules
e. average kinetic energy	fluorine	butane	they are equal	SAME TEMP so SAME KE
f. root-mean-square speed	fluorine	butane	they are equal	$\frac{v_{F_2}}{v_{C_4H_{10}}} = \left(\frac{58.123}{37.9968}\right)^{1/2} = 1.2368$ Fluorine's avg speed is 1.2368 x faster.

6. For the two samples of gas in the above problem:
a. What temperature (in °C) would be needed to double the root mean square speeds of the gases?

$KE = \frac{1}{2}mv^2$ Must quadruple the Kelvin Temp.

If v increases by a factor of 2, KE must increase by a factor of 4. (2^2)

5 + 273 = 278 K
278 x 4 = 1112 K, so 839 °C

b. What temperature (in °C) would be needed to double the average kinetic energies of the gases?

To double the KE, must double the Kelvin temp.

(278 K) (2) = 556 K so 283 °C

c. How could you change the conditions listed above (they are still fluorine and butane, though) to cause the gases to increase their densities?

$D = \frac{m}{V} = \frac{m}{\left(\frac{nRT}{P}\right)} = \frac{Pm}{nRT} = \frac{PM}{RT}$

- increase Pressure
- decrease Temperature

d. What is the rms speed of fluorine at the conditions listed in #1?

(1 Joule = 1 kg m²/s²)

$v_{rms} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314\ J/mol\cdot K)(278\ K)}{.0379968\ kg/mole}} = 427\ m/s$

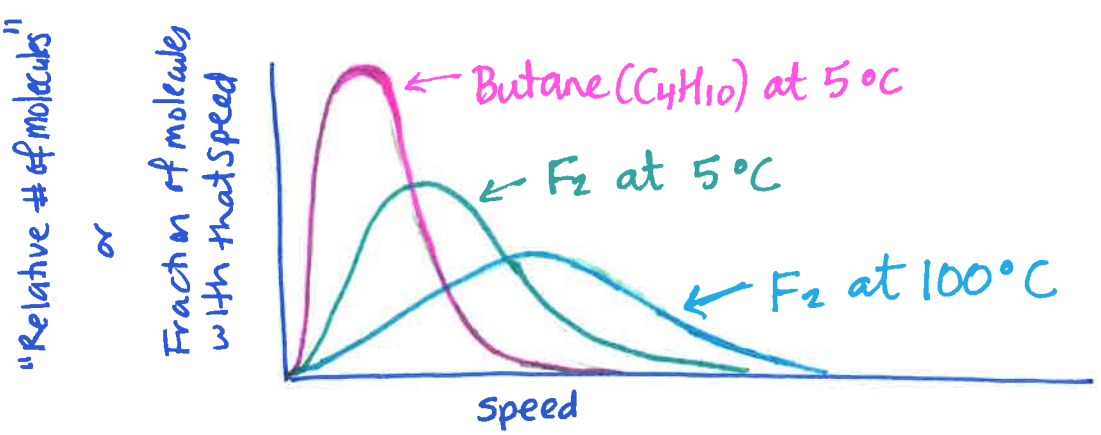
must do molar mass in kg/mole since Joules have kg in them

e. How could you change the temperature and pressure of the gases to cause them to deviate more from ideal behavior? In what ways will the gases act less "ideal" and more "real" if you do this?

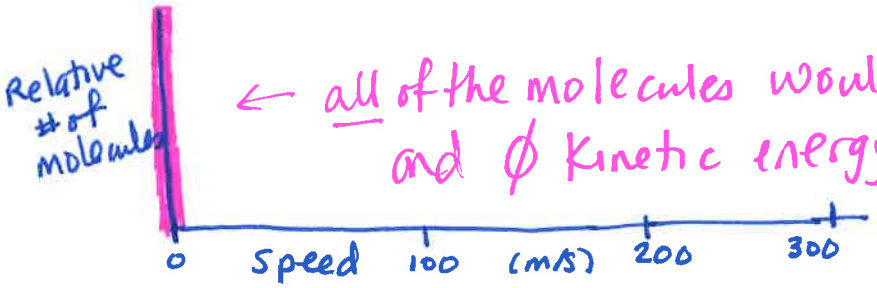
increase pressure and/or decrease temperature

(Both of these will increase $\frac{n}{V}$: the # of moles or molecules per unit volume.

- Intermolecular Forces will be more significant
- The volume occupied by the molecules themselves will be more significant compared to the volume of the container.



6f. Sketch a graph showing the distribution of speeds for the gases in #1. (so, for fluorine and butane, both at 5 °C) (Put the Boltzmann distribution for both gases on the same graph).
g. Still on the same graph, sketch the curve for fluorine again, but at 100 °C
h. What would the Boltzmann distribution look like if the gases could be cooled to zero Kelvin?



← all of the molecules would have ∅ speed and ∅ kinetic energy if at absolute zero. so just a single "spike" at ∅ m/s.

i. Even at the same T and P, butane gas is less ideal than fluorine, because butane molecules are larger (occupy more volume) and also have stronger attractive forces to each other. Suppose you calculated the pressure of a different sample of butane according to the ideal gas law (P = nRT/V) and then you measured the pressure of the sample.

i1. If intermolecular forces were present (but the space occupied by molecules was still negligible), how would the actual pressure compare to the calculated ("ideal") pressure?

actual pressure would be less. If molecules stick to each other, they'll act as if fewer molecules are present, so fewer collisions.

i2. If the molecules themselves occupy significant volume (but had no significant attractive forces), how would the actual pressure compare to the calculated ("ideal") pressure?

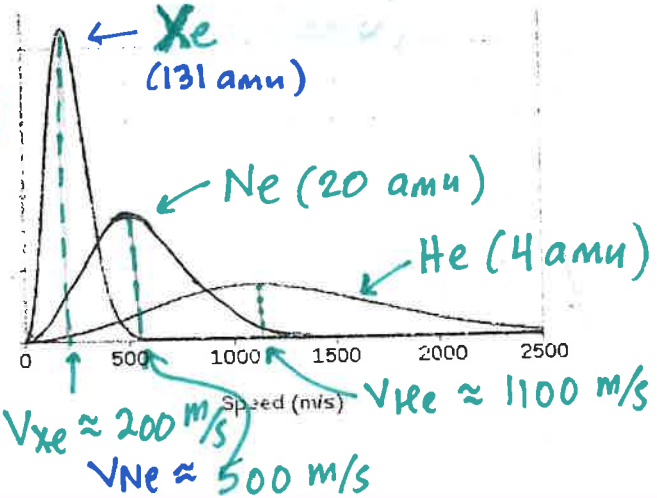
actual pressure would be more than the "ideal" pressure. If molecules occupy significant volume, the actual volume "available" to them will be less than the volume of the container, so pressure will increase.

7. The Boltzmann distributions for Ne, Xe, and He are shown below/right. All gases are at 200 Kelvin.

a. Identify which gas corresponds to which curve, and use the graph to estimate the rms speed of each gas.

see graph: $v_{rms}(Ne) \approx 500$ m/s

$v_{rms}(Xe) \approx 200$ m/s $v_{rms}(He) \approx 1100$ m/s



b. How does the average kinetic energy of Neon compare to the average kinetic energy of Xenon?

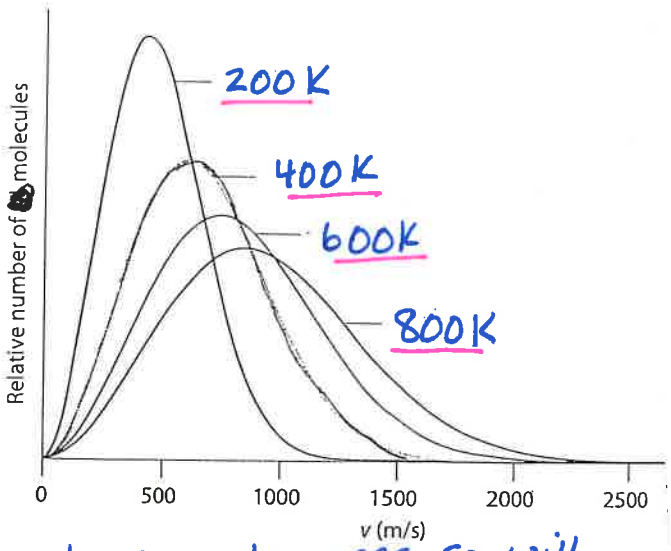
They are the same!!! since both at 200 Kelvin.

8. Boltzmann distributions for ethane gas (C_2H_6) are shown below/right.
The curves correspond to 400 Kelvin, 800 Kelvin, 600 Kelvin, and 200 Kelvin.

a. How does the average kinetic energy of ethane at 400 K compare to the average kinetic energy of ethane at 200 K?

The kinetic energy is higher at 400 Kelvin (twice as high!)

b. Determine which curve corresponds to which temperature.



c. Calculate the rms speed of ethane at 400. Kelvin.

$$v_{rms} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3(8.314\ \frac{J}{mol \cdot K})(400.\ K)}{0.0300694\ \frac{kg}{mole}}}$$

$$= \boxed{576\ m/s}$$

d. How would the effusion rates of ethane and methane compare, if both are at 400. K? (methane is CH_4)

ethane: 30.0694 g/mole

methane: 16.0426 g/mole

$$\frac{r_{CH_4}}{r_{C_2H_6}} = \sqrt{\frac{30.0694}{16.0426}} = 1.3691$$

Methane has a lower molar mass so will have a faster speed at a given temp, so methane will have a higher effusion rate. methane's effusion rate is 1.3691 times greater.

9. A flask contains 100. g of Xenon and 100. g of Krypton (and no other gases) at room temperature.
The total pressure in the flask is 550. mmHg.

a. Without doing any math yet, which gas should be present at a higher partial pressure? Why?

Xe: 131 amu
Kr: 84 amu

Since Kr has a smaller molar mass, 100 g Kr has more moles of gas than 100 g Xe. So Kr's partial pressure will be greater.

b. Without doing any math yet, which gas should have a higher average kinetic energy? Why?

They have the same avg. kinetic energy. Both are at room temp. If temp is the same, kinetic energies are same.

c. Calculate the partial pressure of each gas.

$$(100.\ g\ Xe)\left(\frac{1\ mole}{131.294\ g}\right) = 0.76165\ mole\ Xe$$

$$(100.\ g\ Kr)\left(\frac{1\ mole}{83.798\ g}\right) = 1.19335\ mole\ Kr$$

$$0.76165 + 1.19335 = 1.954996\ moles\ total$$

$$X_{Kr} = \frac{1.19335}{1.954996} = 0.61041$$

$$P_{Kr} = P_{TOTAL} \cdot X_{Kr} = (550.\ mmHg)(0.61041)$$

$$\boxed{P_{Kr} = 336\ mmHg}$$

$P_{Xe} = 550. - 336$

$\boxed{P_{Xe} = 214\ mmHg}$

d. Another flask contains ethanol gas at a partial pressure of 24 mmHg, diethyl ether at a partial pressure of 400. mmHg, and propane at a partial pressure of 227 mmHg. Calculate the mole fraction of propane in this mixture.

$$P_{TOTAL} = P_{ethanol} + P_{ether} + P_{propane} = (24 + 400. + 227)\ mmHg = 651\ mmHg$$

$$P_{propane} = P_{TOTAL} \cdot X_{propane}$$

$$227\ mmHg = 651\ mmHg \cdot X_{propane}$$

$$\boxed{X_{propane} = 0.349}$$

In other words, since 34.9% of the pressure is caused by propane molecules, 34.9% of the molecules must be propane!

10. $2\ H_2O_{2(aq)} \text{------>} 2\ H_2O_{(l)} + O_{2(g)}$ Looked up VP @ 35 °C : 42.2 mmHg.

a. If 2.00 grams of hydrogen peroxide decompose, and the oxygen produced is collected by water displacement at ~~38~~ 35 °C and a room pressure of 748 mmHg, what volume of gas will be collected?

$P_{O2} = P_{TOTAL} - VP_{H2O} = 748\ mmHg - 42.2\ mmHg = 705.8\ mmHg$

$(2.00\ g\ H_2O_2)\left(\frac{1\ mole}{34.0146\ g}\right)\left(\frac{1\ mole\ O_2}{2\ mole\ H_2O_2}\right) = 0.029399\ moles\ O_2$

$V = \frac{nRT}{P} = \frac{(.029399\ mole)(.0821\ \frac{L\cdot g}{mK})(308\ K)}{(705.8/760.\ atm)} = 0.80049\ L \rightarrow$ **0.800 L**

or **800. mL**

b. What was the mole fraction of oxygen in the gas that was collected?

$P_{O2} = P_{TOTAL} \cdot X_{O2}$

$X_{O2} = \frac{P_{O2}}{P_{TOTAL}} = \frac{705.8\ mmHg}{748\ mmHg} =$ **0.944**

c. Household hydrogen peroxide is typically about 0.88 Molar H_2O_2.
If 250. mL of this peroxide solution decomposed completely (so that all the H_2O_2 solute has reacted), what volume of gas should be collected? Assume the oxygen gas is collected by water displacement at 60. °C at a room pressure of 1.00 atm.

VP at 60.°C = 149.4 mmHg acc to chart.

$(0.88\ \frac{mole}{L})(0.250\ L) = 0.22\ moles\ H_2O_2$

$(0.22\ moles\ H_2O_2)\left(\frac{1\ mole\ O_2}{2\ mole\ H_2O_2}\right) = 0.11\ mole\ O_2$

$P_{O2} = 760.\ mmHg - 149.4\ mmHg$

$P_{O2} = 610.6\ mmHg$

$V = \frac{nRT}{P} = \frac{(.11\ mole)(.0821\ \frac{L\cdot atm}{mol\cdot K})(333K)}{(\frac{610.6}{760.}\ atm)} = 3.7431 \rightarrow$ **3.7 L**

(or 3700 mL)

11. A gas with empirical formula CH_2 has an effusion rate that is 71% as high as nitrogen's effusion rate at the same temp. Find the molar mass and molecular formula of the gas.

$\frac{r_x}{r_{N2}} = \sqrt{\frac{M_{N2}}{M_x}}$ $\quad \frac{r_x}{r_{N2}} = \sqrt{\frac{28.0134}{M_x}} = 0.71$ $\quad \frac{28.0134}{M_x} = (0.71)^2$

$M_x = 55.571 \rightarrow$ **56 g/mole** $\quad 56/14 = 4$

CH_2: empirical mass is $\approx$ 14 amu. $\quad 4(CH_2) =$ **C_4H_8**

12. Another gas with the empirical formula CH_2 has a density of 3.02 g/L at 100 °C and 1.1 atm.
Find the molar mass and the molecular formula of the gas.

Assume we have 1 mole of gas:

$V = \frac{nRT}{P} = \frac{(1\text{ mole})(.0821 \frac{L \cdot g}{mol \cdot K})(373K)}{1.1\text{ atm}} = 27.839\text{ L}$ occupied by 1 mole gas

$\left(\frac{3.02\text{ g}}{L}\right)\left(\frac{27.839\text{ L}}{\text{mole}}\right) = \boxed{84\ \frac{g}{mole}}$

or: assume you have 1 L of gas. solve for and get .03592 moles corresponds to 1 Liter.

$\frac{(3.02 g/L)}{(.03592\ mol/L)} = 84\ \frac{g}{mole}$

$\frac{84}{14} = 6$ $\quad 6(CH_2) = \boxed{C_6H_{12}}$

13. Cerium sulfate dihydrate is a solid with a heat of solution of -80 kJ/mole. $Ce_2(SO_4)_3 \cdot 2H_2O$
Sodium nitrate is a solid with a heat of solution of 20 kJ/mole. $NaNO_3$
(Both values are for forming an aqueous solution; the compounds are dissolving into water.)

a. Which substance be more soluble into water as temperature of water increases, and which will be more soluble into water as temperature decreases?

$NaNO_3$ will be more soluble as temp increases, since it has a positive (endothermic) ΔH_{soln}; it needs heat to dissolve.

$Ce_2(SO_4)_3 \cdot 2H_2O$ will be more soluble as temp decreases. It has a negative (exo) ΔH_{soln}, so it must release heat to dissolve.

b. If cerium sulfate dihydrate crystals at room temperature are added to water at room temperature, will the water temperature increase or decrease as the crystals dissolve?

The crystals will release heat, so the H_2O will absorb heat, so H_2O temperature will **increase.**

c. If sodium nitrate crystals at room temperature are added to water at room temperature, will the water temperature increase or decrease as the crystals dissolve?

The crystals will absorb heat, so the H_2O will release heat, so the H_2O temp will **decrease.**

d. A substance's $\Delta H_{solution}$ is determined by $\Delta H_{solvent}$, ΔH_{solute}, and ΔH_{mix}.
For each ΔH term, Explain what process is happening on a molecular/atomic level, and whether the term is exothermic or endothermic, and why. Use sodium nitrate as an example.

$\Delta H_{solvent}$: positive / endothermic. This involves H_2O molecules separating from each other to make "room" for the solute particles. Since the H_2O molecules are attracted to each other, energy is needed (must be absorbed by H_2O) to separate them from each other.

ΔH_{solute} : positive / endothermic. This involves Na^{+1} and NO_3^{-1} ions separating from each other so the solid can dissolve. Since the ions are attracted to each other, they must increase (absorb) potential energy to be separated from each other.

ΔH_{mix} : negative / exothermic. This involves the solvent (H_2O) and solute particles (Na^{+1} and NO_3^{-1}) mixing together, so each ion is surrounded by several H_2O molecules. Since the H_2O and ions are attracted to each other, they release potential energy as they get closer to each other.

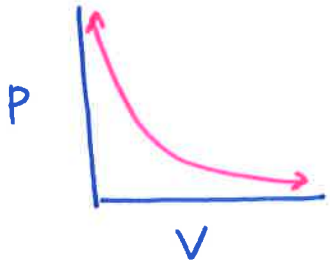
e. Based on the value of sodium nitrate's $\Delta H_{solution}$ (given above part (a)), how do the $\Delta H_{solvent}$, ΔH_{solute}, and ΔH_{mix} compare in magnitude? (how do they compare to each other?)

$\Delta H = +20$ KJ/mole

Since ΔH is positive, more energy is absorbed than released when it dissolves.

So $|\Delta H_{mix}| > |\Delta H_{solute}| + |\Delta H_{solvent}|$

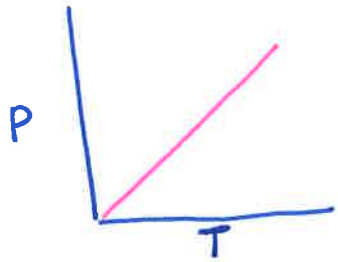
14a. Sketch a graph showing gas pressure vs. volume, assuming that the temperature and moles of gas are constant.



b. Is this a direct or inverse relationship?

INVERSE

c. Sketch a graph showing gas pressure vs temperature, assuming that the volume and moles of gas are constant.



d. Is this a direct or inverse relationship?

DIRECT

e. Explain, in terms of kinetic theory, why gas pressure increases/decreases as temperature increases. (when moles and volume are constant).

As temperature increases, the gas molecules increase kinetic energy; molecules will be moving faster.

gas pressure is caused by collisions of molecules. So as the molecules increase speed, they will collide more frequently with a given surface, and they will collide with more force per collision (the faster they are moving, the more force they must "push off" with, to change direction). So since the molecules collide more frequently and with more force per hit, they will exert more pressure!