

# EXAM 2

# **PRACTICE PROBLEMS**

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CHEMISTRY 161A // FALL 2019

## PRACTICE PROBLEM 1.1

Complete the following chart by calculating the concentration specified in square brackets [ ] in units of M.

— answer —

	Information	[Species]	Concentration (M)
A)	64.7 g LiCl in 250.0 mL water	[LiCl]	6.11 M
B)	0.200 mol CaCl <sub>2</sub> in 2.00 L water	[Cl <sup>-</sup> ]	0.200 M
C)	6.42 × 10 <sup>-2</sup> mM KNO <sub>3</sub>	[KNO <sub>3</sub> ]	6.42 × 10 <sup>-5</sup> M
D)	0.200 mol Na <sub>2</sub> CO <sub>3</sub> in 200.0 g water	[Na <sub>2</sub> CO <sub>3</sub> ]	1.00 M
		[Na <sup>+</sup> ]	2.00 M
		[CO <sub>3</sub> <sup>2-</sup> ]	1.00 M
		[ions]	3.00 M
E)	1.25 M H <sub>3</sub> PO <sub>4</sub>	[H <sup>+</sup> ]	3.75 M



## PRACTICE PROBLEM 1.1

Complete the following chart by calculating the concentration specified in square brackets [ ] in units of M.

— answer —

### Part (D)

To find the molarity (M) of  $[\text{Na}_2\text{CO}_3]$ , I need moles of  $\text{Na}_2\text{CO}_3$  and the volume of water in liters:

$$\begin{aligned} [\text{Na}_2\text{CO}_3] &= \frac{\# \text{ mol Na}_2\text{CO}_3}{\text{Volume (L)}} \\ &= \frac{0.200 \text{ mol Na}_2\text{CO}_3}{200.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mL H}_2\text{O}}{1 \text{ g H}_2\text{O}} \times \frac{1 \text{ L}}{1000 \text{ mL}}} \end{aligned}$$

$$[\text{Na}_2\text{CO}_3] = 1.00 \text{ M}$$

### REMEMBER

The solution is water ( $\text{H}_2\text{O}$ ),

and its density is:

$$D_{\text{H}_2\text{O}} = \frac{1 \text{ g}}{1 \text{ mL}} = \frac{1 \text{ kg}}{1 \text{ L}}$$

To find the concentration of  $[\text{Na}^+]$ ,  $[\text{CO}_3^{2-}]$ , and [total ions]:

$$\frac{1.00 \text{ mol Na}_2\text{CO}_3}{1.00 \text{ L}} \times \frac{2 \text{ mol Na}^+}{1 \text{ mol Na}_2\text{CO}_3} = 2.00 \text{ M Na}^+$$

$$\frac{1.00 \text{ mol Na}_2\text{CO}_3}{1.00 \text{ L}} \times \frac{1 \text{ mol CO}_3^{2-}}{1 \text{ mol Na}_2\text{CO}_3} = 1.00 \text{ M CO}_3^{2-}$$

$$\frac{1.00 \text{ mol Na}_2\text{CO}_3}{1.00 \text{ L}} \times \frac{3 \text{ mol ions}}{1 \text{ mol Na}_2\text{CO}_3} = 3.00 \text{ M ions}$$

## PRACTICE PROBLEM 1.1

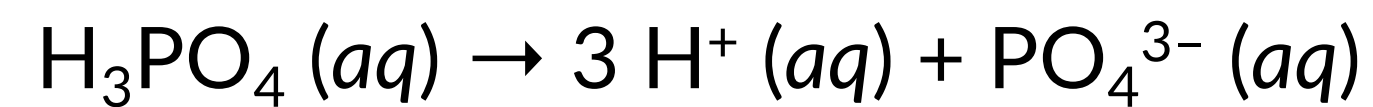
Complete the following chart by calculating the concentration specified in square brackets [ ] in units of M.

— *answer* —

### Part (E)

First, identify that  $\text{H}_3\text{PO}_4$  is a strong acid, which means that it will completely dissociate in water. Moreover, it is a polyprotic (more specifically, a triprotic acid), so it will dissociate into 3  $\text{H}^+$  ions per  $\text{H}_3\text{PO}_4$  unit.

In other words,



To find the concentration of  $[\text{H}^+]$ :

$$\frac{1.25 \text{ mol H}_3\text{PO}_4}{1.00 \text{ L}} \times \frac{3 \text{ mol H}^+}{1 \text{ mol H}_3\text{PO}_4} = 3.75 \text{ M H}^+$$

## PRACTICE PROBLEM 1.2

You have 2.50 mL of a 0.559 M solution of  $\text{Cl}^-$ . How much water do you need to add to make a solution that is 116 mM  $\text{Cl}^-$ ?

— *answer* —

First, find how many moles of  $\text{Cl}^-$  ions you have in 2.50 mL of the 0.559 M  $\text{Cl}^-$  solution:

$$[\text{Cl}^-] = \frac{\# \text{ mol Cl}^-}{\text{Volume (L)}}$$

$$0.559 \text{ M} = \frac{x \text{ mol Cl}^-}{2.50 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}}$$

$$x = 0.00139_8 \text{ mol Cl}^-$$

Now, figure out the total volume of a 116 mM  $\text{Cl}^-$  solution containing this many moles of  $\text{Cl}^-$  ions:

$$[\text{Cl}^-] = \frac{\# \text{ mol Cl}^-}{\text{Volume (L)}}$$

$$116 \text{ mM} \times \frac{1 \text{ M}}{1000 \text{ mM}} = \frac{0.00139_8 \text{ mol Cl}^-}{V}$$

$$V = 0.0120_5 \text{ L}$$

∴ This is total volume!

So subtract to find the amount of water added:  $V_{\text{added}} = 12.0_5 \text{ mL} - 2.50 \text{ mL} = 9.6 \text{ mL water added}$ .

### ALTERNATIVE SOLUTION:

You can use the following equation:

$$\begin{aligned} M_1 V_1 &= M_2 V_2 \\ (0.559 \text{ M})(2.50 \text{ mL}) &= (0.116 \text{ M})V_2 \\ V_2 &= 12.0_5 \text{ mL} \end{aligned}$$

But you still have to subtract at the end to find the volume of water added to the 2.50 mL of the original solution.

## PRACTICE PROBLEM 1.3

Which of the following has the greatest concentration of total dissolved ions?

0.25 M NaBr

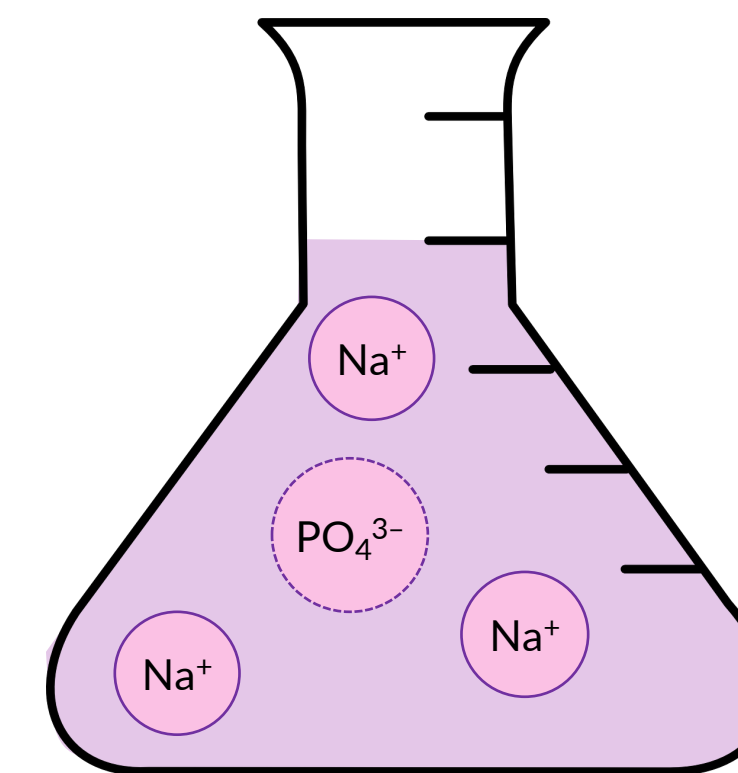
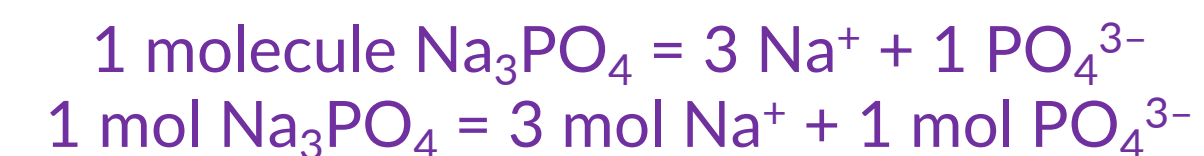
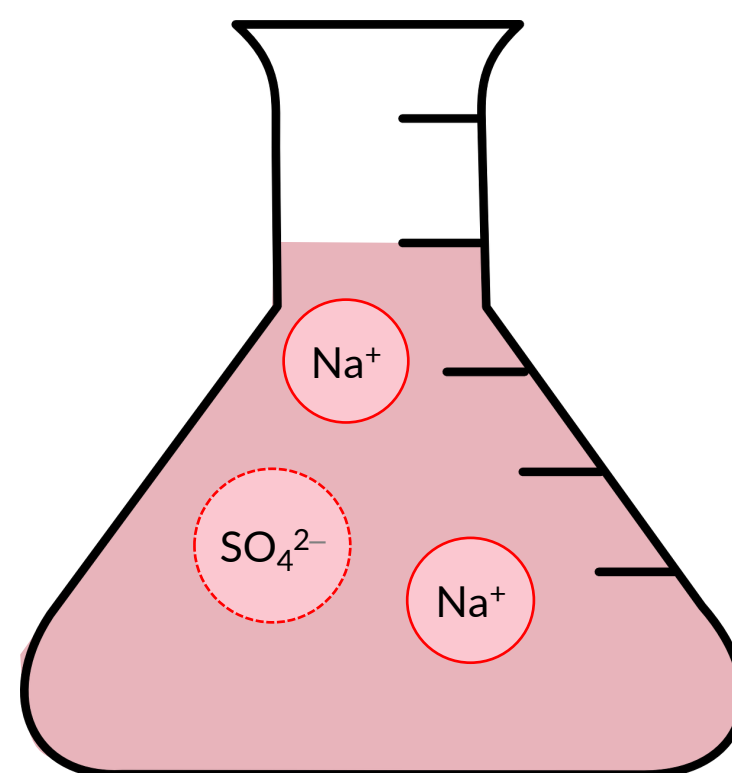
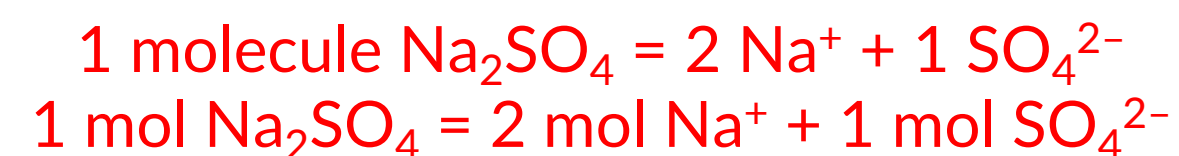
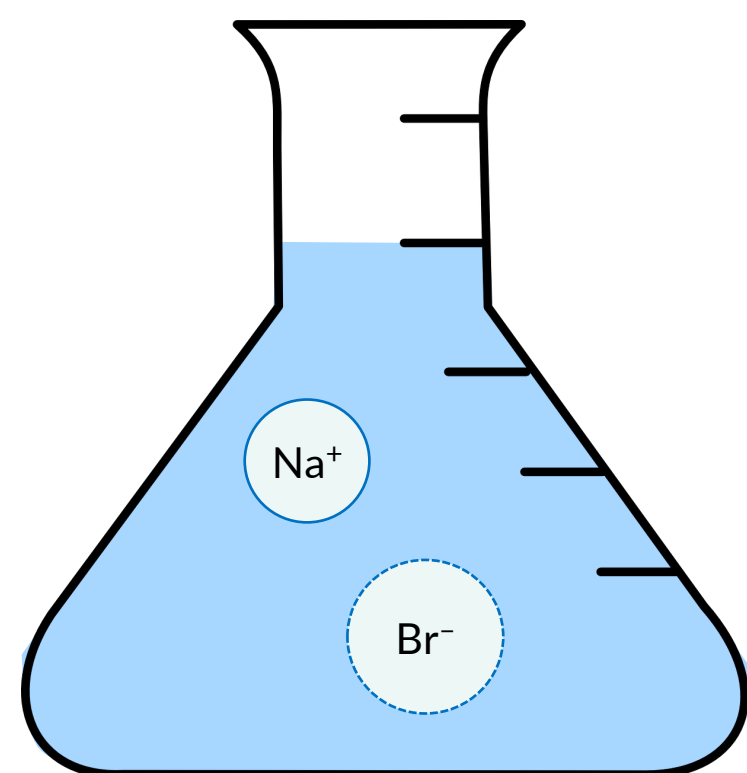
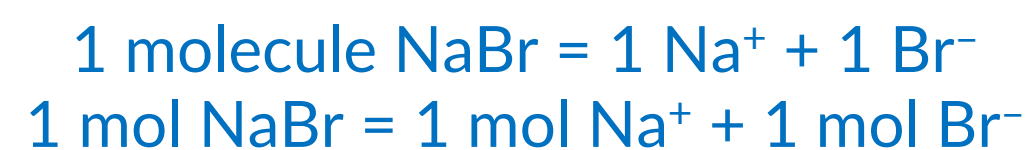
0.25 M Na<sub>2</sub>SO<sub>4</sub>

0.25 M Na<sub>3</sub>PO<sub>4</sub>

— answer —

Firstly, these are all soluble salts.

We can then represent the dissociation of each salt into its ions:



Understand that the concentration of *ions* would be:



$$[\text{ions}] = 2 \times 0.25 \text{ M} = 0.50 \text{ M}$$



$$[\text{ions}] = 3 \times 0.25 \text{ M} = 0.75 \text{ M}$$



$$[\text{ions}] = 4 \times 0.25 \text{ M} = 1.00 \text{ M}$$

## PRACTICE PROBLEM 2.1

For each reaction, write out the balanced molecular, overall/total/complete ionic, and net ionic equations:

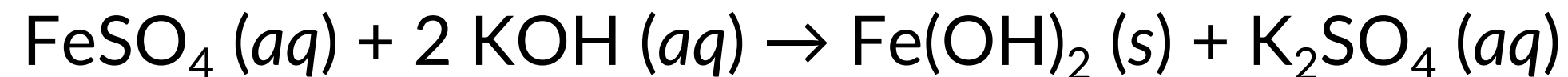
(a) An iron(II) sulfate solution is mixed with a potassium hydroxide solution.

(b) A lead(II) nitrate solution is mixed with a solution of potassium bromide.

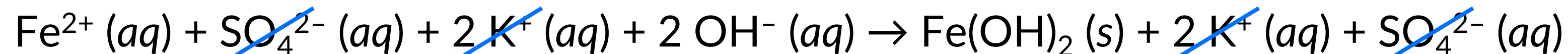
(c) Aqueous nitric acid is mixed into a solution of sodium hydroxide.

— answer —

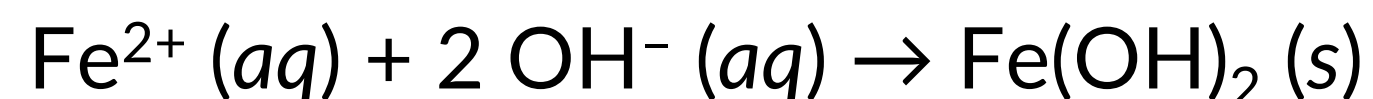
(a) An iron(II) sulfate solution is mixed with a potassium hydroxide solution.



*molecular equation*

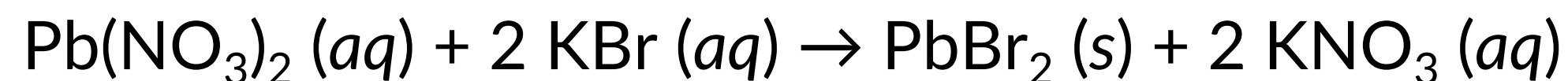


*overall ionic equation*

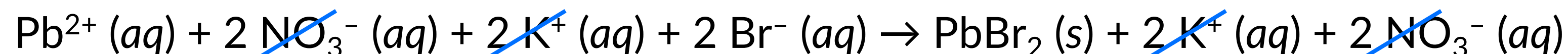


*net ionic equation*

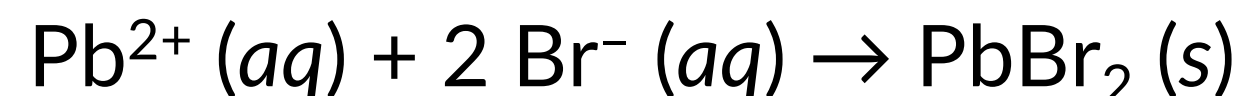
(b) A lead(II) nitrate solution is mixed with a solution of potassium bromide.



*molecular equation*

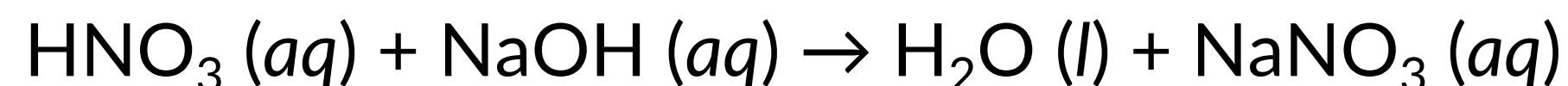


*overall ionic equation*

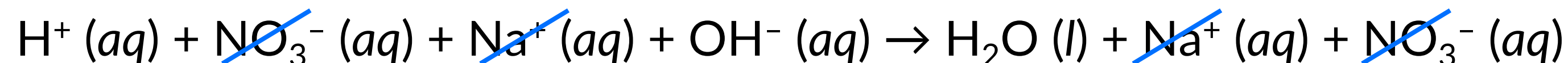


*net ionic equation*

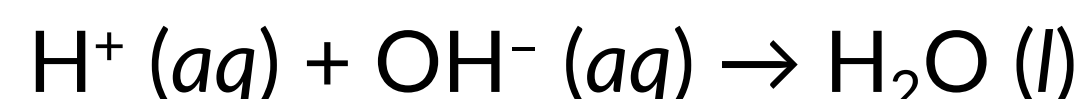
(c) Aqueous nitric acid is mixed into a solution of sodium hydroxide.



*molecular equation*



*overall ionic equation*



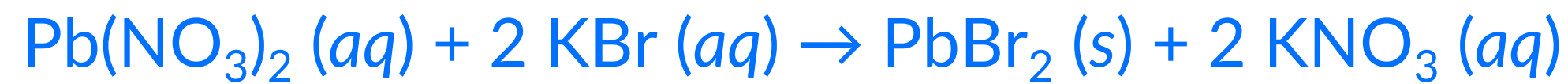
*net ionic equation*

## PRACTICE PROBLEM 2.2

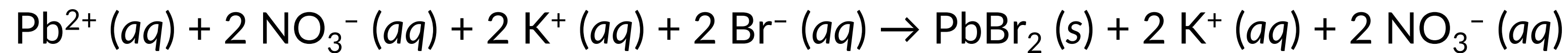
You mix a 1.00 L of a 0.174 M solution of lead(II) nitrate with 1.00 L of a 0.130 M solution of potassium bromide. What mass (in grams) of solid precipitate will be formed?

— answer —

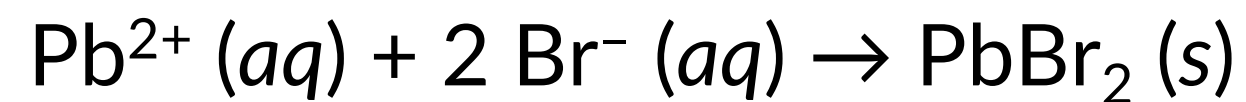
From the previous slide, we know the following balanced chemical equations apply in this precipitation reaction:



*molecular equation*



*overall ionic equation*



*net ionic equation*

It's easiest to apply the molecular equation in stoichiometry of precipitation reactions. First, determine the number of moles of the starting reactants to determine which is the limiting reactant:

$$0.174 \text{ M Pb}(\text{NO}_3)_2 = \frac{x \text{ mol}}{1.00 \text{ L}}$$
$$x = 0.174 \text{ mol Pb}(\text{NO}_3)_2$$

$$0.130 \text{ M KBr} = \frac{y \text{ mol}}{1.00 \text{ L}}$$
$$y = 0.130 \text{ mol KBr}$$

∴ The limiting reactant is KBr because we need a 1:2 mole ratio for  $\text{Pb}(\text{NO}_3)_2$ :KBr, but have a 1:0.747 mole ratio.

Now, figure out how much barium carbonate you can make from this amount of limiting reactant:

$$0.130 \text{ mol KBr} \times \frac{1 \text{ mol PbBr}_2}{2 \text{ mol KBr}} \times \frac{367.0 \text{ g}}{1 \text{ mol PbBr}_2} = 23.9 \text{ g PbBr}_2$$

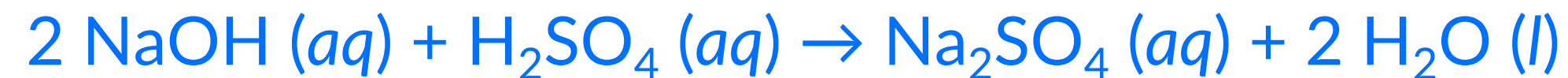


## PRACTICE PROBLEM 2.3

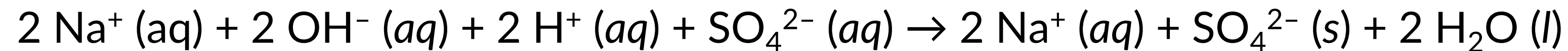
What volume of a 0.500 M NaOH solution would be required to completely neutralize 40.0 mL of 0.400 M H<sub>2</sub>SO<sub>4</sub>?

— answer —

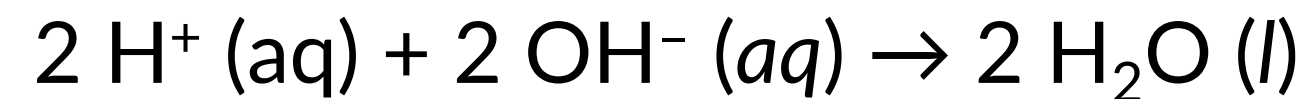
We can write the following balanced chemical equations for the neutralization reaction:



*molecular equation*



*overall ionic equation*



*net ionic equation*

It's easiest to apply the molecular equation in stoichiometry of neutralization reactions.

First, determine the number of moles of the H<sub>2</sub>SO<sub>4</sub>:

$$0.400 \text{ M H}_2\text{SO}_4 = \frac{x \text{ mol}}{0.0400 \text{ L}}$$
$$x = 0.0160 \text{ mol H}_2\text{SO}_4$$

To neutralize all of the H<sub>2</sub>SO<sub>4</sub>, we need an twice the amount NaOH:

$$0.0160 \text{ mol H}_2\text{SO}_4 \times \frac{2 \text{ mol NaOH}}{1 \text{ mol H}_2\text{SO}_4} = 0.0320 \text{ mol NaOH}$$

Now, figure out what volume of 0.500 M NaOH solution gives this many moles:

$$0.500 \text{ M NaOH} = \frac{0.0320 \text{ mol NaOH}}{V}$$
$$V = 0.0640 \text{ L}$$

## PRACTICE PROBLEM 3.1

Give the oxidation state/number for each element in each flask.

— *answer* —

**A**

0.010 mol SF<sub>6</sub>  
{146.1 g/mol}  
300 K

S = +6

F = -1

**B**

0.5 mol CO<sub>2</sub>  
{44.0 g/mol}  
600 K

C = +4

O = -2

**C**

3.0 atm F<sub>2</sub>  
{38.0 g/mol}  
273 K

F = 0

**D**

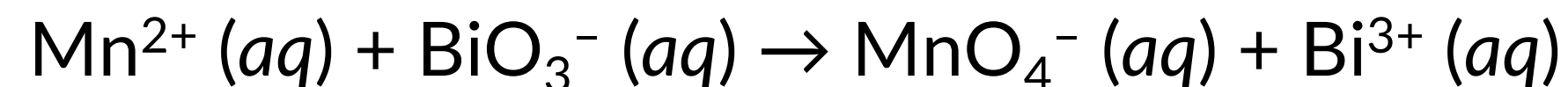
1.0 atm NO  
{30.0 g/mol}  
200 K

N = +2

O = -2

## PRACTICE PROBLEM 3.2

Balance the following redox equation using the half-reactions method in aqueous acidic medium.

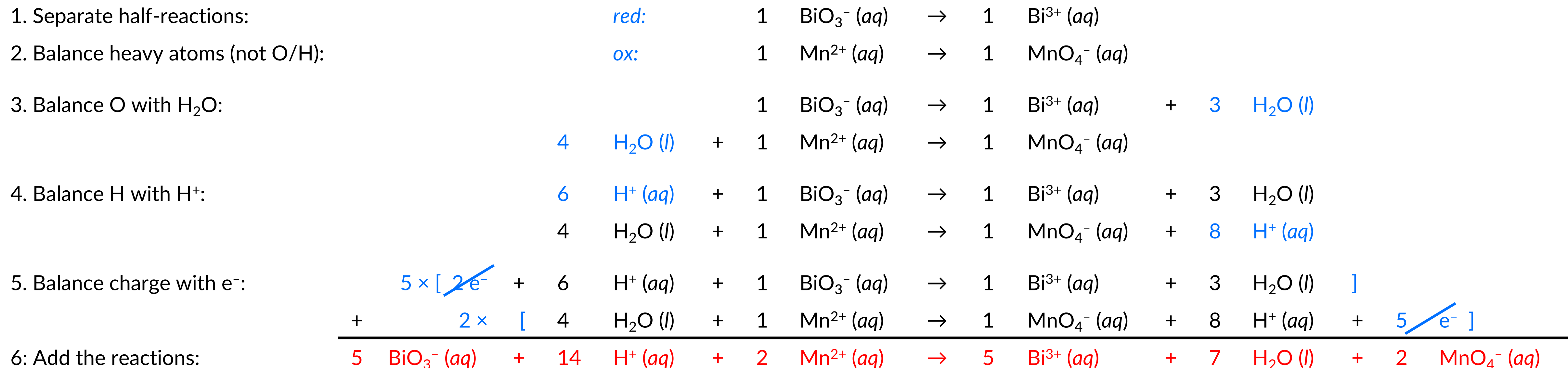


— answer —

We can start by assigning the oxidation states, so that we can determine the corresponding reduction and oxidation half-reactions.



Now, we can balance the redox equation using the half-reactions method through the following steps.



## PRACTICE PROBLEM 4.1

In a calorimeter at constant pressure and 25.2 °C, you mix solutions of potassium carbonate and barium nitrate, which releases 196 kJ of heat. If the total volume of the solution is 1.20 L, what is the final temperature of the solution after the reaction is complete?

$$\text{Molar heat capacity H}_2\text{O (l)} = 75.3 \frac{\text{J}}{\text{mol}\cdot\text{K}}$$

— *answer* —

Understand the heat of the reaction occurring is absorbed by the surrounding solution (the water), so

$$q_{\text{water}} = -\Delta H_{\text{rxn}} = 196 \text{ kJ}$$

Now, figure out what happens to the water if we were to heat it with 196 kJ of heat:

$$q_{\text{water}} = nc_p\Delta T$$

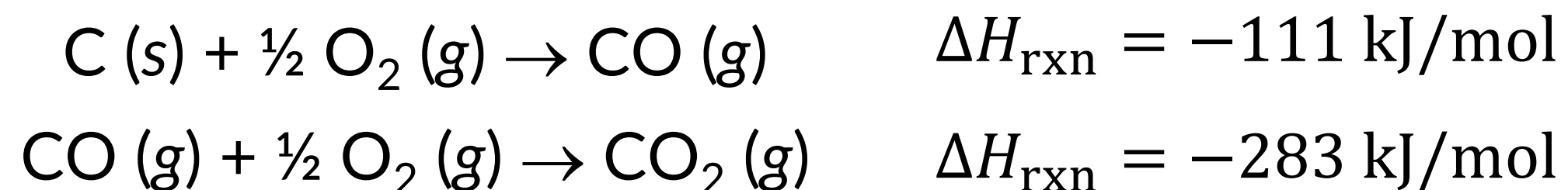
$$196 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = \left( 1.20 \text{ L H}_2\text{O} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left( 75.3 \frac{\text{J}}{\text{mol}\cdot^\circ\text{C}} \right) (T_f - 25.2 \text{ }^\circ\text{C})$$

$$T_f = 64.3 \text{ }^\circ\text{C}$$



## PRACTICE PROBLEM 4.2

Determine the standard heat of formation ( $\Delta H_f^\circ$ ) of  $\text{CO}_2(g)$  using the following thermochemical data.



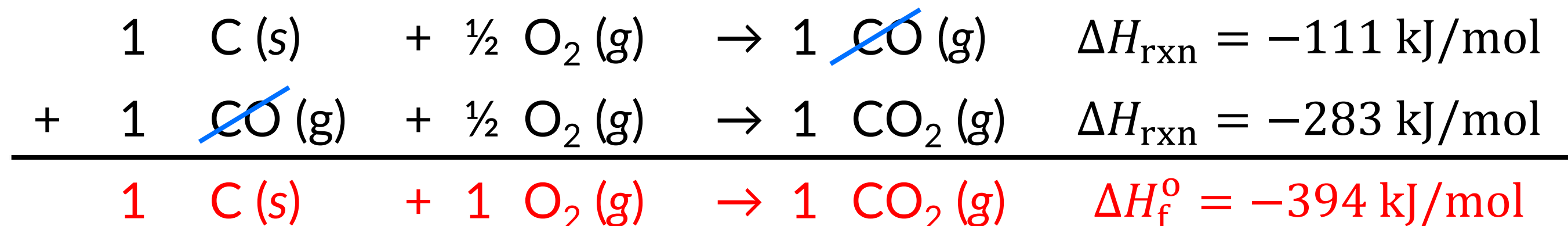
— answer —

The standard heat of formation ( $\Delta H_f^\circ$ ) of  $\text{CO}_2(g)$  can be expressed as the following reaction:



Since enthalpy is a state function, if we add up these two reactions (algebraically) to get the formation reaction of interest, then we can also get the heat of formation through the same algebraic manipulations.

Let's simply add reactions (1) and (2) to get reaction (3) below:



## PRACTICE PROBLEM 5.1

Consider the combustion of ammonia in air:  $4 \text{NH}_3 (g) + 5 \text{O}_2 (g) \rightarrow 4 \text{NO} (g) + 6 \text{H}_2\text{O} (g)$

Calculate the heat of the reaction ( $\Delta H_{\text{rxn}}^{\circ}$ ) using the following standard heats of formations.

$$\Delta H_{\text{f}}^{\circ} [\text{NH}_3 (g)] = -46.1 \text{ kJ/mol}$$

$$\Delta H_{\text{f}}^{\circ} [\text{H}_2\text{O} (g)] = -241.8 \text{ kJ/mol}$$

$$\Delta H_{\text{f}}^{\circ} [\text{NO} (g)] = 90.3 \text{ kJ/mol}$$

— answer —

The heat of the reaction ( $\Delta H_{\text{rxn}}^{\circ}$ ) can be determined using the following expression:

$$\Delta H_{\text{rxn}}^{\circ} = \sum n_{\text{products}} \Delta H_{\text{f,products}}^{\circ} - \sum n_{\text{reactants}} \Delta H_{\text{f,reactants}}^{\circ}$$

Plugging in the values above, we get:

$$\Delta H_{\text{rxn}}^{\circ} = n_{\text{NO}} \Delta H_{\text{f}}^{\circ} [\text{NO} (g)] + n_{\text{H}_2\text{O}} \Delta H_{\text{f}}^{\circ} [\text{H}_2\text{O} (g)] - n_{\text{NH}_3} \Delta H_{\text{f}}^{\circ} [\text{NH}_3 (g)] - n_{\text{O}_2} \Delta H_{\text{f}}^{\circ} [\text{O}_2 (g)]$$

$$= 4 \text{ mol} \times 90.3 \frac{\text{kJ}}{\text{mol}} + 6 \text{ mol} \times \left( -241.8 \frac{\text{kJ}}{\text{mol}} \right) - 4 \text{ mol} \times \left( -46.1 \frac{\text{kJ}}{\text{mol}} \right) - 5 \text{ mol} \times 0 \frac{\text{kJ}}{\text{mol}}$$

$$= 361.2 \frac{\text{kJ}}{\text{mol}} - 1450.8 \frac{\text{kJ}}{\text{mol}} + 184.4 \frac{\text{kJ}}{\text{mol}}$$

$$\Delta H_{\text{rxn}}^{\circ} = -905.2 \frac{\text{kJ}}{\text{mol}}$$

## PRACTICE PROBLEM 5.2

Consider the combustion of ammonia in air:  $4 \text{ NH}_3 (\text{g}) + 5 \text{ O}_2 (\text{g}) \rightarrow 4 \text{ NO} (\text{g}) + 6 \text{ H}_2\text{O} (\text{g})$

4.00 moles of  $\text{NH}_3$  are combusted in a 10.0 L water bath and the temperature of the water rises by 28.0 °C. Calculate the heat of the reaction ( $\Delta H_{\text{rxn}}^\circ$ ) from the calorimetry data if the molar heat capacity  $\text{H}_2\text{O} (\text{l})$  is 75.3 J/mol·K.

— answer —

The heat of the reaction ( $\Delta H_{\text{rxn}}^\circ$ ) can be determined by figuring out how much heat ( $q$ ) the water absorbed based on calorimetry:

$$q_{\text{water}} = -\Delta H_{\text{rxn}}^\circ$$

Therefore,

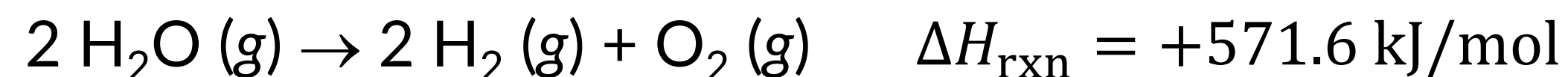
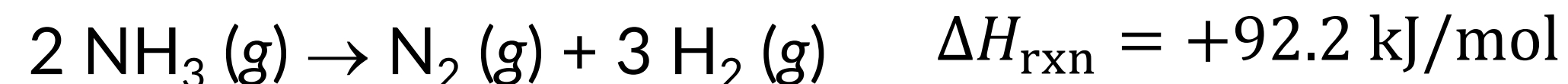
$$\begin{aligned}\Delta H_{\text{rxn}}^\circ &= -q_{\text{water}} \\ &= -n_{\text{H}_2\text{O}} c_p \Delta T \\ &= -\left(10.0 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}}\right) \left(75.3 \frac{\text{J}}{\text{mol} \cdot \text{K}}\right) (28.0 \text{ K}) \\ &= -(554.9 \text{ mol}) \left(75.3 \frac{\text{J}}{\text{mol} \cdot \text{K}}\right) (28.0 \text{ K}) \\ \Delta H_{\text{rxn}}^\circ &= -1.17 \times 10^6 \text{ J}\end{aligned}$$

## PRACTICE PROBLEM 5.3

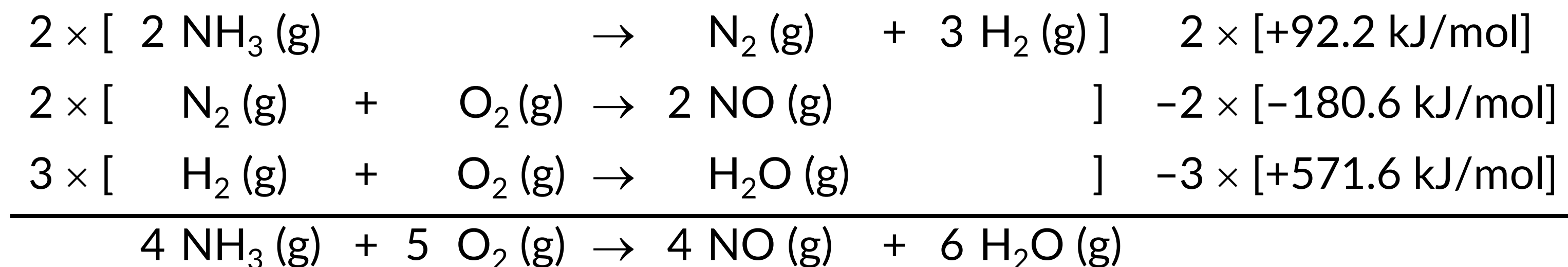
Consider the combustion of ammonia in air:  $4 \text{NH}_3 (\text{g}) + 5 \text{O}_2 (\text{g}) \rightarrow 4 \text{NO} (\text{g}) + 6 \text{H}_2\text{O} (\text{g})$

Calculate the heat of the reaction ( $\Delta H_{\text{rxn}}^{\circ}$ ) using Hess's Law and following thermochemical data for related reactions.

— answer —



We can manipulate the above three equations to formulate our reaction of interest.



Therefore, the heat of the reaction is:

$$\begin{aligned} \Delta H_{\text{rxn}}^{\circ} &= 2 \times \left( +92.2 \frac{\text{kJ}}{\text{mol}} \right) - 2 \times \left( -180.6 \frac{\text{kJ}}{\text{mol}} \right) - 3 \times \left( +571.6 \frac{\text{kJ}}{\text{mol}} \right) \\ &= (184.4 + 361.2 - 1714.8) \text{ kJ/mol} \end{aligned}$$

$$\Delta H_{\text{rxn}}^{\circ} = -1169.2 \frac{\text{kJ}}{\text{mol}}$$



## PRACTICE PROBLEM 6.1

A bulb with a volume of 500.0 mL is filled with a gas at STP. How many moles of gas are in the bulb?

— *answer* —

Start by recognizing that the contents of the bulb are at STP (1 atm and 0 °C).

This means that 1 mole of the gas will occupy 22.4 L of volume.

We can use the following dimensional analysis to find the total number of moles:

$$500.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.02232 \text{ mol gas}$$

*Alternatively, use the Ideal Gas Law:*

$$\begin{aligned} PV &= nRT \\ n &= \frac{PV}{RT} \\ &= \frac{(1.0 \text{ atm}) \left( 500.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \right)}{\left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273.15 \text{ K})} \\ n &= 0.02232 \text{ mol gas} \end{aligned}$$

## PRACTICE PROBLEM 6.2

A 5.00 L rigid container is initially filled with 0.2044 mol of  $\text{C}_9\text{H}_8\text{O}$  gas at 1.00 atm and 25.0 °C. What is the pressure inside the container if the temperature is lowered to 5.0 °C?

— *answer* —

Notice that the number of moles ( $n$ ) and the volume ( $V$ ) of the container are constant, and we are only changing the pressure ( $P$ ) and temperature ( $T$ ).

Therefore, we can set up the following gas relationship to determine the new pressure ( $P_2$ ):

$$\begin{aligned}\frac{P_1}{T_1} &= \frac{P_2}{T_2} \\ P_2 &= \frac{P_1 T_2}{T_1} \\ &= \frac{(1.00 \text{ atm})(5.0 + 273.15)\text{K}}{(25.0 + 273.15)\text{K}} \\ P_2 &= 0.933 \text{ atm}\end{aligned}$$

## PRACTICE PROBLEM 7.1

Which of the following 1.00 L containers has the largest number of moles of gas?

— answer —

**A**

0.010 mol SF<sub>6</sub>  
{146.1 g/mol}  
300 K

**B**

0.5 mol CO<sub>2</sub>  
{44.0 g/mol}  
600 K

**C**

3.0 atm F<sub>2</sub>  
{38.0 g/mol}  
273 K

**D**

1.0 atm NO  
{30.0 g/mol}  
200 K

$$n_{\text{SF}_6} = 0.010 \text{ mol SF}_6$$

$$n_{\text{CO}_2} = 0.5 \text{ mol CO}_2$$

$$n_{\text{F}_2} = \frac{(3.0 \text{ atm})(1.0 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (273 \text{ K})} = 0.13 \text{ mol F}_2$$

$$n_{\text{NO}} = \frac{(1.0 \text{ atm})(1.0 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (200 \text{ K})} = 0.061 \text{ mol NO}$$

To convert each of these mole amounts, we would do the following ( $N_A$  = Avogadro's Number):

$$n_X \times \frac{N_A \text{ molecules}}{1 \text{ mol}} \times \frac{y \text{ atoms}}{1 \text{ molecule}} = (n_X \cdot N_A \cdot y) \text{ atoms}$$

So, **Flask B** has the largest number of atoms (largest  $n_X \cdot y$ ).

## PRACTICE PROBLEM 7.2

Which of the following 1.00 L containers has the greatest pressure?

— answer —

**A**

0.010 mol SF<sub>6</sub>  
{146.1 g/mol}  
300 K

**B**

0.5 mol CO<sub>2</sub>  
{44.0 g/mol}  
600 K

**C**

3.0 atm F<sub>2</sub>  
{38.0 g/mol}  
273 K

**D**

1.0 atm NO  
{30.0 g/mol}  
200 K

$$P_{\text{SF}_6} = \frac{(0.010 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (300 \text{ K})}{1.0 \text{ L}} = 0.25 \text{ atm}$$

$$P_{\text{CO}_2} = \frac{(0.5 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (600 \text{ K})}{1.0 \text{ L}} = 25 \text{ atm}$$

$$P_{\text{F}_2} = 3.0 \text{ atm}$$

$$P_{\text{NO}} = 1.0 \text{ atm}$$

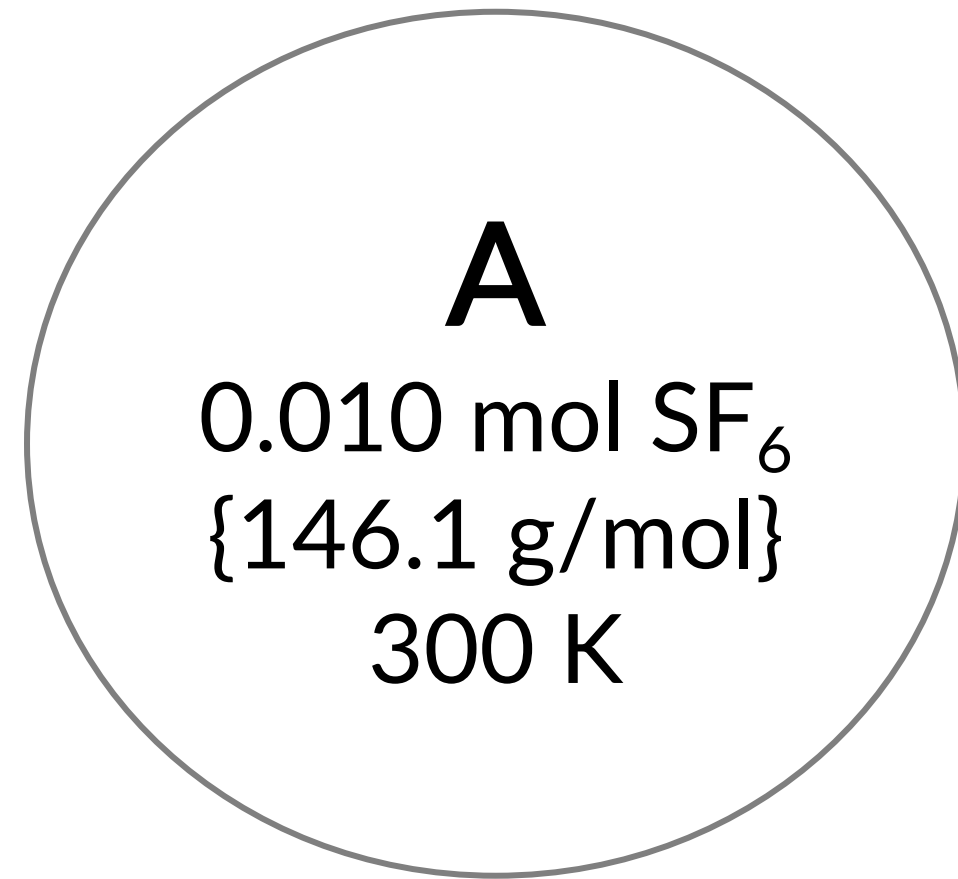
So, **Flask B** has the greatest pressure.



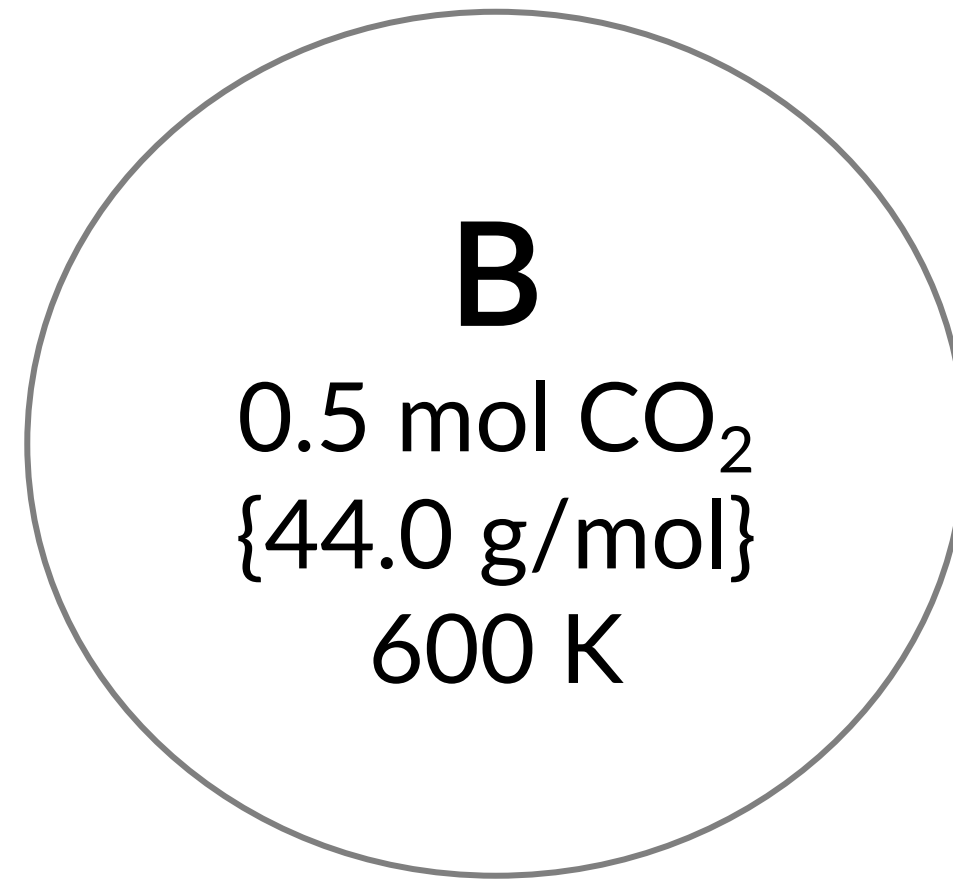
## PRACTICE PROBLEM 7.3

In which container would effusion of the gas the slowest?

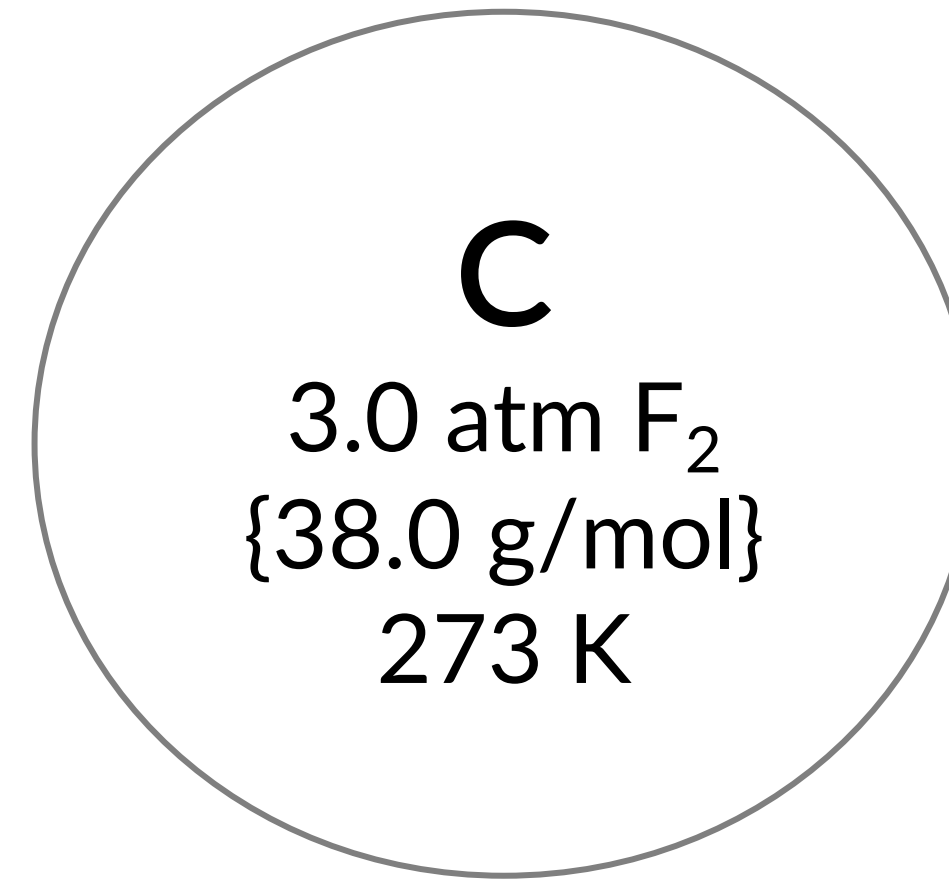
— answer —



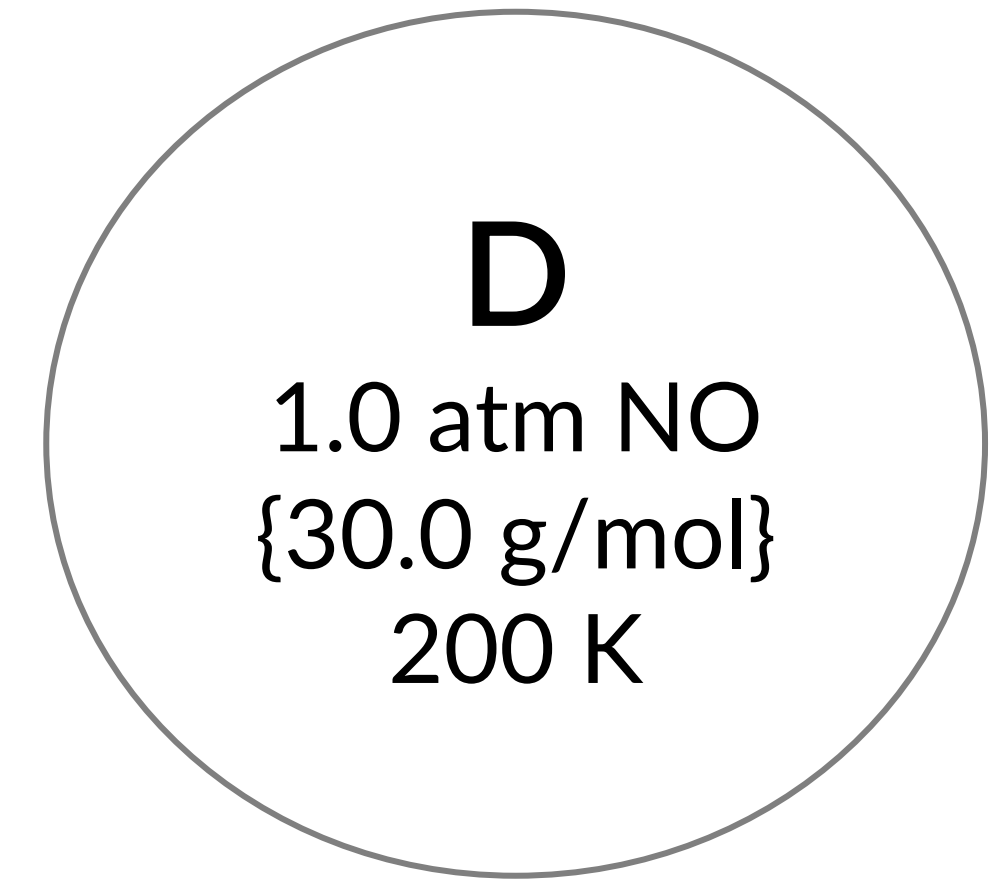
$$\frac{T_A}{MM_A} = \frac{300}{146.1} = 2.1$$



$$\frac{T_B}{MM_B} = \frac{600}{44.0} = 13.6$$



$$\frac{T_C}{MM_C} = \frac{273}{38.0} = 7.2$$



$$\frac{T_D}{MM_D} = \frac{200}{30.0} = 6.7$$

$$v_{\text{rms}} = \sqrt{\frac{3RT}{MM}} \Rightarrow v_{\text{rms}} \propto \frac{T}{MM}$$

The larger the  $v_{\text{rms}}$ , the slower the rate of diffusion.

So, **Flask A** has the slowest rate of diffusion.

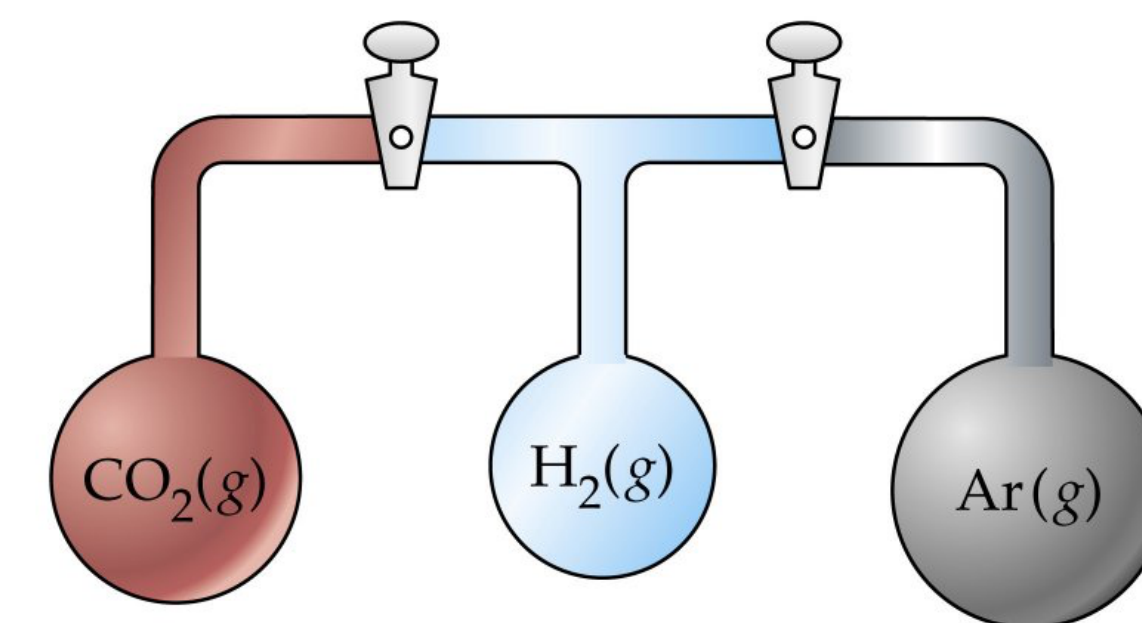
## PRACTICE PROBLEM 8

Consider the diagram to the right, where three containers of gas are connected by stopcocks.

What is the partial pressure of each gas if the stopcocks are opened?

Assume no reaction between the gases.

— answer —



$$\begin{array}{lll} P = 2.13 \text{ atm} & P = 0.861 \text{ atm} & P = 1.15 \text{ atm} \\ V = 1.50 \text{ L} & V = 1.00 \text{ L} & V = 2.00 \text{ L} \end{array}$$

Solve for the number of moles using the pressures of each gas using the ideal gas law.

$$\begin{aligned} n_{\text{CO}_2} &= \frac{PV}{RT} \\ &= \frac{(2.13 \text{ atm})(1.50 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})} \\ n_{\text{CO}_2} &= 0.130_7 \text{ mol} \end{aligned}$$

$$\begin{aligned} n_{\text{H}_2} &= \frac{PV}{RT} \\ &= \frac{(0.861 \text{ atm})(1.00 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})} \\ n_{\text{H}_2} &= 0.0352_1 \text{ mol} \end{aligned}$$

$$\begin{aligned} n_{\text{Ar}} &= \frac{PV}{RT} \\ &= \frac{(1.15 \text{ atm})(2.00 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})} \\ n_{\text{Ar}} &= 0.0940_5 \text{ mol} \end{aligned}$$

If we open the stopcocks, the total volume changes to 4.50 L. Now solve for new pressures:

$$\begin{aligned} P_{\text{CO}_2} &= \frac{nRT}{V} \\ &= \frac{(0.130_7 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{4.50 \text{ L}} \end{aligned}$$

$$P_{\text{CO}_2} = 0.710 \text{ atm}$$

$$\begin{aligned} P_{\text{H}_2} &= \frac{nRT}{V} \\ &= \frac{(0.0352_1 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{4.50 \text{ L}} \end{aligned}$$

$$P_{\text{H}_2} = 0.191 \text{ atm}$$

$$\begin{aligned} P_{\text{Ar}} &= \frac{nRT}{V} \\ &= \frac{(0.0940_5 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{4.50 \text{ L}} \end{aligned}$$

$$P_{\text{Ar}} = 0.511 \text{ atm}$$

## PRACTICE PROBLEM 9

You react 2.00 L of F<sub>2</sub> gas (at 20.0 °C and 1.00 atm) with 1.50 L of H<sub>2</sub> gas (at -20.0 °C and 1.50 atm), which react to form gaseous HF (molar mass = 20.01 g/mol). What is the mass of HF gas that forms?

— answer —

Start by writing a balanced chemical equation:



Determine the number of moles of each reactant:

$$n_{\text{F}_2} = \frac{PV}{RT} = \frac{(1.00 \text{ atm})(2.00 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (293.15 \text{ K})} = 0.0831_4 \text{ mol F}_2$$

$$n_{\text{H}_2} = \frac{PV}{RT} = \frac{(1.50 \text{ atm})(1.50 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (253.15 \text{ K})} = 0.108_3 \text{ mol H}_2$$

Determine that the limiting reactant is F<sub>2</sub> gas using one of the two methods below:

$$0.0831_4 \text{ mol F}_2 \times \frac{1 \text{ mol H}_2}{1 \text{ mol F}_2} = 0.0831_4 \text{ mol H}_2$$

→ We have more H<sub>2</sub> than we need.

$$0.0831_4 \text{ mol F}_2 \times \frac{2 \text{ mol HF}}{1 \text{ mol F}_2} = 0.166_3 \text{ mol HF}$$

$$0.108_3 \text{ mol H}_2 \times \frac{2 \text{ mol HF}}{1 \text{ mol H}_2} = 0.216_6 \text{ mol HF}$$

→ F<sub>2</sub> produces less HF product.

Now determine how much HF gas can be produced from this amount of F<sub>2</sub> gas:

$$0.0831_4 \text{ mol F}_2 \times \frac{2 \text{ mol HF}}{1 \text{ mol F}_2} \times \frac{20.01 \text{ g HF}}{1 \text{ mol HF}} = \mathbf{3.33 \text{ g HF}}$$