

EXAM 2

Review Session

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CHEMISTRY 161
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www.mioy.org/chem161

OUTLINE

1. Concentration & Dilutions
2. Electrolytes, Salts, & Solubility Rules
3. Precipitation Reactions & Acid-Base
4. Redox Reactions
5. Enthalpy, Heat, & Heating Curves
6. Reaction Enthalpies & Hess's Law
7. Ideal Gas Law
8. Partial Pressure
9. Gas: Speed, Diffusion, & Effusion

Different Units of Concentration

- Molarity: $\frac{\text{moles of solute (mol)}}{\text{Volume solution (L)}} = \frac{\text{mol}}{\text{L}} = \text{M}$

- PPM: $1 \text{ ppm} = \frac{1 \text{ mg solute}}{1 \text{ kg solution}}$

- PPB: $1 \text{ ppm} = \frac{1 \mu\text{g solute}}{1 \text{ kg solution}}$

*CAN YOU CONVERT
BETWEEN THESE
UNITS?*

The solution is usually water (H₂O), so remember density is:

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

Determine the concentration of each solution in the correct units:

	INFORMATION	[CONCENTRATION]	M	PPM
A)	64.7 g LiCl in 250.0 mL water	[LiCl]		
B)	0.200 mol CaCl ₂ in 2.00 L water	[Cl ⁻]		
C)	6.42 × 10 ⁻² mM KNO ₃	[KNO ₃]		
D)	0.200 mol Na ₂ CO ₃ in 200.0 g water	[Na ₂ CO ₃]		
		[Na ⁺]		
		[CO ₃ ²⁻]		
		[ions]		n/a
E)	1.25 M HNO ₃	[H ⁺]		

I'll do parts of D for you: 0.200 mol Na₂CO₃ in 200.0 g water

To find the molarity (M) of [Na₂CO₃], I need moles of Na₂CO₃ and the volume of water in liters:

$$\begin{aligned}
 [\text{Na}_2\text{CO}_3] &= \frac{\# \text{ mol Na}_2\text{CO}_3}{\text{Volume (L)}} \\
 &= \frac{0.200 \text{ mol Na}_2\text{CO}_3}{200.0 \text{ g H}_2\text{O} \times \frac{1 \text{ mL H}_2\text{O}}{1 \text{ g H}_2\text{O}} \times \frac{1 \text{ L}}{1000 \text{ mL}}} \\
 [\text{Na}_2\text{CO}_3] &= 1.00 \text{ M}
 \end{aligned}$$

REMEMBER

The solution is usually water (H₂O), 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 convert this concentration from M to PPM, I need the molar mass of Na₂CO₃ = 105.99 g/mol:

$$\frac{1.00 \text{ mol Na}_2\text{CO}_3}{1.00 \text{ L H}_2\text{O}} \times \frac{105.99 \text{ g Na}_2\text{CO}_3}{1 \text{ mol Na}_2\text{CO}_3} \times \frac{1000 \text{ mg}}{1 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mL H}_2\text{O}}{1 \text{ g H}_2\text{O}} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 106000 \text{ ppm Na}_2\text{CO}_3$$

To find the concentration of [Na⁺] in M and PPM:

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

$$\frac{2.00 \text{ mol Na}^+}{1.00 \text{ L H}_2\text{O}} \times \frac{22.99 \text{ g Na}^+}{1 \text{ mol Na}^+} \times \frac{1000 \text{ mg}}{1 \text{ g}} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mL H}_2\text{O}}{1 \text{ g H}_2\text{O}} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 4.60 \times 10^4 \text{ ppm Na}^+$$

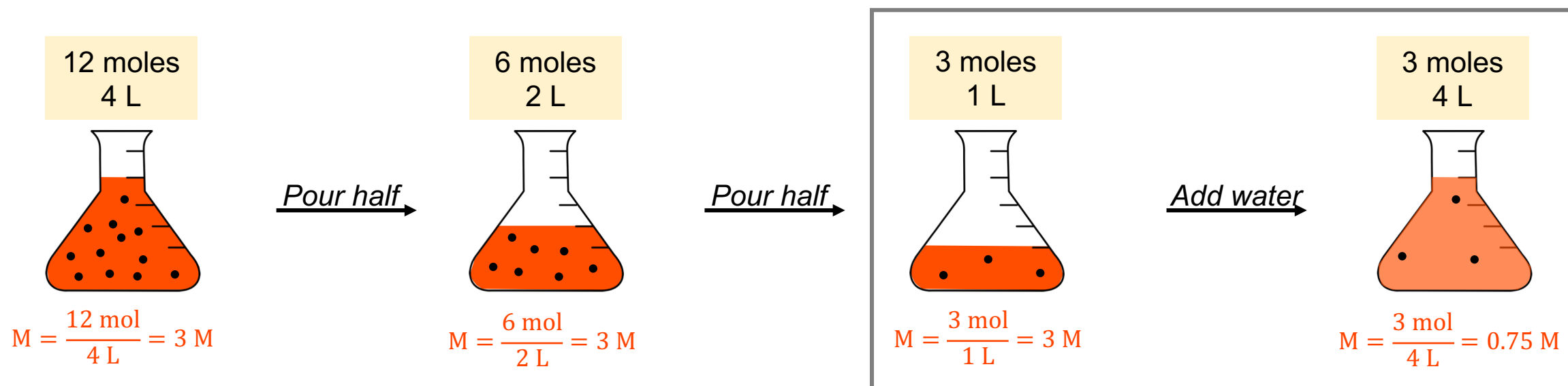
Determine the concentration of each solution in the correct units:

	INFORMATION	[CONCENTRATION]	M	PPM
A)	64.7 g LiCl in 250.0 mL water	[LiCl]	6.11 M	259000 ppm
B)	0.200 mol CaCl ₂ in 2.00 L water	[Cl ⁻]	0.200 M	7090 ppm
C)	6.42×10^{-2} mM KNO ₃	[KNO ₃]	6.42×10^{-5} M	6.49 ppm
D)	0.200 mol Na ₂ CO ₃ in 200.0 g water	[Na ₂ CO ₃]	1.00 M	106000 ppm
		[Na ⁺]	2.00 M	4.60×10^4 ppm
		[CO ₃ ²⁻]	1.00 M	6.00×10^4 ppm
		[ions]	3.00 M	n/a
E)	1.25 M HNO ₃	[H ⁺]	1.25 M	1260 ppm

Dilutions: usually adding water/changing volume

$$\text{Concentration} = \frac{\text{moles of solute}}{\text{Volume (L) of solution}} \quad ; \quad M = \frac{\text{mol}}{\text{L}}$$

Think about what concentration means before getting into the math.



Each black dot represents a mole (the quantity/amount)

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

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

First, find how many moles of Cl^- ions you have in 2.50 mL of the 0.559 M 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 Cl^- solution containing this many moles of Cl^- ions:

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

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

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

\therefore This is total volume! So subtract to find the amount of water added: $V_{\text{added}} = 12.05 \text{ mL} - 2.50 \text{ mL} = 9.55 \text{ mL water added}$.

ALTERNATIVE SOLUTION:

You can use the following equation:

$$M_1 V_1 = M_2 V_2$$

$$(0.559 \text{ M})(2.50 \text{ mL}) = (0.116 \text{ M})V_2$$

$$V_2 = 12.0_5 \text{ mL}$$

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.

Salts & Solubility Rules

A salt is an ionic compound: metal + nonmetal

Some salts are **soluble** in water (*aqueous*, *aq* = dissociates into ions in water).

Some salts are **insoluble** in water (precipitate, solid, *s*).

MEMORIZE SOLUBILITY RULES:

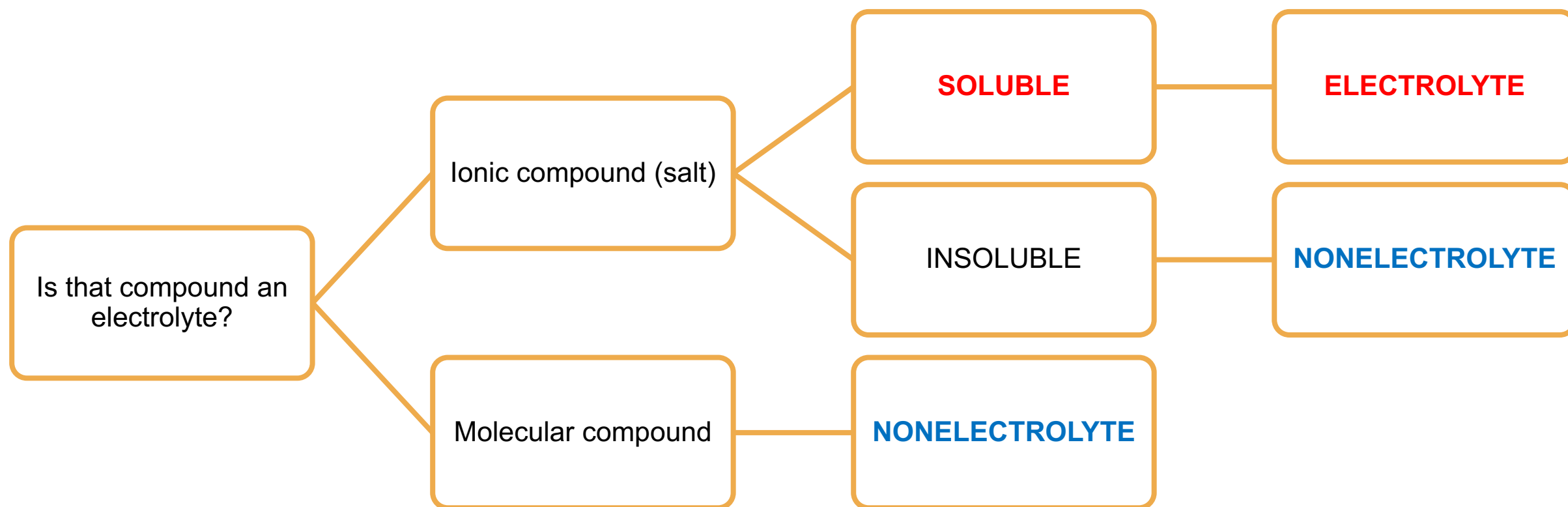
→ Table 4.4 or 4.5

You should be able to quickly identify the ions that comprise a salt!

		<i>Exceptions</i>	
		Group 1 cations	
SOLUBLE	NH ₄ ⁺		
	NO ₃ ⁻		
	CH ₃ COO ⁻		
	Cl ⁻ , Br ⁻ , I ⁻	Ag ⁺ , Hg ₂ ²⁺ , Pb ²⁺ , Cu ⁺	
	SO ₄ ²⁻	Hg ₂ ²⁺ , Pb ²⁺ , Ba ²⁺ , Ca ²⁺ , Sr ²⁺ ,	
INSOLUBLE	OH ⁻	Group 1 cations, Ba ²⁺ , Ca ²⁺ , Sr ²⁺ , NH ₄ ⁺	
	S ²⁻	Group 1 cations, Ba ²⁺ , Ca ²⁺ , Sr ²⁺ , NH ₄ ⁺	
	CO ₃ ²⁻ , PO ₄ ³⁻ , F ⁻	Group 1 cations, NH ₄ ⁺	

Electrolytes or Not?

Electrolytes are solutions that contain ions.



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

0.25 M NaBr

0.25 M Na₂SO₄

0.25 M Na₃PO₄

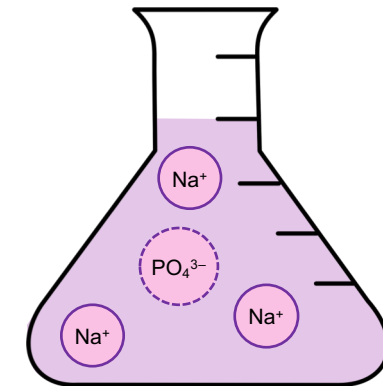
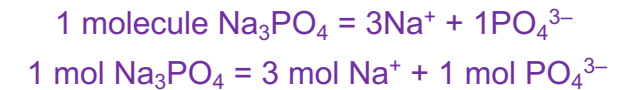
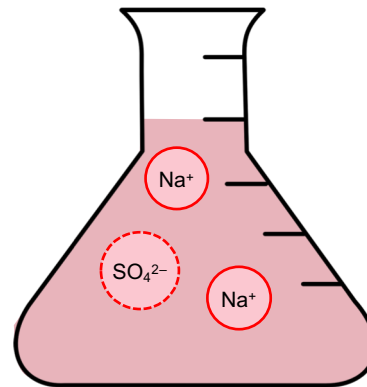
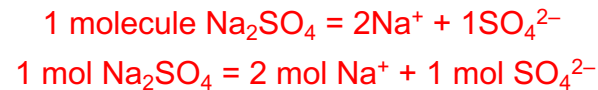
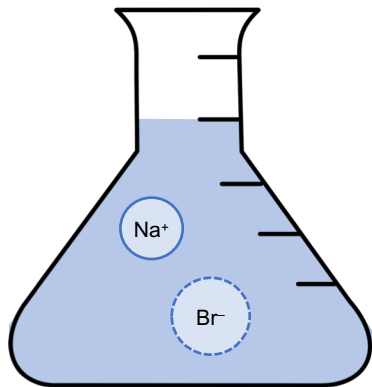
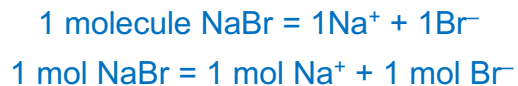
Which of the following has the greatest concentration of dissolved ions?

0.25 M NaBr

0.25 M Na₂SO₄0.25 M Na₃PO₄

These are all soluble salts!

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



Understand that the concentration of *ions* would be:

$$\begin{aligned} 1 \text{ NaBr} &: 2 \text{ ions} \\ [\text{ions}] &= 2 \times 0.25 \text{ M} = 0.50 \text{ M} \end{aligned}$$

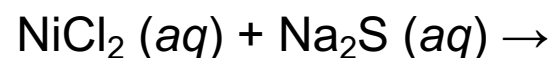
$$\begin{aligned} 1 \text{ Na}_2\text{SO}_4 &: 3 \text{ ions} \\ [\text{ions}] &= 3 \times 0.25 \text{ M} = 0.75 \text{ M} \end{aligned}$$

$$\begin{aligned} 1 \text{ Na}_3\text{PO}_4 &: 4 \text{ ions} \\ [\text{ions}] &= 4 \times 0.25 \text{ M} = 1.00 \text{ M} \end{aligned}$$

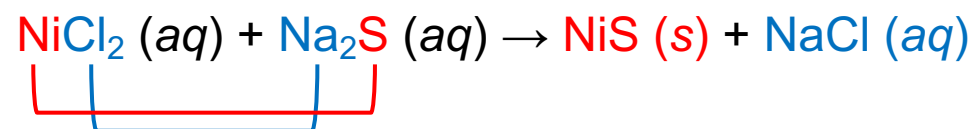
Guided Introduction: Precipitation Reaction

Say we mix an aqueous solution of nickel(II) chloride with an aqueous solution of sodium sulfide.

We can write the reactants of this chemical reaction:



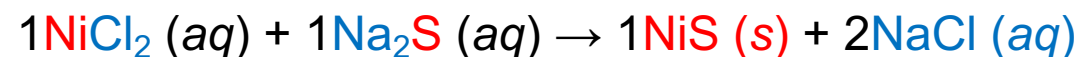
We can predict the products by exchanging the ions:



Determine the products first.

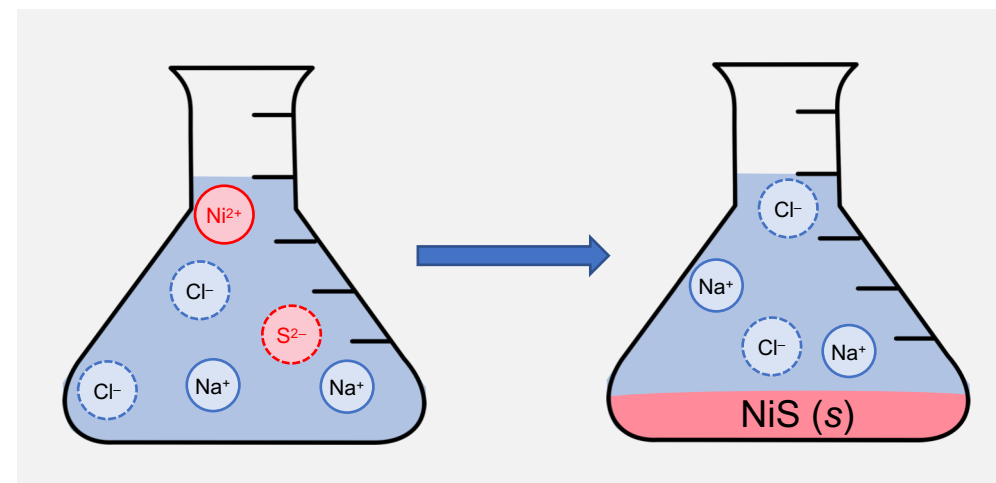
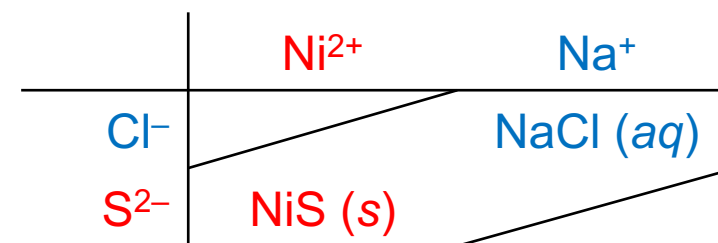
Then determine if each new product is soluble (*aq*) or insoluble (*s*)!

Don't forget to balance the chemical reaction!



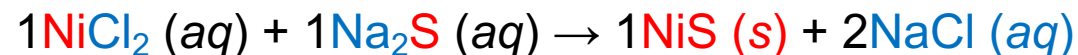
Always go back to the microscopic picture of what's happening!

→ **NiS** precipitate (solid) forms & other ions float in solution



Guided Introduction: Precipitation Reaction

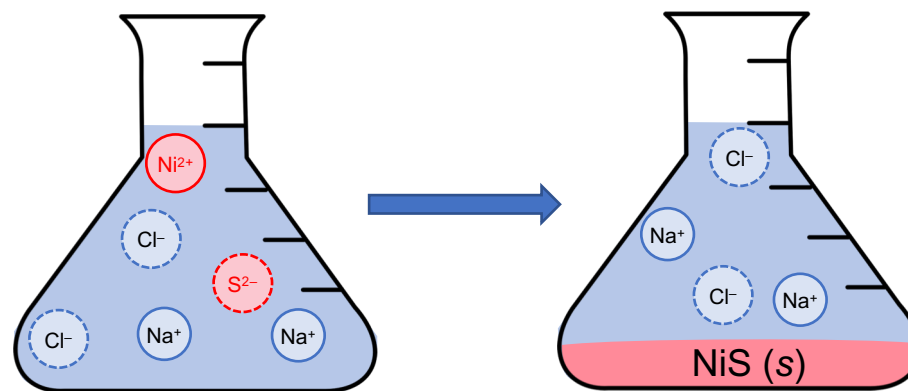
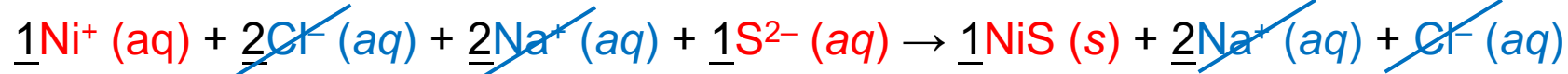
This is called a molecular equation because we keep the salts as neutral “molecules”:



But we know that soluble salts (aqueous) exist as dissociated ions in solution!

It would be more accurate to dissociate the *soluble* salts.

We can transform the molecular equation into an complete ionic equation:

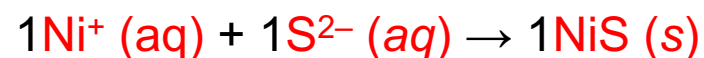


SPECTATOR IONS

You may notice that some of these ions (Na^+ and Cl^-) don't actually chemically react.

We can ignore these!

We can write a net ionic equation by eliminating all the **spectator ions**:



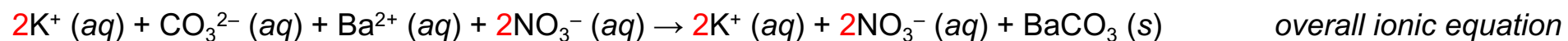
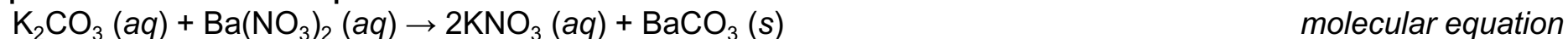
For each reaction, write out the molecular, overall ionic, and net ionic equations:

1. An aqueous solution of potassium carbonate is added to a barium nitrate solution.

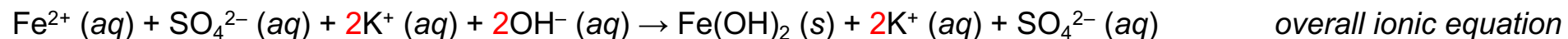
2. An iron(II) sulfate solution is mixed with a potassium hydroxide solution.

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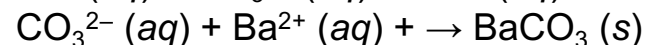
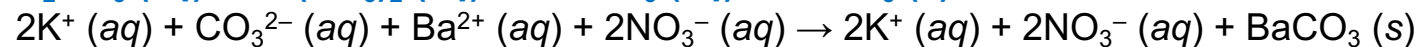
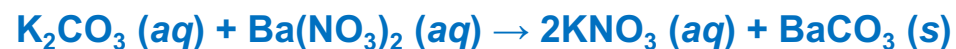
Precipitation Reaction Stoichiometry

A 250.0 mL solution of a 0.15 M potassium carbonate is added to 175 mL of a 0.200 M barium nitrate solution.
What mass of precipitate is formed?

Precipitation Reaction Stoichiometry

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From 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 and determine which is the limiting reactant:

$$0.15 \text{ M K}_2\text{CO}_3 = \frac{x \text{ mol}}{0.2500 \text{ L}}$$

$$x = 0.0375 \text{ mol K}_2\text{CO}_3$$

$$0.200 \text{ M Ba}(\text{NO}_3)_2 = \frac{y \text{ mol}}{0.175 \text{ L}}$$

$$y = 0.0350 \text{ mol Ba}(\text{NO}_3)_2$$

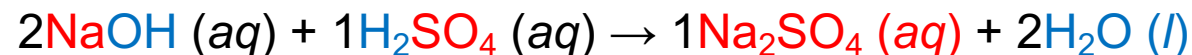
∴ The limiting reactant is $\text{Ba}(\text{NO}_3)_2$. *Why?*

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

$$0.0350 \text{ mol Ba}(\text{NO}_3)_2 \times \frac{1 \text{ mol BaCO}_3}{1 \text{ mol Ba}(\text{NO}_3)_2} \times \frac{137.31 \text{ g BaCO}_3}{1 \text{ mol BaCO}_3} = 6.91 \text{ g BaCO}_3$$

Guided Introduction: Acid-Base Titrations

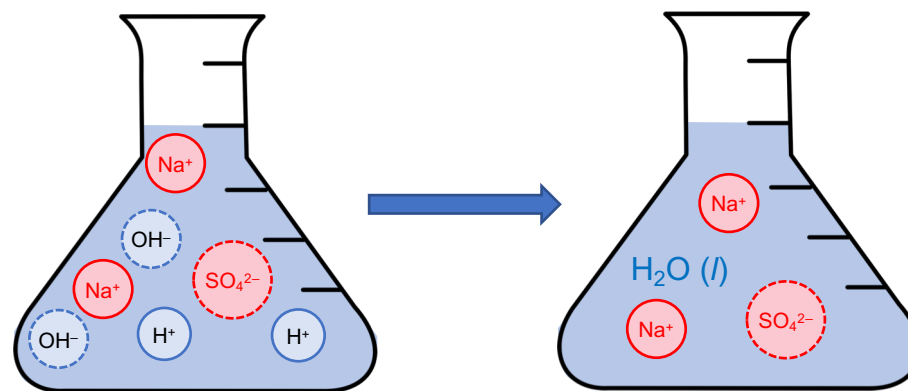
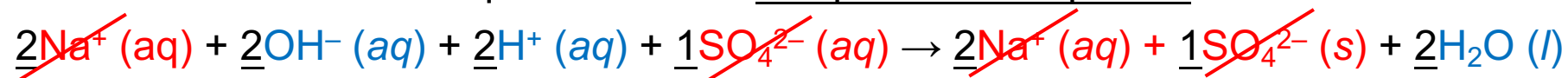
Think of acid-base reactions special precipitation reactions called a neutralization reaction:



But we know that soluble salts (aqueous) exist as dissociated ions in solution!

It would be more accurate to dissociate the *soluble* salts.

We can transform the molecular equation into an complete ionic equation:



SPECTATOR IONS

You may notice that some of these ions (Na^+ and SO_4^{2-}) don't actually chemically react.

We can ignore these!

We can write a net ionic equation by eliminating all the **spectator ions**:



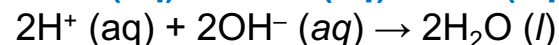
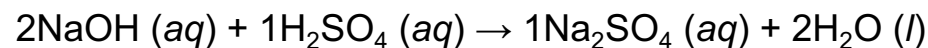
Acid-Base Stoichiometry

What volume of a 0.50 M NaOH solution would be required to neutralize 40.0 mL of a 0.400 M H₂SO₄ solution?

Acid-Base Stoichiometry

What volume of a 0.50 M NaOH solution would be required to neutralize 40.0 mL of a 0.400 M H₂SO₄ solution?

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



molecular equation

overall ionic equation

net ionic equation

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

First, determine the number of moles of the H⁺:

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

$$x = 0.0160 \text{ mol H}_2\text{SO}_4$$

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

To neutralize this many moles of H⁺, we need an equivalent number of moles of OH⁻:

$$n_{\text{OH}^-} = 0.0320 \text{ mol OH}^- = n_{\text{NaOH}}$$

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

$$0.50 \text{ M NaOH} = \frac{0.0320 \text{ mol NaOH}}{V}$$

$$V = 0.0640 \text{ L}$$

Oxidation States/Numbers

The oxidation state is a hypothetical charge of an element IF it were 100% ionic.

Take-home message: Treat oxidation number/state like a charge.

Some general rules:

- | | |
|---|--|
| 1. Pure elements have oxidation number = 0 | Ex) Br ₂ : each Br is 0
Ex) Mg: Mg is 0 |
| 2. Hydrogen usually has an oxidation number = +1 | Ex) H ₂ O; CH ₄ |
| 3. Oxygen usually has an oxidation number = -1 | Ex) H ₂ O; CO ₂ |
| 4. Everything else: follow the rules for ionic charges! | Ex) NaCl: Na = +1; Cl = -1
Ex) CF ₄ : C = +4; F = -1
Ex) CO ₂ : C = +4; O = -2 |

Some reactions involve a change in the oxidation state of an atom/element!

For each flask, give the oxidation state of each element:

A

0.010 mol SF₆
{146.1 g/mol}
300 K

B

0.5 mol CO₂
{44.0 g/mol}
600 K

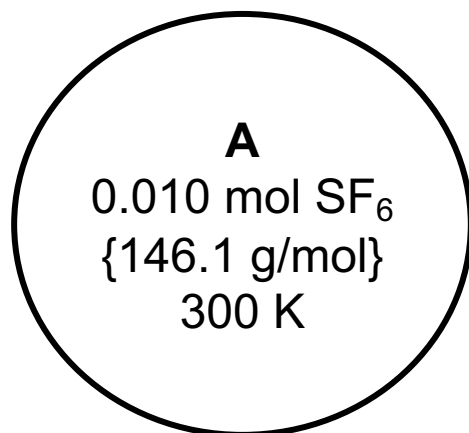
C

3.0 atm F₂
{38.0 g/mol}
273 K

D

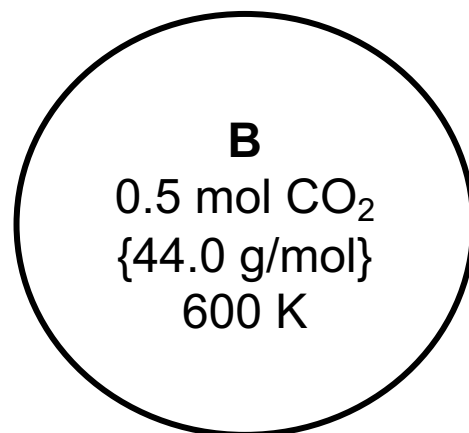
1.0 atm NO
{30.0 g/mol}
200 K

For each flask, give the oxidation state of each element:



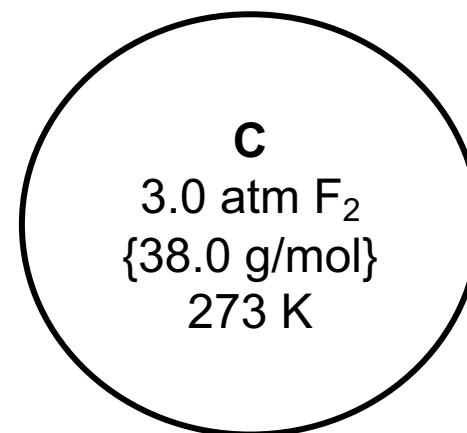
$$\text{S} = +6$$

$$\text{F} = -1$$

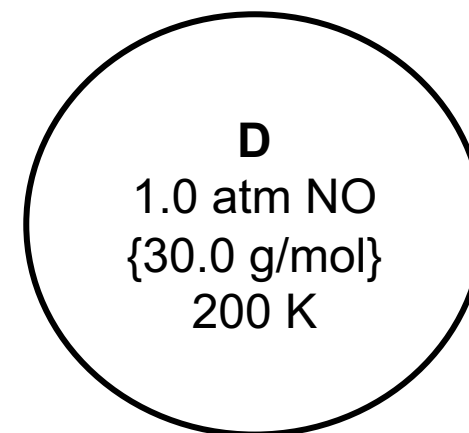


$$\text{C} = +4$$

$$\text{O} = -2$$



$$\text{F} = 0$$



$$\text{N} = +2$$

$$\text{O} = -2$$

What is a redox reaction?

A reaction in which the oxidation states of some elements change.

There are two types:

1. **Reduction**: Gain of an electron; becomes more negatively charged
2. **Oxidation**: Loss of an electron; becomes less negatively charged

LEO = Lose Electron Oxidation



OIL = Oxidation Is Losing

says



GER = Gain Electron Reduction

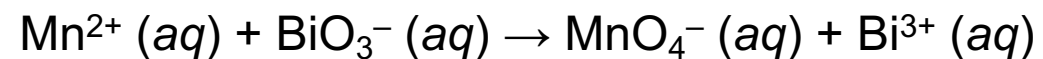


RIG = Reduction Is Gaining



These are called half-reactions because we typically have a both reactions occurring, or a **reduction-oxidation** (redox) reaction.

Assign the oxidation states for each element in the redox reaction below:



Mn²⁺ :

Mn in MnO₄⁻ :

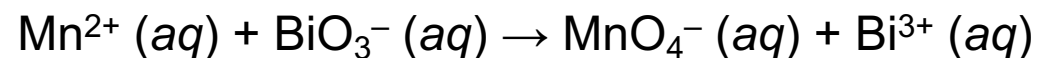
Bi in BiO₃⁻ :

Bi³⁺ :

O in BiO₃⁻ :

O in MnO₄⁻ :

Assign the oxidation states for each element in the redox reaction below:



$\text{Mn}^{2+} : +2$

$\text{Mn in MnO}_4^- : +7$

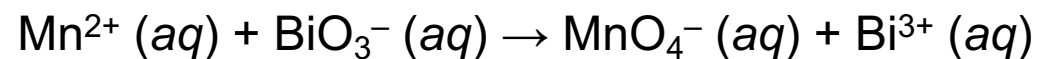
$\text{Bi in BiO}_3^- : +5$

$\text{Bi}^{3+} : +3$

$\text{O in BiO}_3^- : -2$

$\text{O in MnO}_4^- : -2$

Balance the following reaction using the half-reactions method in acidic solution:



$$\text{Mn}^{2+} : +2$$

$$\text{Mn in MnO}_4^- : +7$$

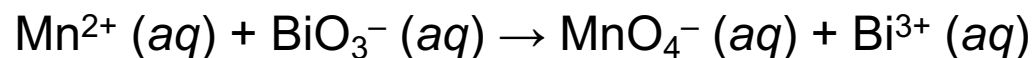
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$$\text{Mn}^{2+} : +2$$

$$\text{Mn in MnO}_4^- : +7$$

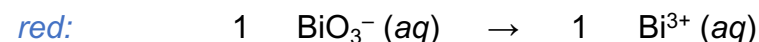
$$\text{Bi in BiO}_3^- : +5$$

$$\text{Bi}^{3+} : +3$$

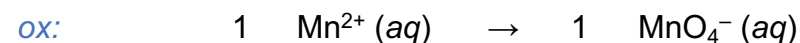
$$\text{O in BiO}_3^- : -2$$

$$\text{O in MnO}_4^- : -2$$

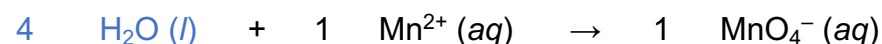
1. Separate half-reactions:



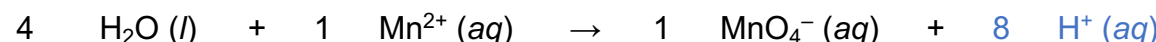
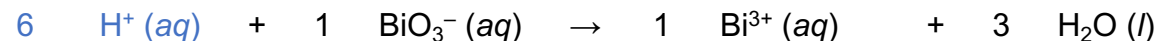
2. Balance heavy atoms:



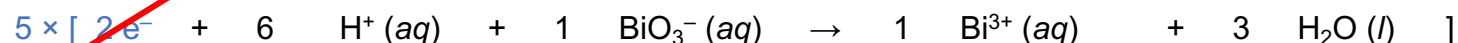
3. Balance O with H₂O:



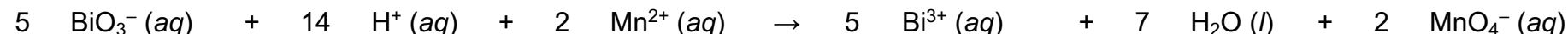
4. Balance H with H⁺:



5. Balance charge with e⁻:



6. Add the reactions:



Exothermic vs. Endothermic

Because of the law of conservation of energy (energy cannot be created or destroyed), any energy *lost* from the system is energy *gained* by the surroundings.

$$q_{\text{surr}} = -q_{\text{sys}}$$

Because of the law of conservation of energy (energy cannot be created or destroyed), any energy *gained* by the system is energy *lost* from the surroundings.

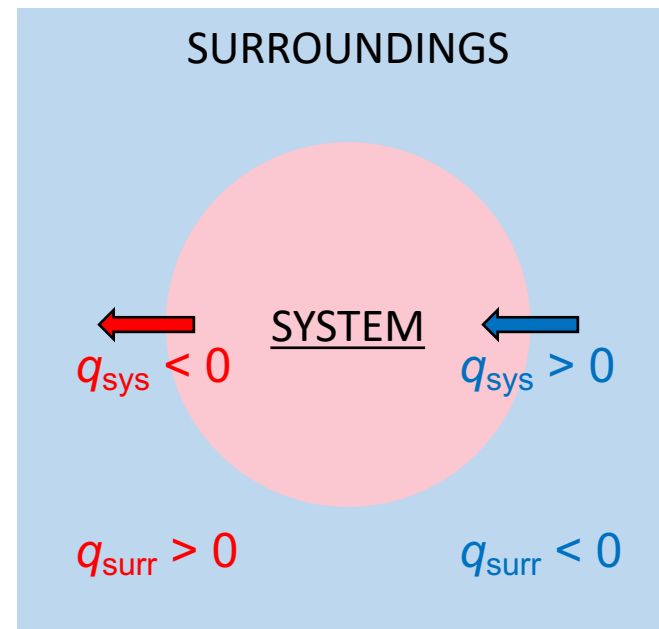
$$q_{\text{surr}} = -q_{\text{sys}}$$

Q1: Is this endothermic or exothermic relative to the system?

A1: Exothermic.

Q2: Will the system be warmer or cooler than it was?

A2: Cooler, lost energy.



Q1: Is this endothermic or exothermic relative to the system?

A1: Endothermic.

Q2: Will the system be warmer or cooler than it was?

A2: Warmer, gained energy.

How do we measure enthalpy of a reaction (ΔH_{rxn})? Calorimetry.

Typically, we don't measure the energy of a system directly (it's hard to do so).

Usually, we measure the change in the energy (or heat) as the system undergoes a change, like a chemical reaction. This is called the change in enthalpy or heat of a reaction, denoted ΔH .

Also, we usually measure the surroundings and not the system itself.

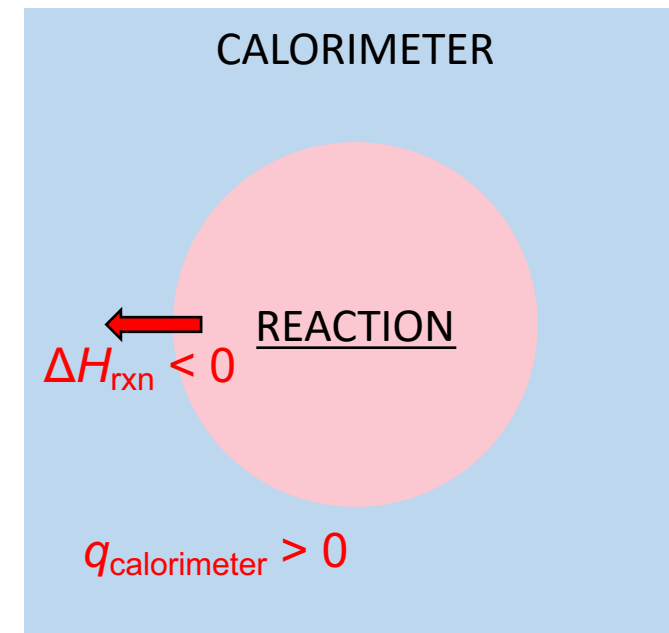
Remember, that $q_{\text{surr}} = -q_{\text{sys}}$.

More specifically, we use a calorimeter as the surroundings.

And our system is a particular reaction taking place.

A calorimeter is a well-insulated water bath, such that:

$$\Delta H_{\text{rxn}} = -q_{\text{calorimeter}}$$



In a calorimeter at constant pressure and 25.2 °C, you mix solutions of potassium carbonate and barium nitrate, which releases $\Delta H_{\text{rxn}} = -196 \text{ kJ}$. If the total volume of the solution is 1.2 L, what is the final temperature of the solution after the reaction is complete?

The molar heat capacity of liquid water is $75.3 \frac{\text{J}}{\text{mol}\cdot^{\circ}\text{C}}$.

In a calorimeter at constant pressure and 25.2 °C, you mix solutions of potassium carbonate and barium nitrate, which releases $\Delta H_{\text{rxn}} = -196 \text{ kJ}$. If the total volume of the solution is 1.2 L, what is the final temperature of the solution after the reaction is complete?

The molar heat capacity of liquid water is $75.3 \frac{\text{J}}{\text{mol} \cdot ^\circ\text{C}}$.

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

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

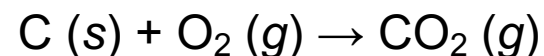
Now, figure out what happens to the water if we were to heat it with 196 kJ of heat:

$$196 \text{ kJ} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = \left(1.2 \text{ L H}_2\text{O} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1 \text{ g}}{1 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \right) \left(75.3 \frac{\text{J}}{\text{mol} \cdot ^\circ\text{C}} \right) (T_f - 25.2 ^\circ\text{C})$$

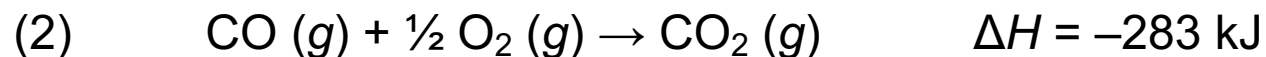
$T_f = 64 ^\circ\text{C}$

Hess's Law: Guided Example

Let's say you want to calculate the heat of reaction for the following reaction:

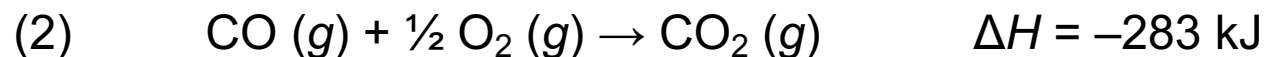


You don't want to carry out the actual reaction, but you have access to the heats of reactions of some related reactions:



Since enthalpy is a state function, if you add up (in some manner) these two reactions to get the reaction you are interested in, then you can also get the heat of reaction.

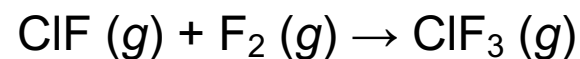
Let's simply add reactions (1) and (2) to get reaction (3) below:



Reaction (3) is exactly the reaction we are interested in. Now let's add the numbers in!

Another Example

Let's say you want to calculate the heat of reaction for the following reaction:

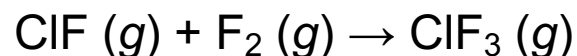


You have access to the heats of reactions of some related reactions:

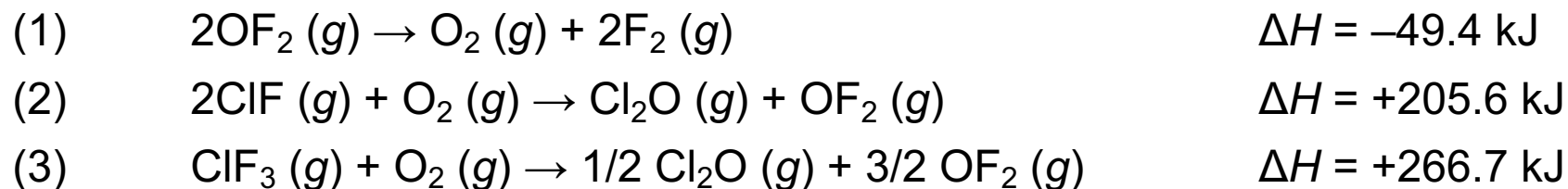
- | | | |
|-----|---|--------------------------------|
| (1) | $2\text{OF}_2 (g) \rightarrow \text{O}_2 (g) + 2\text{F}_2 (g)$ | $\Delta H = -49.4 \text{ kJ}$ |
| (2) | $2\text{ClF} (g) + \text{O}_2 (g) \rightarrow \text{Cl}_2\text{O} (g) + \text{OF}_2 (g)$ | $\Delta H = +205.6 \text{ kJ}$ |
| (3) | $\text{ClF}_3 (g) + \text{O}_2 (g) \rightarrow 1/2 \text{Cl}_2\text{O} (g) + 3/2 \text{OF}_2 (g)$ | $\Delta H = +266.7 \text{ kJ}$ |

Another Example

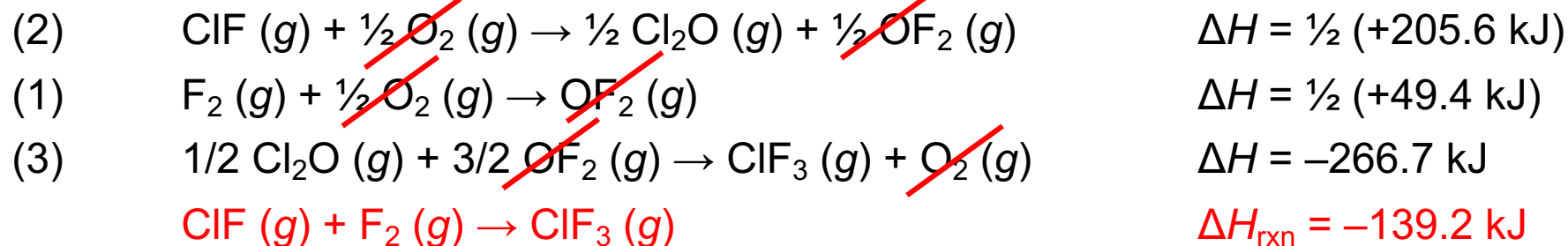
Let's say you want to calculate the heat of reaction for the following reaction:



You have access to the heats of reactions of some related reactions:



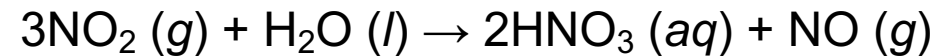
Add up reactions (2), reverse of (1), and reverse of (3):



Note: Flip sign when you invert reaction!

Heats of Formation (ΔH_f)

Let's say we want to calculate the standard heat of reaction for the following reaction:



Instead of Hess's Law, you could also use the standard heats of formations (ΔH_f°) to calculate ΔH_{rxn} .

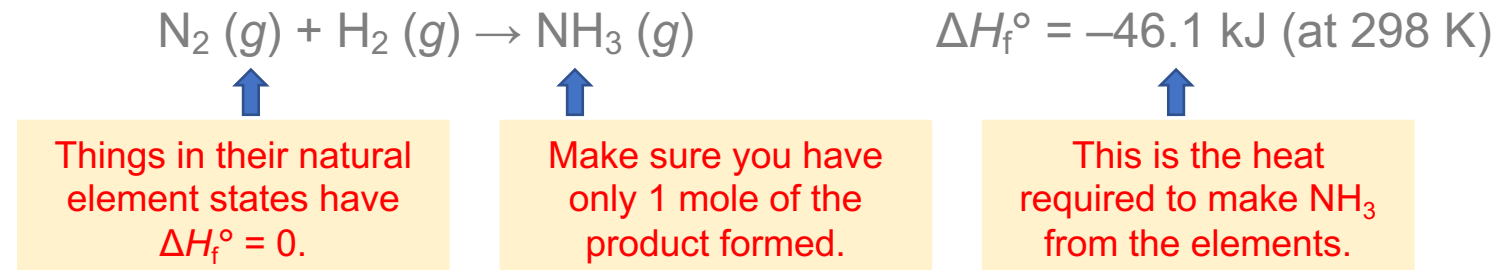
Q: What is a heat of formation anyway?

A: *It's the amount of heat required to form 1 mole of a substance from its elements in standard states.*

Q: What is a standard state?

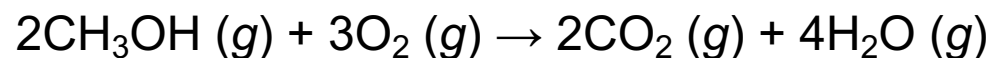
A: *1 bar (~1 atm) of pressure and 25 °C. (This is NOT STP for gases!)*

For example:



How to use heats of formation (ΔH_f)

A heat of combustion (ΔH_{comb}) is the same thing as the heat of reaction but pertains to a combustion reaction, such as:



PAY ATTENTION TO THE STOICHIOMETRIC AMOUNTS WHEN GIVEN HEATS OF REACTION!

Calculate the heat of this reaction (combustion):

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

$$\Delta H_{\text{rxn}}^{\circ} = n_{\text{CO}_2} \Delta H_{\text{CO}_2}^{\circ} + n_{\text{H}_2\text{O}} \Delta H_{\text{H}_2\text{O}}^{\circ} - n_{\text{CH}_3\text{OH}} \Delta H_{\text{CH}_3\text{OH}}^{\circ} - n_{\text{O}_2} \Delta H_{\text{O}_2}^{\circ}$$

$$\Delta H_{\text{rxn}}^{\circ} = 2 \text{ mol} \left(-393.5 \frac{\text{kJ}}{\text{mol}} \right) + 4 \text{ mol} \left(-241.8 \frac{\text{kJ}}{\text{mol}} \right) - 2 \text{ mol} \left(-200.7 \frac{\text{kJ}}{\text{mol}} \right) - 3 \left(0 \frac{\text{kJ}}{\text{mol}} \right)$$

$$\Delta H_{\text{rxn}}^{\circ} = -1353 \text{ kJ}$$

THIS IS HOW MUCH HEAT RELEASED IF 2 MOLES OF CH₃OH ARE BURNED!

When are gases ideal? When are they not ideal?

Gases behave ideally when:

- Pressure is low
- Temperature is high
- Volume of container is much larger than volume of actual gas particles

Gases behave non-ideally when:

- Pressure is high
- Temperature extremely low
- Volume of the gas becomes important
- Intermolecular or van der Waals interactions between the gas particles cannot be ignored

The Ideal Gas Law

$$PV = nRT$$

P = absolute pressure (units: atm)

V = volume of container (units: L)

n = number of moles of gas (units: mol)

T = absolute temperature (units: K)

R = universal gas constant $\left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)$

Reference Points for Gases

Standard Temperature and Pressure (STP)

- $P = 1 \text{ atm}$
- $T = 0 \text{ }^\circ\text{C} = 273 \text{ K}$

Molar Volume: volume occupied by one mole of any ideal gas at STP = 22.4 L

ALWAYS WORK IN ABSOLUTE TEMPERATURE SCALE (K)

ALWAYS WORK IN ABSOLUTE PRESSURE SCALE (ATM)

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?

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?

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:

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

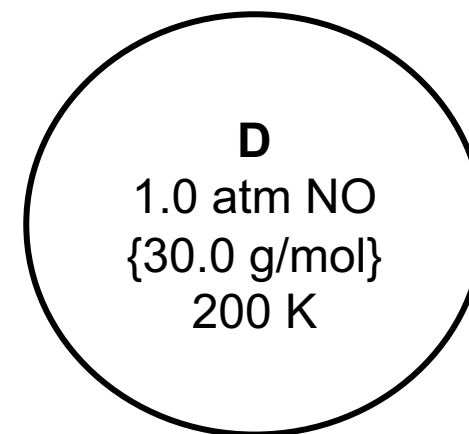
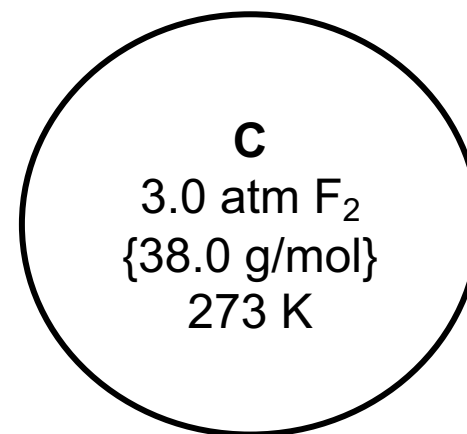
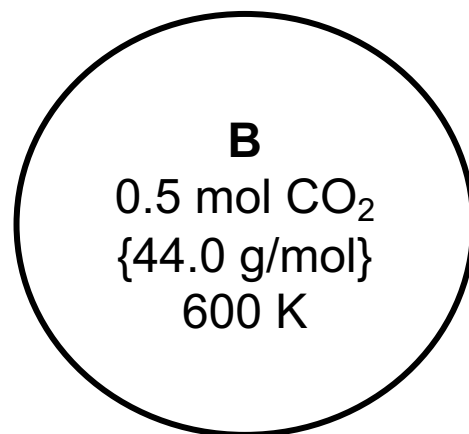
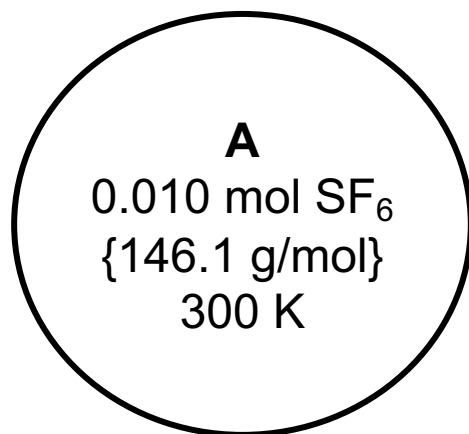
Alternatively, use the Ideal Gas Law: $PV = nRT$

$$n = \frac{PV}{RT}$$

$$= \frac{(1.0 \text{ atm}) \left(500.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \right)}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (273.15 \text{ K})}$$

$$n = 0.02232 \text{ mol gas}$$

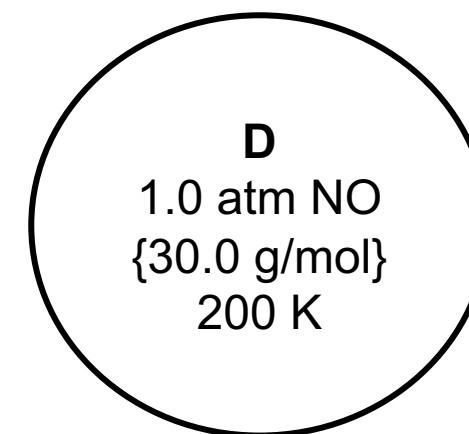
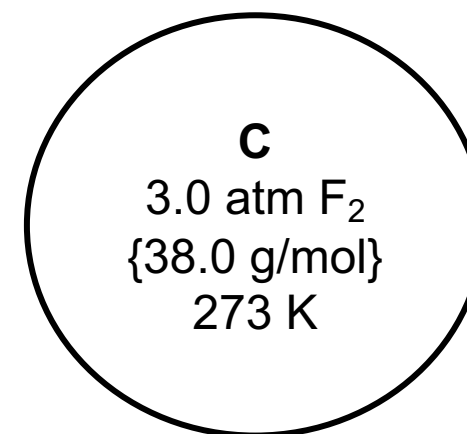
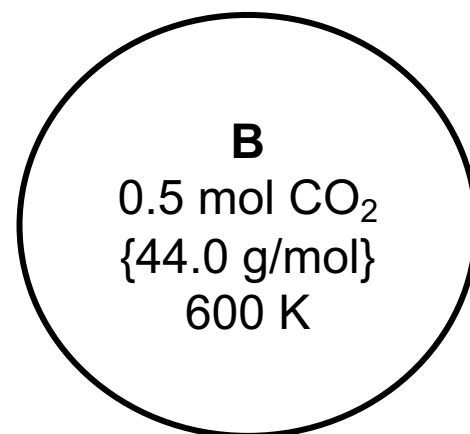
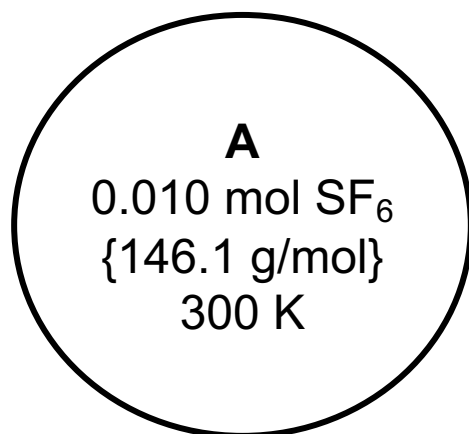
Consider the following four identical 1.0 L flasks:



A) Which flask has the largest number of atoms?

B) Which flask has the highest pressure?

Consider the following four identical 1.0 L flasks:



A) Which flask has the largest number of atoms?

$$n_{\text{SF}_6} = 0.010 \text{ mol SF}_6 \quad n_{\text{F}_2} = \frac{(3.0 \text{ atm})(1.0 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (273 \text{ K})} = 0.13 \text{ mol F}_2$$

$$n_{\text{CO}_2} = 0.5 \text{ mol CO}_2 \quad n_{\text{NO}} = \frac{(1.0 \text{ atm})(1.0 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (200 \text{ K})} = 0.061 \text{ mol NO}$$

To convert each of these mole amounts, we would do the following:

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

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

B) Which flask has the highest pressure?

$$P_{\text{SF}_6} = \frac{(0.010 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (300 \text{ K})}{1.0 \text{ L}} = 0.25 \text{ atm} \quad P_{\text{F}_2} = 3.0 \text{ atm}$$

$$P_{\text{CO}_2} = \frac{(0.5 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (600 \text{ K})}{1.0 \text{ L}} = 25 \text{ atm} \quad P_{\text{NO}} = 1.0 \text{ atm}$$

Given 6.0 L of nitrogen gas at $-25\text{ }^{\circ}\text{C}$, what volume will the nitrogen gas occupy at $72\text{ }^{\circ}\text{C}$? *Assume constant pressure.*

Given 6.0 L of nitrogen gas at $-25\text{ }^{\circ}\text{C}$, what volume will the nitrogen gas occupy at $72\text{ }^{\circ}\text{C}$? *Assume constant pressure.*

Start by collecting the information we know:

$$V_1 = 6.0\text{ L}$$

$$V_2 = ?$$

P and n are constant.

$$T_1 = -25\text{ }^{\circ}\text{C} = 248.15\text{ K}$$

$$T_2 = 72\text{ }^{\circ}\text{C} = 345.15\text{ K}$$

Set up the gas law and solve for final volume (V_2):

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$V_2 = \frac{V_1 T_2}{T_1}$$

$$= \frac{(6.0\text{ L})(345.15\text{ K})}{248.15\text{ K}}$$

$$V_2 = 8.3\text{ L}$$

How to calculate partial pressures of a mixture of gases?

If you have a mixture of gases, and each gas behaves ideally, then the gases do not interact with each other and you can apply the ideal gas law to each gas independently.

This means that you can simply add the partial pressure of each gas together to find the total pressure:

$$P_{\text{total}} = P_A + P_B + P_C \dots$$

$$n_{\text{total}} = n_A + n_B + n_C \dots$$

How to calculate partial pressure of gas A in a mixture:

- Determine moles of the gas
- Determine the mole ratio
- Multiply mole ratio by total pressure
- Ptotal can be found by using the ideal gas law

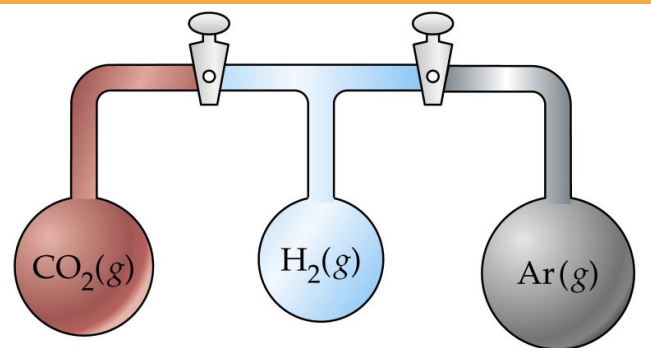
$$n_A$$

$$X_A = n_A/n_{\text{total}}$$

$$P_A = X_A P_{\text{total}}$$

$$P_{\text{total}} = n_{\text{total}}RT/V$$

- Or apply the ideal gas law on gas A only to find P_A



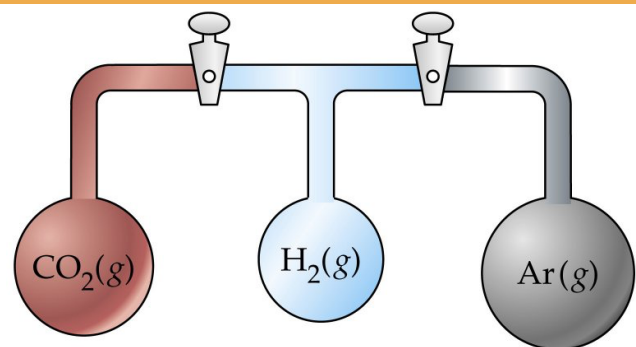
$P = 2.13 \text{ atm}$
 $V = 1.50 \text{ L}$

$P = 0.861 \text{ atm}$
 $V = 1.00 \text{ L}$

$P = 1.15 \text{ atm}$
 $V = 2.00 \text{ L}$

Imagine that both stopcocks were opened so that the gases mix at 298 K.

What is the partial pressure of each gas after opening?



Imagine that both stopcocks were opened so that the gases mix at 298 K.

What is the partial pressure of each gas after opening?

$$P = 2.13 \text{ atm} \quad P = 0.861 \text{ atm} \quad P = 1.15 \text{ atm}$$

$$V = 1.50 \text{ L} \quad V = 1.00 \text{ L} \quad V = 2.00 \text{ L}$$

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

$$n_{\text{CO}_2} = \frac{PV}{RT}$$

$$= \frac{(2.13 \text{ atm})(1.50 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}$$

$$n_{\text{CO}_2} = 0.130_7 \text{ mol}$$

$$n_{\text{H}_2} = \frac{PV}{RT}$$

$$= \frac{(0.861 \text{ atm})(1.00 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}$$

$$n_{\text{H}_2} = 0.0352_1 \text{ mol}$$

$$n_{\text{Ar}} = \frac{PV}{RT}$$

$$= \frac{(1.15 \text{ atm})(2.00 \text{ L})}{\left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}$$

$$n_{\text{Ar}} = 0.0940_5 \text{ mol}$$

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

$$P_{\text{CO}_2} = \frac{nRT}{V}$$

$$= \frac{(0.130_7 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{4.50 \text{ L}}$$

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

$$P_{\text{H}_2} = \frac{nRT}{V}$$

$$= \frac{(0.0352_1 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{4.50 \text{ L}}$$

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

$$P_{\text{Ar}} = \frac{nRT}{V}$$

$$= \frac{(0.0940_5 \text{ mol}) \left(0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (298 \text{ K})}{4.50 \text{ L}}$$

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

The root-mean-square speed (u_{rms})

ROOT-MEAN-SQUARE (RMS) SPEEDS

$$u_{\text{rms}} = \sqrt{\frac{3RT}{MM}} \quad ; \quad \frac{u_{\text{rms,A}}}{u_{\text{rms,B}}} = \sqrt{\frac{MM_B}{MM_A}} \quad ; \quad R = 8.314 \frac{\text{kg} \cdot \text{m}^2}{\text{mol} \cdot \text{K} \cdot \text{s}^2}$$

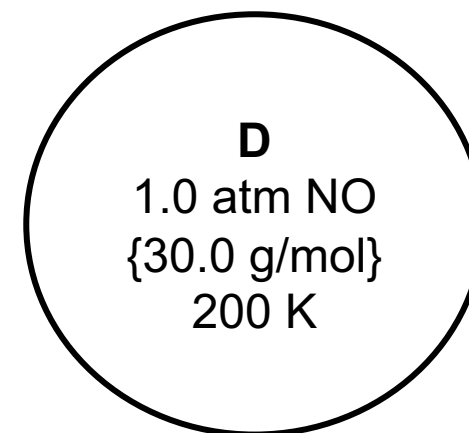
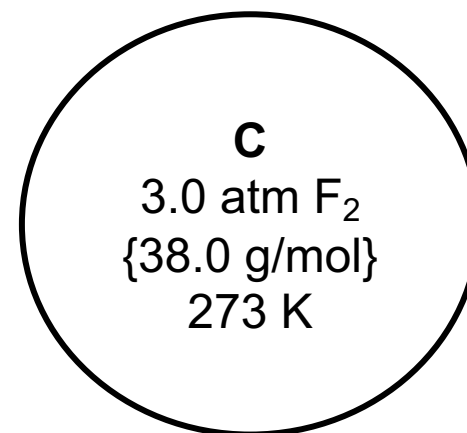
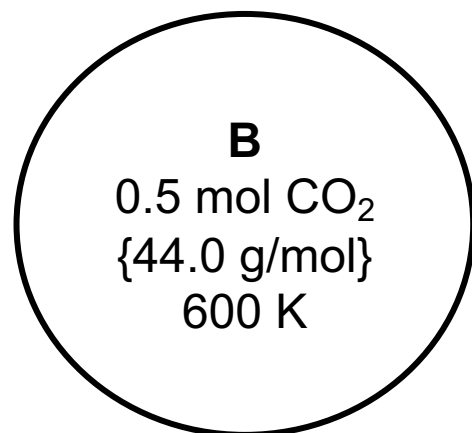
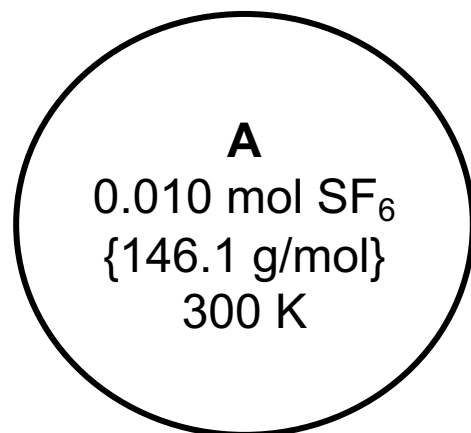
At a given temperature, gases with larger molar masses (MM) move more slowly.

EFFUSION RATE (r_{eff})

$$\frac{r_{\text{eff,A}}}{r_{\text{eff,B}}} = \frac{u_{\text{rms,A}}}{u_{\text{rms,B}}} = \sqrt{\frac{MM_B}{MM_A}}$$

At a given temperature, gases with smaller MMs have a higher chance of escaping.

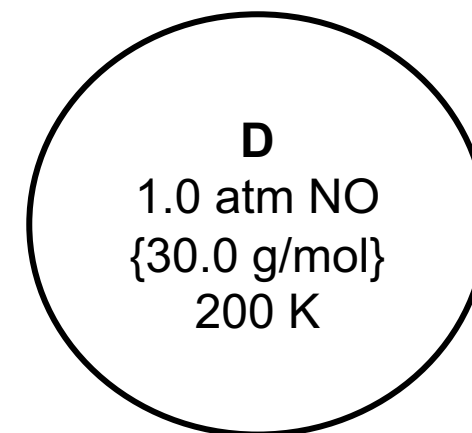
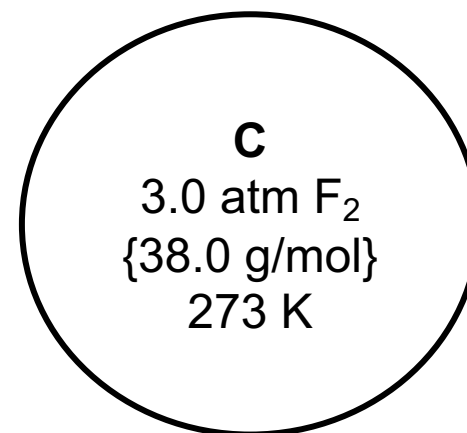
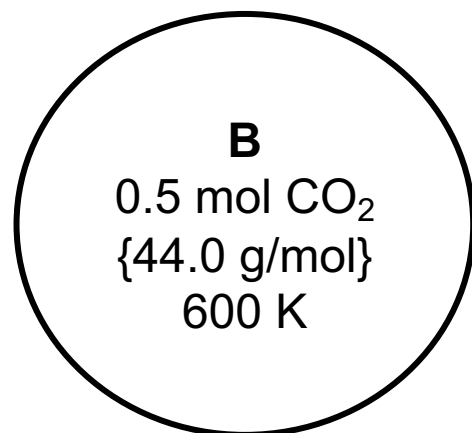
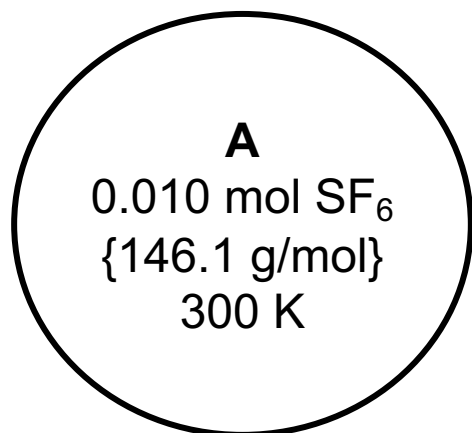
Consider the following four identical 1.0 L flasks:



A) In which flask will the diffusion of the gas be the slowest?

B) Which flask contains the gas with the smallest average velocity?

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A) In which flask will the diffusion of the gas be the slowest?

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

$\frac{T_A}{MM_A} = \frac{300}{146.1} = 2.1$	$\frac{T_C}{MM_C} = \frac{273}{38.0} = 7.2$
$\frac{T_B}{MM_B} = \frac{600}{44.0} = 13.6$	$\frac{T_D}{MM_D} = \frac{200}{30.0} = 6.7$

The larger the u_{rms} , the slower the rate of diffusion.

So, Flask A has the slowest rate of diffusion.

B) Which flask contains the gas with the smallest average velocity?

$$u_{\text{avg}} = \sqrt{\frac{8RT}{\pi \cdot MM}} \Rightarrow u_{\text{avg}} \propto \frac{T}{MM}$$

$\frac{T_A}{MM_A} = \frac{300}{146.1} = 2.1$	$\frac{T_C}{MM_C} = \frac{273}{38.0} = 7.2$
$\frac{T_B}{MM_B} = \frac{600}{44.0} = 13.6$	$\frac{T_D}{MM_D} = \frac{200}{30.0} = 6.7$

Flask A has the smallest average velocity.