



**Common Mistakes and Misconceptions in Chemistry:
How Not to Be an Easy Target For Exam-Writers Who
are Out to Stump You**



New Topics To Chemistry V


K_{sp}

1. Writing the wrong ionic equation for something like CaF_{2(s)}

CaF₂ forms when Ca gives its two valence electrons to different F atoms, each of which can only accommodate 1 more electron. This creates 1 Ca²⁺ ion and *two separate* F⁻ ions. So the equation becomes:



2. Placing solubility in the wrong place in ice chart and writing the wrong equilibrium expression.

	CaF _{2(s)}	 Ca ²⁺	+ 2F ⁻
I		0	0
C	Solubility goes here: example 2.05 X 10 ⁻⁴ mol/L	Stoichiometry applies 2.05 X 10 ⁻⁴ mol/L	Stoichiometry applies 2:1 4.10 X 10 ⁻⁴ mol/L
E		2.05 X 10 ⁻⁴ mol/L	4.10 X 10 ⁻⁴ mol/L

$$K_{sp} = [2.05 \times 10^{-4} \text{ mol/L}][4.10 \times 10^{-4} \text{ mol/L}]^2 = 3.45 \times 10^{-11}$$

Notice the solid does not go into the denominator; it's a constant!

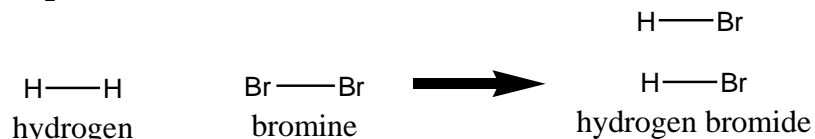
Estimating ΔH From Bond Energies

3. **Forgetting that the energy invested to break bonds is $\Delta H_{bb} = (+)$**

And that the energy released when bonds are formed is $\Delta H_{bf} = (-)$. The enthalpy is then obtained by adding the above two values.

Example:

Estimate the ΔH per mole of H_2 for the following reaction: $H_2 + Br_2 \rightarrow 2 HBr$



progress

ΔH_{bb}	ΔH_{bf}	$\Delta H = \Delta H_{bb} + \Delta H_{bf}$
$436 \text{ kJ} + 193 \text{ kJ} = 629 \text{ kJ}$	$-366 \text{ kJ} * 2 = -732 \text{ kJ}$	$-732 \text{ kJ} + 629 \text{ kJ} = -103 \text{ kJ}$

Table of Bond Energies to consult:

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Bond	Energy(kJ/mole)	Bond	Energy(kJ/mole)
C—H	413	C≡C	839
H—H	436	C—O	358
H—Br	366	C=O	745
Br—Br	193	C=C	607
O—H	460	O=O	498
C—C	340		

Gases

4. *Using n in PV = nRT for a Substance that is Not in the Gaseous State.*

Example: If 2.0 moles of H₂O₂ liquid decomposed and produced oxygen at 101.3 kPa at what temperature was the gas collected in a 2.0 L vessel?



Solution **PV = nRT only works for ideal gases, not for real gases or liquids or solids. So in this case, we cannot use 2.0 as a value for n. Instead, we have to consider the stoichiometric relationship between H₂O₂ and O₂. Since the ratio is 2 to 1, and if 2.0 moles of hydrogen peroxide reacted, only 1.0 mole of oxygen is produced.**



$$T = PV/(nR) = \frac{101.3 \text{ kPa}(2.0 \text{ L})}{1.0 \text{ mole} \left(\frac{8.31 \text{ L kPa}}{\text{K mole}} \right)} = 24 \text{ K}$$

5. *Unit-Butchery → Ignoring R's kPa L/(K mole)*

Example: What is the total mass of a 250.0 mL flask with a tare weight of 200.00 g if it contains hydrogen gas at 22.5 °C and at 888 mm of Hg? Note 101.3 kPa = 760 mm of Hg

Solution **Use the factor label method with the given conversions, and they can throw any gorilla-units at you, and you will not even flinch.**

$$n = PV/RT = \frac{888 \text{ mm Hg} \left[\frac{101.3 \text{ kPa}}{760 \text{ mm Hg}} \right] (0.250 \text{ L})}{\frac{8.31 \text{ L kPa}}{\text{K mole}} (22.5 + 273.15 \text{ K})} = 0.0120439 \dots \text{moles of H}_2$$

Now convert to grams and then add the tare weight of the flask:

0.0120439 moles (2.00 g of H₂/mole) = 0.0241 g

Total mass = 200.02 g with sig figs → 200. g

6. *Confusing an Ideal gas with a Real gas.*

Example: Four boxes, each at STP with a volume of 22.4 L contain Ar, O₂, He and HCl. Which is least likely to contain exactly 1.00 moles? Will it contain more than 6.02 X 10²³ molecules?

Solution HCl is the gas that will not act as an ideal gas because there will be some intermolecular attractions between the H of one molecule and the Cl of a neighbouring molecule. The attractions will cause one mole of HCl at STP to have a volume that is actually less than 22.4 L. So if we consider a 22.4 L box, it will contain proportionally more HCl molecules.

Not convinced? Let's say that one mole of HCl at STP = 22.0 L . Then:

22.4 L (1 mole HCl/22.0 L) (6.02 X10²³ molecules) > 6.02 X10²³ molecules.

7. *Applying Charles Law when Pressure is not Constant.*

Example: A steel container containing 20.0L of an ideal gas is cooled from 30.0 °C to 10.0 °C. What happens to the volume? Why? Does anything else change?

Solution The volume remains constant. It's an inflexible steel container. In order for the volume to shrink, you need a collapsible container like a piston and cylinder, or just a water-trapped gas within a glass cylinder. In our steel case, the molecules will slightly slow down, but they will still reach as far as the boundaries of their container. The collisions will be less frequent, and pressure will decrease. So, more specifically, this is actually a Gay Lussac problem, with $P_1/T_1 = P_2/T_2$.

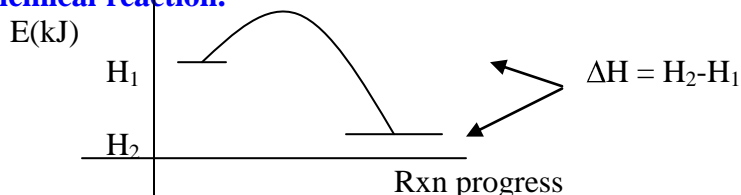
$$P_2 = (10.0 + 273.15)/(30.0 + 273.15) P_1 = 0.934 P_2.$$

Thermochemistry

8. *Confusing Enthalpy with Change in Enthalpy*

Example: Show the difference on a graph

Solution Enthalpy, H, is the total heat content of a substance, which includes the sum of its potential and kinetic energies. Enthalpy change or ΔH, is the difference in enthalpy between the products and reactants of a physical or chemical reaction.



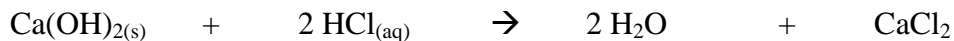
9. *Confusing Enthalpy Change with Molar Enthalpy*

Example: Given that the reaction $2A + B \rightarrow A_2B + 202 \text{ kJ}$, what is the molar enthalpy for A?

Although the molar enthalpy for B or A_2B is -202 kJ/mole , the molar enthalpy for A is, because of the coefficient in the equation, $-202 \text{ kJ}/2 \text{ moles} = -101 \text{ kJ/mole}$.

10. *Inserting the Wrong Mass into $Q = mc\Delta T$ for Neutralization Reactions and Using Wrong n*

Example: If adding 2.0 g of solid $\text{Ca}(\text{OH})_2$ to 100.0 mL of dilute aqueous HCl completely neutralized the acid while causing a $6.6 \text{ }^\circ\text{C}$ increase in temperature, what is the molar heat of neutralization of HCl?



What goes on in this reaction is that a small amount of base dissolved in relatively lots of water is neutralizing a small amount of acid, which is also dilute. Overall there is a negligible volume change. Basically the mass absorbing the heat (Q) of neutralization is $100.0 \text{ ml}(1.0\text{g/ml}) = 100.0 \text{ grams}$ of water, with a specific heat capacity of about $4.19 \text{ J}/(\text{g}^\circ\text{C})$, even though there is a bit of salt and water produced almost instantaneously.

Do not use 102.0 g! The 2.0 g of base is solid in this case.

$$Q = mc\Delta T = 100.0 \text{ g} [4.19 \text{ J}/(\text{g}^\circ\text{C})] (6.6 \text{ }^\circ\text{C}) = 2765.4 \text{ J}$$

But although the water absorbs the heat, the energy was released by the reaction between hydroxide and H^+ ions, so $\Delta H = -Q = -2765.4 \text{ J} = -2.7654 \text{ kJ}$

How much of this heat is released per mole of HCl that react? Well we need moles of HCl = twice the moles of calcium hydroxide(see the ratio in the equation). $2 \text{ moles HCl/mole of Ca}(\text{OH})_2 (2.0 \text{ g}) (\text{mole of Ca}(\text{OH})_2/74.0 \text{ g}) = 0.0541 \text{ moles} = n$

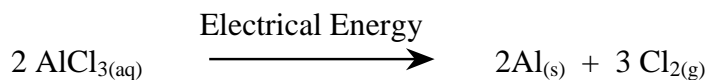
$$\Delta H/n = -2.7654 \text{ kJ}/0.0541 \text{ moles} = -51 \text{ kJ/mole of HCl}$$

Rates

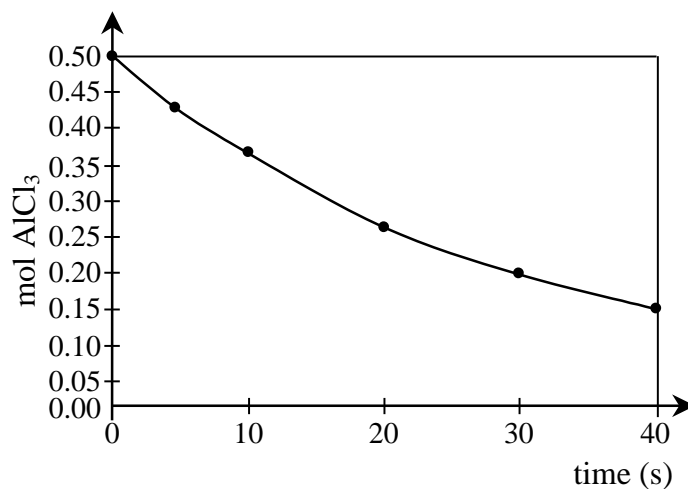
11. *Confusing Leftover Moles of a Reactant With the Amount that is Actually Reacting and/or With What is Actually being Produced*

Example:

A student was studying the electrolysis of 1 L of $\text{AlCl}_{3(\text{aq})}$ according to:



Moles of AlCl_3 versus time



Using the graph above, calculate the average rate of formation of chlorine gas from 10 s to 20 s.

Construct the following table, noting the ratio from the equation:

Time(seconds)	Remaining Moles of AlCl_3	Moles of AlCl_3 that React	Moles of Cl_2 Produced
0	0.50	$0.50 - 0.50 = 0$	0
10	0.37	$0.50 - 0.37 = 0.13$	$(3/2)0.13 = 0.195$
20	0.27	$0.50 - 0.27 = 0.23$	$(3/2)0.23 = 0.3495$

Avg rate = $(0.3495 - 0.195) / (20 - 10) = 0.015$ moles Cl_2 per second.

12. **Average Rate: Do not add up and divide by 2.** On a curve, rate changes continuously. So to get an approximate average, we find the slope of an imaginary straight line running through the endpoints.

Example: After 2.0 minutes we produced 10 moles of H₂. After 8.0 minutes we produced 24 moles. What is the average rate of production of hydrogen gas.

Solution: avg rate = $\Delta n/\Delta t = (24 - 10) \text{ moles} / (8.0 - 2.0 \text{ minutes})$
=2.3 moles/min

13. ***Not Being Able to Identify the Relevant Component of the Fire Triangle***

Example: Firemen dig a circular trench around a forest fire. Why?

Solution: This eliminates fuel. There is only dirt in the trench, which unlike grass and trees, is not flammable. Surface area, oxygen and heat are not factors in this case.

14. ***Not Understanding the Connection between Spontaneity and Rates.***

Example: When phosphorus is taken out of oil it spontaneously bursts into flames, forming P₂O₅. How and why does this happen? Show a graph with chemical equation.

Solution: It happens because a very small activation energy is needed to burn phosphorous. In other words, the substance formed in between phosphorus and its oxide almost has the same enthalpy as the original, so that phosphorus can absorb the small difference in energy from the colliding air molecules in its environment.

Equilibrium

15. ***Incorrect Use of IRFE(ICE) Chart in Calculating K By Ignoring Volume/Stoichiometry***

Example: Given: $2 \text{ NO}_{2(g)} = \text{N}_2\text{O}_{4(g)}$

At a certain temperature 6.0 moles of NO_2 and 1.0 moles of $\text{N}_2\text{O}_4(\text{g})$ are introduced into a 3.0 L container. At equilibrium we find only 4.0 moles of NO_2 along with some N_2O_4 . Calculate K.

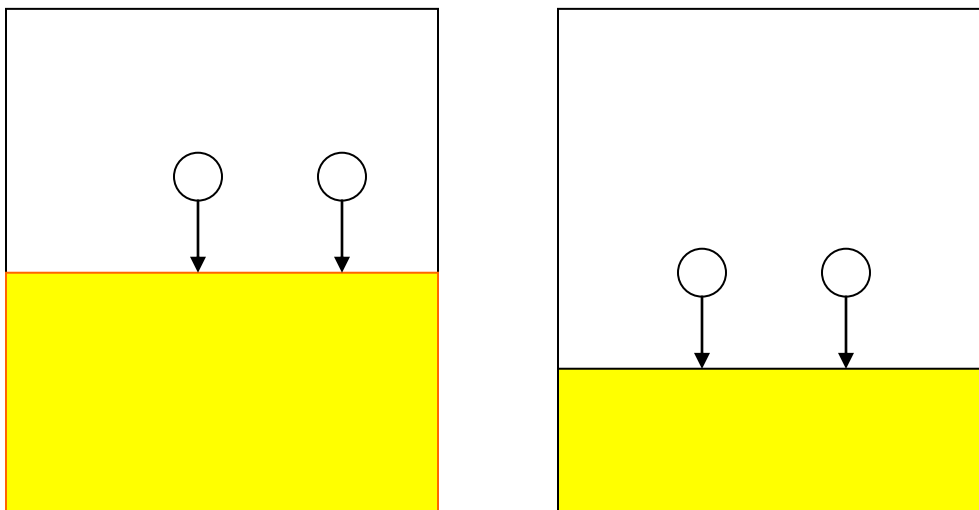
	$2 \text{NO}_2(\text{g}) =$	$\text{N}_2\text{O}_4(\text{g})$
initial	6.0 moles	1.0 moles
Changing (reacting/forming)	6.0 moles – 4.0 = 2.0 moles	2.0/2 = 1.0 moles
equilibrium	4.0 moles	1 + 1 = 2.0 moles

$$K = [\text{N}_2\text{O}_4(\text{g})]/[\text{NO}_2(\text{g})]^2 = (2.0/3.0 \text{ L})/(4.0/3.0 \text{ L})^2 = 0.38$$

16. *Assuming that solids and liquids disturb equilibrium*

In the same way that we exclude solids and liquids from K_{eq} expressions, adding more liquid or solid will not disturb equilibrium.

Look at the diagram below to understand why: The second one has less liquid but there are still two collisions going on (circles are aqueous ions whose concentration will affect equilibrium)



17. *Inability to Combine pH, pOH concepts with Ka and Kb*

Example: What is the Ka of an acid with a pOH of 10.0 if the acid was prepared by dissolving 30.0 g in 2.0 L? Molar mass of HX = 100.0.

Solution: $\text{pH} = 14.0 - \text{pOH} = 14.0 - 10.0 = 4.0$. This leads to the equilibrium concentration of acid:

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-4} \text{ moles/L}$$

We then run backwards to trace the amount of acid that dissociated in order to get its equilibrium value.

	HX	H ⁺	X ⁻
initial	30.0g/(100g/mole)/2.0 L = 0.15 M	0	0
Changing (reacting/forming)	10 ⁻⁴ moles/L	10 ⁻⁴ moles/L	10 ⁻⁴ moles/L
equilibrium	0.15-10 ⁻⁴ moles/L	10 ⁻⁴ moles/L	10 ⁻⁴ moles/L

$$K_a = \frac{[\text{H}^+][\text{X}^-]}{[\text{HX}]} = \frac{[10^{-4}][10^{-4}]}{[0.15-10^{-4}]} = 6.6 \times 10^{-8}$$

18. *Treating Standard Reduction Potentials like ΔH.*

Example: Find the overall voltage for an electrochemical cell consisting of lithium and gold electrodes in standard solutions.

Although the sign of an E value changes when we reverse the reaction, we do **not** multiply the voltage by the coefficient needed by the chemical reagents.

